Evaluation of Ground-State Electronic Energy from Fluorine Nuclear Magnetic Resonance Shielding. I. Prediction of Enthalpies of Reaction for Formation of Lewis Acid Adducts of Benzophenones[1]

C. S. Giam and R. W. Taft

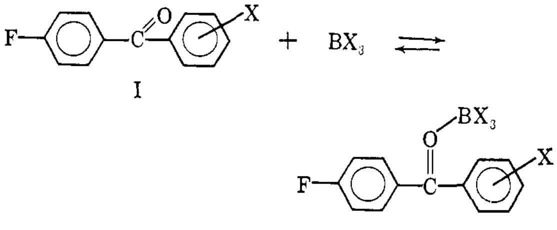
Contribution from the Department of Chemistry, University of California, Irvine, California 92640. Received December 3, 1966

Abstract: The previously proposed linear fluorine nuclear magnetic resonance shielding electronic energy (SEE) relationship for *p*-fluorophenyl derivatives is shown to provide satisfactory prediction of the standard enthalpy of reactions of *m*'- and *p*'-substituted *p*-fluorobenzophenones with the Lewis acids BCl_3 and BBr_3. This result indicates that *p-fluoro*phenyl shielding in appropriate systems provides to useful approximation the evaluation of the ground-state electronic energy change needed for *a priori* prediction of rather complex organic equilibria from spectroscopic data. Using Arnett's calorimetric procedure, standard enthalpies of reaction have been measured for the formation of a substantial series of the Lewis acid adducts in methylene chloride. The observed values, ranging from -11.5 to -20.6 kcal/mole, are in good agreement with the predicted values. The linear SEE relationship has been used to obtain a novel analysis of the contributions of substituent perturbations in reactant and product states to the observed thermodynamic substituent effects.

In a previous paper,[2] it was proposed that substituent effects on fluorine nuclear magnetic resonance (F nmr) shielding in related *p*-fluorophenyl derivatives are linear with the substituent effect on the corresponding standard electronic energy change. This relationship has a theoretical basis in linear relationships between conjugative substituent effects on: (a) *p*-fluorophenyl nmr shielding, (b) the LCAO-MO theory π-charge density at the *para* carbon atom (bonded to F), (c) the molecular ionization potential.[2]

The proposed linear shielding electronic energy (SEE) relationship has been confirmed experimentally for the very large substituent effects involved in the oxidation of hexarylethanes to 4,4''-disubstituted *p*-fluorotriphenylmethyl (*p*-fluorotrityl) cations[2] and in their reduction to 3'- and 4'-substituted *p*-fluorotrityl anions.[3] The substituent effects on the standard free energy change of these processes are taken to be close estimates of the corresponding electronic energy effects.[2,3] The mean value of the energy-shielding coefficient obtained from these two reaction series is 1.0 kcal/mole ppm,[2,3] in accord with the order of magnitude value anticipated by the theory.[2]

The formation of complexes with strong Lewis acids, *e.g.*, BCl_3, is a reaction which has been investigated in detail.[4] The standard enthalpy of adduct formation apparently may be taken as a satisfactory estimator of the standard electronic energy change.[4] In reaction 1, the formation of BCl_3 and BBr_3 adducts of 3'- and 4'-substituted *p*-fluorobenzophenones (series I ketones) bears a close analogy to the formation of the above substituted *p*-fluorotrityl cations.



(1)

In the companion paper[5] the shielding parameters for series I ketones and their Lewis acid adduct in methylene chloride solution (relative to internal fluorobenzene) are reported. In this paper we report the determination of the standard enthalpies of adduct formation of series I ketones in methylene chloride solution using the calorimetric procedure recently described by Arnett.[6] The predicted values of the enthalpies of complex formation obtained from the linear SEE relationship are found to be in quite satisfactory agreement with the experimental values. A novel analysis of the origin of the substituent effects on the standard enthalpies of reaction 1 has been obtained from the shielding parameters for series I ketones and their Lewis acid adducts through the use of linear SEE relationship.

Experimental Section

Materials. *Cf.* ref 5.

Procedure. The calorimeter used was essentially that described by Arnett, *et al.*,[6] and was purchased from the Guild Corp., Bethel Park, Pa. A Leeds and Northrup Model 8687 potentiometer and a Model 153 Honeywell 1-mv recorder were utilized.

