

DETERMINATION OF THERMODYNAMIC FUNCTIONS BY MEANS OF INFRARED SPECTROSCOPY. THE ACCURACY IN $-\Delta H$ DETERMINATION; FORMATION OF COMPLEXES PHENOL-PYRIDINE AND ITS HOMOLOGUES

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Values of thermodynamic functions $-\Delta H$, $-\Delta G$, and $-\Delta S$ of the complex formation between molecules of pyridine (and its methyl derivatives) and phenol in CCl_4 solutions were determined using infrared spectroscopy. After optimizing experimental conditions the error in $-\Delta H$ determination does not exceed 0.2%; this accuracy is sufficiently high to investigate the influence of substituents on the pyridinic nitrogen basicity. The values of $-\Delta H$ determined in this way correlate linearly with values of ionization constants of pyridine bases.

Evaluation of relative basicities of organic molecules by means of thermodynamic functions of the formation of complexes, formed between their molecules and molecules of standard acids and connected by a hydrogen bridge, has been well known for a long time¹. The same holds for the physico-chemical background of this operation. The method of determining the thermodynamic functions of the complex formation by means of infrared spectroscopy is relatively laborious: therefore, it has been applied mainly in those cases, where it is irreplaceable: in investigating the complex formation in weak bases, difficult to be protonized, in which the basicity centers are atoms of O, S, P, N, halogens, π -electron systems (double bonds, aromatic rings) *etc.* If strong bases are involved, this method can be used rather in special cases, as, *e.g.*, in investigating the basicity in non-aqueous solutions, or the basicity with respect to a very weak acid *etc.*

The measure of a compound basicity is usually the value $-\Delta H$ of the complex formation between the base and a standard acid. This paper deals with the accuracy of $-\Delta H$ values, as determined by infrared quantitative analytical procedures. The main source of usual errors in determining the $-\Delta H$ values is in an incorrect determination of the temperature of the equilibrium mixture and in an incorrect choice of the temperature interval over which the equilibrium is investigated; a proper selection of these parameters makes it possible to derive such procedures of the determination of the thermodynamic functions whose precision enables discerning reliably even small differences in the basicity of the molecules under study.

‡ The aim of our paper is to decide, if the attainable accuracy in determining $-\Delta H$ values is sufficient to describe the relationship between the structure and the basicity of molecules of pyridine bases, and if the difficulties connected with obtaining these results are still acceptable from the point of view of a usual laboratory practice.

Therefore, we have thoroughly analyzed all sources of errors in determining values of the equilibrium constant K_1 and in calculating $-\Delta H$ values from them. We have suggested an optimum way of applying the procedure.

EXPERIMENTAL

All studied compounds were of analytical purity. Pyridine and its derivatives (2-methyl-, 2,4-dimethyl-, 2,6-dimethyl-, 2,4,6-trimethylpyridine) were dehydrated by solid KOH, distilled, and stored in small quantities in sealed ampules under an inert atmosphere. The purity of the compounds was checked by gas-liquid chromatography; the impurity content was always lower than 0.2%. Tetrachloromethane, used as the solvent, was also thoroughly dehydrated, distilled, and stored in sealed ampules in an inert atmosphere. Phenol (used as a proton donor) was distilled and kept in the presence of a molecular sieve before being used.

Ternary mixtures used in infrared spectrum measurements were prepared by weighing the particular components into 5 ml flasks; concentrations of bases were kept between 0.02–0.07M, the concentration of phenol was about 0.003M. As standard solutions binary mixtures containing the base and CCl_4 in the same concentrations as in measured solutions were used. All manipulations with phenol and pyridine were carried out in a dry box. The ternary mixtures investigated were always prepared immediately before the measurements.

Infrared spectra were measured in the range of the fundamental frequency of the O—H bond stretching vibration, using a double-beam spectrophotometer UR-10 (Zeiss, Jena). The spectral slit width in this region was about 4 cm^{-1} , when using a LiF prism.

The photometric scale of the device was calibrated by means of a rotating segment diaphragm; the values measured were plotted into a correction calibration graph. The intensity data used in quantitative evaluations were always corrected by means of this graph.

Infrared spectra of the ternary systems were measured in a thermostated cell with CaF_2 windows, with the adsorption path length 10 mm. The measurements were carried out between 5°C and 50°C in 5°C intervals. The temperature of the studied mixture in the cell was measured during recording the spectrum. The thermistor used was immersed into the measured sample in such a way that it did not perturb the light beam and measured the temperature in its close vicinity.

The method of evaluating the spectra and calculating the values of the thermodynamic functions was essentially the same as in ref.^{2,3}.

