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SELF-CONSISTENT ACIDITY SCALES OF NEUTRAL AND CATIONIC BRØNSTED ACIDS IN ACETONITRILE AND TETRAHYDROFURAN

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Dissertation is accepted for the commencement of the degree of Doctor of Philosophy in Chemistry on December 23, 2002, by the Doctoral Committee of the Department of Chemistry, University of Tartu.

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Commencement: March 10, 2003 at 2 Jakobi St., room 430

Publication of this dissertation is granted by University of Tartu

Tartu Ülikooli Kirjastus
Tiigi 78, Tartu 50410
Tellimus nr. 71

To my parents

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LIST OF ORIGINAL PUBLICATIONS

This thesis consists of four articles listed below and a review. The articles are referred in the text by Roman numerals I–IV. The review summarizes and supplements the articles.

I. Spectrophotometric Acidity Scale of Strong Neutral Brønsted Acids in Acetonitrile. Leito, I.; Kaljurand, I.; Koppel, I. A.; Yagupolskii, L. M.; Vlasov, V. M. *J. Org. Chem.* **1998**, *63*, 7868–7874.

II. Self-consistent Spectrophotometric Basicity Scale in Acetonitrile Covering the Range Between Pyridine and DBU. Kaljurand, I.; Rodima, T.; Leito, I.; Koppel, I. A.; Schwesinger, R. *J. Org. Chem.* **2000**, *65*, 6202–6208.

III. Sitting-Atop Complex Formation of 2,3,7,8,12,13,17,18-Octaethylporphyrin with Copper (II) Ion in Acetonitrile. Inamo, M.; Kohagura, T.; Kaljurand, I.; Leito; I.; *Inorg. Chim. Acta*, **2002**, *340*, 87–96.

IV. Acid-Base Equilibria in Nonpolar Media 2. Self-Consistent Basicity Scale in THF Solution Ranging from 2-methoxypyridine to EtP_1(pyrr) Phosphazene. Rodima, T.; Kaljurand, I.; Pihl, A.; Leito, I.; Koppel, I. A.; *J. Org. Chem.* **2002**, *67,* 1873–1881.

ABBREVIATIONS

AN	acetonitrile
AN	acceptor number
CIP	contact ion-pair
D	dielectric constant
DBU	2,3,4,6,7,8,9,10-octahydropyrimido[1,2-a]azepine (1,8-diazabicyclo[5,4,0]undec-7-ene)
dma	N,N-dimethylamino
DMAP	N,N-dimethylpyridin-4-amine (4-(N,N-dimethylamino)pyridine)
DMSO	dimethylsulfoxide
DN	donor number (kcal/mol)
EPA	electron pair acceptor
EPD	electron pair donor
HBD	hydrogen bond donor
N	number of points in statistical analysis
NMR	nuclear magnetic resonance
$R'P_n(R'')$	phosphazene (iminophosporane)
PhTMG	N,N,N',N'-tetramethyl-N''-phenylguanidine
pyrr	N-pyrrolidino
s	standard deviation in statistical analysis
SSIP	solvent-separated ion-pair
T	temperature in Kelvin
TBD	1,3,4,6,7,8-hexahydro-2H-pyrimido[1,2-a]pyrimidine (1,5,7-triazabicyclo[4.4.0]dec-5-ene)
Tf	trifluoromethanesulfonyl (CF_3SO_2–)
THF	tetrahydrofuran
TMG	N,N,N',N'-tetramethylguanidine
Tos	4-toluenesulfonyl (4-$CH_3C_6H_4SO_2$–)
UV	ultraviolet
Vis	visible

INTRODUCTION

Compounds, which function as Brønsted acids and/or bases are of extreme importance. These form the backbone of biological life and have found technological applications since very early times of human civilization. Nowadays they find very wide applications in practical (reagents in organic synthesis,[1–4] catalysts in industry[5]) and theoretical[6–9] fields of chemistry. Exact and reliable quantitative data describing the acid-base properties of compounds and dependence of these properties from medium are very important in applying acids and bases in various fields of chemistry and in designing new compounds with desired properties.

The intrinsic acidity or basicity can be measured in gas-phase experiments or theoretically calculated in simpler cases. Up to the present a lot of gas-phase acid-base data[10,11] have been collected for acids and bases. But in this field there is enough space for improvement — for example basicities of only a few superbases (by definition the superbases have *GB* (gas–phase basicity) over 239 kcal/mol)[12] have been determined. The reason is the lack of suitable reference bases,[13] low volatility of strong bases, long stabilization periods and tendency of strong bases to undergo fragmentation in the experiments.

In condensed media the history of investigation of acid-base properties extends back to the last decades of 19th century and a huge amount of compounds have been studied. Although a vast number of acidity and basicity data in condensed media have been collected,[14–17] the situation is still far from ideal. Contrary to the success in creating new acids and bases with improved properties, significant experimental difficulties have been met here and to date little consistent data is available especially for strong acids and bases. In some cases the situation is almost ridiculous. For the pK_a of perchloric acid in aqueous solution, for example, the values found in the literature range from –1.6 to –14.[6] The uncertainty exceeds 12 orders of magnitude!

While studying the acid-base properties of compounds in condensed media one should always keep in mind that there is no such thing as "*universal solvent*" which is suitable for all compounds without exception. Every known solvent has influences and limitations that one should consider. The acid and base strength of solvent, various association processes depending on solvent and solute polarity, decomposition of compounds or solvent are only some effects that have to be considered while measuring the acid-base properties in condensed media. In water, the most common media for acid-base studies, for example, both strong acids and bases are leveled up and only the acid-base properties of moderately strong acids and bases can be reliably measured. Stronger acids and bases can be reliably investigated in solvents that have lower basicity or acidity, respectively, and are inert if exposed to these compounds. In such solvents the acid-base strengths are not leveled and the intrinsic properties of bases are expressed to higher extent. Several solvents, including dimethyl-

sulfoxide, acetonitrile, THF etc. have found wide application as media for studies of strong acids and/or bases, a vast number of papers and several compilations have been published, truly consistent acidity scales in DMSO[14,15,17] and THF[18–20] exists. Although AN has been also very popular solvent for acid-base studies and a vast number of measurements in AN have been carried out, the situation is still not good. The data obtained by different authors even by the same method often lack consistency and contain significant gaps in the region of strong bases and acids. The measured values deviate often by up to one or more pK_a units, that is far more than stated with the experimental errors.

The goal of this study was to improve the situation in this field by building reliable continuous self-consistent and sufficiently wide spectrophotometric acidity and basicity scales in AN and a basicity scale in THF and to include different types of compounds to bridge earlier results together.

1. CONCEPTS

1.1. Acid-base equilibria in condensed media

Brønsted and Lowry created a theory that describes the proton transfer from acid to solvent with forming of conjugate base. Further studies of this reaction have shown that this acid dissociation equilibrium could be divided into several steps. Thus following consecutive equilibria will set up when neutral ($z = -1$) acid HA^{z+1} reacts with solvent S to give away proton:

$$HA^{z+1} + S \underset{}{\overset{K_1}{\rightleftarrows}} [HA^{z+1}..S] \overset{K_2}{\rightleftarrows} [A^z SH^+] \overset{K_3}{\rightleftarrows} [A^z]\,[SH^+] \overset{K_4}{\rightleftarrows} A^z + SH^+ \qquad (1)$$

Here A^z is its conjugate anionic base. The overall product of these consecutive equilibrium constants: $K_1 \cdot K_2 \cdot K_3 \cdot K_4 = K_a$ expresses acid strength of the compound. To make numerical data comparable, as a measure of basicity of neutral base A^z ($z = 0$) is generally presented the proton transfer equilibrium constants from its conjugate cationic acid HA^{z+1} to solvent molecule described with the same equation. Here an important point must be noticed: on dissociation of neutral acid there is a formation of two charged particles of opposite sign, whereas on dissociation of cationic acid there is charge transfer. In fact, to be correct and to maintain electroneutrality a negative ion A_1^{z-1} should be introduced, when $z = 0$. This enables several new equilibria to appear on the scene:

$$HA^{z+1} + A_1^{z-1} \overset{K_5}{\rightleftarrows} [HA^{z+1}]\,[A_1^{z-1}] \overset{K_6}{\rightleftarrows} [HA^{z+1}..\,A_1^{z-1}] \overset{K_7}{\rightleftarrows}$$

$$\overset{K_8}{\rightleftarrows} [A^z HA_1^z] \rightleftarrows A^z + HA_1^z \qquad (2)$$

The acid dissociation constant K_a of equilibrium in eq 1 expresses:

$$K_a = \frac{a(SH^+) \cdot a(A^z)}{a(HA^{z+1})} \qquad (3)$$

and the negative logarithm pK_a is presented as a measure of the strength of an acid HA^{z+1} in particular medium:

$$pK_a = -\log \frac{a(SH^+) \cdot a(A^z)}{a(HA^{z+1})} \qquad (4)$$

In media of good ability to separate and stabilize ions, the situation where these are separated to infinity, is favored. In media of poor ability to separate and stabilize ions these tend to aggregate or form conjugate complexes. The two important side reactions are homo- (eq 5) and heteroconjugation (eq 6) reactions:

$$A_1^z + HA_1^{z+1} \rightleftarrows [A_1 \cdot\cdot HA_1]^{2z+1}, \quad (5)$$

$$A_1^z + HA_2^{z+1} \rightleftarrows [A_1 \cdot\cdot HA_2]^{2z+1}. \quad (6)$$

It is generally observed that the extent of homoconjugation increases with increasing number of acidic hydrogen atoms in the molecule. Also poor steric hindrance of the protonation/deprotonation center is a source of increased homoconjugation.[21] The easiest way to decrease the influence of the association processes on acid-base equilibria without changing media, is to use research methods that allow use of very dilute solutions. Ion-pairing is third important type of side reactions, which extent depends on the solvent properties, size of the ions and the charge distribution in ions; hydrogen bonding and specific solvation possibility. The general trend is that small ions tend to form solvent-separated ion-pairs (SSIP) (eq 7) while large ions with delocalised charge tend to form contact ion-pairs (CIP) (eq 8).

$$HB^+ + A^- \rightleftarrows HB^+_s \cdot A^-_s \quad (7)$$

$$HB^+ + A^- \rightleftarrows [HBA]_s \quad (8)$$

Increase in ion-solvating and ions separating power of the solvent favors formation of the solvent-separated ion-pairs. Formation probability of higher aggregates must be also considered.

1.2. The role of solvent on acid-base reactions

As it already appeared in previous section, several medium properties have an important influence on the acid-base reactions in condensed media. From these, in combination of the solute molecule properties (charge and its location, polarity, polarizability etc.) and on the mole ratio (concentration) of solvent(s) and the solute(s) depends the character and extent of the results (solubility, ionization, dissociation, aggregation etc.) of its interactions with solvent and other solutes.[22]

One of the most important solvent properties in acid-base chemistry — *ionizing* power of the solvent depends mainly on its ability to be an electron pair acceptor (EPA) or electron pair donor (EPD). This property is empirically described with donor number (*DN*) and acceptor number (*AN*), higher values mean that solvent has higher ability to ionize neutral ionogen molecules and to stabilize the ions formed. Ionizing power is coordination dependent; thus it expresses on full degree on particular solute-solvent complex if it has no steric restrictions.

Dielectric constant (*D*) of solvent represents its ability to reduce the coulombic interaction between charged particles in solution, to separate or to *dissociate*

and to orient dipole molecules. Solvents with large D are called *polar* contrary to ones with low D are called *apolar* or *nonpolar*.

Brønsted *acid-base properties* of solvent have also important influence on the strengths of acids and bases. These properties are a special case of EPD-EPA properties. Solvent's ability to donate (eq 9) or accept (eq 10) a proton characterizes these properties:

$$SH \rightleftarrows H^+ + S^- \tag{9}$$

$$SH + H^+ \rightleftarrows SH_2^+ \tag{10}$$

A sum (eq 11) of these processes is described quantitatively by autoprotolysis constant (K_{auto}) of the solvent:

$$2\,SH \overset{K_{auto}}{\rightleftarrows} SH_2^+ + S^- \tag{11}$$

The lower the K_{auto}, the wider is the range of acid and/or base strengths, which can exist in solvent. If solvent molecule has no protons at all or the reaction in eq 9 is unfavored then the solvent is called *aprotic*, if the acid dissociation of solvent is present then the solvent is called *protic*. In *amphiprotic* solvent both reactions (eqs 9 and 10) are present. Low acid-base properties make the solvent good *differentiating* solvent. Significantly acidic or basic nature of solvent makes it *leveling* solvent for basic and acidic solutes respectively. Hydrogen bonding plays a particularly important role in the interactions between ions and solvents. Protic solvents stabilize anions better and aprotic solvents stabilize cations better.

The combination of these three solvent properties — ionizing power, dissociative power and acid-base properties and the nature of solute molecules determine the suitability of solvent for acid-base studies of particular compounds. For example, water with its high EPD-EPA ability, high dielectric constant (D = 78.4) and relatively high acid-base properties (pK_{auto} = 14) is good ionizing and dissociating solvent, but also leveling solvent for both strong acids and bases. It is very eager to form hydrogen-bonded complexes and thus it largely masks the intrinsic acid-base properties of superacids and -bases. If solvent is good EPD but poor EPA it stabilizes well cations but anions poorly and these tend to form aggregates described in previous section.

It appears, that generally appropriate solvent to study acid-base reactions over the wide range, should have very low acid-base properties, but be relatively polar to separate and stabilize ions.

1.3. Experimental methods of pK_a determination in condensed media

There is no universal experimental method that is able to describe quantitatively all the processes appearing in the studies of acid-base equilibria in condensed media. Several methods have been developed and described,[23,24] their applicability and advantages over others depend on properties of solvent and compounds under study. It is obvious that the combined use of different methods gives more and reliable information, thus the combined use, if possible, is preferred. On choosing the appropriate method or combination of methods thorough analysis of the particular system should be carried out on keeping in mind an ultimate goal. Undesirable side-reactions and effect of impurities are only few aspects to consider besides the others mentioned in previous sections. Here only the main methods used for studying of acid-base equilibria in aprotic dipolar media are briefly discussed.

Potentiometry has been by far most popular experimental technique for studies of acid-base equilibria both in water and other protic media. It has an advantage over other methods that it is specific for the measurement of solvated proton activity and it is applicable over wide range of proton activities. Disadvantages of the method are: need to calibrate the electrode system with buffer solutions that make it indirect method; need to use relatively high concentrations in which ion association processes will play particularly important role in media of poor ability to stabilize free ions (weakly solvating media); need to know the activities of analyte and its conjugate compound that are problematic to determine in nonaqueous media at relatively high concentrations; and potential drift electrode system in nonaqueous solvents.

Conductimetric methods base on evaluation of the limiting equivalent conductivity of an appreciably dissociated electrolyte from the dependence of equivalent conductivity from concentration. These methods are non-specific and thus do not distinguish between simple acid dissociation and homoconjugation. These methods are sensitive towards the ionizable impurities of solvent and analyte.

