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Basicities of the 2-, 4-, 2,4-Di-, and 2,6-Disubstituted *tert*-Butyl Pyridines in the Gas Phase and Aqueous Phase: Steric Effects in the Solvation of *tert*-Butyl-Substituted Pyridines and Pyridinium Cations

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Abstract: A complete analysis is presented for the thermodynamics of protonation, both in the gas phase and aqueous phase, for 2-*tert*-butylpyridine (2-TBP), 4-*tert*-butylpyridine (4-TBP), 2,4-di-*tert*-butylpyridine (2,4-DTBP), and 2,6-di-*tert*-butylpyridine (2,6-DTBP). Gibbs free energy, enthalpy, and entropy changes for the transfer from the gas phase to the aqueous phase are reported for the neutral and protonated species for all the *tert*-butyl-substituted pyridines in this study. The aqueous basicity of 2,6-DTBP is anomalous, compared to that of the other alkyl-substituted pyridines; but this is due primarily to entropy rather than enthalpy effects. The enthalpy change for the transfer of the protonated 2,6-DTBP species (2,6-DTBPH$^+$) from the gas phase to the aqueous phase is normal; however, the corresponding entropy change is abnormal. These results are qualitatively similar to the recent thermodynamic parameters reported for the attachment of one water molecule to substituted pyridinium ions in the gas phase. All the data presented here and that currently in the literature are consistent with the following conclusions: (1) The aqueous 2,6-DTBPH$^+$ cation is hydrogen bonded to a water molecule via the N–H bond. (2) In the aqueous solvation complex of the 2,6-DTBPH$^+$ cation, several internal rotations are substantially restricted because of steric interactions which cause a loss of entropy that is not seen for the other pyridinium ions in aqueous solution. (3) In pure water a significant fraction of aqueous 2,6-DTBP is also probably hydrogen bonded to a water molecule at the pyridine nitrogen.

During the past 30 years several groups of workers have investigated the thermodynamic and spectroscopic properties of *tert*-butyl-substituted pyridines because of the perceived unusual behavior of the 2,6-DTBP member of this series. Brown and Kanner[1,2] first synthesized 2-TBP and 2,6-DTBP, but because of the low solubility of these compounds in water, they determined the basicities (pK_a values for the corresponding cations) in an alcohol–water mixture. As expected, 2-TBP was more basic than pyridine; however, 2,6-DTBP was reported to be substantially less basic than pyridine. This phenomenon was correctly perceived to be due to the steric requirements of the bulky *tert*-butyl groups, but the exact nature of the effect could not be discerned solely from the pK_a measurements. Consequently, they postulated three possibilities: (1) the N–H bond in the 2,6-DTBPH$^+$ cation is compressed or strained by the steric crowding around the nitrogen; (2) steric interactions interfere with the hydrogen bonding of a solvent molecule to the 2,6-DTBPH$^+$ cation; or (3) the overall electrostatic solvation of the 2,6-DTBPH$^+$ cation is reduced by steric hindrance. Because of the large deviation of the basicity of 2,6-DTBP from the expected trend, they preferred one of the first two arguments based on a direct steric interactions at or near the N-H bond in the 2,6-DTBPH$^+$ cation. Additional studies on the basicities of these pyridines in a variety of alcohol–water mixtures were performed by McDaniel and Ozcan,[3] who concluded that the lower than expected basicity of 2,6-DTBP probably was due to steric inhibition to solvation of the 2,6-DTBPH$^+$ cation. Condon[4] independently arrived at the same conclusion from a detailed analysis of the basicity of sterically hindered bases. Furthermore, the first possibility has been eliminated because the preliminary data of two groups of workers[5,6] indicate that the gas-phase basicity of 2,6-DTBP is close to what one predicts from the gas-phase basicities of the other *tert*-butyl-substituted pyridines. Thus, the abnormally low basicity of 2,6-DTBP in solution must be due to solvent effects that remain to be defined.

Arnett and Chawla[7] included 4-TBP and 2,6-DTBP in their analysis of the thermodynamic data for the protonation of a series of alkyl- and chloro-substituted pyridines in the gas and aqueous phases. When they combined their very precise and accurate vapor pressure data on five substituted pyridines with a variety of thermodynamic parameters, either estimated or taken from the literature, they were able to connect the gas-phase results for 12 substituted pyridines, via a thermodynamic cycle, to those in the aqueous phase. Unfortunately, they had to estimate some of the data for 2,6-DTBP in the aqueous phase because of its extremely low solubility in water; nevertheless, their results appeared to them to be consistent with the conclusion that 2,6-DTBP has a lower than expected aqueous basicity because steric hindrance totally precludes hydrogen bonding of a water molecule to both the cation and neutral species. They also noted that similar conclusions had been stated previously by Wolf, Harsch, and Taft.[5]

Two other groups of investigators[8,9] reported entropy and volume changes for the protonation of 2,6-DTBP in solution that cannot be simply interpreted in terms of steric hindrance to electrostatic solvation or hydrogen bonding. When Le Noble and Asano[8] measured the pK_a and volume changes ($\Delta V_i°$) in pure methanol for the ionization of a series of 2,6-disubstituted pyridinium cations, they also found the pK_a for the 2,6-DTBPH$^+$ cation to be lower than expected. But quite surprisingly, they found the $\Delta V_i°$ value for the 2,6-DTBPH$^+$ cation to be substantially larger than the corresponding values for the other pyridinium cations in the series. In their discussion, they noted that the $\Delta V_i°$ value for the 2,6-DTBPH$^+$ cation is the "largest known volume change for a proton transfer" and that this value is not consistent with a greatly diminished electrostriction around the 2,6-DTBPH$^+$ cation. Because of this observation they stated that "one must conclude that there is no evidence for a conspicuous lack or absence of electrostriction around the cation". A similar conclusion was reached by Hopkins and Ali[9] from their analysis of the entropy

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changes ($\Delta S_p°$) found in alcohol–water mixtures for the protonation of 2-TBP, 4-TBP, 2,4-DTBP, and 2,6-DTBP. These workers were able to establish the entropy of the cations relative to the neutral pyridines, $\bar{S}_{py} - \bar{S}_{pyH^+}$. In 90% methanol–water this difference is larger for 2,6-DTBP than for the other pyridines in the series and is comparable to that for the other pyridines in 40.9% ethanol–water. If the 2,6-DTBPH$^+$ cation is not as well solvated as the other pyridinium cations, then the difference in entropies should be substantially smaller for 2,6-DTBP than for the other pyridines. The fact that this is not the case prompted Hopkins and Ali[9] to state, just as did Le Noble and Asano,[8] that steric inhibition to solvation cannot be the sole cause for the unusually low basicity of 2,6-DTBP; however, because the enthalpy change for the protonation ($\Delta H_p°$) of 2,6-DTBP in both solvent systems was greater than that for the other pyridines in the series, they could not entirely dismiss the possibility of steric hindrance to solvation for the 2,6-DTBPH$^+$ cation. These workers also suggested that one or more of the barriers to internal rotation of the *tert*-butyl groups on 2,6-DTBP could conceivably increase upon protonation because of steric interactions caused by solvation around the cation. They also showed that such a phenomenon would produce a decrease in entropy that could be responsible for the observed entropy changes.

In an effort to determine if the proposed increases in the rotational barriers could be detected by spectroscopic means, Hopkins and Ali[10,11] also performed a detailed study of the carbon-13 NMR spectra of a series of alkyl-substituted pyridines in alcohol–water solvents. For pyridines with methyl groups in the 2-, 3-, and 4-position, they found the barrier to internal rotation of the methyl group to be approximately zero for both the neutral and cation species. In similar carbon-13 NMR studies[11] on 4-TBP and 2-TBP, they calculated the barrier to the internal rotation of the methyl groups to be near 2.5 kcal mol^{-1} for both the neutral and cation species. However, the barrier for the methyl groups on 2,6-DTBP increased nearly 1 kcal mol^{-1} upon protonation in a methanol–water mixture. The relaxation time for the methine carbon of the *tert*-butyl groups in 2,6-DTBP also was found to be reduced upon protonation, which is consistent with a significant decrease in the rotational motion of the *tert*-butyl groups on the 2,6-DTBPH$^+$ cation. Both of these observations were interpreted by Hopkins and Ali[11] as support for their previous postulate that the rotational motions of the *tert*-butyl groups on 2,6-DTBP are more hindered in the 2,6-DTBPH$^+$ cation than in the neutral pyridine.