The calorimeter consisted of a 300-ml dewar flask (4.5 in. deep and 2.5 in. i.d.). It was fitted tightly with a Teflon stopper which held in place in the dewar a 5000-ohm thermistor and a 100-ohm electric heating coil. Additional holes in the stopper were used to accommodate syringe ports and a sealed-glass tubing stirrer with a Teflon blade. The stirrer was driven by a constant speed motor. Plastic insulating tape was wrapped around the seal between the Teflon stopper and the dewar to reduce any heat leakage.

---

(1) This work was supported in part by the National Science Foundation. A preliminary presentation of this work was made in Abstracts of the 49th Canadian Chemical Conference, Saskatoon, June 7, 1966, p 62.

(2) R. W. Taft and L. D. McKeever, *J. Am. Chem. Soc.*, **87**, 2489 (1965).

(3) L. D. McKeever and R. W. Taft, *ibid.*, **88**, 4544 (1966).

(4) F. G. A. Stone, *Chem. Rev.*, **58**, 101 (1958); H. C. Brown, D. H. McDaniel and O. Hafliger in "Determination of Organic Structure by Physical Methods," E. A. Braude and F. C. Nachod, Ed., Academic Press Inc., New York, N. Y., 1956, p 567.

(5) R. G. Pews, Y. Tsuno, and R. W. Taft, *J. Am. Chem. Soc.*, **89**, 2391 (1967).

(6) E. M. Arnett, W. G. Bentrude, J. J. Burke, and P. McDuggleby, *ibid.*, **87**, 1541 (1965).

Methylene chloride (210 ml, dried by distilling from CaH_2) was placed in the dewar. In order to minimize hydrolysis of BCl_3 or BBr_3, the dewar was first flushed with dry nitrogen followed by the boron halide and then dry nitrogen. A known volume of BCl_3 was measured with a gas buret and introduced in the gas washing bottle containing the dry methylene chloride. The resulting solution was then transferred under N_2 to the dewar. Two or more syringes containing the samples for heat measurements were firmly wedged in the holes in the Teflon stopper so that the injection ends of the syringes were just below the liquid in the dewar.

When equilibrium conditions were established (25 ± 1°), heating curves were taken. Typically, using a heating period of 15 sec with the heater current at 0.1 amp, ~1-mv, deflection on the recorder was obtained. When two or more successive heating curves did not vary more than 2–3%, samples were introduced for determination of heats of solution or of reaction. Following sample injection, heating curves were again repeated which usually did not vary more than 5% from the initial results.

Reaction heats were determined by successive injections of base into methylene chloride solutions containing on the order of 100-fold excess of the acid. The initial acid concentration was 0.1–0.3 *M*. The base was injected either as pure liquid or solid or a methylene chloride solution. In the former cases, the reaction heats were obtained by correcting for the heat of solution of the base in methylene chloride as measured in separate experiments. Liquid samples were injected by hypodermic-type syringes. Solid samples were placed in polypropylene syringes. The injection ends of these syringes were sliced off and replaced by Teflon plugs in order to obtain a wide opening through which the solid could be forced. In injection of solutions, typically 0.2–1.0 ml of a 0.1–0.3 *M* stock solution of base in methylene chloride was used. The reaction heat obtained in this manner was corrected, if necessary, for the small heat of dilution of the acid solution. After the first or second injection of base into the acid solution, it was found that heats of acceptable precision and reproducibility were obtained (*cf.* Table I). The result of the first or second injection was generally erratic and was not considered in obtaining the mean value. The $\Delta H°_{f(exptl)}$ values of Table I as the mean of a least two (generally 3–5) acceptable values.

