to that of furan.[3,11] From this it could be inferred that the $C_4H_4O\cdot^+$ ions have the furan structure I.

Ions of formula $C_4H_4O\cdot^+$ in the spectra of I, II, and III decompose to yield two "metastable peaks" corresponding to the transitions

$$C_4H_4O\cdot^+ \xrightarrow{m^*_1} C_3H_4\cdot^+ + CO$$

$$C_4H_4O\cdot^+ \xrightarrow{m^*_2} C_2H_2O\cdot^+ + C_2H_2$$

Abundance ratios m^*_1/m^*_2 (Table I) of approximately 8.4, 3.7, and 2.3 are found for the $C_4H_4O\cdot^+$ metastables from furan, 2-pyrone, and 4-pyrone, respectively.

Table I

Compound	$[m_1^*]/[C_4H_4O\cdot^+]$	$[m_2^*]/[C_4H_4O\cdot^+]$
Furan	$(2.0 \pm 0.1) \times 10^{-3}$	$(2.4 \pm 0.2) \times 10^{-4}$
2-Pyrone	$(1.7 \pm 0.1) \times 10^{-2}$	$(4.6 \pm 0.2) \times 10^{-3}$
4-Pyrone	$(1.3 \times 0.1) \times 10^{-2}$	$(5.7 \pm 0.2) \times 10^{-3}$

Following arguments presented earlier,[6] these metastable ion characteristics indicate that a substantial part of the $C_4H_4O\cdot^+$ ions from the 2- and 4-pyrones must *not* correspond in structure to the $C_4H_4O\cdot^+$ ions from furan.[12]

Similarly coumarin (IV) on electron impact loses CO to yield abundant $C_8H_6O\cdot^+$ ions for which the benzofuran structure V has been inferred.[2b] Ions of formula $C_8H_6O\cdot^+$ in the spectra of IV and V decompose to give metastable peaks corresponding to the transitions

$$C_8H_6O\cdot^+ \xrightarrow{m^*_3} C_7H_6\cdot^+ + CO$$

$$C_8H_6O\cdot^+ \xrightarrow{m^*_4} C_6H_4O\cdot^+ + C_2H_2$$

Abundance ratios $[m^*_3]/[m^*_4]$ (Table II) of approximately 28 and 164 are found for the $C_8H_6O\cdot^+$ metastables from coumarin and benzofuran, respectively. These metastable ion characteristics indicate that a substantial part of the $C_8H_6O\cdot^+$ ions in the mass spectrum of coumarin must not correspond in structure to those of benzofuran.[12]

Table II

Compound	$[m^*_3]/[C_8H_6O\cdot^+]$	$[m^*_4]/[C_8H_6O\cdot^+]$
Coumarin	$(9.7 \pm 0.2) \times 10^{-3}$	$(3.4 \pm 0.2) \times 10^{-4}$
Benzofuran	$(1.4 \pm 0.1) \times 10^{-2}$	$(8.5 \pm 0.2) \times 10^{-5}$

These results raise questions concerning the structures of other $(M - CO)\cdot^+$ ions in which formation by expulsion of CO has been assumed to be accompanied by ring closure.[2] We are presently extending our studies to such systems.

Metastable ion abundances were determined in the first drift region of a CEC 21-110 double-focusing mass spectrometer using the defocusing technique.[13]

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(14) Postdoctoral Fellow, 1966–1967.

W. T. Pike,[14] F. W. McLafferty
Department of Chemistry, Purdue University
Lafayette, Indiana 47907

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Hydrogen Bonding. I. Two Approaches to Accurate Heats of Formation[1]

Sir:

Despite the crucial role of hydrogen bonding in chemical and biochemical theory,[2] relatively few determinations of heats of hydrogen-bond formation (ΔH_f) have been described previously for well-characterized systems. Most of these have employed a temperature variation of equilibrium constant, a procedure which is usually less reliable than calorimetry.[3] Hydrogen-bond energies (ΔG_f and ΔH_f) commonly fall in the 3–7-kcal/mole range.[2] Cited error limits are often close to ±1 kcal/mole, and discrepancies between published results may exceed 100%.

In the hope of developing a definitive collection of reliable values suitable for testing or refuting theories, we have pooled our resources[4] to study well-characterized hydrogen-bonding systems by several independent methods (infrared, ^{19}F nmr, and solution calorimetry).