From the previous discussion and analysis, we concluded that the phenomena responsible for the unusually low basicity of 2,6-DTBP in solution still were not well understood. Consequently, as a continuation of our study of the thermodynamics of the protonation of pyridines in the gas phase and aqueous phase, we endeavored to determine for 2-TBP, 4-TBP, 2,4-DTBP, and 2,6-DTBP the data needed to complete the thermodynamic cycle:

$$\begin{tikzcd}
\mathrm{BH^+(gas)} \arrow[r] & \mathrm{B(gas)} \arrow[d] & + & \mathrm{H^+(gas)} \arrow[d] \\
\mathrm{BH^+(aq)} \arrow[u] \arrow[r] & \mathrm{B(aq)} & + & \mathrm{H^+(aq)}
\end{tikzcd}$$

Since the solubility of these pyridines in water is low, the pK_a values could only be determined by a spectroscopic method in which the concentrations were maintained in the 10^{-5} M range. Once these measurements were completed, we were able to devise a new and simple procedure for measuring the solubilities of these pyridines in water from pH measurements. Direct calorimetric measurements could not be performed on either the protonation or the dissolution process for these pyridines, but the enthalpy changes for these processes could be evaluated by indirect procedures. The enthalpy changes for protonation ($\Delta H_p°$) were determined from the temperature dependence of the pK_a values, whereas the enthalpy changes for the dissolution process were calculated from the $\Delta H_p°$ values and the calorimetrically determined enthalpy changes for dissolutions of the pyridines into aqueous HCl.

In this paper we have summarized all the thermodynamic data for 2-TBP, 4-TBP, 2,4-DTBP, and 2,6-DTBP that is needed to complete the cycle shown above. After this presentation and a comparison of our results with those of Arnett and Chawla,[7] we discuss the variations found for the thermodynamic parameters. Our results and conclusions concerning the solvation of the aqueous 2,6-DTBPH$^+$ cation are in excellent agreement with the recent gas-phase studies conducted by Meot-Ner and Sieck[12] on the attachment of a water molecule to pyridinium cations; i.e., the 2,6-DTBPH$^+$ cation hydrogen bonds to a water molecule, but several internal rotations in the complex are severely restricted. We are also able to show that 2,6-DTBP is probably hydrogen bonded in aqueous solution to a water molecule at the pyridine nitrogen.

Experimental Procedures

pK_a Measurements, A spectrophotometric procedure, similar to one that previously has been described,[9,13] was used to determine the pK_a values for the pyridinium ions. For each of the cations, the spectra in the UV region were dominated by an intense absorption band, which had a maximum near 270 nm. As NaOH was added to solutions of these cations, the band near 270 nm decreased in intensity and a new one, but much weaker in intensity, appeared at longer wavelengths. Well-defined isobestic points also appeared in the spectra for 4-TBP, 2-TBP, and 2,4-DTBP during this systematic addition of NaOH. Since we chose to perform the studies at the maximum in the absorbance for each cation, extinction coefficients for the cations and the neutral bases were determined at this wavelength in the usual manner. These were found to vary little with temperature, and the values determined at 25 °C were used at all temperatures in the calculations. The pK_a values for the pyridinium ions were determined in an acetate buffer from the absorbances at 270 nm by a procedure first descried by Bates and Gary.[14] This procedure was followed at each temperature, which was determined to the nearest 0.05 °C with a thermometer calibrated against a platinum resistance thermometer certified by the National Bureau of Standards.

For 2,6-DTBP a well-defined isobestic point was not found when NaOH was added to the solution of the cation. Instead, the absorbance at the maximum in the UV region merely diminished continuously as NaOH was added, and there was not a point where the absorbance remained constant during the conversion of the cation to its conjugate base. When small quantities of ethanol were added to an aqueous solution of this cation, however, an isobestic point was present during the addition of NaOH. Consequently, we decided to determine the extinction coefficients for the 2,6-DTBPH$^+$ cation and its conjugate base at the maximum in the spectra for the cation and use these values in the same procedure described for the other pyridines.

When the absorbance for the 2,6-DTBPH$^+$ cation at the maximum was determined at high concentrations of HCl or $HClO_4$, a plot of absorbance vs. concentration of the cation was nonlinear. This curvature disappeared when the concentration of the acid was below 0.002 M; thus, a precise value for the extinction coefficient for the cation was determined below this concentration by the normal least-squares method. Because the solubility of 2,6-DTBP is less than 5×10^{-5} M, we could not determine a precise value for the extinction coefficient of the base in pure water; however, we were able to determine the extinction coefficient for the base by the normal method in a 5% by volume ethanol–water mixture. Since this value was not much different from that estimated from the experiments in pure water and was much more precise, we used it in all the calculations.

Calorimetric Measurements. All enthalpy changes for the dissolution of the pyridines into aqueous HCl were measured calorimetrically with a Tronac Model 450 titration calorimeter, in which the voltages were measured with a Hewlett-Packard digital voltmeter (Model 3450). Since the liquid bases were added directly into the aqueous HCl from a syringe in a water bath outside the Dewar flask, the heat of the reaction was calculated from the overall temperature change and the heat capacity determined after the introduction. A correction was also made for the small heat, determined in a separate experiment, that accompanied the transfer of liquid from the water bath to the Dewar flask. When we

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studied 2-TBP, 4-TBP, and 2,4-DTBP, a rapid evolution of heat was observed when 0.5 to 0.7 × 10^{-3} mol of one of these bases was introduced into 25 mL of 0.5 M HCl; however, results from the same type of experiment with 2,6-DTBP were erratic. We concluded from a subsequent analysis of the actual time vs. temperature curves that the rate of dissolution of 2,6-DTBP into aqueous HCl is much slower than that for the other pyridines in this study. In 0.2 and 0.3 M HCl the rate was even slower, but when the time for equilibration in the computer program was increased appropriately, reproducible enthalpy changes for the dissolution were found for all three molarities of HCl.

Vapor-Pressure Measurements. A specially constructed cell, which has been previously described,[15] was used in the determination of the vapor pressures by the Knudsen effusion method at different temperatures. The measurements on each pyridine were performed in a temperature interval, located between −25 and 35 °C, in which the mean free paths of the gaseous molecules were always larger than the diameter of the hole in the top of the cell.

Solubility Determinations. Low solubilities of liquid organic bases in water are not easily established by conventional methods involving simple analytical procedures. In order to avoid the problems inherent in determining by analytical means the true concentration of the aqueous base in equilibrium with the liquid base, we developed a method that relies on knowing the pK_a for the ionization of the cation,

$$\mathrm{BH^+(aq)} \rightarrow \mathrm{B(aq)} + \mathrm{H^+(aq)} \quad K_a = A_B A_{H^+}/A_{BH^+}$$

In this method the pH of an aqueous solution of HCl is measured before and after the addition of an excess of a base, and the solubility of the base is calculated from the pH values and the pK_a of the cation.