UV-Vis spectrophotometric methods base on light absorption difference of acid (or base) and conjugate base (or acid). Generally the exact analytical concentrations of compounds are needed and the extent of association processes, if present, should be determined with other methods. Advantages of these methods are that very low concentrations can be used and thus in several solvents of poor ability to stabilize ions the picture of association processes is simpler than with the above methods. Undesirable side reactions generally affect the UV-Vis spectra of the system and their presence is easily determinable. Disadvantages of these methods are that compounds must have difference in light absorption spectra of acid (or base) and conjugate base (or acid). Solvent

must be transparent in the analytical wavelength region. The results of these methods can be sensitive to the acidic and basic impurities.

Several other experimental techniques have found application, let only mention some of these: NMR (side reactions are easy to determine, high concentrations are needed, indirect method), voltammetry, IR (applicable for hydrogen bonding and association studies) and so on.

Besides the experimental techniques intensive work is in progress to describe or at least to estimate the processes and their extent arising on acid-base equlibria in condensed media with quantum chemistry and correlation methods. The crucial problem of quantum chemistry methods is the very high computing power necessary to take into account all electronic, steric and medium effects. The situation of theoretical gas-phase acid-base studies is much better as this medium is virtually free from medium effects. These methods supplement the overall picture and there is no doubt that in future these methods take an equivalent place beside or probably override the traditional methods in describing the acid-base equilibria in condensed media.

1.4. Pure UV-Vis spectrophotometric method for ΔpK_a determination

It appeared in previous section that UV-Vis spectrophotometric methods have several advantages over others. To exclude the necessity for measuring the hydrogen ion activity (see eq 1) we have developed a "pure" UV-Vis spectrophotometric method. In this method proton distribution equilibrium between two compounds HA_1^{z+1} and HA_2^{z+1} respectively, is studied:

$$HA_2^{z+1} + A_1^z \overset{K}{\rightleftarrows} A_2^z + HA_1^{z+1} \qquad (12)$$

The negative logarithm of the equilibrium constant K measures the difference of acidities of the acids HA_1^{z+1} and HA_2^{z+1} at given conditions:

$$\Delta pK_a = pK_a(HA_2^{z+1}) - pK_a(HA_1^{z+1}) = -\log K = \log \frac{a(A_1^z) \cdot a(HA_2^{z+1})}{a(HA_1^{z+1}) \cdot a(A_2^z)} \qquad (13)$$

The relationships of basic and acidic form of compounds in eq 13 are known as indicator ratios:

$$I_1 = \frac{a(A_1^z)}{a(HA_1^{z+1})} \quad \text{and} \quad I_2 = \frac{a(A_2^z)}{a(HA_2^{z+1})}. \qquad (14)$$

From the eq 13 it appears that then measuring the relative acidities of two compounds there is no need to measure the activity of the hydrogen ion $a(SH^+)$.

Advantages and disadvantages of the "pure UV-Vis spectrophotometric method". The presence of the two compounds in the same solution eliminates many possible sources of error or reduces their influence: (1) The disturbing effects (traces of water in the medium, concentration errors) affect both compounds with the same magnitude and are expected to partially cancel out. Thus one can expect that the effect of traces of water on the measurements will be less pronounced than in such measurements where pK_a of a single compound is measured at a time. (2) The solutions can be very dilute in certain cases (π-electron rich systems conjugated with acidity center) to lower the extent of unwanted competing equilibria or to clarify the picture. (3) Low concentrations allow use of very weak buffering with this method. If a minor acidity change of the solution occurs, it will affect both of the compounds and its effect will cancel out. (4) An important point is that the method eliminates the need for quantitative measurement of acidity of the medium. (5) Concentrations of solutions must be known only approximately, for maximum absorbances to fall into the range of 0.5 to 1.2 AU on applicable wavelengths.

A disadvantage of the method is that acid-base properties of only such compound can be measured that absorb in the UV-Vis spectral region and for which the spectra of the acid and the base forms are different at least in details. Also the pK_a-s of the compounds must not be very different from each other (preferably not more than 1.5–2 pK_a units).

2. EXPERIMENTAL

2.1. Calculation methods for pure UV-Vis spectrophotometric ΔpK_a determination method

General essence of calculation methods of pure UV-Vis spectrophotometric method is as follows. According to the Lambert-Beer law the absorbance A_x^{λ} of compound X in a layer of unit thickness at wavelength λ expresses as follows:

$$A_X^{\lambda} = [X]\varepsilon_X^{\lambda} \tag{15}$$

where ε_x^{λ} is a molar absorbance coefficient at given wavelength. If in solution is two partially dissociated acids (that is four separate compounds), then at wavelength λ expresses the net absorbance assuming the solvent absorbance is compensated:

$$A^{\lambda} = [HA_1^{z+1}]\varepsilon_{HA_1^{z+1}}^{\lambda} + [A_1^{z}]\varepsilon_{A_1^{z}}^{\lambda} + [HA_2^{z+1}]\varepsilon_{HA_2^{z+1}}^{\lambda} + [A_2^{z}]\varepsilon_{A_2^{z}}^{\lambda} \tag{16}$$

If we take a net concentration of acid and base form equal to one, we can express the concentration of acid forms:

$$[HA_1^{z+1}] = 1 - [A_1^{z}] \tag{17}$$

$$[HA_2^{z+1}] = 1 - [A_2^{z}]. \tag{18}$$

and the eq 16 can be rearranged:

$$A^{\lambda} = \varepsilon_{HA_1^{z+1}}^{\lambda} + \varepsilon_{HA_2^{z+1}}^{\lambda} + [A_1^{z}](\varepsilon_{A_1^{z}}^{\lambda} - \varepsilon_{HA_1^{z+1}}^{\lambda}) + [A_2^{z}](\varepsilon_{A_2^{z}}^{\lambda} - \varepsilon_{HA_2^{z+1}}^{\lambda}) \tag{19}$$

If we take the molar absorbance coefficients of pure acid forms to the left side of the equation and divide both sides of the equation with $(\varepsilon_{A_2^{z}}^{\lambda} - \varepsilon_{HA_2^{z+1}}^{\lambda})$, we get:

$$\frac{A^{\lambda} - \varepsilon_{HA_1^{z+1}}^{\lambda} - \varepsilon_{HA_2^{z+1}}^{\lambda}}{(\varepsilon_{A_2^{z}}^{\lambda} - \varepsilon_{HA_2^{z+1}}^{\lambda})} = [A_1^{z}]\frac{(\varepsilon_{A_1^{z}}^{\lambda} - \varepsilon_{HA_1^{z+1}}^{\lambda})}{(\varepsilon_{A_2^{z}}^{\lambda} - \varepsilon_{HA_2^{z+1}}^{\lambda})} + [A_2^{z}] \tag{20}$$

This equation describes the line with a slope $[A_1^{z}]$ and intercept $[A_2^{z}]$. At given wavelength all the ε^{λ}-s are constant and are easy to determine from the spectra of solutions of pure acid and base forms of compounds. At different wavelengths of solution of certain acidity are all the members except concentrations in equation 20 variables and with regression analysis the latter ones are determinable. Using the normalized concentrations from eq-s 17 and 18 and equation 13 we get:

$$\Delta pK_a = \log\frac{[H^+][A_1^{z}]}{(1-[A_1^{z}])} - \log\frac{[H^+][A_2^{z}]}{(1-[A_2^{z}])} = \log\frac{[A_1^{z}](1-[A_2^{z}])}{(1-[A_1^{z}])[A_2^{z}]} \tag{21}$$

The method is universal, but to employ it one needs spectra of pure acid and base forms of the compounds to determine corresponding ε^{λ}-s. Also one needs to know the ratio of concentrations of compounds in mixture and pure forms.

Sometimes, depending on spectra of compounds, it is possible to use very simple and elegant calculation methods to get indicator ratios. On the simpler cases one needs only the spectra of acid and base form of the mixtures and a set spectra of mixtures of variable acidity. Some examples of different cases are given here.

a) If there is a wavelength λ, at which neither forms of one compound do not absorb and basic form of second compound absorb

$$\varepsilon^{\lambda}_{HA_1^{z+1}} = \varepsilon^{\lambda}_{HA_2^{z+1}} = \varepsilon^{\lambda}_{A_2^{z}} = 0 \quad \text{and} \quad \varepsilon^{\lambda}_{A_1^{z}} \neq 0$$

then the eq 19 simplifies:

$$A^{\lambda} = [A_1^z]\varepsilon^{\lambda}_{A_1^z} \tag{22}$$

and the indicator ratio expresses then:

$$\frac{[A_1^z]}{[HA_1^{z+1}]} = \frac{A^{\lambda}}{A^{\lambda}_{A^z} - A^{\lambda}} \tag{23}$$

A^{λ} is absorbance of solution where both compounds are in acid and base form and $A^{\lambda}_{A^z}$ is absorbance of solution having both compounds in base form.

b) If one compound has an isosbestic point at certain wavelength that is both acid an base form have same absorbance coefficients and only basic form of second compound absorbs

$$\varepsilon^{\lambda}_{HA_2^{z+1}} = \varepsilon^{\lambda}_{A_2^{z}} \neq 0 \neq \varepsilon^{\lambda}_{A_1^{z}} \quad \text{and} \quad \varepsilon^{\lambda}_{HA_1^{z+1}} = 0$$

then eq 19 simplifies:

$$A^{\lambda} = [A_1^z]\varepsilon^{\lambda}_{A_1^z} + \varepsilon^{\lambda}_{HA_2^{z+1}}. \tag{24}$$

The indicator ratio of second compound expresses then as follows:

$$\frac{[A_1^z]}{[HA_1^{z+1}]} = \frac{A^{\lambda} - A^{\lambda}_{HA^{z+1}}}{A^{\lambda}_{A^z} - A^{\lambda}} \tag{25}$$

$A^{\lambda}_{HA^{z+1}}$ is here a net absorbance of mixture solution where both compounds are in acid forms.

c) If there is a spectrum of the mixture where both compounds are in base form then the net absorbance from eq 16 we get:

$$A^{\lambda} = [A_1^z]\varepsilon^{\lambda}_{A_1^z} + [A_2^z]\varepsilon^{\lambda}_{A_2^z} \tag{26}$$

The rightsided members of this eq can be expressed as absorbances of pure compounds in base forms multiplied by coefficients b_n:

$$[A_1^z]_{mixture}\,\varepsilon^{\lambda}_{A_1^z} = b_1^{A^z}\,A^{\lambda}_{A_1^z\,pure} \tag{27}$$

and

$$[A_2^z]_{mixture}\,\varepsilon^{\lambda}_{A_2^z} = b_2^{A^z}\,A^{\lambda}_{A_2^z\,pure} \tag{28}$$

These coefficients $b_n^{A^z}$ are constant over the wavelength range, where ε^{λ}-s do not equal to zero. From the combination of eq-s 26, 27 and 28 is possible to calculate from the spectrum of the mixture of compounds in base form, and from the spectra of both compounds in base form coefficients $b_1^{A^z}$ and $b_2^{A^z}$ using least squares minimization over a wavelength range by minimizing S_p:

$$S_p = \sum_{\lambda}\left[A^{\lambda} - \left(b_1^{A^z}\,A^{\lambda}_{A_1^z\,pure} + b_2^{A^z}\,A^{\lambda}_{A_2^z\,pure}\right)\right]^2 \tag{29}$$

These coefficients $b_1^{A^z}$ and $b_2^{A^z}$ show the ratio of concentrations in mixture and pure forms. Analogously the absorbance of the mixture solution where both compounds are in acid and base form by combining eq-s 19, 15, 27, 28 and introducing for both compounds a dissociation level α_n, which shows the ratio of base form to analytical concentration we get:

$$A^{\lambda} = b_1^{A^z}A^{\lambda}_{HA_1^{z+1}\,pure} + b_2^{A^z}A^{\lambda}_{HA_2^{z+1}\,pure} + \alpha_1 b_1^{A^z}\left(A^{\lambda}_{A_1^z\,pure} - A^{\lambda}_{HA_1^{z+1}\,pure}\right) + \alpha_2 b_2^{A^z}\left(A^{\lambda}_{A_2^z\,pure} - A^{\lambda}_{HA_2^{z+1}\,pure}\right) \tag{30}$$

from this eq using the least squares minimization over the wavelength range described in eq 29 by minimizing S_s respective α_1 and α_2 for compounds at different acidities are found

$$S_s = \sum_{\lambda}\left\{A^{\lambda} - b_1^{A^z}A^{\lambda}_{HA_1^{z+1}\,pure} - b_2^{A^z}A^{\lambda}_{HA_2^{z+1}\,pure} - \left[\alpha_1 b_1^{A^z}\left(A^{\lambda}_{A_1^z\,pure} - A^{\lambda}_{HA_1^{z+1}\,pure}\right) + \alpha_2 b_2^{A^z}\left(A^{\lambda}_{A_2^z\,pure} - A^{\lambda}_{HA_2^{z+1}\,pure}\right)\right]\right\}^2 \tag{31}$$

These α_1 and α_2 are substituted to eq 13 and the ΔpK_a expresses then:

$$\Delta pK_a = \log\frac{\alpha_1(1-\alpha_2)}{(1-\alpha_1)\alpha_2}. \tag{32}$$

This so called least-squares of linear combination method is universal and can be used when compounds have overlapping absorbances, the only limit is that, the spectra must not be identical. From our experiments we have concluded that if the compounds have similar shape absorbance spectra but the difference of absorbance maxima is at least 6 nm, then this calculation method is usually well applicable. From UV-Vis spectrophotometric data it is possible to calculate with good confidence level ΔpK_a values up to 2.5 units. Usually the ΔpK_a values ob-

tained using different data treatment methods agreed well. The raw spectrophotometric data was imported to and calculations were done in spreadsheet calculation program MS EXCEL.

2.2. Chemicals and solvents

The origin or synthesis and purification of acids and bases used, is described in detail or references are given in publications I–IV. Special thanks to Dr Toomas Rodima and Dr Ivo Leito for synthesizing and purification of the aryl phosphazenes and sulfonimides-sulfonic acids, respectively. The experimental method used set only few demands on the criteria of purity of the chemicals. The most important is that these must not contain impurities that change their UV-Vis absorption spectra in the region of wavelengths where calculations are carried out during the change in acidity of medium. This can be easily checked by titrating solution containing pure acid or base with an acidic or basic titrant, respectively. Any irregular behavior on the set of spectra or deviation of isosbestic points proves the presence of unwanted impurities. Reversibility of proton change processes could be easily checked by back titration. The presence of unwanted side reactions could be also determined from UV-Vis spectra as usually these reactions captivate the compounds under study and thus cause change in their spectra. The presence of acidic or basic impurities has little effect on the overall equilibria, as they influence both compounds simultaneously and in a similar way. Of course it was still preferred to keep the concentrations of active (acidic and basic impurities as well as compounds that promote association processes) and inert impurities as low as possible. The acidic and basic titrants must not have absorption spectra in UV-Vis region where the calculations have made.