Results

In Table I are listed the standard enthalpies (and their precision measures) of complex formation, $\Delta H°_f$, obtained for series I ketones with BCl_3 and BBr_3 in methylene chloride solution. Values are also listed for the corresponding change in the F nmr shielding parameter, Δ, between complexed and uncomplexed ketone, *i.e.*, $\Delta = \int_{H(c)}{}^{p\text{-}X} - \int_{H(u)}{}^{p\text{-}X} = \int_{H}{}^{p\text{-}X_0} + \Delta\int_{(c)} - \Delta\int_{(u)}$, where the latter shieldings are as designated and listed in Tables I–III of ref 5. The predicted values of $\Delta H°_f$ obtained from the linear SEE relationship (eq 7, below) are given also. In general the agreement between calculated and observed values of $\Delta H°_f$ is within the combined uncertainties. The only exception occurs for the enthalpy figure for the 3′-nitro-*p*-fluorobenzophenone–BBr_3 adduct. With BBr_3 the precision of $\Delta H°_f$ was generally poorer than for BCl_3, but the deviation in this case is beyond experimental error. In view of the generally satisfactory agreement in Table I, the only suggestion which we can offer for the larger observed than predicted enthalpy change is the possibility that there is a contribution to $-\Delta H°_{f(obsd)}$ resulting from interaction of BBr_3 with the nitro group in the adduct which occurs at the very high acid–base ratio (>100) employed in the enthalpy determination.

Table I. Experimental and Predicted Values of the Standard Enthalpy of Formation (Reaction 1) of BCl_3 and BBr_3 Adducts or *m*′- and *p*′-Substituted *p*-Fluorobenzophenones in Methylene Chloride Solution at 25°

	$-\Delta H°_{f(exptl)}$, kcal	$-\Delta$, ppm	$-\Delta H°_{f(calcd)}$,[a] kcal
	BCl_3 Adducts		
p-OCH_3	19.0 ± 0.1	10.40	19.3
p-OC_6H_5	18.4 ± 0.1	11.30	18.4
p-*t*-C_4H_9	16.7 ± 0.2	13.07	16.6
H	15.1 ± 0.2	14.51	15.2
p-Cl	14.5 ± 0.2	15.04	14.7
m-F	13.6 ± 0.2	16.20	13.5
m-Br	13.5 ± 0.1	16.42	13.3
m-CF_3	12.6 ± 0.2	17.04	12.7
p-CF_3	12.3 ± 0.2	17.44	12.3
m-NO_2	11.5 ± 0.3	18.05	11.7
	BBr_3 Adducts		
p-OCH_3	...	11.74	21.3
p-OC_6H_5	20.6 ± 0.3	12.56[b]	20.5
p-*t*-C_4H_9	18.8 ± 0.2	14.50	18.6
H	17.2 ± 0.3	16.16	16.9
p-Cl	17.1 ± 0.3	16.49	16.6
m-F	...	17.66	15.4
m-Br	...	17.85	15.2
m-CF_3	14.9 ± 0.4	18.66	14.4
p-CF_3	...	19.04	14.0
m-NO_2	14.8 ± 0.1	19.98	13.1

[a] Calculated from eq 7. [b] Calculated from eq 9 of ref 5 using $\sigma^+ = -0.49$.

Discussion

The proposed linear SEE relationship is expressed by

$$\int_{H}^{p\text{-}X} = mE° + b' \qquad (2)$$

where $\int_H{}^{p\text{-}X}$ is the shielding parameter for the *p*-fluorophenyl derivative relative to internal fluorobenzene, $E°$ is the molecular electronic stabilization energy (*cf.* subsequent Discussion) and *m* and *b* are constants characteristic of the related series of *p*-fluorophenyl derivatives. For reaction 1 application of eq 2 gives

$$\Delta E° = E°_{(c)} - E°_{(u)} = m\int_{H(c)}^{p\text{-}X} - m\int_{H(u)}^{p\text{-}X} + b = m\Delta + b \qquad (3)$$

where Δ is the difference in the shielding parameter between complexed and uncomplexed ketone. It is assumed for present purposes that for reaction 1 $\Delta E° \cong \Delta H°_f$ (298°K).

In the present application of eq 3 two types of structural variations are being considered, *i.e.*, variable Lewis acid and variable *m*′ and *p*′ substituents. With *p*-fluorobenzonitrile it was found[7] (qualitatively) that *m* has positive sign, *i.e.*, increasing $-\Delta H°_f$ is accompanied by increasing $-\Delta$. A similar sign of *m* is expected for the Lewis acid adducts of *p*-fluorobenzophenone. On the other hand, with increased electron withdrawal by the *m*′ or *p*′ substituent, $-\Delta$ increases but $-\Delta H°_f$ decreases (as expected); *i.e.*, *m* has a negative sign for the substituent effect.