RESULTS

The formation of complexes in the systems pyridine base-phenol- CCl_4 can be described by the equilibrium equation



As a base the following compounds were used: pyridine, 2-methyl-, 4-methyl-, 2,4-dimethyl-, 2,6-dimethyl-, 2,4,6-trimethylpyridine. Several experiments were carried out with 2,6-dimethylphenol as a proton donor.

For the system pyridine-phenol in CCl_4 the values of the association constant determined at 25°C (K_{25}) together with the values of $-\Delta H$, $-\Delta G$, and $-\Delta S$ were compared with the available published values (Table I).

TABLE I
Values K_{25} , $-\Delta G$, $-\Delta H$ and $-\Delta S$ of Complex Formation Phenol-Pyridine in CCl_4 Solution

K_{25}	$-\Delta G$, kcal/mol (kJ/mol)	$-\Delta H$, kcal/mol (kJ/mol)	$-\Delta S$, cal/mol $^\circ\text{C}$ (J/mol $^\circ\text{C}$)	Ref.
—	—	5	—	4
—	2.30	6.50 ± 0.40	14.4 ± 1.3	5
59.8 ^a	2.4	7.0	15.6	6
48.5	2.3 (9.62)	5.85 (24.49)	11.9 (49.82)	this paper

^a At 20°C .

Table II summarizes the values of K_{25} and the values of the thermodynamic functions $-\Delta H$, $-\Delta G$, and $-\Delta S$ for all the studied systems pyridine base-phenol in CCl_4 . The data in Table II are arithmetic means of at least three independent measurements. The values S show the relative error in determining the $-\Delta H$ values, calculated according to ref.⁷.

TABLE II
Equilibrium Constants and Thermodynamic Functions $-\Delta G$, $-\Delta H$ and $-\Delta S$ of Complex Formation Phenol-Pyridine Bases

Pyridine base	K_{25}	$-\Delta G$, kcal/mol (kJ/mol)	$-\Delta H$, kcal/mol (kJ/mol)	$-\Delta S$, cal/mol $^\circ\text{C}$ (J/mol $^\circ\text{C}$)	$\pm S^a$	pK_a^b
Pyridine	48.4	2.30 (9.62)	5.85 (24.49)	11.90 (49.82)	0.10	5.23
2-Methylpyridine	66.8	2.50 (10.46)	6.38 (26.71)	13.01 (54.47)	0.07	5.96
4-Methylpyridine	71.3	2.53 (10.59)	6.54 (27.38)	13.44 (56.27)	0.09	6.05
2,4-Dimethylpyridine	83.2	2.62 (10.97)	7.01 (29.35)	14.72 (61.62)	0.03	6.79
2,6-Dimethylpyridine	107.0	2.77 (11.59)	7.34 (30.73)	15.32 (64.14)	0.01	6.62
2,4,6-Trimethylpyridine	125.3	2.86 (11.97)	7.39 (30.94)	15.20 (63.64)	0.15	7.45

^a Relative error of $-\Delta H$ determination; ^b Ref.².

Fig. 1 is an example of the temperature dependence of the equilibrium constant values for the system 2,6-dimethylpyridine-phenol in CCl_4 . Fig. 2 shows the relationship between the $-\Delta H$ values from Table II and the ionization constant values² of the corresponding pyridine bases. The linear relationship between these quantities can be described by equation (1)

$$\text{p}K_a = C_1 \Delta H + C_2 \quad (1)$$

Coefficients C_1 (1.44) and C_2 (-3.28) were determined from the experimental data by means of the least square method.

DISCUSSION

In studying the protonization and ionization of pyridine and its homologues by classical methods⁸, simple relations were discovered between the structure of the studied compounds and their basicity. Thus, *e.g.*, the substitution of the pyridine ring by the methyl group in the α -position increases the value of $\text{p}K_a$ by 0.73, and the substitution in the γ -position by 0.82. In the substitution of both α -positions a correction on the steric hindrance of the pyridine nitrogen must be taken into consideration². The following empirical relationship was derived for determining the basicity of pyridine methyl derivatives the molecules of which contain methyl groups in various amounts in the positions α , β , and γ

$$\text{p}K_a = 5.22 + 0.74a + 0.41b + 0.76c \quad (2)$$

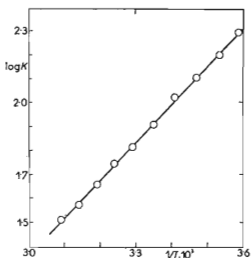


FIG. 1
The Temperature Dependence of the Equilibrium Constant Values for the System 2,6-Dimethylpyridine-Phenol in CCl_4

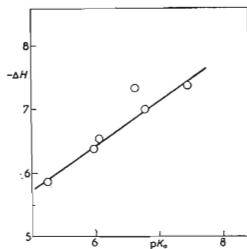


FIG. 2
The Relation between the $-\Delta H$ Values (Table II) and the Ionization Constant Values of the Corresponding Pyridine Bases

where a , b , and c are the numbers of methyl groups in the positions α , β , and γ , respectively. A series of similar equations has been (or could be) derived which take into consideration the size of the ensemble, the nature of substituents, methods employed, solvents *etc.* These equations differ in proportionality constants; in principle, however, their applicability shows that the accuracy in determining the pK_a values is sufficient to discern the relations between structural factors and the basicity of pyridine derivatives.