Before the addition of base, the chloride concentration is related to the pH:

$$[\mathrm{Cl^-}]_1 = [\mathrm{H^+}]_1 - [\mathrm{OH^-}]_1 = (10^{-\mathrm{pH}_1}/\gamma_{\mathrm{H^+}}) - (K_w/(10^{-\mathrm{pH}_1}\gamma_{\mathrm{OH^-}}))$$

After the addition of base, the equation for electrical neutrality is

$$[\mathrm{Cl^-}]_1 = [\mathrm{H^+}]_2 + [\mathrm{BH^+}]_2 - [\mathrm{OH^-}]_2$$

Since the ionic strength of the solution is not affected by the addition of the base, $\gamma_{\mathrm{H^+}}$ and $\gamma_{\mathrm{OH^-}}$ remain constant and are approximately equal. If one now makes the reasonable assumption that $\gamma_{\mathrm{BH^+}} = \gamma_{\mathrm{H^+}}$, the activity of $\mathrm{BH^+}$ is

$$A_{\mathrm{BH^+}} = \gamma_{\mathrm{H^+}}[\mathrm{BH^+}] = 10^{-\mathrm{pH}_1} - 10^{-\mathrm{pH}_2} + K_w(10^{\mathrm{pH}_2} - 10^{\mathrm{pH}_1})$$

This value can be substituted into the equilibrium expression for the ionization of $\mathrm{BH^+}$ to yield the activity of the aqueous base, A_B. Addition of an inert electrolyte will not affect appreciably the values calculated in this manner, because the ionic strength of the solution is not changed by the dissolution of the base into the HCl solution. So long as there is an excess of liquid base in the vessel and equilibrium is established, the activity of the base calculated in this manner is also the activity of the aqueous base in equilibrium with the liquid base. Therefore, one can conveniently determine by this procedure the aqueous solubility of a base at the ionic strength of the solution. This calculated solubility will depend to some extent on the ionic strength[16,17] because of the well-known salting-in and salting-out phenomena. If the ionic strength is kept low ($I < 0.1$), the measured solubilities usually are not appreciably different from the values in pure water.[16,17]

All pH measurements were made with a Fisher Accumet (Model 114) or Orion (Model 701A) pH meter that was calibrated before each measurement with a standard buffer. In the actual experiments, an excess of a liquid pyridine was added to an aqueous solution of HCl (concentration near 0.01 or 0.001 M) for which the pH had been already measured. This mixture was stirred until we observed a new and steady value for the pH, which usually appeared in less than 24 h. The precision of the measurements increased when KCl was present in the solution and most of the measurements were performed in a standard solution, purchased from Fisher Scientific Co., which had a stated pH of 2.00 ± 0.01 and an ionic strength of 0.05 M. For 2,6-DTBP a turbid solution was observed at high rates of stirring and the pH values were erratic. When the rate of stirring was reduced, the aqueous solution in contact with liquid 2,6-DTBP did not appear to be turbid and the pH values were stable after 24 h.

Gas-Phase Protonation Measurements. The equilibrium constants for the proton transfer from protonated references bases to 2,4-DTBP were determined by drift and trapped cell ion cyclotron resonance techniques.[18,19] The experimentally determined Gibbs free energy changes at 25 °C for the proton transfer to 2,4-DTBP from the di-*sec*-butylammonium ion (GB = 222.7 kcal $\mathrm{mol^{-1}}$), the triethylammonium ion (GB = 223.4 kcal $\mathrm{mol^{-1}}$), and the 2,2,6,6-tetramethylpiperidinium ion (GB = 223.9 kcal $\mathrm{mol^{-1}}$) are −0.8, −0.3, and 0.4 kcal $\mathrm{mol^{-1}}$, respectively. From these values we calculated a GB ($-\Delta G_p°$(gas)) of 223.6 kcal $\mathrm{mol^{-1}}$ for 2,4-DTBP. For pyridine, 2-TBP, and 4-TBP, the Gibbs free energy changes for similar proton transfers and the derived GB's are reported elsewhere.[18] Our attempts to establish the GB for 2,6-DTBP by trapped-cell equilibrium measurements produced a GB of about 221 kcal $\mathrm{mol^{-1}}$, but ion ejection experiments indicated that the rates of proton transfer with 2,6-DTBP were less than 10^{-11} $\mathrm{cm^3}$ $\mathrm{molecule^{-1}}$ $\mathrm{s^{-1}}$ for reference bases with nearby GB values, such as triethylamine and diethylmethylamine, which would cause the equilibria in the trapped cell to be incomplete.

In drift cell double-resonance experiments with $\mathrm{Me_2C{=}NEt}$ (GB = 221.7 kcal $\mathrm{mol^{-1}}$), proton transfer from this protonated base to 2,6-DTBP appeared to occur at an appreciable rate. The rate of proton transfer from the 2,6-dimethylpyridinium ion (GB = 219.3 kcal $\mathrm{mol^{-1}}$) to 2,6-DTBP was measured from the rate of decay of the 2,6-dimethylpyridinium ion as small amounts of 2,6-DTBP were added to the trapped cell. From these experiments we determined the rate constant for the second-order process to be 1.2 × 10^{-9} $\mathrm{cm^3}$ $\mathrm{molecule^{-1}}$ $\mathrm{s^{-1}}$, which is 4 times higher than the value obtained by Jasinski and Brauman[20] by the same technique. When the rate of proton transfer between 2,6-DTBP and the 2,6-dimethylpyridinium ion was monitored by following the rate of formation of the 2,6-DTBPH$^+$ cation, we calculated a rate constant for the second-order process which was identical with the one determined from monitoring the 2,6-dimethylpyridinium ion. Using the same technique, we also found a value of 1.6 × 10^{-9} $\mathrm{cm^3}$ $\mathrm{molecule^{-1}}$ $\mathrm{s^{-1}}$ for the rate constant for the transfer of a proton from the 2,6-dimethylpyridinium ion to pyridine. When this technique was applied to the measurement of the rate of proton transfer from the triethylammonium ion or the diethylmethylammonium ion to 2,6-DTBP, the rates were found to be too small to measure. Thus, the rate of proton transfer to 2,6-DTBP in the gas phase appears to be too slow for one to make accurate equilibrium measurements with bases that have GB values near that of 2,6-DTBP. Only when the proton-transfer reaction is exothermic by 4 kcal $\mathrm{mol^{-1}}$, as is the case for the transfer between the 2,6-dimethylpyridinium ion and 2,6-DTBP, does the rate of proton transfer approach the collision limit.

Because of this observation, we chose to utilize another technique for the indirect determination of the gas-phase basicity of 2,6-DTBP and 2,6-diisopropylpyridine (2,6-DIPP). Utilizing mass spectrometric methods,[21] we measured the relative intensities of the ammonium ion fragments produced from the decomposition of the metastable proton-bound complex ion formed by the 2,6-DTBPH$^+$ cation and 2,2,6,6-tetramethylpiperidine. From these measurements we determined a gas-phase basicity of 223.4 kcal $\mathrm{mol^{-1}}$ for 2,6-DBTP. In a similar manner, we determined the GB for 2,6-DIPP to be the same as that of *N*-methyl-*cis*-3,5-dimethylpiperidine (GB = 222.2 kcal $\mathrm{mol^{-1}}$).

Results

Thermodynamics for Aqueous Protonation and Dissolution. In the procedure followed in the pK_a measurements, we actually determined apparent pK_a values for the ionization of the cations and then estimated the true thermodynamic pK_a values by applying a correction for the activity coefficients for the cations. These corrections, calculated from the Debye-Hückel equation, were in the neighborhood of 0.08 pK_a units for all cases; thus, since the experimental uncertainties were no more than 0.01 pK_a units, we assume that the total uncertainties associated with the thermodynamic pK_a values given in Table I are no more than ±0.05 pK_a units. For each pyridinium ion, the pK_a values were determined in this manner at 5-deg intervals between 15 and 40 °C. In all cases the pK_a values could be fit to a linear equation in $1/T$, which had an associated correlation coefficient greater than 0.98. The slopes derived from the least-squares procedure were used to calculate the $\Delta H_p°$(aq) values and associated standard deviations given in column two in Table I.