Described below are direct calorimetric measurements of molar enthalpies of formation (ΔH_f°) for 1:1 hydrogen-bonded complexes between phenol and *p*-fluorophenol (PFP) with three typical bases in carbon tetrachloride solution. ΔH_f° is defined as the standard enthalpy change for the reaction $A(CCl_4) + B(CCl_4) \rightarrow C(CCl_4)$, where A is the hydrogen-bonding acid, B is the acceptor base, and C is their 1:1 adduct. These results and those in the accompanying communication[4] will be extended to other bases and correlated with related thermodynamic and spectral properties in a complete article which will also contain experimental details and documentation of results. Here we will demonstrate two independent calorimetric methods for determining ΔH_f which give mutually consistent values. We will also show that the ΔH_f for a given hydrogen-bonded complex is rather insensitive to the medium in

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(3) (a) E. J. King, "Acid–Base Equilibria," The Macmillan Co., New York, N. Y., 1965, p 192 ff; (b) J. W. Larson and L. Hepler in "Solvent–Solute Interactions," Interscience Publishers, Inc., New York, N. Y., in press.

(4) D. Gurka, R. W. Taft, L. Joris, and P. von R. Schleyer, *J. Am. Chem. Soc.*, **89**, 5957 (1967).

Table I. *p*-Fluorophenol (PFP) and Phenol as Proton Donors

Base	Equilibrium constant, K^{298}, l. mole^{-1}		Heat of hydrogen bonding, kcal/mole		$\Delta\bar{H}_s$ of solute in pure base, kcal/mole	
	a	*b*	$-\Delta H_f°$ *c*	$-\Delta H_f^B$ *d*	PFP	PFA
	p-Fluorophenol as Proton Donor					
Pyridine	75.0 ± 1.5	74.0 ± 5.0	7.2 ± 0.2	7.40 ± 0.09	−1.01 ± 0.06	+0.06 ± 0.01
THF	17.7 ± 0.5	19.4 ± 1.0	6.0 ± 0.3	5.75 ± 0.08	−0.17 ± 0.01	−0.75 ± 0.01
DMF	116.0 ± 3.0	122 ± 9	6.7 ± 0.2	6.97 ± 0.11	−1.13 ± 0.08	−0.49 ± 0.02
CCl_4	...	...	...	...	+7.00 ± 0.07	+0.67 ± 0.02
	Phenol as Proton Donor					
					Phenol	Anisole
Pyridine	49.7 ± 1.0	48.0 ± 2.0	7.2 ± 0.2	7.34 ± 0.10	−1.41 ± 0.07	+0.03 ± 0.01
THF	13.3 ± 0.4	13.6 ± 0.4	5.7 ± 0.3	5.75 ± 0.08	−0.36 ± 0.10	−0.51 ± 0.02
DMF	75.5 ± 2.0	70.0 ± 6.0	6.3 ± 0.2	6.86 ± 0.08	−1.06 ± 0.02	−0.10 ± 0.03
CCl_4	...	...	...	...	+6.27 ± 0.07	+0.37 ± 0.02

[a] From infrared measurements of L. J. and P. von R. S. [b] Estimated fom ΔH_f in pure base and eq 1 and 2. [c] In dilute CCl_4 at 25° solution by method I. [d] Value in pure base by method II. Errors are standard deviations.

which it is determined, so that many donor–acceptor interactions have a separable, relatively invariant enthalpy.

Infrared spectra showed that PFP is not associated at concentrations below 0.02 *M* in carbon tetrachloride. Furthermore, Figure 1 shows an absence of calorimetrically significant solute–solute interactions for PFP across our working concentration range.

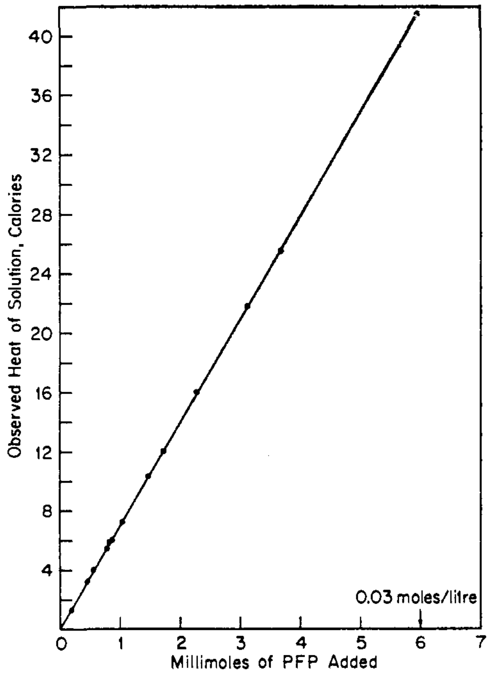

Figure 1. Demonstration, by constancy of $\Delta\bar{H}_s$ (*i.e.*, $\partial H_s/\partial n$), that PFP is not associated at concentrations below 0.03 *M*. Least-squares treatment gives −0.0071 ± 0.05 cal for the intercept and ±0.094 cal for standard deviation of points from the line.