Therefore, it was surprising to find out that the basicity of pyridine derivatives expressed by means of the respective $-\Delta H$ values of the complex formation between these bases and phenol, as obtained from infrared spectra, do not exhibit any relation to the structure of the bases and form a group of randomly scattered values⁶. Absolute values of $-\Delta H$ reported by various authors for the system pyridine-phenol in CCl_4 (Table I) differ considerably and the same holds for the reported values of errors in determining the $-\Delta H$ values. This all could indicate that the error in determining the thermodynamic function of the complex formation between pyridine bases and phenol, as obtained by the infrared quantitative analytical methods is so big that it does not allow for discerning relations between their structure and basicity. The determination of thermodynamic functions of the complex formation from infrared spectra has been described in a series of earlier communications^{1,5}. An analysis of some errors of this procedure has been treated in ref.⁹, a general view of the error analysis in determining the thermodynamic functions has been given in ref.¹⁰.

Equilibrium constants of the studied systems are determined from the mass action law which for the systems under study can be written as

$$K = [C_6H_5OH \dots base] / [base] [C_6H_5OH]. \quad (3)$$

In this equation only molar concentrations of the reacting substances appear so that the error in the equilibrium constant determination is given only by the error in the determination of the concentrations of the components in question. However, because the concentrations of the components in the equilibrium state are not known and can be determined only from infrared spectra, the error of the concentration determination is augmented by the error in the photometric scale of the instrument used.

Practically, the following equation is used to evaluate the equilibrium constant

$$K = (1 - \alpha) / [C_b^0 - C_a^0(1 - \alpha)], \quad (4)$$

where C_b^0 and C_a^0 are analytic concentrations of the base and the acid, respectively, and α is the degree of association defined by

$$\alpha = A_{max} / A_{max}^0. \quad (5)$$

Here, A_{\max} is the absorbance at the $\nu(\text{OH})$ band maximum of the free phenol in the mixture with the base, A_{\max}^0 the calculated absorbance of the $\nu(\text{OH})$ band of phenol of the same concentration in the solution, without the base. Infrared spectroscopy enters the experiment through the quantity α : the value $(1 - \alpha)$ represents the complex concentration. Most of unwanted influence on the resulting value of the equilibrium constant (*e.g.*, the influence of temperature on the component concentrations on the ternary mixture and on the absorption path length) can be removed by introducing the ratio α .

The error in sampling the individual components during the ternary mixture preparation can be minimized by weighing all the components, working at a constant temperature with thermostated glassware, minimizing the evaporation of the solutions as well as the influence of humidity and CO_2 .

In measuring the temperature dependence of the equilibrium constants in a higher temperature range a possibility of a distortion of the measured absorption band intensities is likely, resulting from the escape of the solvent (or of the volatile base) from the system through possible leaks in the measuring cell. In order to check the correct function of the thermostated cell, after a completed measurement at 50°C , the ternary mixture was cooled to 25°C and the infrared spectrum was measured again. Practically the same intensities were obtained as in the first measurement at this latter temperature. This check of the measuring arrangement function is, we believe, very essential for obtaining reliable results.

An important step in obtaining correct values of the association constants is a proper choice of the cell thickness. With the employed cell 10 mm thick, the phenol concentration in an experimental series was about $3 \cdot 10^{-3}\text{M}$. Therefore, it was not necessary to take into consideration the self-association of phenol and to introduce into the fundamental expression for the equilibrium constant calculation any expressions for the phenol dimer or trimer formation (as recommended in ref.⁹ for higher phenol concentrations).

The choice of the optimum acid-base ratios in an experiment has been discussed in detail in ref.¹¹.

The value $-\Delta H$ is determined from equation (6)

$$\ln(K_1/K_2) = (\Delta H/R)/[(1/T_1) - (1/T_2)], \quad (6)$$

where K_1 and K_2 are the equilibrium constant values measured at two temperatures T_1 and T_2 , respectively.