The pK_a values at 25 °C for the aqueous 2-TBPH$^+$ and 2,6-DTBPH$^+$ cations are within experimental error the same as those

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Table I. Thermodynamics at 25 °C for the Aqueous Protonation and Dissolution of 2-TBP, 4-TBP, 2,4-DTBP and 2,6-DTBP

	pK_a	$-\Delta H_p°(aq)$, kcal mol^{-1}	$-\Delta H_s°(HCl)$, kcal mol^{-1}	$\Delta H_s°(H_2O)$, kcal mol^{-1}	solubility, mol L^{-1}	$\Delta G_s°(H_2O)$, kcal mol^{-1}
2-TBP	5.69	6.4 ± 0.2	8.85 ± 0.01	-2.4 ± 0.2	$(1.8 \pm 0.1) \times 10^{-2}$	2.4 ± 0.05
4-TBP	5.66	6.4 ± 0.3 6.2 ± 0.2^a	8.13 ± 0.03	-1.7 ± 0.3	$(5.4 \pm 0.1) \times 10^{-2}$	1.7 ± 0.05
2,4-DTBP	6.70	9.3 ± 0.4	9.03 ± 0.05	0.3 ± 0.4	$(1.8 \pm 0.2) \times 10^{-4}$	5.1 ± 0.05
2,6-DTBP	4.95	8.5 ± 0.4	7.45 ± 0.05 6.7 ± 0.3^b	1.0 ± 0.4	$(3.7 \pm 0.3) \times 10^{-5}$ $9.4 \times 10^{-3\,b}$	6.0 ± 0.05 2.8^b

[a] 2% ethanol–water. [b] Reference 7.

reported by McDaniel and Ozcan[3] in 2% by volume alcohol–water solutions. Our $\Delta H_p°(aq)$ value for 4-TBP, derived from the temperature dependence of the pK_a values, is within experimental error the same as that determined calorimetrically in a 2% by volume ethanol–water solution. Nevertheless, our value for $\Delta H_p°(aq)$ for 2,6-DTBP, derived from the temperature dependence of the pK_a values, is nearly 2 kcal mol^{-1} more negative than the estimate of Arnett and Chawla.[7] This value was estimated from a combination of their calorimetric measurements on the dissolution of 2,6-DTBP into aqueous HCl and their thermal titration of the aqueous 2,6-DTBPH$^+$ cation with NaOH to form the liquid base. Since the enthalpy changes for both of these processes were approximately the same, Arnett and Chawla assumed that $\Delta H_s°(H_2O)$ was 0, and $\Delta H_p°(aq)$ was equal to the $\Delta H_s°(HCl)$ for 2,6-DTBP.

In our calorimetric experiments on the dissolution of the *tert*-butyl-substituted pyridines into aqueous HCl, we performed at least three calorimetric determinations for each base; however, nine independent determinations were performed for 2,6-DTBP: three in 0.5 M HCl, three in 0.3 M HCl, and three in 0.2 M HCl. These values could be converted into the corresponding values at infinite dilution, if the enthalpies of dilution were available for HCl and the chloride salts of the protonated pyridines. The former are known, but the latter are not; however, only the difference between these quantities is needed to convert the experimental values to the ones at infinite dilution. Thus, we have assumed that these differences are zero and have listed the experimental values as the $\Delta H_s°(HCl)$ values in Table I. Our value for $\Delta H_s°(HCl)$ for 2,6-DTBP is 0.8 kcal mol^{-1} more negative than that reported by Arnett and Chawla,[7] but this is probably due to the slow rate of dissolution observed by us for the dissolution of 2,6-DTBP into aqueous HCl.

The $\Delta H_s°(HCl)$ values were combined in the usual manner with those for $\Delta H_p°(aq)$ to yield the enthalpy changes for the dissolution of the bases into pure water ($\Delta H_s°(H_2O)$). Arnett and Chawla[7] also reported a value for $\Delta H_s°(H_2O)$ for 2,6-DTBP, which they estimated from their experiments on the dissolution of 2,6-DTBP into HCl and the neutralization of the 2,6-DTBPH$^+$ cation with aqueous NaOH. Their value for this parameter is 0.0 ± 0.6 kcal mol^{-1}, which must be considered the same as our value of 1.0 ± 0.4 kcal mol^{-1}, if one applies standard statistical analysis. Since Arnett and Chawla[7] arrived at their value by a different route than we used, the agreement found between the two values is encouraging.

For each base, the solubility was determined several times by the pH procedure, and the average and standard deviation for these measurements are given also in Table I. The standard Gibbs free energy changes ($\Delta G_s°(aq)$) for the dissolution of the pyridines into water were calculated from these values and the standard relationship, assuming that the activity coefficient for each base is 1.0. Pyridine is infinitely miscible with water, but adding a *tert*-butyl group at the 2- or 4-position reduces the solubility to near 10^{-2} M. As expected, adding a second *tert*-butyl group substantially further reduces the solubility. These variations are consistent with the well-known observation that hydrocarbon moieties usually avoid the aqueous environment.

Nevertheless, our value for the aqueous solubility of 2,6-DTBP is more than 200 times smaller than the value report by Arnett and Chawla.[7] They attempted to prepare a saturated aqueous solution of 2,6-DTBP by stirring a mixture of the liquid base and water for a long period of time. At the end of this period, two layers formed and the concentration of 2,6-DTBP in the aqueous layer was calculated from spectroscopic measurements in the UV region. Because the absorbance of this solution was high, the absorbance of the aqueous solution was not determined directly but inferred from the absorbance of a diluted solution, in which the concentration of 2,6-DTBP was reported to be in the 10^{-5} M range. Following this procedure, they could have easily overlooked the formation of an emulsion of 2,6-DTBP in water, a phenomenon that we have seen to occur at high rates of stirring. This is probably the reason that their reported solubility for 2,6-DTBP is much higher than the value given here.

Table II. Thermodynamic of Vaporization for 2-TBP, 4-TBP, 2,4-DTBP and 2,6-DTBP at 25 °C

	$\Delta G_v°$, kcal mol^{-1}	$\Delta H_v°$, kcal mol^{-1}	$\Delta S_v°$, cal mol^{-1} deg^{-1}
2-TBP	3.42 ± 0.06	10.7 ± 0.4	24.4
4-TBP	4.23 ± 0.05 4.32^a	12.5 ± 0.2 12.4^a	27.7
2,4-DTBP	4.07 ± 0.07	11.7 ± 0.1	25.6
2,6-DTBP	4.68 ± 0.07 4.75^a	13.5 ± 0.1 13.5^a	29.6

[a] Reference 7.

Thermodynamics for Vaporization. For each of the pyridines in this study, the vapor pressures could be fitted to the Clausius–Clapeyron equation with a correlation greater than 0.98. The standard enthalpy and Gibbs free energy changes for vaporization ($\Delta H_v°$ and $\Delta G_v°$) were calculated from this relationship, assuming that the gases behave ideally. A summary of the derived values and the entropy changes calculated from these values are given in Table II, along with the standard deviations derived from the statistical analysis. Arnett and Chawla[7] determined very precise vapor pressure values for 4-TBP and 2,6-DTBP at several temperatures and have reported $\Delta H_v°$ and $\Delta G_v°$ values at 25 °C for the vaporization of these liquids that are in perfect accord with our values.

Thermodynamics for Gas-Phase Protonation. The gas-phase basicity for each pyridine (GB), defined as $-\Delta G_p°(gas)$, was determined from equilibrium measurements with reference bases that have known GB values on a scale fixed by the GB of methylamine (205.7 kcal mol^{-1}).[19] Since the neutral base and the corresponding cation have the same symmetry, the entropy of protonation is approximately equal to minus the standard entropy[22] of the proton (−7.75 cal mol^{-1} deg^{-1}). This approximation is applicable so long as the vibrational frequencies and the barriers to internal rotation of the base and the cation do not differ appreciably.[19] Using this approximation, we calculated the proton affinities (PA), defined as $-\Delta H_p°(gas)$, from the GB values.

Thermodynamics for the Transfer from the Gas Phase to the Aqueous Phase. In Table III we have summarized all the thermodynamic data determined by us for the *tert*-butylpyridines: columns one and two contain the GB and PA parameters derived from the gas-phase data; the next three columns contain the parameters derived from the pK_a data in water; the next three columns contain the parameters for the transfer of the neutral base from the gas phase to the aqueous phase derived from the calorimetric data, the solubility data, and the vapor pressure data; the last three columns contain the parameters for the transfer of

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[22] Altshuller, A. P. *J. Am. Chem. Soc.* **1955**, *77*, 3480.