Method I. When B is added to a solution of A (*e.g.*, PFP) to produce C, the observed heat of reaction, ΔH_{obsd}, is related to $\Delta H_f°$ by the expression[5]

$$\Delta H_{obsd} = \Delta H_f° cv \qquad (1)$$

where *v* is the volume in liters of the solution in the calorimeter and *c* is the equilibrium molar concentration of C. *c* is obtained by inserting measured values for *a*, *b*, and the molar equilibrium constant K_f (from infrared) into the quadratic equation

$$c^2 - c\left(a + b + \frac{1}{K_f}\right) + ab = 0 \qquad (2)$$

where *a* = initial molar concentration of PFP and *b* = initial molar concentration of the base. Expression 1 yields $\Delta H_f°$ when ΔH_{obsd} is plotted *vs.* *cv* (see Figure 2). Excellent linearity based on 15 measurements of ΔH_{obsd} for PFP concentrations of 0.02, 0.01, and 0.005 *M* and pyridine concentrations from 3.44 × 10^{-3} to 18.57 × 10^{-3} *M* gives strong support to the 1:1 complex treatment.

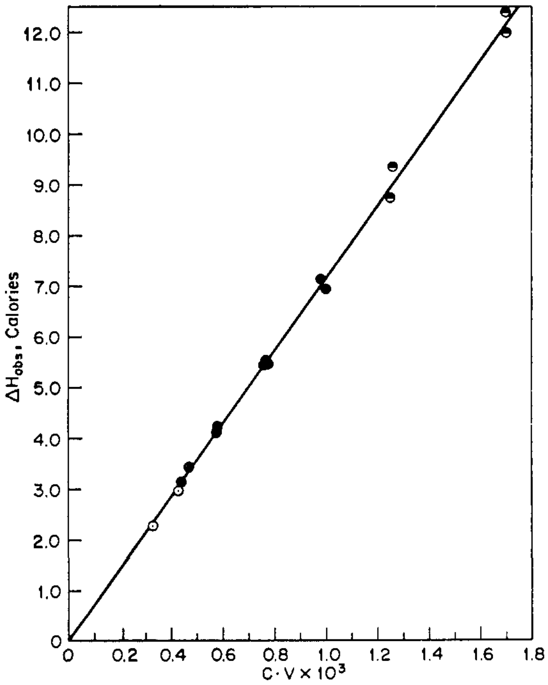

Figure 2. Estimation of $\Delta H_f°$ for *p*-fluorophenol–pyridine complex using eq 1. ΔH_{obsd} is corrected for $\Delta\bar{H}_s$ of pyridine in carbon tetrachloride alone. PFP concentrations are: ◓ = 0.02 *M*, ● = 0.01 *M*, ○ = 0.005 *M*; K^{298} = 75 ± 1.5[6]; $\Delta H_f°$ = −7.2 ± 0.2 kcal/mole.

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(5) ΔH_{obsd} is determined experimentally by subtracting the heat of solution of the base $(\Delta\bar{H})_s$ into pure solvent from that of the same quantity of base into a solution containing acid of molarity, *a*.

Method II. Values of $\Delta H_f°$ determined through eq 2 depend strongly on the reliability of K_f. In view of the difficulty in obtaining K_f with high precision (and, more importantly, acceptable *accuracy*), a totally different approach was also used. Assume that at high dilution in pure aprotic B *as solvent*, each molecule of PFP is hydrogen bonded to one molecule of B. If we could correct for other interactions between PFP and the surrounding B molecules, the interaction energy due to hydrogen bonding alone could be isolated. To do this, anisole and *p*-fluoroanisole (PFA) are used as models. The measurement of heat of formation of C in pure B, which we shall call ΔH_f^B, is then simply a matter of measuring $\Delta \bar{H}_s$ for the phenol and its corresponding methyl ether, first in pure base and then in an inert solvent (carbon tetrachloride) to provide a suitable reference state. Then, for example

$$(\Delta H_f^{B})_{PFP\cdots B} = (\Delta \bar{H}_s^{PFP} - \Delta \bar{H}_s^{PFA})_B - (\Delta \bar{H}_s^{PFP} - \Delta \bar{H}_s^{PFA})_{CCl_4} \quad (3)$$

If all PFP molecules are converted to the 1:1 complex, no estimate of K_f is needed.

This approach assumes that there is not a large medium effect on the heat of hydrogen-bond formation (*i.e.*, $\Delta H_f° \approx \Delta H_f^B$). Comparison of columns *c* and *d* in Table I provides striking evidence that this need not be a serious problem. For the three cases reported here, $\Delta H_f°$ values, at high dilution in CCl_4, and ΔH_f^B, in pure B, agree within experimental error. When allowance is made for uncertainties in K_f, the second method may actually be the more accurate and even be used to estimate K_f through eq 1 and 2. We find that agreement between the two methods is surprisingly general, but breaks down slightly for highly associated solvents such as dimethyl sulfoxide. The relative invariance of most hydrogen-bond strengths to environmental factors suggests that $\Delta H_f°$ for an A–H···B bond determined at high dilution in CCl_4 might be used with little error as a model for a structurally similar situation under other conditions as, for example, within a protein helix.