Thus the following two relationships are generated from eq 3: for various Lewis acid adducts of *p*-fluorobenzophenone

$$\Delta H°_f = \cong \Delta E_0° = m\Delta_0 + b \qquad (4)$$

and for the substituent effect (relative to unsubstituted *p*-fluorobenzophenone)

$$\int_R \Delta H°_f \cong \int_R \Delta E° = -m\int_R \Delta = m(\Delta_0 - \Delta) \qquad (5)$$

---

[7] R. W. Taft and J. W. Carten, *J. Am. Chem. Soc.*, **86**, 4199(1964).

where Δ refers to the substituted fluorobenzophenone and Δ_0 to unsubstituted *p*-fluorobenzophenone. Combining eq 4 and 5 gives

$$\Delta H^\circ_f = \Delta E_0^\circ + \int_R \Delta E^\circ = m(2\Delta_0 - \Delta) + b \quad (6)$$

on the assumption that the magnitude of m is the same (1.00 kcal/ppm; *cf.* introductory section) for both types of structural dependence.

In order to obtain an independent evaluation of the constant b in eq 3 we have determined Δ and ΔH°_f for reaction 1 for the formation of the BCl_3 adduct of *p*-fluoroacetophenone. We believe the value of b obtained with the use of the latter ketone will be essentially correct for use with *p*-fluorobenzophenone adducts since acetophonone and benzophenone are of essentially identical base strength.[8] Values of $\Delta = -18.20 \pm 0.06$ ppm and $-\Delta H^\circ_f = 18.9 \pm 0.2$ kcal/mole have been obtained with *p*-fluoroacetophenone from which $b = -0.7$ kcal/mole. Equation 6 may be reduced then to the following simplified explicit form for reaction 1

$$\Delta H^\circ_f = 2\Delta_0 - \Delta - 0.7 \text{ (in kcal/mole)} \quad (7)$$

The success of eq 7 is displayed in Table I in the comparison between observed and predicted values of $-\Delta H^\circ_f$ for reaction 1. It is to be emphasized that the predictions of eq 7 *depend in no way upon the correlations* made in the companion paper.[5] Further, the *correlation* of Δ values with σ^0 is limited to substituents which involve no direct conjugative effects[5] whereas the confirmed *predictions* of eq 7 include both strongly conjugating and nonconjugating substituents.

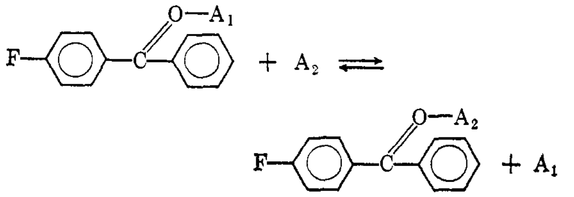
Equation 7 and similar relationships derived from the linear SEE relationship promise to be of great utility in obtaining values of ΔH°_f which are difficult or impossible to obtain by direct experimental determination. In our hands, Δ has been found generally to be a more precisely and easily measured quantity than ΔH°_f.

Equation 7 may be combined with the polar substituent effect relationship,[5] $\Delta = \Delta_0[1 + (\sigma^0/2.75)]$, to obtain the (limited) relationship

$$-\Delta H^\circ_f = \Delta_0[\sigma^0/2.75] + 0.7 \quad (8)$$

Equation 8 may be used (where applicable) to calculate ΔH°_f values for any of the Lewis acid adducts for the series I ketones for which Δ_0 (*cf.* Table I, ref 5 for acids, *e.g.*, $AlCl(C_2H_5)_2$, BI_3, etc.) and σ^0 values are available. While these uses of F nmr shielding for predictions of reactivities must be regarded as having substantial consequences, the greatest significance of the linear SEE relationship is probably in the unique extrathermodynamic analysis of substituent effects on reactivities which it permits.