Requirements on solvents are generally similar to these of chemicals but stringent on some points. Solvent must be transparent and must not have change in its absorbance upon addition of transparent acidic or basic titrant in UV-Vis wavelength region, where the ΔpK_a calculations are carried out. High concentrations of water and other impurities should be avoided, as these may change significantly some important medium properties (see Section 1.2). Carefully house purified or commercial AN and THF of extra purity and dryness (both Romil, water content <0.005%) were used in all of experiments. The water content of pure AN and titrated solutions of AN was checked with classical Karl Fischer titration. For details see Experimental sections of publications I–IV.

2.3. Experimental setup

Due to the possible sensitivity of the acids and bases to the moisture, oxygen and carbon dioxide special care was taken to avoid or minimize the contact with these. In publication I the vials, solvent bottle and the spectrophotometer cell (all capped with rubber septa) were interconnected with thin PTFE tubing, through which the solutions were siphoned by means of aspiration with a syringe. The air, which replaced the solutions, was dried with anhydrone. All standard solutions were made fresh daily.

Weighing operations (except the weighing of TfOH for the standard acid solution in AN and for some THF experiments), preparation of all solutions (all solutions were made daily), titration and the spectrophotometric measurements of publication II, III and partially IV were carried out in Mecaplex glovebox in an atmosphere of dry nitrogen. For continuous drying and purifying the atmosphere in the glovebox from volatile basic and acidic contaminants was used molecular sieves (Aldrich, 4 Å), powdered P_2O_5 and KOH pellets. Remaining titration experiments of publication IV were carried out in professional MBraun glovebox in the atmosphere of argon that was constantly circulated through a purification system containing activated carbon, molecular sieves and activated copper for removal of volatile organics, water vapor and oxygen, respectively. The residual concentrations of water and oxygen in the atmosphere of the glovebox during the measurements were constantly monitored and were generally below 1 ppm.

The computer controlled Perkin-Elmer Lambda 2S and Lambda 40 UV-Vis spectrophotometers were used for all spectral measurements. For working in gloveboxes an external sample compartment (ESC) was used. The 2 meters long quartz-fiber light conductor cables of the ESC were guided through the walls of the gloveboxes and fastened air-tightly. A reference cell with pure solvent was placed in the spectrophotometer cell holder.

When working in gloveboxes the spectrophotometer cell was closed with hollow PTFE stopper with PTFE/silicone septa and open-top screw-cap. The concentrations of individual acids or bases were usually in the 10^{-5} M range and their total concentration in our experiments generally did not exceeded 2.2×10^{-4} M. Higher concentrations (up to 0.23 M) were used in THF for studying relative ion-pair basicity dependence from concentrations. The acidic and basic titrants were added into the cell through the septa using Hamilton gastight microliter syringes. The concentrations of standard solutions of both titrants were generally in the mM range.

The mixture solution of acids or bases as well as both solutions of acids or bases separately was titrated with an optically transparent acid and/or base solution and the data for ΔpK_a calculations was obtained from UV-Vis spectra (corrected for dilution). From each titration experiment, the ΔpK_a was

determined as the mean of 5–25 values (see Figure 1), calculated from spectra of varied medium acidity.

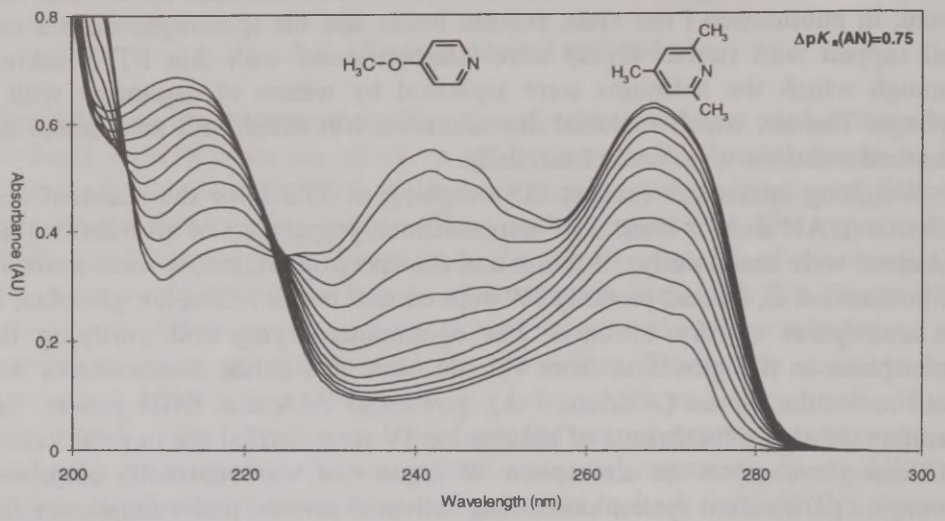

Figure 1. UV-Vis spectra of titrated mixture solution of two compounds in AN.

On several cases experiments were repeated after varying some experimental details. In all the cases the agreement between the ΔpK_a-s obtained before and after variations was good to very good.

3. ACIDITY SCALE OF STRONG NEUTRAL BRØNSTED ACIDS IN ACETONITRILE

3.1. Introduction

Strong acids and their derivatives are receiving increasing attention from both practical (reagents in organic synthesis,[1–3] catalysts in industry[5]) and theoretical[6,7] points of view. Therefore substantial theoretical and experimental effort has been devoted to development of strongly acidic media and molecules with high intrinsic acidity.[1,8] Several families of interesting acids have been created: cyanocarbon acids,[25,26] Kuhn's acidic hydrocarbons,[27] acids with Yagupolskii's substituents,[28,29] superacidic metal hydrides,[30] etc.

Above (in Section 1.2) are described the requirements to the media for acid-base studies. Acetonitrile has many properties that make it suitable for this work. It has low basicity and very low ability to solvate anions.[23] The low basicity gives AN an advantage over the other very popular solvent for acid-base studies — DMSO — which is considerably more basic (stronger acceptor of hydrogen bond). AN has high dielectric constant (D=36.0[23]) and hence favors the dissociation of ion pairs into free ions. The autoprotolysis constant of AN is very low: $pK_{auto} \geq 33$,[31] (even values of pK_{auto} as high as 44 have been suggested[32,33]) this makes it a good differentiating solvent. All these properties put together make it a good differentiating solvent for strong acids. Additional advantages of AN are its transparency down to 190 nm and relative ease of purification.

In solvents of intermediate dielectric constant (D = 15 ... 40) the ratio between free and associated ions depends on the structure of the solvent and as well as compounds to dissociate (e.g. ion size and its concentration, charge distribution, hydrogen-bonded ion pairs, specific ion solvation etc.).[22]

Starting from the classical works of groups of Kolthoff and Coetzee considerable amount of acidity data for various compounds in acetonitrile have been accumulated. Analysis of literature shows that a rather continuous and self-consistent acidity scale in the pK_a range of 14–27 exists in acetonitrile.[15,23,34–36] Measurements in the lower pK_a range have been made too,[25,37,38] but here the things are far from satisfactory both in terms of the amount of data available and its self-consistency*. The present work was undertaken to improve the situation by building a unified self-consistent scale of acidity in AN in the range of 4–16 pK_a units which would be a logical extension of the pK_a scale for the relatively weak acids into the domain of strong and very strong neutral Brønsted acids. In publication I the building of self-consistent acidity scale in AN is described.

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* Because AN solvates ions more weakly than water, various association processes have to be considered.[23,39] These processes may have an important role in explaining the fact, that different results are obtained with different methods and conditions by different authors.

Table 1. Results of the acidity measurements of neutral Brønsted acids in AN together with the absolute pK_a values of the acids.

No	Acid	Measured ΔpK_a	pK_a(AN)
1	2,4-dinitrophenol		16.66
2	$(4\text{-}CF_3C_6F_4)_2CHCN$	0.54	16.14
3	$3\text{-}CF_3C_6H_4CH(CN)_2$	1.43	14.72
4	Saccharin	0.15	14.58
5	$4\text{-}CH_3C_6F_4CH(CN)_2$	0.71	13.88
6	$C_6F_5CH(CN)_2$	0.87	13.01
7	$4\text{-}HC_6F_4CH(CN)_2$	0.03 0.04	12.98
8	$2\text{-}C_{10}F_7CH(CN)_2$	0.74	12.23
9	Tos_2NH	1.38 0.26	11.97
10	$4\text{-}NO_2C_6H_4CH(CN)_2$	0.28 1.21	11.61
11	$(C_6H_5SO_2)_2NH$	0.60 0.60	11.34
12	$4\text{-}ClC_6H_4SO_2NHTos$	1.43 0.10	11.10
13	Picric acid	0.79 0.91	11.00
14	$(4\text{-}ClC_6H_4SO_2)_2NH$	-0.01	10.20
15	$4\text{-}CF_3C_6F_4CH(CN)_2$	0.14	10.19
16	$4\text{-}NO_2C_6H_4SO_2NHTos$	1.05	10.06
17	$4\text{-}Cl\text{-}3\text{-}NO_2C_6H_3SO_2NHTos$	0.53	9.69
18	$4\text{-}NO_2C_6H_4SO_2NHSO_2C_6H_4\text{-}4\text{-}Cl$	0.56 2.3	9.15
19	TosOH	1.21 0.23	8.6
20	$(4\text{-}NO_2C_6H_4SO_2)_2NH$	0.25	8.31
21	$1\text{-}C_{10}H_7SO_3H$	0.19 1.04	8.00
22	$C_6H_5CHTf_2$	0.54 1.25	7.83
23	$4\text{-}ClC_6H_4SO_3H$	0.53	7.3
24	$3\text{-}NO_2C_6H_4SO_3H$	0.44 0.75	6.76
25	$4\text{-}NO_2C_6H_4SO_3H$	0.36	6.71
26	TosNHTf	0.98	6.29
27	$C_6H_5SO_2NHTf$	0.77	6.01
28	$4\text{-}ClC_6H_4SO_2NHTf$	0.53	5.46
29	$2\text{-}NO_2C_6H_4NH_3^+$	0.38 0.75	5.30
30	$4\text{-}ClC_6H_4SO(=NTf)NHTos$	0.41 0.94	5.27
31	$2,4,6\text{-}Tf_3C_6H_2OH$	0.05 1.1	4.93
32	$4\text{-}NO_2C_6H_4SO_2NHTf$	1.10 1.15	4.53
33	$4\text{-}ClC_6H_4SO(=NTf)NHSO_2C_6H_4\text{-}4\text{-}Cl$	0.74 0.20	4.48
34	$4\text{-}Cl\text{-}2\text{-}NO_2C_6H_3NH_3^+$	0.50	4.36
35	2,3,5-tricyanocyclopentadiene		4.16
36	$4\text{-}ClC_6H_4SO(=NTf)NHSO_2C_6H_4\text{-}4\text{-}NO_2$		3.75

Tos = $4\text{-}H_3CC_6H_4SO_2-$, Tf = F_3CSO_2-

3.2. Results

The results of the measurements are presented in Table 1. Each arrow represents the mean ΔpK_a value from one titration experiment. To make the results more reliable and to be able to estimate the consistency of the results, multiple overlapping measurements were carried out. The entire acidity range covered involves at least two independent pathways of measurements and the relative acidity of any two acids can be obtained by combining at least two independent sets of measurements.

In order to assign absolute pK_a values to the acids of Table 1 the scale has to be anchored to a reference compound with pK_a reliably established. This compound should be a "well-behaved" compound. That is, it should not undergo homo- or heteroconjugation reactions to a marked extent and it should have favorable UV-Vis spectral characteristics. For several of the acids belonging to the scale, pK_a value in AN has been reported in the literature. Analysis of the data in literature shows that picric acid is the compound of choice as it is a well-behaved compound and its pK_a in acetonitrile has been determined with great care by Kolthoff et al[40] using three different experimental methods and has been found to be 11.0 ± 0.1.

The pK_a values for individual acids were found minimizing the sum of squares of differences between directly measured ΔpK_a values and the assigned pK_a values:

$$u = \sum_{i=1}^{n_m} (\Delta pK_a^i - (pK_a(HA_2) - pK_a(HA_1)))^2 \qquad (33)$$

The sum is taken over all the measurements whereby ΔpK_a^i is the result of a relative acidity measurement of acids HA_1 and HA_2 (HA_2 is the acid whose pK_a is higher). $pK_a(HA_1)$ and $pK_a(HA_2)$ are the absolute pK_a values for the two acids as found by the least squares procedure. The precision and the consistency of the results can be assessed using a standard deviation as defined by eq 34:

$$s = \sqrt{\frac{u}{n_m - n_c}} \qquad (34)$$

where n_m=74 is the number of measurements, n_c=36–1=35 is the number of pK_a-s determined. For our results s=0.04 pK_a units. This is a low enough value for the scale to be considered self-consistent. This value was also taken as the basis for giving the absolute pK_a values with 2 decimal digits in Table 1.

3.3. Discussion

Sulfonimides and sulfonic acids. pK_a values for aromatic sulfonimides are almost lacking from the literature. One reason is that sulfonimides are strong acids and their acidity cannot be measured in nondifferentiating solvents. On

the other hand, aromatic sulfonimides totally lack volatility and therefore no gas-phase data are available. To the best of our knowledge there has been only one work[41] where pK_a data of some aromatic sulfonimides in water have been reported.

Applying the Hammett equation to the aromatic sulfonimides from Table 1, the following relationship was found:

$$pK_a = (11.25 \pm 0.05) - (1.97 \pm 0.07)(\sigma_1 + \sigma_2) \quad (35)$$
$$n = 7;\ r^2 = 0.993;\ s = 0.11,$$

where σ_1 and σ_2 are the Hammett constants of the corresponding substituents (taken from ref. 42).

The aromatic sulfonic acids from Table 1 show a similar sensitivity towards substitution

$$pK_a = (8.0 \pm 0.2) - (1.9 \pm 0.4)\,\sigma \quad (36)$$
$$n = 4;\ r^2 = 0.92;\ s = 0.3,$$

although the correlation is poorer. It is of interest to compare these data to those for substituted benzoic acids. Applying the Hammett equation to a set of substituted benzoic acids (4-H, 4-Br, 4-NO_2, 4-COOH, pK_a values from ref. 15) gives:

$$pK_a = (20.8 \pm 0.1) - (2.6 \pm 0.2)\,\sigma \quad (37)$$
$$n = 4;\ r^2 = 0.987;\ s = 0.12.$$

It can be seen that the acidity of benzoic acids is about 1.3 times more sensitive towards substitution in the aromatic ring than the acids with acidity center SO_2XH. The probable cause is that -SO_2- fragment in the anion is bigger, more polarizable and more electronegative than the corresponding -CO- fragment in carboxylates, and can therefore "hold" more charge and has lesser tendency to delocalize it into the aromatic ring.