Table III. Summary of the Thermodynamic[a] Data for Pyridine, 2-TBP, 4-TBP, 2,4-DTBP, and 2,6-DTBP at 25 °C: Protonation in the Gas Phase and Aqueous Phase and Transfer of the Cation[b,c] and Base from the Gas Phase to the Aqueous Phase

	GB	PA	$-\Delta G_p°$(aq)	$-\Delta H_p°$(aq)	$\Delta S_p°$(aq)	$\Delta G°_{g\to s}$(B)	$-\Delta H°_{g\to s}$(B)	$-\Delta S°_{g\to s}$(B)	$-\Delta G°_{g\to s}$-(BH$^+$)	$-\Delta H°_{g\to s}$-(BH$^+$)	$-\Delta S°_{g\to s}$-(BH$^+$)
pyridine	212.6	220.4	7.11	4.8	7.7	−2.80[d]	11.7[d]	29.8	57.8	65.9	27.2
2-TBP	218.6	226.4	7.8	6.4	4.7	−1.0	13.1	40.6	50.7	62.9	40.9
4-TBP	218.3	226.1	7.7	6.4	4.4	−2.5	14.2	39.2	52.4	64.3	39.9
2,4-DTBP	223.6	231.4	9.1	9.3	−0.7	1.0	11.4	41.6	45.0	59.1	47.3
2,6-DTBP	223	231	6.8	8.5	−5.7	1.3	12.5	46.3	43	60	56

[a] GB, PA, $\Delta G°$, and $\Delta H°$ values are in kcal mol^{-1}; $\Delta S°$ values are in cal mol^{-1} deg^{-1}. These values are based on the ideal gas and 1 atm standard state in the gas phase and the infinite dilution and 1 M one in the aqueous phase. [b] Calculated by using $\Delta G°_{g\to s}(H^+) = -260.5$ kcal mol^{-1}: Noyes, R. M. *J. Am. Chem. Soc.* **1962**, *84*, 513. [c] Calculated by using $\Delta H°_{g\to s}(H^+) = -269.8$ kcal mol^{-1}: Friedman H.; Krishnan C. V. "Water: A Comprehensive Treatise"; Franks, F. F., Ed.; Plenum Press: New York, 1973; Vol. 3, Chapter 1. [d] Andon, R. J. L.; Cox, J. D.; Herington, E. F. G. *J. Chem. Soc.* **1954**, 3188. Their values have been converted to the same standard states used in this paper.

the cations from the gas phase to the aqueous phase derived, via the thermodynamic cycle, from the parameters in the other columns. In all the calculations performed to derive these parameters, we have chosen to use the 1 atm and ideal gas standard state for the gas phase and the hypothetical 1 M and infinite dilution standard state for the aqueous phase. These parameters are not directly comparable to those derived by Arnett and Chawla,[7] who chose to use the infinite dilution and 1 mol fraction standard state based on Henry's Law for the aqueous phase. Transforming our values for $\Delta G°_{g\to s}$(B) for 4-TBP and 2,6-DTBP to their standard states yields −0.12 and 3.68 kcal mol^{-1}, respectively, compared to −0.19 and 0.40 kcal mol^{-1} given by them[7] for these parameters. The discrepancy between our value and their value for 2,6-DTBP is entirely due to the much higher solubility that they report for this compound.

Further direct comparisons between our values for the other parameters given in this paper for 2,6-DTBP and 4-TBP and those of Arnett and Chawla[7] cannot be made because they reported thermodynamic parameters for the substituted pyridines relative to those for pyridine. If the same standard states are used for each pyridine, these differences should not depend on the choice of standard states, and the differences calculated from our data can be compared without applying a correction to those reported by Arnett and Chawla.[7] The differences calculated from our GB, PA, and $\Delta G_p°$(aq) parameters are in excellent agreement with those tabulated by Arnett and Chawla.[7] Nevertheless, the relative $\Delta S°_{g\to s}$ values for 2,6-DTBP and the 2,6-DTBPH$^+$ cation calculated from our data, i.e., −16.5 and −29 cal mol^{-1} deg^{-1}, respectively, are substantially different from their values of −8.4 and −16.1 cal mol^{-1} deg^{-1}. Again these differences are due primarily to the difference between our and their solubility for 2,6-DTBP. The discrepancies found between our and their values for 4-TBP are much smaller, i.e., in the 3.0 cal mol^{-1} deg^{-1} range. However, the values calculated from our numbers for the differences between the enthalpy changes ($\Delta H°_{g\to s}(BH^+)$) for the transfer of the 2,6-DTBPH$^+$ and 4-TBPH$^+$ cations from the gas phase to the aqueous phase are within experimental error the same as those of Arnett and Chawla.[7] This is because these values can be calculated from the $\Delta H_s°$(HCl), $\Delta H_v°$, and PA parameters, for which our values and those of Arnett and Chawla[7] are in reasonable agreement.

Discussion

Gas-Phase Basicities. In previous gas-phase studies[6,19,23] on the protonation of alkyl-substituted pyridines and amines, the gas-phase basicities and proton affinities were found to increase as one increases the number and size of the alkyl groups. As can be seen from an inspection of the GB and PA values in Table III, such a trend also exists for the *tert*-butyl-substituted pyridines in this study. Presumably, this is due to a stabilization of the cation via a polarization of the bonds in the alkyl groups;[19] and as expected, the mean increase of 5.6 kcal mol^{-1} per *tert*-butyl group in the GB and PA values is larger than that found for methyl (2.9 ± 0.5 kcal mol^{-1}) or ethyl (4.0 ± 0.5 kcal mol^{-1}) groups by Aue and Bowers.[19] Also, the effect of adding two *tert*-butyl groups is slightly smaller than twice that found for one. Thus, the magnitude and near additivity of this substituent effect is in accord with what has been observed in previous gas-phase studies on the effect of methyl, ethyl, and isopropyl groups[12,19] on the basicities of amines and pyridines. Similar trends also have been recently reported by Meot-Ner and Sieck[12] in a series of gas-phase basicities for alkyl-substituted pyridines which were measured at higher temperatures and by a high-pressure mass spectrometric technique. The trends and additivity relationships derived from their data are similar to those seen here, but the substituent effects derived from their data are somewhat larger.

Further analysis of the gas-phase results in Table III provides the conclusion that one should discard the postulate of Brown and Kanner[2] that the size of the *tert*-butyl group might interfere with the attachment of a proton to the pyridine nitrogen of 2,6-DTBP and cause a steric compression of the N–H bond. Since 2,4-DTBP and 2,6-DTBP have essentially the same gas-phase basicities, we conclude that, at least in a thermodynamic sense, the N–H bond in the 2,6-DTBPH$^+$ cation is normal. On the other hand, our ion cyclotron resonance studies on 2,6-DTBP and also those of Brauman and Jasinski[20] are consistent with the conclusion that rates of proton transfer between cations and bases involving 2,6-DTBP are much slower than those for the other substituted pyridines. Even in the solution phase, the transfer of the proton from H_3O^+ to 2,6-DTBP is 50–70 times slower[24] than expected for a diffusion-controlled reaction. From these observations we conclude that the N–H bond in the 2,6-DTBPH$^+$ cation is normal, but apparently there does exist, in both the gas phase and solution phase, a steric hindrance to the rate of approach of the proton to the equilibrium position.

Aqueous Basicities. In the aqueous phase, $\Delta G_p°$(aq) for 2,6-DTBP is much lower than one would predict from the gas-phase results for other substituted pyridines[6,19,23] or the trends in the parameters for the aqueous phase found for the other pyridines in this series. A similar trend can be seen in alcohol-water solvents[2,3,9] and also in pure methanol.[8] Because 2,6-DTBP has the expected basicity in the gas phase, its abnormally low basicity in these solvents must be due to solvation effects. Recently, Arnett and Chawla[7] concluded that the low basicity of 2,6-DTBP was probably due primarily to a blockage of a water molecule from hydrogen bonding to the sterically hindered 2,6-DTBPH$^+$ cation. This should preclude the normal solvation of the 2,6-DTBPH$^+$ cation and cause its solvation to be nonspecific, as is thought to be the case with carbocations.[5] If such an interpretation is correct for the aqueous phase, then one would expect to see both unusual enthalpy and entropy effects for the protonation of 2,6-DTBP. This hypothesis can now easily be tested by comparing the gas-phase thermodynamic values for the alkyl-substituted pyridines with those determined in the aqueous phase.