Procedures for determining $\Delta \bar{H}_s$ and ΔH_{obsd} at 25° have been described previously.[6] The insensitivity of ΔH_f to temperature[7] was demonstrated (for the first time, to our knowledge) by a complete determination of $\Delta H_f°$ for the pyridine–PFP complex at high dilution in carbon tetrachloride at 12°. The value at this temperature to be compared with that in Table I is 7.2 ± 0.2 kcal/mole.[8]

[6] E. M. Arnett, W. G. Bentrude, J. J. Burke, and P. McC. Duggleby, *J. Am. Chem. Soc.*, **87**, 1541 (1965).

[7] ΔH_{obsd} obtained by method I has a large temperature coefficient through *K*.

[8] *K* in this case was 130.6, calculated from the value at 25° by the van't Hoff equation.

Edward M. Arnett, T. S. S. R. Murty
Department of Chemistry, University of Pittsburgh
and Mellon Institute of Carnegie-Mellon University
Pittsburgh, Pennsylvania 15213

Paul von R. Schleyer, L. Joris
Department of Chemistry, Princeton University
Princeton, New Jersey 08540

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Regarding Proton Transfer in Hydrogen-Bonded Complexes as Measured by Fluorine Nuclear Magnetic Resonance[1]

Sir:

We wish to report the use of fluorine nuclear magnetic resonance (F nmr), with the *p*-fluorophenyl label, in the quantitative study of hydrogen bonding. The F nmr shielding of *p*-fluorophenyl derivatives provides a very sensitive probe of the interaction between the benzene ring and the functional substituent.[2] This probe has been used previously to investigate quantitatively Brønsted and Lewis acid–base equilibria,[3] as well as much weaker interactions of the dipole–dipole type.[4]

In the present study *p*-fluorophenol is used as the reference acid toward three typical bases. The work constitutes part of a collaborative investigation with Professor E. M. Arnett (*cf.* companion paper[5]), in which definitive hydrogen-bond parameters are being obtained through a combination of independent infrared, nmr, and calorimetric determinations.

The F nmr shielding results were obtained (D. G. and R. T.) with the use of a Varian HA-60 spectrometer equipped with a Varian 1024 computer of average transients and an NMR Specialties heteronuclear decoupler, Model SD60-B. The temperature-control probe was maintained at 25.0°. By decoupling the ring protons, F nmr shielding measurements at less than 0.01 *M* $p\text{-FC}_6H_4OH$ were obtained. Accordingly, all measurements were made under conditions for which infrared analysis indicates the existence of only the monomeric phenol. For base concentrations up to 0.02 *M*, the same F nmr shift (±0.02 ppm, the experimental error) of $p\text{-FC}_6H_4OH$ is obtained with the use of an external reference as with the internal reference, *p*-fluoroanisole (0.01 *M*). From 0.02 to 0.5 *M* base concentrations only the internal reference gives shifts which meet the analysis conditions discussed below. Since shielding effects of 2–4 ppm are measured to ±0.02 ppm, the F nmr method has sufficient sensitivity for a critical analysis.

The F nmr shifts, δ, for CCl_4 solutions with variable $p\text{-FC}_6H_4OH$ and base concentrations (covering a range of 30–80% complex formed) have been analyzed in terms of the assumed formation of a 1:1 complex to give values for the formation constant, K_f, and the corresponding change in shielding, Δ, between complexed and uncomplexed *p*-fluorophenol.[4] It is initially assumed that both Δ and K_f are independent of concentration effects. Our δ values upon such analysis yield fits to within the experimental error (condition 1). However, the precision of fit of the nmr shift alone is not a sufficient condition for accuracy of K_f and Δ values. Either experimental error or concentration effects could accumulate systematic errors in both parameters even though

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[1] This work was supported in part by the Public Health Service, Project GM14078. We gratefully acknowledge the support of the National Science Foundation which made available to the U.C.I. Chemistry Department the nmr spectrometer.

[2] R. W. Taft, *et al.*, *J. Am. Chem. Soc.*, **85**, 3146 (1963).

[3] (a) R. W. Taft and P. L. Levins, *Anal. Chem.*, **34**, 436 (1962); (b) R. W. Taft and J. W. Carten, *J. Am. Chem. Soc.*, **86**, 4199 (1964); (c) R. G. Pews, Y. Tsuno, and R. W. Taft, *ibid.*, **89**, 2391 (1967).

[4] R. W. Taft, G. B. Klingensmith, and S. Ehrenson, *ibid.*, **87**, 3620 (1965).

[5] E. M. Arnett, T. S. S. R. Murty, P. von R. Schleyer, and L. Joris, *ibid.*, **89**, 5955 (1967).