A correct determination of the equilibrium mixture temperature is obviously the most sensitive step in determining the equilibrium constant values and — consequently — of the $-\Delta H$ values. In those cases where the temperature of the cell jacket or that of the thermostating medium is used instead of the temperature of the measured system, a considerable error is introduced into the association constant

values and the $-\Delta H$ values due to non-respecting the temperature gradient between the measured system and its environment. This fact was confirmed on the system 2,4-dimethylpyridine-phenol in CCl_4 : in measuring the temperature dependence of the equilibrium constants the ternary mixture temperature was measured either by means of a thermistor placed within the measured system or by a thermometer of the ultrathermostat used to thermostat the cell. Fig. 3 shows both dependences of $\log K - 1/T$ obtained in this way. When the temperature inside of the ternary mixture studied was measured, the value $-\Delta H = 7.02 \text{ kcal/mol}$ resulted, while in the latter case $-\Delta H = 5.80 \text{ kcal/mol}$ was calculated. At the temperature 5°C of the ternary mixture the temperature difference cell - bath is largest, as the temperature of the thermostating medium has to be kept at -4°C which makes the over-all difference 9°C . At 50°C this difference is substantially smaller, only 1.5°C . It is evident from these data that by measuring the temperature directly in the cell, *e.g.*, by a thermistor immersed into the measured solution, the temperature determination is assured over the entire range studied¹².

Also, the determination of the $-\Delta H$ value is strongly influenced by the temperature span over which the equilibrium is studied, and by the number of measurements in this chosen temperature interval. In most studies the authors have measured the association constant values at two, not too much differing temperatures (*e.g.*, 25°C and 50°C). Fig. 4 shows the temperature dependence of the association constant values for the system 2,6-xylenol-pyridine in CCl_4 which exhibited a somewhat larger scatter of experimental points, evidently due to random error in the measurements. Table III summarizes the $-\Delta H$ values obtained from the measurements in arbitrarily chosen temperature intervals; for the intervals $5^\circ\text{C} - 50^\circ\text{C}$ and $10^\circ\text{C} - 50^\circ\text{C}$ the course of the temperature dependence was evaluated by the least square method.

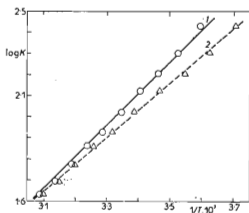


FIG. 3
Dependences at $\log k - 1/T$

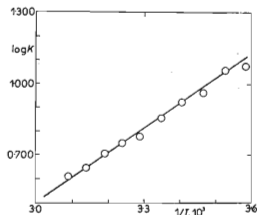


FIG. 4
The Temperature Dependence of the Association Constants Values for the System 2,6-Xylenol-Pyridine in CCl_4

If we base our considerations on the fully justified assumption that the correct $-\Delta H$ value was obtained when the measurements were carried out over the widest temperature interval and with the largest number of experimental points, then Table III clearly shows that under other conditions the results obtained were either loaded with a considerable error or an agreement was quite fortuitous (*cf.* interval 10°C to 45°C). It can be concluded that the laborious and time-consuming character of the necessary measurements cannot be avoided, as the reduction of the number of experimental points leads to results often loaded with a large error.

If the $-\Delta H$ values are determined by the procedure described in this paper or by another procedure which respects the problems of the equilibrium mixture temperature determination, a very good correlation with the values of the ionization constants of pyridine bases (Fig. 2) can be obtained. It is interesting that this correlation holds with such a remarkable precision for the values determined both in the tetrachloromethane solution ($-\Delta H$) and in aqueous or alcoholic - aqueous solutions (pK_a). An explanation of this fact may be based on the assumption that a proportionality exists between the solvation of all the bases studied, by water and by other solvating media (*e.g.*, tetrachloromethane).

It can be seen from Fig. 2 that the $-\Delta H$ value of the complex 2,6-dimethylpyridine-phenol does not correlate very well with the pK_a value - the only case where this correlation fails. Because pK_a of all pyridine bases correspond very well to each other and because the deviation of $-\Delta H$ cannot be explained by, *e.g.*, a steric hindrance of a hydrogen bridge of two methyl groups in the vicinity of the nitrogen atom (the $-\Delta H$ value for 2,4,6-trimethylpyridine corresponds exactly to the corresponding value of pK_a), one has to assume that the 2,6-dimethylpyridine used was not quite pure or that the entire experiment was influenced by another gross error. This is why this point was not included into the calculation of the correlation straight line (1).

TABLE III

Values of $-\Delta H$ of the System 2,6-Dimethylpyridine-Phenol in CCl_4 Solution

Temperature range °C	$-\Delta H$ kcal/mol (kJ/mol)	Number of points in temperature interval
5-50	4.76 ± 0.04 (19.92)	10
10-50	4.61 (19.30)	5
15-50	4.18 (17.50)	2
20-50	4.46 (18.67)	2
10-45	4.79 (20.05)	2

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