Yagupolskii's substituents. Compounds **30**, **33** and **36** from Table 1 can be considered as derivatives of **12**, **14** and **18** respectively where an =O fragment of a sulfonyl group adjacent to the NH acidity center is replaced by =N-Tf. The acidifying effects of the substitution are 5.8, 5.7, 5.4 pK_a units for **12**, **14** and **18** respectively. The following values of pK_a have been found for $C_6H_5SO_2NH_2$, 4-$CH_3C_6H_4S(O)(=N\text{-}Tf)NH_2$ and 4-$CH_3C_6H_4S(=N\text{-}Tf)_2NH_2$ in DMSO: 16.0,[43] 8.0,[29] 3.4.[29] It can be seen that the acidity increase is not additive: the first substitution increases the acidity by 8 pK_a units while the second substitution by 4.6 pK_a units (the small effect of the 4-methyl group can be neglected here). It has not yet been possible to measure the acidities of these compounds or the sulfonimides in the gas phase[6] but there is a value of gas phase acidity available for a "superacidic" aniline 4-$(CF_3SO(=N\text{-}Tf))$-$C_6H_4NH_2$ ΔG_{acid} = 313.4 kcal/mol.[6] This compound is 13 orders of magnitude (!) more acidic in the gas

phase than the corresponding unmodified aniline 4-Tf-$C_6H_4NH_2$ (ΔG_{acid} = 331.3 kcal/mol[10]).

The sulfonimides **26**, **28** and **32** can be considered as derivatives of sulfonic acids in which an =O fragment of the sulfonyl group is replaced with =N-Tf. These compounds can exist in two tautomeric forms:

a b (38)

Evidence (NMR) has been presented, that **a** is the dominating form in acetone and chloroform.[28] The differences in acidities between **26**, **28**, **32** and the corresponding sulfonic acids **19**, **23** and **25** are 2.3, 1.8 and 2.2 pK_a units respectively. These results can be regarded as evidence in favor of the structure **a** in AN, as otherwise the differences should be similar to the ones obtained for sulfonimides.

Phenols. pK_a value –1.0 for **31** has been reported in aqueous H_2SO_4.[44] This is about 1.3 pK_a units lower than the pK_a of picric acid.[44] In AN, according to Table 1, **31** is about 6 pK_a units more acidic than picric acid. This qualitatively higher difference leads to a conclusion that some solvent effect is in operation here. 2-nitrophenols are known to give intramolecular hydrogen bonding[45] in AN. In water, on the other hand, this hydrogen bonding is absent due to the competition from water.[45] This intramolecular hydrogen bond causes considerable extra-stabilization of the neutral in AN compared to water. There are no data in the literature on hydrogen bond acceptor properties of trifluoromethanesulfonyl group but it is likely that these are weak in comparison with nitro group because the hydrogen bond basicity of sulfones is generally very low. Another factor might be that the picrate anion, due to its nitro groups, is likely to be more solvated in water than the deprotonated **31**.

Phenylmalononitriles. An interesting result with this class of compounds is that the pK_a-s of **6** and **7** in AN are practically equal or that of **7** is even slightly lower. The same behavior of these two compounds has been observed also in dimethoxyethane.[46] This is not completely unexpected: a similar effect is seen when comparing phenol to 4-fluorophenol: the latter is by only 0.2 pK_a units more acidic in water than the former.[14] The reason for this might be that although fluorine is an electronegative substituent, it is also a weak resonance donor.[42] The F is in the 4 position to the acidity center, which means that the inductive/field effect (but not the resonance effect) is weakened by the distance. The final factor is the strong electron-deficiency of the ring that still weakens the inductive/field effect by saturation. These arguments have to be treated with caution however because the same is not observed in the gas phase where **6** is more acidic by about 2 kcal/mol.

The correlation between the pK_a-s in AN and the gas phase acidities[6] is poor:

$$pK_a=(-108 \pm 36) + (0.40 \pm 0.12)\ \Delta G_{acid} \qquad (39)$$
$$n = 7;\ r^2 = 0.69;\ s = 0.91.$$

This is surprising, because of all the compounds' classes in this scale the phenylmalononitriles should be relatively weakly influenced by solvation and they are very suitable for measurements both in AN and in the gas phase.

4. BASICITY SCALE OF NEUTRAL AND CATIONIC BRØNSTED BASES IN ACETONITRILE

4.1. Introduction

Neutral organic bases have found wide field of applications in the organic synthesis as reagents in base-mediated transformations and are often irreplaceable.[2,4] They have many substantial advantages as compared to ionic bases, such as milder reaction conditions, enhanced reactivity of the more naked anions in the poorly associating ion pairs formed, better solubility.[47–49] Several new and very promising families of strong neutral bases — "proton sponges",[50,51] guanidines,[52] amidines, phosphazenes,[12,47–49,9] phosphorus ylides,[53,54] have emerged. Their quantitative basicity data must be known as these largely determine applicability of these bases in practice.

Some of the advantages of acetonitrile over other aprotic solvents as a medium for acid-base studies are described in previous section. AN is a weak electron-pair donor and totally lacks the HBD ability. Hence, it solvates cations better than anions.[23,22]

AN has the disadvantage that very strong superbases tend to oligomerize this solvent.[23,32] The pK_a of conjugated acid of EtP_1(pyrr) in AN, the strongest base involved in this experiment, is 28.6[55] being sufficiently low not to decompose the solvent in a short time but still sufficiently high to be able to deprotonate the conjugate acids of the bases under study. However, if the standard solution of EtP_1(pyrr) was left to stand for several days some discoloration of the solution was observed. Publications II and III are devoted to the building of self-consistent basicity scale in AN.

A spectrophotometric titration method of previous works[I,56] was modified. Due to the possible sensitivity of the bases to the moisture and oxygen all weighing operations (except the weighing of TfOH for the standard acid solution), preparation of all solutions, titration and the spectrophotometric measurements were carried out in a glovebox in an atmosphere of dry nitrogen.

The calculation methods for ΔpK_a determination were similar to previous works[I,56]. For polybasic bases (see Table 3) with pK_a differences smaller than 3-4 pK_a units a different pK_a calculation method was developed, for details see III.

4.2. Results

All in all 79 individual relative acid-base equilibrium measurements between 42 neutral and cationic bases were carried out to give a continuous basicity scale presented in Tables 2 and 3 ranging over 17 pK_a units. Multiple overlapping measurements make the results more reliable and help to estimate their self-consistency. The entire basicity range covered involves at least two independent

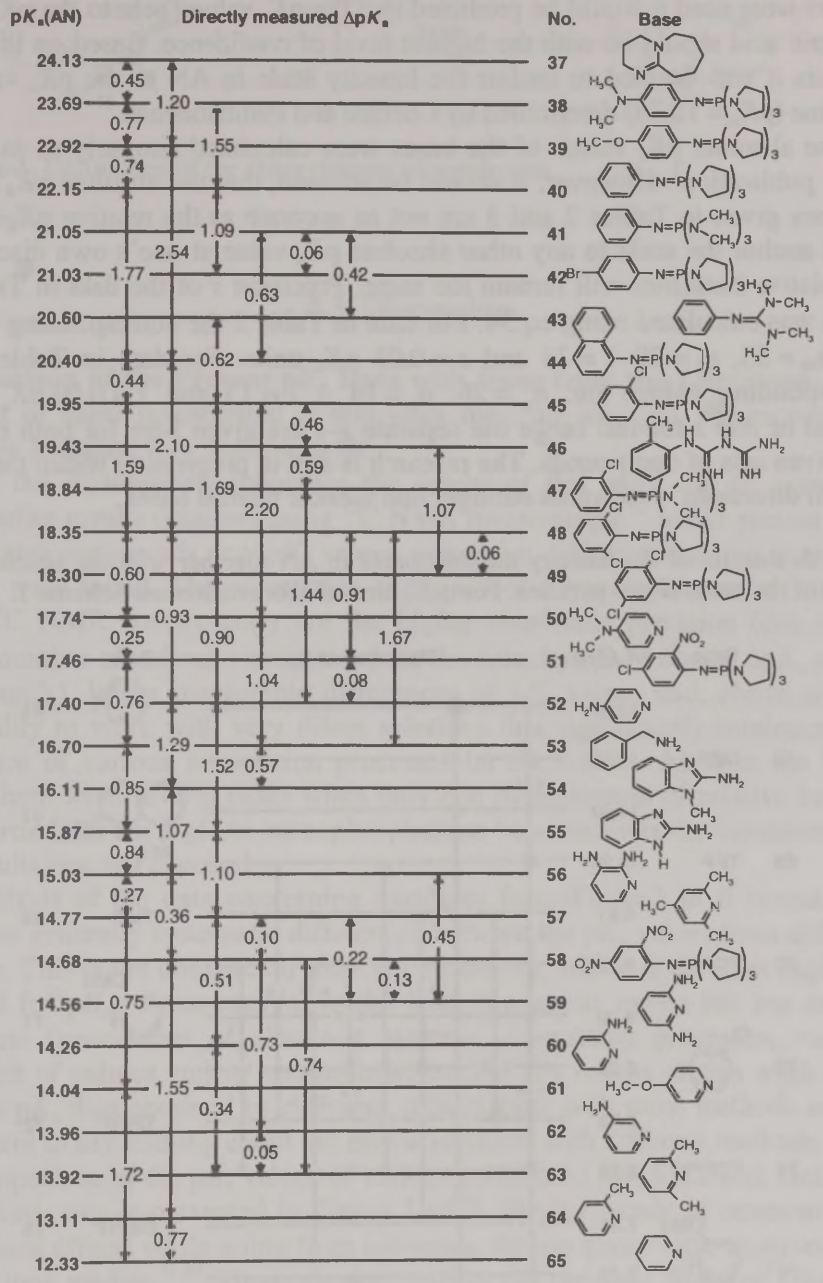
pathways of measurements and the relative basicity of any two bases can be obtained by combining at least two independent sets of measurements. Reversibility of protonation/deprotonation process of all bases was checked. All equilibria were reached in minutes and were stable.

In AN homo- and heteroconjugation reactions (see eqs 5 and 6) of bases must be taken into account, if the concentration of bases is higher than 10^{-3} M or if the homoconjugation constant K_{BHB} is high.[57] Pyridines have the K_{BHB} in AN between 4 and 100,[58,59,21] benzylamine 15,[21] and DBU 35[60] L/mol. It is evident that phosphazene bases have lower homoconjugation constants due to their higher degree of charge delocalization and steric hindrance of the protonation center. In our experiments the concentration of the bases was usually in the 10^{-5} M range and total concentration of bases never exceeded 2.2×10^{-4} M. Therefore it was assumed that there is no need to consider homo- and heteroconjugation in the calculation procedures.

To assign the absolute pK_a values for the conjugate acids of the bases, the scale has to be anchored to a reference compound or compounds for which the pK_a value(s) are known. Direct anchoring of the present scale to the "well-behaved" picric acid[40] — an anchor compound of acidity scale in AN[I], is probably not the best solution since it has a different charge type. In addition, the introduction of another anion (besides the TfO^-) to the solution may have some effect to the pK_a values of bases through the possible difference in solvation and ion-pairing as compared to the TfO^-. Minor influence of counteranion on the basicity of bases has been observed for potentiometric and conductometric measurements at higher concentrations.[21] The dissociation constant for pyridinium picrate ion-pair was measured in AN and was found to be 3.0×10^{-3} mol/L,[61] this is sufficiently low to have some influence upon to the acid-base equilibria. Anchoring the scale to some other neutral acid with highly delocalized charge in the anionic form and with reliably measured pK_a in AN (for example with 2-(pentafluorophenyl)malono-1,3-dinitrile (pK_a in AN 13.01)[I] or 2-(2-perfluoro-naphthyl)malono-1,3-dinitrile (pK_a in AN 12.23),[I] etc.) is also not preferred since these compounds are not very common and the ion-pairing reactions of these compounds have not been investigated. However, the majority of basicity values in AN given in the literature that have been obtained from potentiometric measurements have indirectly been measured relative to picric acid, because picrate buffers have commonly been used to calibrate the glass electrode.

Although the pK_a values in AN have been reported for several compounds (see Table 2 in II) from the present scale in earlier works, there are often disagreements between the results from different authors that are higher than the stated experimental errors. Anchoring the scale to more than one point may distort the final results because some values are more influenced by possible erroneous values of anchoring points than others. Anchored to one point, all absolute values of the given scale are influenced by same extent and relative values, calculated from the overlapping measurements, remain unaffected.

Table 2. Results of the basicity measurements in AN together with the absolute pK_a values of the bases above pyridine. Pyridine (**65**) was taken as anchor compound.

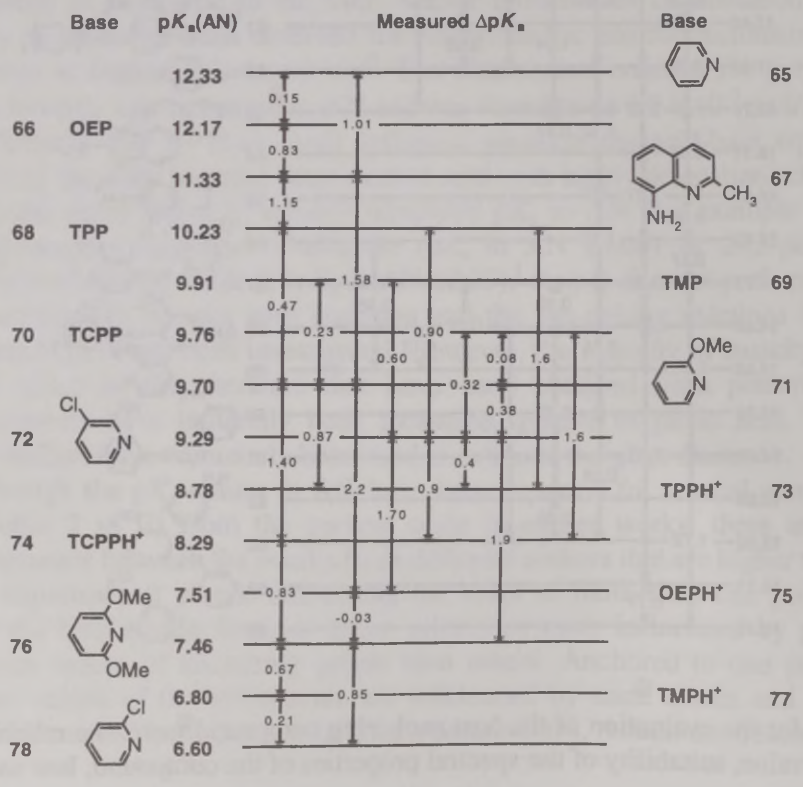


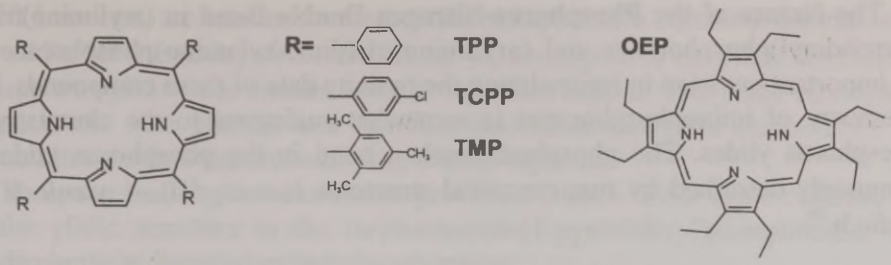
Criteria for the evaluation of the best anchoring compound were: the reliability of the pK_a value, suitability of the spectral properties of the compound, low extent of

association processes. Number of parallel runs and consistency with results from other authors was also considered. Since in most works the potentiometric method was used where for the calibration of the glass-electrode the picrate buffers were used it would be predicted that the pK_a values near to the pK_a value of picric acid should be with the highest level of confidence. Based on literature analysis it was decided to anchor the basicity scale in AN to the pK_a value of pyridine (pK_a = 12.33) determined by Coetzee and Padmanabhan.[21]

The absolute pK_a values of the bases were calculated similarly as in a previous publication[1]. However, it should be stressed, that the absolute pK_a values of bases given in Tables 2 and 3 are not as accurate as the relative pK_a-s. One could anchor the scale to any other absolute pK_a value at one's own discretion, the relative basicities will remain the same. Precision s of the data in Tables 2 and 3 was calculated using eq.34. For data in Table 2 the corresponding values are: $n_m = 53$, $n_c = 29-1 = 28$ and $s = 0.03$ pK_a units. For data in Table 3 the corresponding values are: $n_m = 26$, $n_c = 14 - 1 = 13$ and $s = 0.09$ pK_a units. Instead of one s for full range the separate s–s are given here for both regions and given sets of compounds. The research is still in progress to widen the scale in both directions — towards stronger and weaker neutral bases.