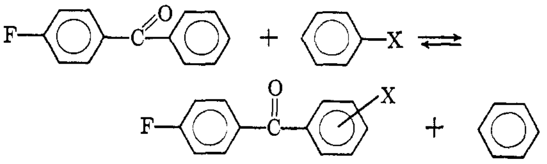
The electronic stabilization energy which is measured according to eq 2 by the F nmr shielding parameter is extrathermodynamic.[9] That is, in terms of the theoretical models used to derive eq 2 the perturbation on the π electronic energy levels of *p*-fluorobenzophenone is measured *but the energy content of the substituent is not included.* The relevant changes in E° may be approximated (for conceptual purposes) by the following model thermodynamic exchange reactions:


(9)

for variable Lewis acid

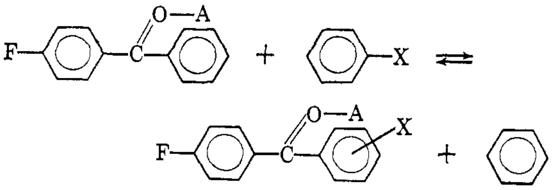
$$\Delta E^\circ_{(9)} \cong E^\circ_{A_2} - E^\circ_{A_1} = \int_R E^\circ$$

for variable substituent in base


(10)

$$\Delta E^\circ_{(10)} \cong E^\circ_{X(u)} - E^\circ_{H(u)} = \int_R E^\circ_{(u)} \equiv -(SE)_{(u)} = -\Delta\int_{(u)}$$

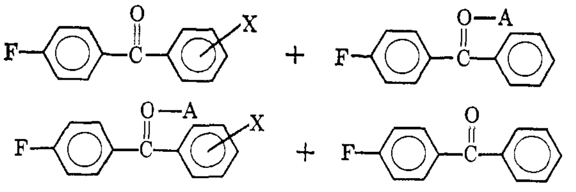
for variable substituent in adduct


(11)

$$\Delta E^\circ_{(11)} \cong E^\circ_{X(c)} - E^\circ_{H(c)} = \int_R E^\circ_{(c)} \equiv -(SE)_{(c)} = -\Delta\int_{(c)}$$

Although values of $\Delta E^\circ_{(9)}$, $\Delta E^\circ_{(10)}$, and $\Delta E^\circ_{(11)}$ represent fundamental quantities, there is little immediate prospect that precise values of these quantities will become available by thermodynamic means.

The thermodynamic substituent effects measured in this investigation, $\int_R \Delta H^\circ_f$, correspond to the following reaction.


(12)

$$\int_R \Delta H^\circ_{f(1)} = \Delta H^\circ_{f(12)} \cong \Delta E^\circ_{(12)} = (SE)_u - (SE)_c = \Delta E^\circ_{(11)} - \Delta E^\circ_{(10)}$$

Values of the substituent stabilization energies (as defined above) for uncomplexed and complexed ketone obtained from the linear SEE relationship therefore permit a novel analysis of the origin of the thermodynamic substituent effects. Specifically, the relative contributions of the substituent perturbations on the reactant and product states of reaction 1 is obtained.

---

[8] E. M. Arnett, *Progr. Phys. Org. Chem.*, 1, 378 (1963).

[9] J. E. Leffler and E. Grunwald, "Rates and Equilibria of Organic Reactions," John Wiley and Sons, Inc., New York N. Y., 1963, Chapter 6.

The results of this analysis for formation of the BCl_3 adducts are given in Table II. Similar analysis may be made with other data from Table I and Tables II and III of ref 5.

Table II. Analysis of the Contributions of Substituent Effects on the Stabilization Energy of *m*′- and *p*′-Substituted *p*-Fluorobenzophenones and Their BCl_3 Adducts to the Thermodynamic Substituent Effects (kcal/mole)

Subst	Reactant state $SE_{(u)} = -\Delta\int_{(u)}$	Product state $SE_{(c)} = -\Delta\int_{(c)}$	State change $-\int_R\Delta E° = \int_R\Delta$	State change measured $= -\int_R\Delta H°_{f(obsd)}$ ($\pm$0.2–0.3)
p′-OCH_3	+0.99	+4.90	+3.91	+3.9
p′-OC_6H_5	0.45	+3.66	+3.21	+3.3
p′-*t*-C_4H_9	0.41	+1.85	+1.44	+1.6
H	(0.00)	(0.00)	(0.00)	(0.0)
p′-Cl	−0.52	−1.05	−0.53	−0.6
m′-F	−0.75	−2.44	−1.69	−1.5
m′-Br	−0.89	−2.80	−1.91	−1.6
m′-CF_3	−1.17	−3.70	−2.53	−2.5
p′-CF_3	−1.32	−4.25	−2.93	−2.8
m′-NO_2	−1.79	−5.33	−3.54	−3.6