For most substituted pyridines, the solvation of the neutral species and the protonated base is related to the magnitude of the proton affinity in such a way that the $\Delta G_p°$(aq) and $\Delta H_p°$(aq) values vary linearly with the gas-phase basicities.[6,19,23] As can be seen in Figure 1, where PA is plotted vs. $\Delta G_p°$(aq) and

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(23) Arnett, E. M.; Chawla, B.; Bell L.; Taagepera, M.; Hehre, W. J.; Taft, R. W. *J. Am. Chem. Soc.* **1977**, *99*, 5729.

(24) Bernasconi, C. F.; Carre, D. J. *J. Am. Chem. Soc.* **1979**, *101*, 2707.

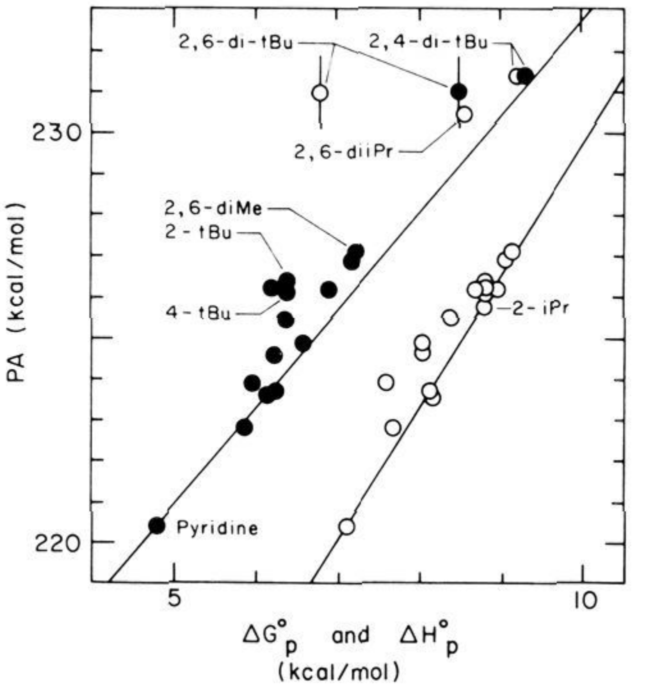

Figure 1. PA plotted vs. $\Delta G_p°$(aq) (○) and $\Delta H_p°$(aq) (●) in pure water for a series of alkyl-substituted pyridines.

$\Delta H_p°$(aq), such a linear variation is observed for all alkyl-substituted pyridines except 2,6-DTBP, 2,4-DTBP, and 2,6-diisopropylpyridines. In the plot of the aqueous Gibbs free energy data vs. PA, the point for 2,6-DTBP deviates from the line by 3.5 kcal mol^{-1}, whereas the points for 2,4-DTBP and 2,6-diisopropylpyridine deviate by nearly 1.2 and 1.5 kcal mol^{-1}, respectively. In the corresponding plot for $\Delta H_p°$(aq) vs. PA, all the alkyl-substituted pyridines, including 2,6-DTBP, 2,4-DTBP, and 2,6-diisopropylpyridine, have $\Delta H_p°$(aq) values that are linearly related to PA; i.e., none of the points for the alkyl-substituted pyridines fall substantially away from the line defined by the points for the other alkyl-substituted pyridines. From this analysis, we conclude that the magnitude of the enthalpy parameters associated with the solvation of both 2,6-DTBP and the 2,6-DTBPH^+ cation are not affected very much by the steric requirements and that the abnormally low basicity of 2,6-DTBP is primarily due to entropy effects. In the next sections, which contain our analysis of the thermodynamic parameters for the transfer of pyridines and corresponding pyridinium cations from the gas phase to the aqueous phase, we will attempt to discover the nature of the phenomena responsible for these entropy effects.

Transfer of Pyridines from the Gas Phase to the Aqueous Phase. According to the scaled particle theory, which is based on a hard-sphere model of liquids, transferring a molecule from the gas to the liquid phase requires an expenditure of energy because of the work needed to create a cavity in which to place the solute.[25] There are, however, many types of interactions that also must be included in the analysis: London dispersion forces; electrostatic interactions of dipoles, quadrupoles, ions, and induced dipoles; in the case of pyridines, hydrogen bonding to the pyridine nitrogen. In aqueous solutions there is also another type of interaction, called the hydrophobic interaction,[26,27] which is thought to exist by most investigators. This is now thought to be a structuring of the aqueous solution in the immediate vicinity[26,27] of the wall of the cavity containing the hydrocarbon moieties. Any discussion of the thermodynamic parameters for the transfer from the gas phase to the aqueous phase of molecules, like the pyridines included in this study, must include all of these; however, one might expect hydrophobic and hydrogen-bonding interactions to be responsible for the variations found for these parameters for the *tert*-butyl-substituted pyridines in this series.

Normally, hydrophobic interactions for alkyl groups produce large enthalpy and entropy terms that compensate each other, resulting in little or no change in the Gibbs free energy parameter. For alkanes, alcohols, and amines, such a compensation[28] has been observed in the thermodynamic parameters for the transfer from the gas to aqueous phase. This is also the case for 4-TBP, for which $\Delta H°_{g\to s}$(B) is 2.5 kcal mol^{-1} more exothermic than that for pyridine; but the $T\Delta S°_{g\to s}$(B) term compensates for this change and the Gibbs free energy parameter for 4-TBP is virtually the same as that for pyridine. Similar changes occur in the enthalpy and entropy terms of 2-TBP, but the compensation is incomplete: the enthalpy parameter is less exothermic than expected, and the entropy parameter is also slightly more negative than expected. For 2,4-DTBP and 2,6-DTBP the compensation is also incomplete and the enthalpy parameters are not as negative as expected from hydrophobic effects alone. Because this is the first systematic investigation on these two parameters for *tert*-butyl groups and the theoretical basis[26,27] for a discussion of hydrophobic effects is incomplete, we are not able to make a direct comparison of our data with other studies or theoretical calculations. Nevertheless, assuming that the variations are due to changes in (1) hydrophobic interactions around the *tert*-butyl groups and (2) hydrogen bonding of water molecules around the nitrogen, we probably should not be surprised to find deviations from the patterns seen for simpler systems. When *tert*-butyl groups are adjacent to a hydrogen-bonding center, two different regions of solvation overlap and this overlap could actually diminish the total solvation.

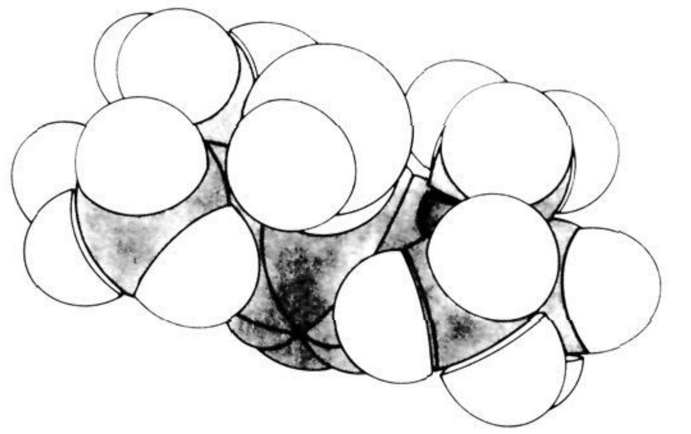

Figure 2. Tracing of a photograph of the complex formed by H_2O and 2,6-DTBP from CPK models, in which H_2O is hydrogen bonded to the pyridine nitrogen.