Table 3. Results of the basicity measurements in AN together with the absolute pK_a values of the bases below pyridine. For definition of abbreviation see **Scheme 1.**





Scheme 1. Definitions of the abbreviations of porphyrins.

4.3. Discussion

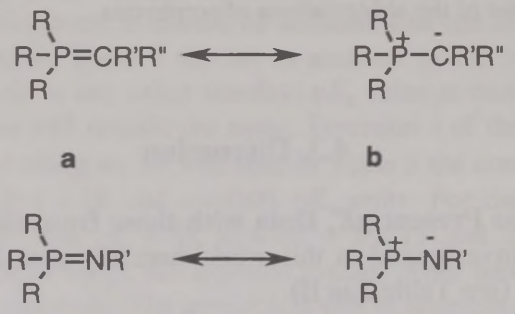
Comparison of the Present pK_a Data with those from the Literature. For a number of bases investigated in this work there are also pK_a values available from the literature (see Table 2 in II).

For the aryliminophosphoranes the results of this work can be compared with earlier results obtained using ^{13}C NMR spectroscopy.[62] In the present work for several compounds new pK_a values, somewhat different from the ones from ref 62, were found. The advantages of the UV-Vis spectrophotometric method over ^{13}C NMR spectroscopy are the higher obtainable precision (due to the larger number of indicator ratios used for the calculations to obtain ΔpK_a of two compounds), larger measurable differences of pK_a values and, above all, the possibility to work with very dilute solutions that significantly minimizes the influence of various association processes on the results. Also, in the NMR study there were several cases when only one measurement of relative basicity was carried out for a given iminophosphorane base and thus the consistency of the results was not checked.

Analysis of the data concerning pyridines from Table 2 in II reveals that there are generally systematic differences between the pK_a values from different authors. This is not unexpected since potentiometry, which is the most exploited method for pK_a measurements, works well in aqueous media but has several problems (restrictions of electrode systems, association processes, variable activities of solutes etc) in non-aqueous media. The results of this work show that the pK_a data obtained by different authors even with same methods are not consistent to say nothing about the results obtained with different methods.

Comparison of the pK_a values of various substituted pyridines with their gas-phase basicities is presented in Figure 1 of II. Similar trends of attenuation of substituent effects while going from solvent to the gas phase were observed also in previous studies.[51,63] These results confirm the previous findings[51] that the basicities of neutral bases are significantly less sensitive towards structural effects than the anionic bases.

The Nature of the Phosphorus-Nitrogen Double Bond in (arylimino)tris(1-pyrrolidinyl)phosphoranes and (arylimino)tris(dimethylamino)phosphoranes is an important question in rationalizing the basicity data of these compounds. The chemistry of iminophosphoranes is somewhat analogous to the chemistry of phosphorus ylides. The phosphorus-carbon bond in the phosphorus ylides is commonly described by two canonical structures (see eq 40) — *ylenic* **a** and *ylidic* **b**.[64]



(40)

Modern theoretical calculations and experimental physical methods have shown, that the ylidic structure has higher contribution to the phosphorus ylides. Interestingly, the nature of the phosphorus — nitrogen double bond has received far less attention. It is most commonly represented as a formal double bond and in the case simple of alkyl and aryl iminophosphoranes this approach is supported by X-ray crystallography and electron diffraction experiments.[53]

From the correlation of the pK_a values in AN for ring-substituted (unsubstituted, 4-MeO, 4-Br, 2-Cl, 2,5-Cl_2) (arylimino)tris(1-pyrrolidinyl)phosphoranes and ring-substituted PhTMGs from ref 52 the following equation is obtained:

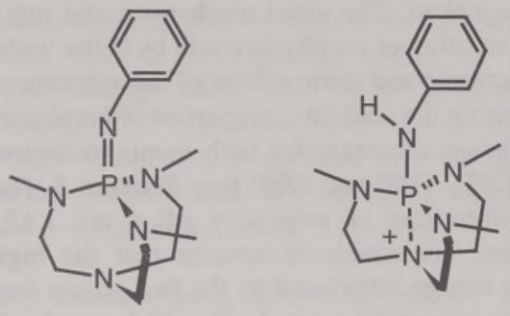
$$pK_a(PhP_1(pyrr)) = (-5.7 \pm 0.9) + (1.36 \pm 0.05)\ pK_a(PhTMG) \qquad (41)$$
$$n = 5;\ r^2 = 0.996;\ s = 0.127.$$

It appears that the basicities of the mentioned phosphoranes are 1.3 ... 1.4 times more sensitive towards substitution in the phenyl group. The second interesting feature is that the pK_a difference (6.5 pK_a units) is far higher between MeP_1(dma) (pK_a = 27.55[48]) and PhP_1(dma) than between pentamethylguanidine (pK_a = 25.00[48]) and PhTMG (4.4 pK_a units). At first these findings seem surprising because the degree of delocalization of the positive charge in the protonated iminophosphoranes is expected to be higher due to larger number of NMe_2 fragments (3 vs 2) and the more electropositive character of phosphorus atom in phenyliminophosphoranes as compared to carbon atom in PhTMGs. If one assumes that the phosphorus-nitrogen bond is a double bond, then no appreciable delocalization of the lone electron pair from the imino nitrogen takes place into the aromatic ring in either group of compounds.

If, however, we assume in these iminophosphoranes certain contribution of the ylidic (zwitterionic) structure (see eq 40) analogous to the ylidic structure in phosphorus ylides, the situation seems far more logical. The ylidic structure is isoelectronic with phenols and therefore delocalization of the electrons on the imino nitrogen into the aromatic ring can be expected.

Thus, our basicity measurements provide evidence about some contribution of the ylidic structure in the (arylimino)tris(1-pyrrolidinyl)phosphoranes and (arylimino)tris(dimethylamino)phosphoranes.

Tang et al.[49] have found the polycyclic phosphazene **a** (see **Scheme 2**) to be weaker base than DBU and stronger than PhP_1(dma) in deuterated AN. In other words, its basicity is quite similar to the basicity of PhP_1(pyrr), although the exact pK_a is not given. They have suggested the transannulation in the polycyclic cage as shown in **b** (see **Scheme 2**) to significantly enhance the stability of the protonated **a.**



a **b**

Scheme 2

However, this base strengthening effect cannot be considered very large (maybe around 2 pK_a units) because even substitution of three dimethylamino groups in PhP_1(dma) with quite similar three pyrrolidino groups already gives a base-strengthening effect of about one pK_a unit.

Comparison of the two sets of compounds: PhP_1(pyrr), PhP_1(dma), PhTMG, **a** on **Scheme 2** vs. phenyl-substituted (arylimino)tris(1-pyrrolidinyl)phosphoranes makes clear that the latter series (changing the substituents in aryl group) has a much larger potential in differentiation of basicities, with pK_a values quite well predictable from analogs. This knowledge is very useful for creating new reference base series for basicity measurements in regions where good references are lacking.

Basicity of porphyrins. Porphyrins and their complexes with metals are very important natural and synthetic compounds. Naturally occurring porphyrins constitute the basis of the respiratory systems for both *flora* and *fauna,* also well known vitamin B_{12} is Co(II) complex of porphyrin. Synthetic porphyrins and their complexes with metals have found several applications in industry – paint protectors against UV radiation, colouring pigments (porphyrins are ones with

highest molar absorptivities in UV-Vis region), ingredients in polymers, solar energy transformers; and in medicine as for the identification of carcinogenic tissues in the human body and in chemotherapy.[65, III]

The insertion mechanism of cation into porphyrin core and the exact role of catalysts have not been fully established not only in biological systems but also in solution. The overall metalloporphyrin formation reaction consists of the coordination of pyrrole nitrogen atoms to metal ion and the release of two pyrrole protons. Concerning the pyrrole protons, a typical reaction intermediate called a sitting-atop complex was proposed.[66] This is intermediate, in which two pyrrolenine nitrogen atoms of the porphyrin coordinate to the metal ion and two pyrrole protons still remain. The presence of this intermediate depends strongly on solvent acid-base properties and of presence of bases. The basicity of acetonitrile is very low, and therefore, the pyrrole protons of porphyrin core are not easily released even after forming the intermediate with the cation without additional proton acceptors. The exact mechanism and rate of the sitting-atop complex formation reaction of porphyrin could be better understood after taking into account the electronic and steric effects of the substituents of the porphyrin peripherals and knowing the acid-base properties of porphyrin.

The acid dissociation constants for both monoprotonated and diprotonated porphyrins TPP, TCPP, TMP and OEP (see **Scheme 1.** for definitions) were determined. The differences of respective pK_a-s are 1.45, 1.47, 3.11, 4.66 respectively. It seems reasonable to assume that the higher the degree of delocalization of the charge introduced by the first proton from the center of the porphine moiety, the closer the second pK_a will be to the first. This explains why OEP has the greatest ΔpK_a of this quartet — the delocalizing power of the phenyl rings of TPP is greater than that of the ethyl groups of OEP. Although the phenyl groups in the free base of TPP are almost perpendicular to the porphyrin core, this does not hold for the mono- and diprotonated forms, for which the porphyrin core is also quite distorted. Such a situation can be seen in the molecular structure of the porphyrin diacid species (see ref-s 41–48 in III). The phenyl groups can rotate to some extent about their symmetry axis so that they will be more coplanar with the porphyrin core and thus enabling some charge delocalization by resonance. On the other hand TMP is in different situation than other phenylporphyrins, because it has two *ortho*-methyl substituents on each phenyl group. These two methyl substituents sterically hinder the rotation of phenyl group about the symmetry axis and thus cause smaller charge delocalization by the resonance from phenyl group. As a result, the ΔpK_a of TMP is intermediate between OEP on one hand and TPP on the other hand.

5. BASICITY SCALE OF NEUTRAL BRØNSTED BASES IN THF

5.1. Introduction

THF is on its differentiating ability (estimated $pK_{auto} = 34.7$[67]) similar to the AN, it is better cation solvator (DN = 20.0) than AN, but it solvates anions even more poorly (AN = 8.0), besides it has low dielectric constant ($D = 7.58$[22]). As a result, in THF the ion-pairing processes are much more favoured than in AN and even at low concentrations the ion-pairing processes (eq-s 7 and 8) must be considered. THF is much more resistant to the superbases, at least 8 orders of magnitude stronger bases can be studied in THF than in AN.[32,55]

Numerous acidity studies, mostly focused on CH-acids, have been carried out in THF.[18,19,20] Alkali metal amides or carbanions with alkali metal counterions have mostly been used as deprotonating agents. On the basis of these measurements ion-pair acidity scales have been developed relative to 9-phenylfluorene or fluorene.[18,19,20] Those scales have been anchored to the pK_a values of these reference compounds in aqueous sulfolane and DMSO, respectively. This approach has also been used in constructing acidity scales in various other nonpolar media — cyclohexylamine[68], benzene[69,70], dimethoxyethane[71], diethyl ether[72], etc.

Contrary to acidity measurements, studies on basicity in THF are scarce. Recently, the Morris' group compiled an acidity scale in THF based on NMR measurements including numerous metal hydrides, phosphines, etc.[73] The observed pK_{ip} values were corrected for ion-pairing using the Fuoss equation.[74] Besides the phosphines the pK_{ip} values for a number of other organic compounds were established and the pK_α value 12.5 of triethylamine was suggested as the secondary standard for anchoring acidity and basicity scales in THF.[73]

In nonpolar media the measured equilibrium constants do not generally reflect the free ion acidity but rather refer to ion pairs. An attempt to suppress the interactions between cations and anions of the CH-acids in nonpolar media was made by the Konovalov group.[75,18] They have used lithium [2.1.1]cryptate as the counterion for the anions of CH-acids. In this cryptate the [2.1.1]cryptand acts as a layer of solvent molecules separating the ions and so eliminating the specific interactions. In a previous work from our lab phosphazene *t*-BuP_4(dma) was used as the deprotonating agent to create an acidity scale in n-heptane.[56] Protonated *t*-BuP_4(dma) is a large cation with a delocalized charge and has very weak interaction with anions.

In the previous section it was shown that arylsubstituted P_1 phosphazenes $ArN{=}P(R)_3$, where R = dma or pyrr — are suitable indicators together with amines to compile a basicity scale in AN.[II,III]

In this part of thesis a self-consistent basicity scale in THF medium that incorporates aryl- and alkyl- P_1 phosphazenes $R'N{=}P(R)_3$, some aryl P_2

phosphazenes R'N=P(R)$_2$–N=P(R)$_3$, various substituted pyridines and several other bases is described. Publication IV is devoted to this subject.