The identification of the substituent stabilization energies for uncomplexed and complexed ketone with the quantities $\Delta\int_{(u)}$ and $\Delta\int_{(c)}$, of course, does involve the assumption that $m = 1.00$ kcal/mole ppm as found for the $\Delta E°_{(12)}$ (Δ) process. This assumption is supported by the well-behaved $\sigma^0\bar{\rho}$ behavior displayed by both the substituted uncomplexed and complexed ketones[5] and the fact that the range in Δ values involved (*cf.* Table I) of about 10 ppm is of similar order of magnitude as the range of $\int_H^{p\text{-}X}$ values between the uncomplexed *p*′-OCH_3 compound and the BBr_3 complexed *m*′-NO_2 compound (about 22 ppm). In view of this and subsequent discussion it is highly unlikely that any errors due to this assumption exceed 20% of the result listed in Table II and smaller error seems probable.

It is clear from the results in Table II that the thermodynamic substituent effects are the resultants of substantial effects in both complexed and uncomplexed ketone, although effects in the former state do predominate. While the substituent effects are qualitatively parallel in both states, there is no fixed factor which quantitatively relates the effects of either state to the observed thermodynamic properties. However, for those substituents which exert only a polar (σ^0) effect, fixed factors are applicable. These are obtained from the ratios of the $\bar{\rho}$ values given in ref 5.

It is apparent from the above discussion and results that the linear SEE relationship provides also the prediction of the reaction constant, ρ for reaction 1. That is, for the relationship $\int_R\Delta H°_f = \sigma^0\rho$ (for polar effects) the predicted values of ρ (in kcal/mole $-\sigma^0$ units) are obtained as $\rho = -\Delta_0/2.75$, where $\Delta_0 = \int_{H(c)}^{p\text{-}X_0} - \int_{H(u)}^{p\text{-}X_0}$ (values of these parameters are given in Table I of ref 5 for the acids $AlCl(C_2H_5)_2$, BF_3, BCl_3, $AlCl_2C_2H_5$, $AlCl_3$, BBr_3, and BI_3).

Limitations of the Linear SEE Relationship. LCAO–MO theory *does not* indicate completely general linear relationships between local charge density and electronic energy.[10] Consequently the SEE relationship, which makes a specialized use of such linearity,[2] is anticipated *to have structural limitations.* These limitations are not well defined by present theory and, in any case, must be subjected to experimental determination.

The results of this and previous investigations,[2,3] leave little doubt that the SEE relationships are satisfactorily linear for *p*-fluorophenyl substitution at a fixed functional group (within the usual context of such usage). It is not necessary, however, that the energy shielding coefficient, *m*, be a universal constant of precisely 1.00 kcal/mole ppm.

In the formation of trityl cations from their carbinols[11] $m = 0.70$ kcal/mole ppm, while in the ethane reduction to trityl anions[3,12] $m = 1.30$ kcal/mole ppm. These results suggest that *m* increases (in magnitude) as the *p*-fluorophenyl label is substituted at substantially more "saturated" (less sensitive) functional centers. Unpublished results of Dr. M. G. Schwartz[13] and Dr. D. Gurka[14] with adducts of *p*-fluorophenol and of Mr. A. A. Grey[14] with adducts of *p*-fluorophenylboron dichloride definitely confirm this trend.

In the application of eq 4 to the present study nearly equivalent results to those listed in Table I are obtained with $m = 1.05$ kcal/mole ppm and $b = 0$ whereas in the application of eq 5 $m = 1.00$ or slightly smaller. However, the differences in the alternate treatments are not significant. Consequently, as a working hypothesis, *m* is taken to be a constant for each broadly defined functional group, (*e.g.*, the carbonyl group in series I ketones and their Lewis acid adducts).