Admittedly, we do not have a complete understanding of all the changes found for the thermodynamic parameters for the transfer of pyridine, 2-TBP, 4-TBP, and 2,4-DTBP from the gas phase to the aqueous phase, but we conclude from our data that the solvation of these pyridines in water is normal. Because of this conclusion, we will assume during the remainder of this discussion that a water molecule is probably hydrogen bonded to the pyridine nitrogen of these pyridines in aqueous solution. If a water molecule is hydrogen bonded to the nitrogen in the aqueous 2,4-DTBP species but not to the nitrogen in the aqueous 2,6-DTBP species, as was suggested by Arnett and Chawla,[7] then $\Delta H°_{g\to s}$(B) for 2,6-DTBP should be substantially less negative than the corresponding value for 2,4-DTBP and $\Delta S°_{g\to s}$(B) for 2,4-DTBP should be more negative than that for 2,6-DTBP. As can be seen in Table III, $\Delta H°_{g\to s}$(B) for 2,6-DTBP is more exothermic than the value for 2,4-DTBP by 1.1 kcal mol^{-1}, and $\Delta S°_{g\to s}$(B) for 2,6-DTBP is 4.7 cal deg^{-1} mol^{-1} more negative than the value for 2,4-DTBP. Both of these observations are counter to predictions based on the postulate that a water molecule does not hydrogen bond to the pyridine nitrogen of 2,6-DTBP in aqueous solution.

Because of the following observations, we suggest that a water molecule can and probably does hydrogen bond to a significant fraction of the 2,6-DTBP molecules in aqueous solutions. First, such a complex can be easily constructed with space-filling models, as can be seen in the photograph (Figure 2) of the complex that we constructed from CPK models (Improved Corey–Pauling models made by The Ealing Corporation). Second, after a water molecule is placed at the hydrogen bonding distance (Figure 2) in the channel formed by the two *tert*-butyl groups, it cannot rotate

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[25] Pierotti, R. A. *J. Phys. Chem.* **1963**, *67*, 1140; **1965**, *69*, 281.

[26] Tanford, C. "The Hydrophobic Effect: Formation of Micelles and Biological Membranes"; Wiley: New York, 1981.

[27] Ben-Naim, A. "Hydrophobic Interactions"; Plenum Press: New York, 1980.

[28] Aue, D. H.; Webb, H. M.; Bowers, M. T. *J. Am. Chem. Soc.* **1976**, *98*, 318.

freely around the hydrogen bond. In addition, crowding in the channel in the complex also appears to hinder the rotation of the *tert*-butyl groups and two of the methyl groups on the *tert*-butyl groups. These reductions in the freedom to rotate around bonds will reduce the entropy of the 2,6-DTBP–H_2O hydrogen-bonded complex and cause the $\Delta S^\circ_{g\to s}(B)$ value for 2,6-DTBP to be substantially more negative than the value for 2,4-DTBP. Such a difference is exactly what is seen in Table III. Third, the existence of such a complex in aqueous solutions would permit the transfer of the proton from H_3O^+ to 2,6-DTBP to occur by the mechanism postulated by Grunwald and Ralph[29] and currently accepted as the correct one for proton transfers from the hydronium ion to amines, i.e.:

$$\geqslant N{:}\cdots HOH + H_3O^+ \rightarrow\ \geqslant N^+{-}H\cdots OH_2 + H_2O$$

Bernasconi and Carre,[24] who studied the kinetics of proton transfer from H_3O^+ to 2,6-DTBP in a 20% dioxane water mixture, concluded that their observations are consistent with this mechanism, even though the rate of transfer is 50–70 times slower than the diffusion-controlled limit found for other tertiary amines.

Transfer of Cations from the Gas Phase to the Aqueous Phase. When one is discussing the variations for the thermodynamic parameters for the transfer of a pyridinium cation from the gas phase to the aqueous phase, one must include in this analysis additional terms due to ion–solvent interactions. In most solvents, the two most important terms are the nonspecific electrostriction around the cation and the hydrogen bonding of a solvent molecule to the hydrogen on the pyridinium nitrogen. According to the Born model for electrostriction,[30–32] in which the ions in solution are represented by conducting, charged spheres in a dielectric continuum, the $\Delta G^\circ_{g\to s}(BH^+)$ and $\Delta H^\circ_{g\to s}(BH^+)$ values should be large and negative and decrease as the effective size of the ion increases. As can be seen, the values in Table III fit this description, at least in a general way, for the *tert*-butyl-substituted pyridines in this study. Hydrogen bonding of a water molecule to the hydrogen on the pyridinium nitrogen will also stabilize the cations in the aqueous solution, but it is difficult to separate this effect from the overall electrostriction of the ions caused by electrostatic interactions. In the gas phase, the attachment[12] of one water molecule to the pyridinium cation has a $\Delta H^\circ = -16.1$ kcal mol^{-1}, but this is only $^1/_4$ of the total enthalpy change for the overall process of transferring the cation from the gas phase to the aqueous phase. As the proton affinity of the pyridine increases, this enthalpy decreases, and for the 2,6-DTBPH$^+$ cation, $\Delta H^\circ = -12.5$ kcal mol^{-1} for the attachment of a water molecule to the cation. Except for the $\Delta G^\circ_{g\to s}(BH^+)$ value for the 2,6-DTBPH$^+$ cation, which is nearly 3 kcal mol^{-1} less negative than one would predict from the general trend, the range and trends seen in Table III for the $\Delta G^\circ_{g\to s}(BH^+)$ and $\Delta H^\circ_{g\to s}(BH^+)$ values are consistent with the predictions derived from the Born model and the known thermodynamics of hydrogen bonding to pyridinium cations in the gas phase.

If, as suggested by Arnett and Chawla[7] and Taft et al.,[5] the 2,6-DTBPH$^+$ cation is precluded from hydrogen bonding to a water molecule in aqueous solution, which certainly is not the case for the other pyridinium cations, then both the $\Delta H^\circ_{g\to s}(BH^+)$ and $\Delta S^\circ_{g\to s}(BH^+)$ values for the 2,6-DTBPH$^+$ cation should deviate substantially from those for the 2,4-DTBPH$^+$ cation. Quite remarkably, the $\Delta H^\circ_{g\to s}(BH^+)$ values for both cations are within experimental error the same. From this observation alone, we conclude that the hydrogen bonding of the solvent to the 2,6-DTBPH$^+$ cation in aqueous solution is similar to that for the other cations in this series. Furthermore, if electrostriction around the 2,6-DTBPH$^+$ cation was less than that for the other cations in the series or if a water molecule was excluded from attachment to the hydrogen on the pyridinium nitrogen, then the $\Delta S^\circ_{g\to s}(BH^+)$ value for the 2,6-DTBPH$^+$ cation would be greater than that for the 2,4-DTPBH$^+$ cation. Since the opposite is observed, we conclude that our analysis of the available thermodynamic data is not consistent with either of the postulates of Brown and Kanner[2] involving the reduced solvation of the 2,6-DTBPH$^+$ cation.

Our results are consistent, however, with the postulate of Hopkins and Ali[9] that a decrease in the rotational motions of the *tert*-butyl groups upon protonation is responsible for the unusually low basicity of aqueous 2,6-DTBP. Such a phenomenon could be at least partially responsible for the observations that $\Delta S^\circ_{g\to s}(BH^+)$ for the 2,6-DTBPH$^+$ cation is 6 cal mol^{-1} deg^{-1} more negative than that for the 2,4-DTBPH$^+$ cation, whereas the corresponding difference in the $\Delta S^\circ_{g\to s}(B)$ values for 2,6-DTBP and 2,4-DTBP is only −4.7 cal mol^{-1} deg^{-1}. As was noted in the introduction, Hopkins and Ali[11] have also presented spectroscopic evidence that is consistent with this conclusion: in aqueous methanol the rotation of the methyl groups and the rotation of the *tert*-butyl groups around the carbon–carbon bond are both more hindered in the 2,6-DTBPH$^+$ cation than in 2,6-DTBP. Both of these effects produce a loss of entropy for the 2,6-DTBPH$^+$ cation that does not occur for the other pyridinium cations.