The convenient and common experimental technique to study ion-pairing processes is combined use with conductometry but also theoretical estimation of ion-pair dissociation constant K_d using Fuoss equation (eq 42) has been used[73].

$$K_d = \frac{3000 \cdot e^b}{(4\pi N a^3)} \quad \text{where} \quad b = \frac{-e^2}{aDkT} \tag{42}$$

$N = 6.02 \times 10^{-23}$ mol^{-1}, a is the distance of ion centres in centimetres ($a = r^+ + r^-$), e is the charge of electron 4.80×10^{-10} esu (1 esu = 1.60×10^{-19} C), Boltzmann constant $k = 1.38 \times 10^{-16}$ erg · K^{-1} (1.38×10^{-23} J · K^{-1})

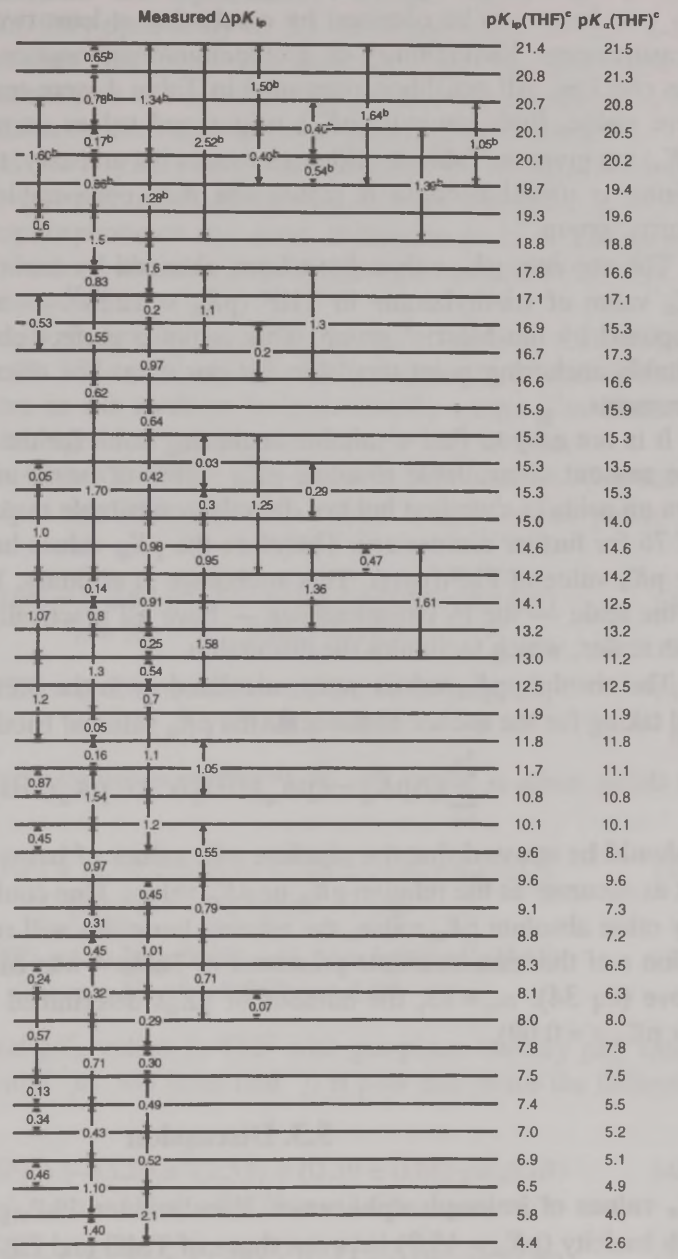
In THF the experimental setup and general principle of the ΔpK calculations is the same as in I, II and III. As UV-Vis spectrophotometry does not make any difference between free ions, solvent separated ion-pairs and loosely bound contact ion-pairs (these last two are the main forms of monocharged ions at concentration below 0.01 M in THF)[73] the experiments yield ΔpK_{ip}-s instead of ΔpK_a-s. The K_d was calculated using the Fuoss equation (eq 42) as described in ref 73. Using the K_d value an estimate of $\Delta pK_a - \Delta pK_\alpha$ is then obtained.

In some cases ("invisible bases", e.g. aliphatic amines (pyrrolidine and triethylamine) and *t*-BuP$_1$(dma), also DBU and TMG versus "visible" aromatic bases) treatment of spectrophotometric data described in Section 2.1 could not be used and the calculations were carried out in the "classical" way — on molar basis. The solution containing a mixture of known amounts (in moles) of "invisible" and visible base was titrated with titrant of known concentration. From the added titrant mass and its concentration the amount (in moles) of titrant in the cell was found. Combining the spectra of solutions containing both bases fully deprotonated, fully protonated and the mixture of protonated and deprotonated forms, the indicator ratio of the visible base was calculated. Knowing the amount of the visible base and the amount of titrant added the indicator ratio for the "invisible" base was calculated. The ΔpK_{ip} calculation is then straightforward. The agreement between relative ion-pair basicities obtained with different calculation methods was satisfactory to good.

5.2. Results

Altogether 95 individual acid-base equilibrium measurements in THF involving 45 bases were carried out using the UV-Vis spectrophotometric or ^{13}C NMR method (for experimental and calculation details see IV). These measurements give a continuous basicity scale in THF as presented in Table 4.

Table 4. Results of the basicity measurements in THF together with the absolute pK_{ip} and pK_{α} values of the bases.



[a] The numbers on the arrows are the direct experimental ΔpK_{ip} values (uncorrected for ion pairing) obtained from UV-Vis spectrophotometric measurements if not indicated otherwise. [b] NMR measurements, for experimental details see IV. [c] Absolute pK_{ip}(THF) and pK_{α}(THF) estimated values for conjugate acids of the respective bases.

Multiple overlapping measurements make the results more reliable and help to estimate their self-consistency. The entire basicity range covered involves at least two independent pathways of measurements and the relative basicity of any two bases can be obtained by combining at least two independent sets of measurements. Reversibility of protonation/deprotonation process of all bases was checked. All equilibria presented in Table 4 were reached in minutes and were stable. Both ion-pair (pK_{ip}) values and values corrected for ion pairing (pK_α) are given in Table 4. Although somewhat arbitrary, the correction for ion-pairing is useful because it makes the data comparable to the data by the Morris' group.[73]

The absolute pK_α values have been obtained by anchoring the scale to the pK_α value of triethylamine in THF (pK_α = 12.5)[73] — a secondary standard proposed by the Morris' group. This is not a perfect choice but is the most suitable anchoring point available for our data. See discussion for additional comments.

It is not easy to find a suitable anchoring point for the ion-pair pK_{ip} values. The amount of available absolute pK_{ip} values of bases in THF is scarce. The data on acids is abundant but not directly comparable to pK_{ip} data of bases (see ref 76 for further discussion). Therefore the pK_{ip} values have been anchored to the pK_α value of PhP_1(pyrr). This anchorage is arbitrary, but this way the core of the scale — the P_1 phosphazenes — have all practically the same values in both scales, which facilitates the discussion.

The absolute pK_α values were calculated as in the previous publicatons[I,II,III] and taking for the anchor of the scale the pK_α value of triethylamine — 12.5.

$$u = \sum_{i=1}^{n_m} (\Delta pK_\alpha^i - (pK_\alpha(HB_2^+A^-) - pK_\alpha(HB_1^+A^-)))^2 \qquad (43)$$

It should be stressed that the absolute pK_α values of bases given in Table 4 are not as accurate as the relative pK_{ip} or pK_α values. One could anchor the scale to any other absolute pK_α value, the relative basicities will remain the same. Precision s of the measurements presented in Table 4 was calculated as described above (eq 34): $n_m = 83$, the number of pK_αs determined $n_c = 43$ and $s = 0.10$ (for pK_{ip} $s = 0.09$).

5.3. Discussion

pK_α values of iminophosphoranes. Unsubstituted PhP_1(pyrr) is a strong base with basicity (pK_α = 15.9) between those of TMG and DBU. By substitution of the phenyl ring, the basicity can be varied over a wide pK_α range. In this work the pK_α values of substituted PhP_1(pyrr) range from 7.5 (2,6-dinitro-) to 17.1 (4-dimethylamino-), that is — by almost ten orders of magnitude. The influence of substituents in the phenyl ring on the basicity of the phosphorane is easily

predictable giving the possibility to conveniently "tune" the basicity of the phosphorane.

Alkyliminophosphoranes are significantly stronger bases than aryliminophosphoranes. Thus, EtP_1(pyrr) is by ca 5.5 and *t*-BuP_1(pyrr) by ca 4 orders of magnitude stronger than PhP_1(pyrr). The inductive effect and some delocalization of the lone electron pair of the imino nitrogen into the aromatic ring (see Section 4.3 for discussion on this topic) are most probably the reasons.

As can be expected, the P_2 phenyliminophosphoranes are stronger bases than the corresponding P_1 phenyliminophosphoranes. The difference is ca 4–5 pK units. For alkyliminophosphoranes the same difference is ca 6 pK units in acetonitrile. We do not have data on P_2 alkyliminophosphoranes in THF, so direct comparison is not possible.

The relatively good predictability of the basicity together with suitable spectral properties in the UV range make the phenyliminophosphoranes convenient neutral indicators in the medium to high basicity range. The choice of neutral indicators in the high basicity range is currently very limited.

Comparison of basicities in THF with those in other media. Correlation of pK_{ip} and pK_α values in THF with pK_a values in acetonitrile yields the following equations:

$$\mathrm{p}K_{ip}(\mathrm{THF}) = (-2.68 \pm 0.55) + (0.83 \pm 0.03)\cdot \mathrm{p}K_a(\mathrm{AN}) \qquad (44)$$
$$n = 39;\ r^2 = 0.959;\ s = 0.89$$

$$\mathrm{p}K_\alpha(\mathrm{THF}) = (-5.08 \pm 0.39) + (0.92 \pm 0.02)\cdot \mathrm{p}K_a(\mathrm{AN}) \qquad (45)$$
$$n = 39;\ r^2 = 0.983;\ s = 0.63$$

Correlation of pK_{ip} and pK_α values in THF with pK_a values in water yields the following equations:

$$\mathrm{p}K_{ip}(\mathrm{THF}) = (1.78 \pm 0.64) + (1.08 \pm 0.08)\cdot \mathrm{p}K_a(\mathrm{H_2O}) \qquad (46)$$
$$n = 17;\ r^2 = 0.926;\ s = 1.05$$

$$\mathrm{p}K_\alpha(\mathrm{THF}) = (-0.31 \pm 0.49) + (1.14 \pm 0.06)\cdot \mathrm{p}K_a(\mathrm{H_2O}) \qquad (47)$$
$$n = 17;\ r^2 = 0.960;\ s = 0.80$$

Correlation of pK_{ip} and pK_α values in THF with gas-phase basicity pK_a values (p$K_a(GB)$ = GB (kcal·mol^{-1})/1.364 (kcal·mol^{-1})) is poor and yields the following equations:

$$\mathrm{p}K_{ip}(\mathrm{THF}) = (-53.29 \pm 12.52) + (0.39 \pm 0.08)\cdot \mathrm{p}K_a(GB) \qquad (48)$$
$$n = 15;\ r^2 = 0.676;\ s = 2.57$$

$$\mathrm{p}K_\alpha(\mathrm{THF}) = (-61.34 \pm 12.48) + (0.43 \pm 0.08)\cdot \mathrm{p}K_a(GB) \qquad (49)$$
$$n = 15;\ r^2 = 0.718;\ s = 2.56.$$

From these correlations it appears that the differentiating ability of THF for basicities is between that of water and AN. One can see that the transfer of this

reaction series from THF to AN increases slightly its sensitivity towards substituent effects both in case of pK_{ip} and pK_{α} whereas the transfer from THF into water leads to the opposite result.

It is interesting to note that in all cases the correlation is better with the pK_{α} than with the pK_{ip} values. This result indirectly validates the method of correction for ion-pairing.

Concentration dependence of pK_{ip} values. If we assume that no larger associates than 1:1 ion pairs exist in the solution then the pK_{ip} values should not show any concentration dependence. The concentrations in the NMR measurements are intrinsically higher than in UV-Vis measurements and the agreement between these two methods serves as a good indicator. According to the data in Table 4 in this thesis and Table 3 in IV the results of the two methods agree well for the ΔpK_{ip} values of arylphosphazenes. With bases of smaller size, however, disagreements are observed. When comparing the results of the measurement of the same equilibrium carried out by different methods, the following is observed: the ΔpK_{ip} between DBU and 4-Me_2N-$C_6H_4P_1$(pyrr) according to the UV-Vis measurements is 0.83, according to NMR it is 0.48; according to the UV-Vis measurements DMAP is by 0.54 pK_{ip} units stronger base than 2-Cl-$C_6H_4P_1$(dma), whereas according to the NMR measurements DMAP is by 0.38 units weaker. The situation is even more serious with triethylamine: according to the UV-Vis measurements it is by 1.58 pK_{ip} units stronger base than 2-Cl-$C_6H_4P_1$(dma), according to the NMR it is by 0.87 pK_{ip} units weaker. These discrepancies are larger than the possible uncertainties of these measurements. On the UV-Vis measurements we did not observe noticeable ΔpK_{ip} dependence (see Table S1 in Supporting information of Publication IV) on concentrations while changing the concentration of Et_3N over the wide range (from 4.5×10^{-5} M to 2.3×10^{-2} M) and keeping the 2-Cl-$C_6H_4P_1$(pyrr) concentration 10^{-4} .. 10^{-5} M.

The disagreements can be due to the formation of aggregates of 1:1 ion pairs at higher concentrations, especially at those used in the NMR method. Due to this concentration dependence, the NMR measurements involving bases of small size were not included in the basicity scale in THF (Table 4).

The Streitwieser method[77] was used to estimate the mean aggregation numbers of the ion pairs. For the UV-Vis data these were around 1 indicating that no significant aggregation of the ion pairs was taking place during the measurements. In principle the method also permits to find the mean aggregation constants, however, due to the very low extent of aggregation it was not possible to get reliable estimates for the aggregation constants. Also, it was not possible to apply that method to the NMR data, because it is necessary that one of the protonated bases would be in the solution only in the form of a monomeric ion-pair. This condition is not met at the concentrations used for the NMR measurements.

Anchoring of the scale. These disagreements cast some doubt on the suitability of triethylamine as anchoring point for basicity data in THF: the pK_{α}

value of triethylamine was determined by the Morris group using NMR measurements (concentrations were in the range of 0.02–0.07 M).[73] The other two compounds common in this work and ref 73 – N,N-dimethylaniline (pK_α = 6.0)[73] and TMG (estimated pK_α = 15)[73] — are not as suitable because they are, respectively, either at the very bottom of the scale or have pK_α value that has only been estimated, not measured. They are of similar size to triethylamine, so that similar concentration dependence problems can be anticipated. The pK_α values of N,N-dimethylaniline and TMG from this work are 4.9 and 15.3, respectively. Thus, in the case of N,N-dimethylaniline there is a disagreement, but the general picture would remain the same if one of these two compounds would be used as the anchoring point. To the best of our knowledge, besides the work of the Morris group,[73] there is no other absolute basicity data in THF available in the literature that could be used to anchor our scale.

6. CONCLUSIONS

As a result of this work three self-consistent acidity scales in condensed non-aqueous media — acetonitrile and THF were built. These scales fill several still existing gaps, cast some new light in the research of acid-base chemistry and underpin future acidity and basicity studies in acetonitrile and THF. Acetonitrile was chosen as solvent as it has good differentiating ability of acids and bases; very low basicity and acidity; high dielectric constant which favors the dissociation of ion pairs into free ions. Tetrahydrofuran is of similar differentiating ability; virtually absent acidity and thus very good medium for studies of strong bases. Both solvents are transparent in UV-Vis wavelength region and find wide applications in various fields of chemistry.

In the first part (Section 3) of this work UV-Vis spectrophotometric self-consistent acidity scale of neutral strong organic Brønsted acids covering the range from 3.75 to 16.66 pK_a units in acetonitrile was created. These results set solid ground for pK_a studies of strong acids in acetonitrile.

In the second part (Section 4) of this work UV-Vis spectrophotometric self-consistent acidity scale of cationic and some dicationic organic Brønsted acids covering the range from 6.60 to 24.13 pK_a units in acetonitrile was created. Acidity trends within various families of cationic acids were ascertained. Results provide evidence about some contribution of the ylidic (zwitterionic) structure in P=N bond of the arylphosphazenes.