Thermodynamic Considerations. The assumed approximation $\Delta H°_f \cong \Delta E°_{elect}$ more accurately must be replaced by the relationship

$$\Delta H°_f = T\Delta[(H° - E_0°)/T] + \Delta E_0°$$

where $\Delta[(H° - E_0°)/T]$ represents the sum of the thermodynamic function for the products of reaction 1 minus that for the reactants, and $\Delta E_0° = \Delta E°_{elect} + \Delta E°_{vibr}$, where $\Delta E°_{vibr}$ is the zero point vibrational energy change for reaction 1.

Since the thermodynamic functions and $\Delta E°_{vibr}$ can be evaluated from vibrational and rotational spectroscopic data, it follows that the linear SEE relationship provides in principle a means of precise prediction of $\Delta H°_f$ from the various spectroscopic measurements and statistical thermodynamic relationships. A similar conclusion applies with respect to the *a priori* prediction of the equilibrium constant through the relationship

$$\Delta F°_f = RT \ln K = T\Delta[(F° - E_0°)/T] + \Delta E°_{elect} + \Delta E°_{vibr}$$

In practice, the application of this precise relationship must await the evaluation of the precision of

---

[10] C. A. Coulson, *Proc. Phys. Soc.* (London), **A65**, 933 (1952); L. Goodman and H. Shull, *J. Chem. Phys.*, **23**, 33 (1955).

[11] For shielding in substituted trityl cations *cf.* ref 2. Substituent shielding effects in the corresponding carbinols have been obtained: L. D. McKeever, Ph.D. Thesis, University of California, Irvine, Calif., Sept 1966. The latter shieldings parallel those obtained in the trityl cations but are only 3% as large.

[12] *Cf.* ref 3. The substituent shielding effects in the ethanes have been approximated as those observed in the corresponding triarylmethanes. The latter shielding effects are approximately 7% as large as for corresponding shieldings in the trityl anions.

[13] M. G. Schwartz, Ph.D. Thesis, The Pennsylvania State University, Dec 1965.

[14] Work in progress at University of California, Irvine, Calif.

$\Delta E^\circ_{\mathrm{elect}}$ obtained by the linear SEE relationship. In the meanwhile, the linear SEE relationship appears to provide an extremely valuable approximate relationship as discussed above.

Acknowledgment. We are pleased to acknowledge the helpful assistance of Professor Arnett, Dr. C. Douty, and Mr. T. Murty in obtaining preliminary results with their calorimeter.

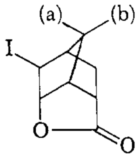
Analysis of the Nuclear Magnetic Resonance Spectra of Some 2,6-Bridged Bicyclo[2.2.1]heptane Derivatives

Kermit C. Ramey,[1a] David C. Lini,[1a] Robert M. Moriarty,[1b] H. Gopal,[1b] and Harold G. Welsh[1b]

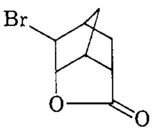
Contribution from the Research and Development Department of the ARCO Chemical Company, a Division of Atlantic Richfield Company, Glenolden, Pennsylvania, and the Department of Chemistry, The Catholic University of America, Washington, D. C. *Received December 15, 1966*