Finally, what is most remarkable is the similarity between our results and those of Meot-Ner and Sieck,[12] which were independently derived from their gas-phase studies on the thermodynamics of hydrogen bonding of a water molecule to alkyl-substituted pyridinium cations. Meot-Ner and Sieck[12] clearly demonstrated that ΔH° for the attachment of a water molecule to an alkyl-substituted pyridinium cation is proportional to the PA of the pyridine and is not affected by steric interactions to any appreciable extent. In contrast to this result, they find a substantial entropy change for the attachment of a water molecule to the 2,6-DTBPH$^+$ cation and conclude from their analysis that 13.5 cal mol^{-1} deg^{-1} of this change is due to steric interactions which cause some of the barriers to internal rotation in the ion–water complex to be large. Both of these conclusions are in accord with the ones derived by us from the data in Table III; i.e., $\Delta H^\circ_{g\to s}(BH^+)$ is not affected much by steric requirements for *tert*-butyl-substituted pyridines, but $\Delta S^\circ_{g\to s}(BH^+)$ is markedly affected by steric interactions.

Furthermore, Moet-Ner and Sieck[12] concluded from a comparison of available pK_a data with their calculated parameters for steric effects that the 2,6-DTBPH$^+$ cation in the aqueous environment also forms a hydrogen-bonded complex that has a substantially reduced entropy. If we assume that barriers to internal rotation in 2,4-DTBP do not increase much upon protonation, as was concluded by Hopkins and Ali[11] from an analysis of their carbon-13 NMR data, then we readily conclude that the difference of 6 cal mol^{-1} deg^{-1} between the $\Delta S^\circ_{g\to s}(BH^+)$ values for the 2,4-DTBPH$^+$ and 2,6-DTBPH$^+$ cations is probably due to a lower than expected entropy for the aqueous 2,6-DTBPH$^+$ cation. From this analysis we conclude that the 2,6-DTBPH$^+$ cation in the aqueous environment has lost approximately 6 cal mol^{-1} deg^{-1} of entropy because of hindered rotations caused by steric interactions in the 2,6-DTBPH$^+$ cation–H_2O complex. The agreement between our calculated decrease in entropy for the 2,6-DTBPH$^+$ cation in water due to steric interactions and that determined by Meot-Ner and Sieck[12] in the gas phase may be fortuitous. But in light of the NMR and enthalpy data, we think this is unambiguous evidence that the same phenomena occur in both the gas phase and aqueous phase.

As is now obvious from the previous discussion, all of the available results are consistent with the following conclusions: (1) The enthalpy parameter for the solvation of the 2,6-DTBPH$^+$ cation is not affected appreciably in magnitude by steric requirements, but the entropy of the cation in the aqueous environment is markedly reduced by steric interactions. (2) Apparently, a hydrogen-bonded complex between a water molecule and the 2,6-DTBPH$^+$ cation does exist, both in the gas and the aqueous phase, but the barriers to internal rotation for the water molecule and the *tert*-butyl groups in the complex are substantial. *These restricted rotations in the complex are primarily responsible*

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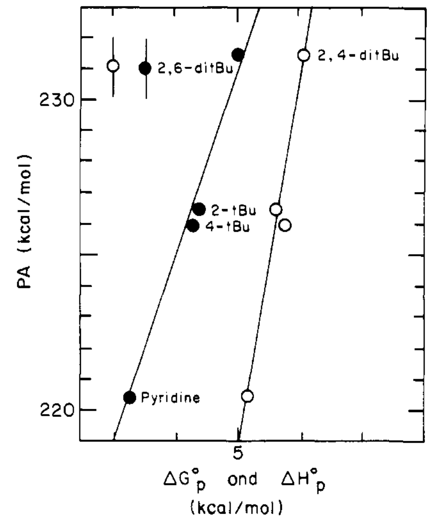

Figure 3. PA plotted vs. $\Delta G_p°$ (○) and $\Delta H_p°$ (●) in 90% methanol-water for a series of *tert*-butyl-substituted pyridines.

for the unusually low basicity of 2,6-DTBP in aqueous solutions.

Correlation of PA and the Basicities in Alcohol-Water Mixtures. Our conclusions derived from the data for the gas phase and aqueous phase can now be compared to a similar analysis of the basicities of *tert*-butyl-substituted pyridines in 90% methanol-water and 40.9% ethanol-water mixtures.[9] Since in these solvent mixtures only pK_a and $\Delta H_p°$ values are available for these compounds, we are restricted to an analysis in which the variations in the basicities in these solvent systems are compared to the corresponding ones in the gas phase. In Figure 3, where PA is plotted vs. both $\Delta G_p°$ and $\Delta H_p°$ for the 90% methanol-water system, one sees that the points for 2,6-DTBP deviate from the linear relationships found for the other pyridines in this study by 1.5 kcal mol^{-1} for $\Delta H_p°$ and by 3.0 kcal mol^{-1} for $\Delta G_p°$. Similar plots for the data in 40.9% ethanol-water have deviations of 2.5 kcal mol^{-1} in $\Delta H_p°$ and only 2.0 kcal mol^{-1} for $\Delta G_p°$. In pure methanol the point for 2,6-DTBP deviates nearly 4.8 kcal mol^{-1} from the linear trend found for PA vs. $\Delta G_p°$, whereas the point for 2,6-diisopropylpyridine deviates from the linear trend by 1.3 kcal mol^{-1}. We note, just as Hopkins and Ali[9] did earlier, that the abnormally low basicity of 2,6-DTBP in these solvents is due apparently to both enthalpy and entropy effects.

Again one is tempted to conclude that the specific solvation of the 2,6-DTBPH$^+$ cation, via hydrogen bonding of a solvent molecule to the pyridine nitrogen, has ceased or at least been reduced substantially. Nevertheless, we easily constructed from CPK models methanol and ethanol hydrogen-bonded complexes for 2,6-DTBP and the 2,6-DTBPH$^+$ cation. In the 2,6-DTBPH$^+$ cation-ethanol complex, steric crowding is much greater than for either the water or methanol complex, which could account for the larger enthalpy effect in the 40.9% ethanol-water mixture. There is, however, uncertainty about which one of the two types of solvent molecules is present in the specific solvation complexes in these mixtures. Actually, both types of hydrogen-bonded complexes could exist in the alcohol-water mixtures for both the neutral pyridines and their cations; thus, there might be different fractions of each type of complex for each of the *tert*-butyl-substituted pyridines. Consequently, further studies must be performed before one can complete a detailed analysis of the basicities of these pyridines in the alcohol-water solvents.

Registry No. 2-TBP, 5944-41-2; 4-TBP, 3978-81-2; 2,4-DTBP, 29939-31-9; 2,6-DTBP, 585-48-8; 2-TBP·H$^+$, 62907-59-9; 4-TBP·H$^+$, 40569-37-7; 2,4-DTBP·H$^+$, 62907-60-2; 2,6-DTBP·H$^+$, 62907-61-3; pyridine, 110-86-1; pyridine·H$^+$, 16969-45-2.

Ab Initio Theoretical Frequencies and Intensities in the Interpretation of Infrared Spectra

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Abstract: A comparison of the ab initio theoretical vibrational spectra with the experimental IR spectra of ethylene oxide and its tetradeuterio derivative shows that such calculated spectra can be of use in interpreting experimental results. However, it is crucial that relative intensities be calculated as well as frequencies.

It has been found that ab initio vibrational calculations can be useful in identifying unstable organic species.[1] The comparison of computed[2-5] and experimental spectra of cyclobutadiene and its deuterated derivatives has been useful in resolving the question of its structure (square or rectangular). Similarly, computed IR spectra of thiirene[6,7] have helped confirm its presence in an argon matrix. Since computed frequencies can be in error by as much as a few hundred wavenumbers at short wavelengths, they alone were not sufficient, but when frequencies were combined with computed intensities, fairly definite assignment of observed bands was possible.

In the two cases mentioned the primary purpose of the theoretical work was to aid in picking bands of the molecule sought from others in a reacting mixture. However, theoretical vibrational spectra can also help in interpreting IR bands of pure and otherwise well-characterized molecules. As an example we consider

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