In the third part (Section 5) of this work UV-Vis spectrophotometric self-consistent acidity scale of cationic organic Brønsted bases covering the range from 2.6 to 19.6 pK_α (4.4 to 19.3 pK_{ip}) units in tetrahydrofuran was created. From results it appears, that the differentiating ability of tetrahydrofuran for basicities is between that of water and acetonitrile.

Besides the "pure" UV-Vis spectrophotometric method developed by us, a "classical" UV-Vis spectrophotometric method of ΔpK_a determination was also successfully applied on several compounds with poor spectral properties. Generally good agreement with the "pure" UV-Vis spectrophotometric method was observed.

The experimental technique used in course of present work opens new quality level in the field of acid-base studies. As multiple overlapping measurements were carried out, direct information about self-consistency of the data is observed. The results of this work show that the pK_a data obtained by different authors even with same methods are too often not consistent to say nothing about the results obtained with different methods.

A novel group of strong bases — substituted arylphosphazenes (aryliminophosphoranes) was successfully used in this work. The relatively good predictability of the basicity together with suitable spectral properties in the UV range make the arylphosphazenes convenient neutral indicators in the medium

to high basicity range. The choice of neutral indicator bases in the high basicity range is currently very limited.

The employed "pure" UV-Vis spectrophotometric experimental technique is a good example of green chemistry, as minimum amounts of analytes are needed and wastes are produced.

SUMMARY IN ESTONIAN

Kooskõlalised neutraalsete ja katioonsete Brønstedi hapete happelisuse skaalad atsetonitriilis ja tetrahüdrofuraanis

Ainete happelis-aluseliste omaduste uurimine on väga oluline füüsikalise keemia uurimissuund. Nendel omadustel on sageli otsustav roll ainete rakendusvõimaluste määramisel. Väitekirjas uuriti eksperimentaalselt UV-Vis spektrofotomeetrilise meetodi abil neutraalsete ja katioonsete Brønstedi hapete happelisusi aprotoonsetes keskondades — atsetonitriilis ja tetrahüdrofuraanis. Atsetonitriil on heade happeid ja aluseid diferentseerivate omadustega; väga madala happelisuse ja aluselisusega; selle kõrge dielektriline konstant soodustab ioonpaaride dissotsieerumist vabadeks ioonideks. Tetrahüdrofuraan on eelnevaga sarnase happeid ja aluseid diferentseeriva võimega; praktiliselt olematu happelisusega, seega on see väga sobiv keskkond tugevate aluste uurimiseks. Mõlemad solvendid on UV-Vis spektrialas läbipaistvad ja leiavad laialdast rakendust erinevates keemia valdkodades. Töö tulemusena koostati kolm kooskõlalist happelisuse skaalat neis kahes solvendis. Skaalad täidavad mitmeid lünki ja neil olevad andmed heidavad uut valgust hape-alus tasakaalude uurimisele.

Väitekirja esimeses osas loodi kooskõlaline UV-Vis spektrofotomeetriline tugevate neutraalsete Brønstedi hapete happelisuse skaala atsetonitriili keskkonnas vahemikus 3.75 kuni 16.66 pK_a ühikut. Tulemused loovad kindla aluse tugevate hapete happelisuse edaspidiseks uurimiseks selles keskkonnas.

Väitekirja teises osas loodi kooskõlaline UV-Vis spektrofotomeetriline katioonsete ja mõnede dikatioonsete Brønstedi hapete happelisuse skaala atsetonitriili keskkonnas vahemikus 6.60 kuni 24.13 pK_a ühikut. Tehti kindlaks mitmete katioonsete hapete perekondade sisesed happelisuse muutuse trendid. Tulemused annavad tõestust üliidse (tsvitterioonse) struktuuri osakaalust arüülfosfasenide P=N sidemes.

Väitekirja kolmandas osas loodi kooskõlaline UV-Vis spektrofotomeetriline katioonsete Brønstedi hapete happelisuse skaala tetrahüdrofuraani keskkonnas vahemikus 2.6 kuni 19.6 pK_α (4.4 kuni 19.3 pK_{ip}) ühikut. Töö tulemustest selgub, et tetrahüdrofuraan on oma võimelt diferentseerida happeid ja aluseid atsetonitriili ja vee vahepeal.

Väitekirja tulemused näitavad, et erinevate autorite poolt saadud eksperimentaalsed andmed ainete happelis-aluseliste omaduste kohta ei ole sageli omavahel kooskõlas. Käesolevas töös leidis edukalt kasutust uudne grupp tugevaid neutraalseid aluseid — arüülfosfasenid. Üsna hästi etteennustatavad arüülfosfasenide aluselisused ja sobivad spektraalsed omadused UV alas teevad neist head indikaatoralused keskmise ja kõrge aluselisusega piirkonna jaoks. Väitekirjas peamiselt kasutatud "puhas" UV-Vis spektrofotomeetriline meetod on hea näide rohelisest keemiast. Eksperimendi läbiviimiseks vajalikud ainete hulgad ja tekkivad jääkide kogused on minimaalsed.

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ACKNOWLEDGMENTS

I would like to thank my supervisors Associate Professor Dr. Ivo Leito and Professor Dr. Ilmar Koppel for their guidance and encouragement.

Also I would like to thank Dr. Toomas Rodima for fruitful collaboration in the area of phosphazenes.

I am very thankful to Dr. Aino Pihl, Dr. Viljar Pihl, Dr. Vahur Mäemets and Ms. Lilli Sooväli for their help with the experiments.

I deeply appreciate the support of my family.

The support from the Estonian Science Foundation (grants no. 3060, 3366, 4376, 3992 and 3370), research scholarship from University of Münster and scholarship from Estonian Students Fund in USA is gratefully acknowledged.

PUBLICATIONS

I

Spectrophotometric Acidity Scale of Strong Neutral Brønsted Acids in Acetonitrile

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Received June 11, 1998

A continuous self-consistent quantitative UV−vis spectrophotometric pK_a scale of strong acids in acetonitrile has been created. The 36 compounds studied include phenols, phenylmalononitriles, sulfonimides, sulfonic acids, and sulfonimides modified with Yagupolskii's superacceptor substituents. The scale spans about 13 pK_a units and consists of 74 independent equilibrium constant measurements, each describing the relative acidity of two acids. The method of measurements eliminates the need for the direct determination of the acidity of the medium. The most acidic compounds studied have pK_a values around 4 in acetonitrile and can be regarded as true superacids. The scale is anchored to the pK_a value 11.0 for picric acid. The status of 2,4,6-$(SO_2CF_3)_3C_6H_2OH$ as the most acidic phenol presently known (pK_a = 4.9) is confirmed. It is shown that the replacement of an =O fragment with $=N-SO_2CF_3$ in a $-SO_2-$ group in aromatic sulfonimides results in an acidity increase of more than 10^5 times. The most acidic compound in the scale is 4-$ClC_6H_4SO(=NSO_2CF_3)NHSO_2C_6H_4$-4-$NO_2$ (pK_a = 3.75). The present results together with those from the other authors furnish a unified scale of acidity in acetonitrile ranging from 4 to 27 pK_a units and set solid ground for pK_a measurements of strong acids in acetonitrile.

Introduction

Strong acids and their derivatives are receiving increasing attention from both practical (reagents in organic synthesis,[1,2] catalysts in industry[3]) and theoretical[4,5] points of view. Therefore, substantial theoretical and experimental effort has been devoted to development of strongly acidic media and molecules with high intrinsic acidity.[1] Several families of interesting acids have been created: cyanocarbon acids,[6,7] Kuhn's acidic hydrocarbons,[8] acids with Yagupolskii's substituents,[9] superacidic metal hydrides,[10] etc.

Quantitative measurements of acidity are a vital part of studies involving superacids. Contrary to the success in creating superacids, significant experimental difficulties have been met here, and up to date, little consistent acidity data is available for strong acids. In some cases, the situation is almost ridiculous. For the pK_a of perchloric acid in aqueous solution, for example, the values found in the literature range from −1.6 to −14.[4] The uncertainty exceeds 12 orders of magnitude!

The acidity of an acid can be measured in different ways. Gas-phase measurements yield intrinsic acidities, free of any medium effects. A truly absolute scale of acidity in the gas phase has been established, ranging for almost 90 orders of magnitude.[4,11] With very strong acids, the gas-phase measurements have often failed, however, because the latter are usually not volatile enough and tend to undergo fragmentation in the course of the experiments.[4]

Significant effort has been devoted to acidity studies in aqueous solutions of strong acids (sulfuric acid, perchloric acid, etc.). These measurements are complicated,[12] however, because of the leveling of acidities.

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10.1021/jo9811241 CCC: $15.00
Published on Web 10/16/1998

The number of acidity measurements of strong neutral Brønsted acids performed in nonaqueous solvents is smaller than that performed in aqueous medium.[6,13-17] This is true even for dimethyl sulfoxide (DMSO) solution, where due to efforts of Bordwell and others more than 2000 pK_a values for different classes of Brønsted acids have been determined.[16,17,18] DMSO is a rather basic leveling solvent and thus is not very suitable for studies of strong acids (however, see ref 15). Several acidity orders of strong acids have also been reported,[4,19] but quantitative acidity data for strong acids continue to be scarce. The main problems arising in most nonaqueous solvents are the various association processes between charged (and also neutral) species in these solvents (see below) and difficulties in reliable and reproducible determination of medium acidity in these media. Hence, the pK_as determined in different laboratories often differ more than the experimental errors stated.

In view of this situation, a pressing need exists for a self-consistent scale of acidity of strong acids in solution. We report here a spectrophotometric pK_a scale of strong acids in acetonitrile (AN).

AN has many properties that make it suitable for our work. It has low basicity and very low ability to solvate anions.[20] The low basicity gives AN an advantage over the other very popular solvent for acid–base studies—DMSO—which is considerably more basic (stronger acceptor of hydrogen bond). AN has a high dielectric constant (D = 36.0[20]) and hence favors the dissociation of ion pairs into free ions. The autoprotolysis constant of AN is very low; pK_{auto} = 33.[21,22] All these properties put together make it a good differentiating solvent for strong acids. Additional advantages of AN are its transparency down to 190 nm and relative ease of purification.

Starting from the classical works of Kolthoff and Coetzee, a considerable amount of acidity data for various compounds in acetonitrile has been accumulated. Analysis of the literature shows that a rather continuous and self-consistent acidity scale in the pK_a range of 14–27 exists in acetonitrile.[18,20,23-25] Measurements in the lower pK_a range have been made too,[6,13,14] but here, the things are far from satisfactory both in terms of the amount of data available and its self-consistency. The present work was undertaken to improve the situation by building a unified self-consistent scale of acidity in AN in the range of 4–16 pK_a units, which would be a logical extension of the pK_a scale for the relatively weak acids into the domain of strong and very strong neutral Brønsted acids.

Acidity of an acid HA in solvent S refers to the equilibrium

$$HA + S \rightleftharpoons A^- + SH^+ \quad (1)$$

and is expressed as the equilibrium constant K_a or its negative logarithm pK_a.

$$K_a = \frac{a(SH^+)\, a(A^-)}{a(HA)} \quad (2)$$

where the a values are the activities of the corresponding species. The acid–base equilibria in weakly solvating solvents such as acetonitrile are more complex than those in water. In addition to the equilibrium 1, there are other equilibria present in the system.[20] In AN, the poorly solvated anions eagerly form hydrogen-bonded complexes with hydrogen-bond donors present in the solution. When the donor is the conjugate acid of the anion, the homoconjugation process takes place.

$$A^- + HA \overset{K_{AHA}}{\rightleftharpoons} A^-\cdots HA \quad (3)$$

K_{AHA} (the homoconjugation constant) is the constant of formation of the homoconjugate complex $A^-\cdots HA$.

$$K_{AHA} = \frac{a(A^-\cdots HA)}{a(A^-)\, a(HA)} \quad (4)$$

If the donor is some other acid HX, then the heteroconjugation process is present.

$$A^- + HX \overset{K_{AHX}}{\rightleftharpoons} A^-\cdots HX \quad (5)$$

These side reactions have to be suppressed or taken into account if the accurate acidity data are to be obtained.

Because of the problems with measuring the acidity of the medium, $a(H^+)$, in nonaqueous solutions, we use a method that eliminates the need for its determination. Our method of acidity measurement gives relative acidities of the acids HA_1 and HA_2 according to the following equilibrium:

$$HA_2 + A_1^- \rightleftharpoons A_2^- + HA_1 \quad (6)$$

The pK of this equilibrium is the relative acidity (ΔpK_a) of the acids HA_1 and HA_2.

$$\Delta pK_a = pK_a(HA_2) - pK_a(HA_1) = \log \frac{a(A_1^-)\, a(HA_2)}{a(A_2^-)\, a(HA_1)} \quad (7)$$

The method consists of UV–vis spectrophotometric titration of a solution, where both of the acids are present, with a transparent acid or base.

Experimental Section

Method of pK_a Determination. Relative acidities were measured using a UV–vis spectrophotometric titration technique that was similar to one applied previously to acidity

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measurements in heptane.[26] An acetonitrile solution containing two acids, HA_1 and HA_2 (or the salts of the acids in some cases, see below), was prepared. The acidity of the solution was varied by adding small amounts of acidic or basic titrant. A spectrum was recorded after each addition of titrant. This way, 10−30 spectra with different acidity of the solution were obtained, including those where both acids were in fully anionic or neutral form. Both of the acids were also titrated separately to obtain the spectra of the neutral and the ionized forms of both acids. To test the reversibility of the protonation/deprotonation process, after the addition of one titrant was finished, several portions of the other one were added. From each titration experiment, the ΔpK_a was determined as the mean of 5−20 values.

The calculation methods for ΔpK_a were essentially the same as those described previously.[26] Two additional features were included, however.

1. With some acids, it was necessary to take the homoconjugation process into account. Two cases have to be treated separately.

a. One of the acids (HA) forms the homoconjugation complex, and the other does not. In this case, the analytical concentration C of the acid HA can be expressed as follows:

$$C = [A^-] + [HA] + 2[A^-\cdots HA] \quad (8)$$

We assume that the molar absorption coefficient ϵ of the homoconjugation complex $HA\cdots A^-$ between acid HA and its anion A^- can be expressed as a sum.

$$\epsilon(A^-\cdots HA) = \epsilon(HA) + \epsilon(A^-) \quad (9)$$

From eqs 8 and 9, it follows that, for the homoconjugating acid, the indicator ratio I (see ref 26) as it is found from the spectrum of the mixture is not equal to the conventional $[A^-]/[HA]$, but

$$I = ([A^-] + [A^-\cdots HA])/([HA] + [A^-\cdots HA]) \quad (10)$$

Using eqs 4, 8, and 10 and assuming that the activities in eq 4 can be replaced with concentrations, the relative concentration of the complex $A^-\cdots HA$ can be found.

$$\frac{[A^-\cdots HA]}{C} = \frac{1 + (1/(CK_{AHA}))}{2} - \sqrt{\left(\frac{1 + (1/(CK_{AHA}))}{2}\right)^2 - \alpha(1-\alpha)} \quad (11)$$

where $\alpha = 1/(1 + (1/I))$. Now $[A^-]/[HA]$ can be calculated.

$$\frac{[A^-]}{[HA]} = \frac{\alpha - ([A^-\cdots HA]/C)}{1 - \alpha - ([A^-\cdots HA]/C)} \quad (12)$$

With known $[A^-]/[HA]$, the calculation of ΔpK_a is straightforward (see ref 26).

b. Both of the acids form homoconjugation complexes. In this case, we assume that they also form heteroconjugation complexes. In this work, we take homoconjugation into account only when dealing with sulfonic acids (see results section). We assume that for all the sulfonic acids the homoconjugation constants are equal and that all the heteroconjugation constants are equal to the homoconjugation constants. In this case, it can be shown that the four species, HA_1, HA_2, A_1^-, and A_2^-, are consumed proportionally to their concentrations for the formation of the homo- and heteroconjugation complexes, and the relative decrease of their concentrations will cancel out so that the formation of the complexes can be ignored when calculating ΔpK_a.