Abstract: The 60- and 100-Mc/sec nuclear magnetic resonance spectra of several bridged bicyclo[2.2.1]heptane derivatives have been analyzed in detail. These compounds possess the common structural feature of a 2,6 oxygenated bridge which may be either a lactone as in 5-*exo*-iodo-6-*endo*-hydroxybicyclo[2.2.1]heptane-2-*endo*-carboxylic acid lactone (**1**) and the 5-*exo*-bromo (**2**), 5-*exo*-acetoxy (**3**), 5-*exo*-tosyloxy (**4**), 2-*exo*-methyl-5-*exo*-iodo (**5**), 2-*exo*-methyl-5-*exo*-bromo (**6**) derivatives; or the 2,6 bridge may be an oxido unit as in 4-*exo*-tosyloxy-6-oxatricyclo[3.2.1.$1^{3,8}$]nonane (**7**) and 4-*exo*-acetoxy-6-oxatricyclo[3.2.1.$1^{3,8}$]nonane (**8**). Chemical shifts for all the protons in these structures have been assigned, and the geminal and vicinal couplings measured. Aromatic solvent shifts observed for compounds **1**, **5**, and **7** are discussed in terms of solvent–solute collisional complexes of defined stereochemistry. Long-range couplings for the proton pairs: 1–4, 3-*endo*–7(a), 5-*endo*–7(b), 2-*exo*–6-*exo*, 1–3-*exo*, 2-*exo*–4, and 6-*exo*–4 in compound **1** were observed and most of these were confirmed using spin-decoupling techniques. For the C_5-*endo* proton doublet in the lactone derivatives it is noteworthy that the principal coupling is with the $C_{7(b)}$ proton; this amounts to about 2.5 cps while the C_5-*endo*–C_6-*exo* vicinal coupling is negligibly small (0.3 cps) and the C_5–*endo*–C_4-vicinal ranges from 0.5 to 1.0 cps. The nmr spectra of the 5-keto and 7-keto derivatives in the 2,6 lactone series are discussed in relationship to the changes in chemical shifts relative to the precursor secondary alcohols. The alcohol–ketone pairs are 5-*exo*,6-*endo*-dihydroxybicyclo[2.2.1]heptane-2-*endo*-carboxylic acid lactone (**9**) and 5-keto,6-*endo*-hydroxybicyclo[2.2.1]heptane 2-*endo*-carboxylic acid lactone (**10**); 6-*endo*,7(b)-dihydroxybicyclo[2.2.1]-heptane-2-*endo* carboxylic acid lactone (**11**) and 7-keto,6-*endo*-hydroxybicyclo[2.2.1]heptane-2-*endo*-carboxylic acid lactone (**12**).

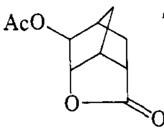
The norbornyl system has served as a substrate for the generation and evaluation of numerous mechanistic hypotheses in modern organic chemistry. Mechanistic conclusions originating from studies in the norbornyl series frequently have been based upon the structures of rearranged products. Detailed nuclear magnetic resonance spectral analyses in this series obviously are of importance in facilitating the elucidation of rearrangement products. Moreover, due to the conformational rigidity of these systems, long-range couplings which are often of an unexpectedly large magnitude may be detected. This paper presents detailed analyses of the spectra of compounds **1–8**. The compounds included in this study are of current and particular interest because of uncertainties of interpretation[2a] and previous erroneous assignments.[2b]



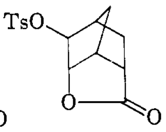
1



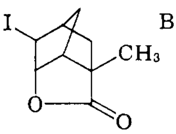
2



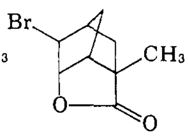
3



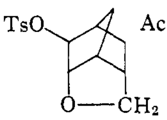
4



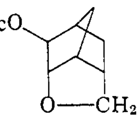
5



6



7



8

Chemical Shifts

In a preliminary communication of a portion of this work,[3] we pointed out that the chemical shifts of the C_1- and C_2-*exo* protons in compound **1** were anomalous in the sense that the C_1 proton appeared at a lower field position relative to the C_2-*exo* proton which is attached to the carbon atom bearing the carbonyl group of the lactone. This assignment was required to explain the magnitude of the couplings associated with the C_6-*exo* proton. An earlier interpretation of the spectrum of **2** used the reverse assignment of the chemical shift of the C_1- and C_2-*exo* protons.[2b] This incorrect assignment was used recently by Jensen and Miller[2a] who corrected stereochemical assignments of Traylor and Factor for the structure of the oxymercuration product derived from 5-norbornene-2-*endo*-car-

---

(1) (a) The ARCO Chemical Co., Glenolden, Pa.; (b) The Catholic University of America, Washington, D. C.

(2) (a) F. R. Jensen and J. J. Miller, *Tetrahedron Letters*, 40, 4861 (1966); (b) E. Crundwell and W. Templeton, *J. Chem. Soc.*, 1400 (1964).

(3) R. M. Moriarty, H. Gopal, H. G. Walsh, K. C. Ramey, and D. C. Lini, *Tetrahedron Letters*, 38, 4555 (1966).

Ramey, et al. / Nmr Analysis of Bicyclo[2.2.1]heptane Derivatives