2. Two acids in the present work are cationic acids. When calculating the relative acidity of a neutral and a cationic acid according to the eq 13, then the assumption that the ratios of

$$HA + B \rightleftharpoons A^- + BH^+ \quad (13)$$

the activity coefficients are equal (see ref 26) is not valid anymore and the ΔpK_a value is:

$$\Delta pK_a = \log \frac{[B][HA]}{[A^-][BH^+]} - 2 \log f \quad (14)$$

where the activity coefficients $f = f(A^-) = f(BH^+)$ and $f(B) = f(HA) = 1$ (these approximations are valid, since very dilute solutions are used). The value of $\log f$ was calculated using the Debye−Hückel equation that follows:[20]

$$\log f = -\frac{1.64Z^2\sqrt{J}}{1 + 0.48a\sqrt{J}} \quad (15)$$

where J is the ionic strength of the solution, Z is the charge of the ion, and a is the size parameter of the ion that was taken to be 6 Å. The $\log f$ is not very sensitive to this parameter.

Chemicals. The following compounds were synthesized according to a procedure from the literature:[27] **9** (mp 173−174. ^{13}C NMR 145.73, 138.71, 130.50, 128.42, 21.53.), **11** (mp 158−159. ^{13}C NMR 142.02, 134.30, 129.97, 128.27.), **12** (mp 162.5−163.3. Anal. Calcd for $C_{13}H_{12}ClNO_4S_2$: C, 45.15; H, 3.50; N, 4.05. Found: C, 45.41; H, 3.53; N, 4.43. ^{13}C NMR: 145.79, 140.55, 140.44, 138.64, 130.52, 130.14, 130.02, 128.40, 21.53.), **14** (mp 208.8−210.4. ^{13}C NMR: 141.48, 140.17, 130.03, 129.95.), **16** (mp 202.4−203.8. Anal. Calcd for $C_{13}H_{12}N_2O_6S_2$: C, 43.81; H, 3.39; N, 7.86. Found: C, 43.58; H, 3.47; N, 8.01. ^{13}C NMR 151.33, 148.33, 142.31, 140.61, 128.43, 127.72, 126.22, 123.28, 20.71.), **17** (mp 151.3−151.8. Anal. Calcd for $C_{13}H_{11}ClN_2O_6S_2$: C, 39.95; H, 2.84; N, 7.17. Found: C, 40.01; H, 2.77; N, 7.47. ^{13}C NMR 148.65, 145.16, 143.20, 139.41, 133.65, 132.60, 131.39, 130.33, 128.24, 125.32, 21.49.) **18**[28] (mp 222.7−223.4. ^{13}C NMR 151.09, 149.73, 142.59, 139.54, 129.81, 129.53, 124.77.), **20**[29] (mp 239−242 dec. ^{13}C NMR 150.96, 150.68, 129.41, 124.70.). The starting materials were sulfonyl chlorides from Aldrich. The compounds were purified by recrystallization from mixtures of ethyl alcohol and concentrated hydrochloric acid.

The sulfonic acids **19**, **23**−**25** were synthesized[30] from the corresponding sulfonyl chlorides (from Aldrich). The sulfonic acid **21** was obtained from REAKHIM. Sulfonic acids are inconvenient to handle and were used as salts. The acid **19** was converted to tetramethylammonium salt according to ref 31. **19a** (mp 255.0−255.9 dec. Anal. Calcd for $C_{11}H_{19}NO_3S$: C, 53.85; H, 7.81; N, 5.71. Found: C, 54.18; H, 7.64; N, 5.57. ^{13}C NMR 144.03, 141.63, 129.85, 126.93, 56.20, 21.30.). The salt was purified by recrystallization from aqueous acetone. The acids, **21** and **23**−**25**, were converted into the corresponding triethylammonium salts according to ref 31. **21a** (mp 149.6−150.3. Anal. Calcd for $C_{16}H_{23}NO_3S$: C, 62.11; H, 7.49; N, 4.53. Found: C, 62.08; H, 7.72; N, 4.41. ^{13}C NMR 142.27, 135.52, 132.27, 130.30, 129.27, 127.74, 127.53, 127.11, 126.52, 125.32, 47.96, 9.12.), **23a** (mp 93.0−95.3. Anal. Calcd for $C_{12}H_{20}ClNO_3S$: C, 49.06; H, 6.86; N, 4.77. Found: C, 49.46; H, 6.93; N, 4.71. ^{13}C NMR: 145.45, 137.07, 129.41, 128.65, 48.2 (obscured by solvent), 9.25.), **24a** (mp 78.2−78.9. Anal. Calcd for $C_{12}H_{20}N_2O_5S$: C, 47.36; H, 6.62; N, 9.20. Found: C, 47.61; H, 6.89; N, 9.03. ^{13}C NMR 149.37, 148.70, 132.99, 131.01, 125.66, 121.89, 48.2 (obscured by solvent), 9.32.), and **25a** (mp 120.9−122.5. Anal. Calcd for $C_{12}H_{20}N_2O_5S$: C, 47.36; H, 6.62; N, 9.20. Found: C, 47.10; H, 6.90; N, 9.46. ^{13}C NMR

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152.65, 150.07, 128.33, 124.59, 48.2 (obscured by solvent), 9.30.). The salts were purified by recrystallization from acetone.

The following compounds were of commercial origin. Some of them were purified prior to use: **1** (REAKHIM, sublimed in vacuo, mp 113.2−113.7), **4** (Aldrich, 99+%), **13** (REAKHIM, recrystallized from EtOH, mp 121.2−122.7), TfOH (Aldrich, 99+%), tetramethylammonium hydroxide (Aldrich, 25% solution in MeOH), Et_3N (REAKHIM, distilled under atmospheric pressure, stored under argon in refrigerator, small amounts for measurements distilled second time as needed), $HClO_4$ (REAKHIM, "special purity"), acetic acid (REAKHIM, "chemically pure").

The synthesis and purification of the following compounds has been described previously: **2**,[32] **3**,[33] **5**,[34] **6**,[35] **7**,[35] **8**,[36] **10**,[33] **15**,[35] **22**,[37] **26**−**28**,[9] **29**,[39] **30**,[9] **31**,[38] **32**,[9] **33**,[9] **34**,[39] **36**.[9] A sample of **35**[6] was donated by the late Prof. R. W. Taft.

^{13}C NMR. The spectra were run on a Bruker AC-200 instrument. The solvent was CD_3OD, except for **16**. DMSO-d_6 was used with this compound. Tetramethylsilane was used as the internal standard.

Solvent. Acetonitrile suitable for our work must be dry and must not contain impurities that absorb UV radiation. Impurities with a UV spectrum that changes with changing acidity of the medium are especially dangerous. Merck "Lichrosolv" AN was used. It was distilled from P_2O_5 through a 1.2 m long 2 cm id column packed with PTFE chips prior to use to reduce its water content and to further purify it. Alternatively, REAKHIM "pure" AN was distilled through the same column from $KMnO_4$ and then from P_2O_5. Preference was given to the acetonitrile from Merck. The solvent was stored in dark bottles in a desiccator over P_2O_5.

Experimental Setup. The setup was very similar to that described previously.[26] Only the significant differences are given here.

A solution of TfOH in AN was used as the acidic titrant in most cases. In a few experiments, a solution of $HClO_4$ was used. This solution was prepared from a 25% solution of $HClO_4$ in acetic acid. A solution of triethylamine in acetonitrile was used as the basic titrant. The standard syringe techniques could not be used due to the strongly acidic media involved. Instead, the vials, solvent bottle, and spectrophotometer cell (all capped with rubber septa) were interconnected with thin PTFE tubing, through which the solutions were siphoned by means of aspiration with a syringe.

The concentrations of the acids were in the following ranges: **2**, **3**, **5**−**8**, **15**, **16**, **18**, **20**, **24**, **25**, **29**, **31**, **32**, **34**−**36**, (1−7) × 10^{-5} M; **1**, **10**, **13**, (1−3) × 10^{-5} M; **4**, **9**, **11**, **12**, **14**, **17**, **22**, **26**−**28**, **30**, **33**, 8 × 10^{-5} to 5 × 10^{-4} M; **19**, **23**, (1−1.5) × 10^{-3} M. The concentrations of the titrants were chosen for each titration experiment according to the concentrations of the acids and ranged from 5 × 10^{-3} to 5 × 10^{-2} M. All solutions were made fresh daily.

The water content of the solution in the cell was determined after the measurement using Karl Fisher titration. This approach ensures that the amount of water in the solution during titration is less than or equal to that determined. The water content of the solution was mostly between 1 and 2 mM and never higher than 3.5 mM.

Spectrophotometry. A Perkin-Elmer Lambda 2S UV−vis spectrophotometer equipped with cell holders thermostated to 25 °C was used for all measurements. The spectrophotometer was controlled from a PC, and the spectra were stored in digital form. Fused silica cells with an optical path length of 1 cm were used. The reference cell contained pure AN.

Table 1. pK_a Values of the Acids Derived from the Results of the ΔpK_a Measurements

No.	Acid	Directly measured ΔpK_a	pK_a
1	2,4-dinitrophenol		16.66
2	(4-$CF_3C_6F_4$)$_2$CHCN	0.54, −1.92	16.14
3	3-$CF_3C_6H_4CH(CN)_2$	1.43	14.72
4	Saccharin	0.15, −0.84	14.58
5	4-$CH_3C_6F_4CH(CN)_2$	0.71	13.88
6	$C_6F_5CH(CN)_2$	0.87, −0.89	13.01
7	4-$HC_6F_4CH(CN)_2$	0.03, 0.04, −0.79	12.98
8	2-$C_{10}F_7CH(CN)_2$	0.74	12.23
9	Tos_2NH	1.38, 0.28, −0.82	11.97
10	4-$NO_2C_6H_4CH(CN)_2$		11.61
11	$(C_6H_5SO_2)_2NH$	0.28, 1.21, −0.90	11.34
12	4-$ClC_6H_4SO_2NHTos$	0.80[a], 0.80, −0.36, 1.43, 0.10	11.10
13	Picric acid		11.00
14	(4-$ClC_6H_4SO_2)_2NH$	0.79, 0.91, −0.82	10.20
15	4-$CF_3C_6F_4CH(CN)_2$	−0.01, −0.13	10.19
16	4-$NO_2C_6H_4SO_2NHTos$	0.14, −0.52, −1.06	10.06
17	4-Cl-3-$NO_2C_6H_2SO_2NHTos$	1.05	9.69
18	4-$NO_2C_6H_4SO_2NHSO_2C_6H_4$-4-Cl	0.53	9.16
19	TosOH	0.56, 2.3, −1.72	8.8
20	(4-$NO_2C_6H_4SO_2)_2NH$	1.21, 0.23, −1.3	8.31
21	1-$C_{10}H_7SO_3H$	0.25, −0.85	8.00
22	$C_6H_5CHTf_2$	0.19, 1.04	7.83
23	4-$ClC_6H_4SO_3H$	0.54, 1.25	7.3
24	3-$NO_2C_6H_4SO_3H$	0.53, −0.51	6.76
25	4-$NO_2C_6H_4SO_3H$	−0.51, 1.26	6.71
26	TosNHTf	0.44, 0.75, −0.62	6.29
27	$C_6H_5SO_2NHTf$	0.36, −0.83	6.01
28	4-$ClC_6H_4SO_2NHTf$	0.88, −0.70	5.48
29	2-$NO_2C_6H_4NH_3^+$	0.77	5.30
30	4-$ClC_6H_4SO(=NTf)NHTos$	0.53[a], 0.35[a]	5.27
31	2,4,6-$Tf_3C_6H_2OH$	0.38, 0.75, −0.82	4.93
32	4-$NO_2C_6H_4SO_2NHTf$	0.41[a], 0.94, 1.17[a]	4.53
33	4-$ClC_6H_4SO(=NTf)NHSO_2C_6H_4$-4-Cl	0.06, 1.1, −0.87	4.48
34	4-Cl-2-$NO_2C_6H_3NH_3^+$	1.10[a], 1.15, −0.31	4.36
35	2,3,5-tricyanocyclopentadiene	0.74, 0.20[a]	4.16
36	4-$ClC_6H_4SO(=NTf)NHSO_2C_6H_4$-4-$NO_2$	0.50	3.75

[a] Perchloric acid was used as the acidic titrant. Tos denotes 4-$CH_3C_6H_4SO_2$−; Tf denotes CF_3SO_2−.

Results

The results of the measurements are presented in Table 1. Each arrow represents the ΔpK_a from one titration experiment. To make the results more reliable and to be able to estimate the consistency of the results, we carried out multiple overlapping measurements. The entire acidity range covered involves at least two independent pathways of measurements, and the relative acidity of any two acids can be obtained by combining at least two independent sets of measurements.

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Table 2. Comparison of the pK_a Values Determined in This Work with Those Reported in the Literature (Values of Obviously Low Quality Have Been Omitted)

acid	pK_a in AN: this work	lit.	difference
2,4-dinitrophenol	16.66	16.0[a]	+0.7
		18.4[b]	−1.7
picric acid		11.0[c]	
TosOH	8.6	8.01[d]	+0.6
		8.73[b]	−0.1
2-$NO_2C_6H_4NH_3^+$	5.30	4.95[e]	+0.35
4-Cl-2-$NO_2C_6H_3NH_3^+$	4.36	4.2[e]	+0.16
2,3,5-tricyanocyclopentadiene	4.16	3.00[f]	+1.16

[a] Reference 40. [b] Reference 41. [c] Reference 42. [d] Reference 13. [e] Reference 43. [f] Reference 6.

To assign absolute pK_a values to the acids of Table 1, the scale has to be anchored to a reference compound with pK_a reliably established. This compound should be a "well-behaved" compound; that is, it should not undergo homo- or heteroconjugation reactions to a marked extent and it should have favorable UV−vis spectral characteristics. For several of the acids belonging to the scale, a pK_a value in AN has been reported in the literature. These acids are presented in Table 2. Analysis of the data in Table 2 shows that picric acid is the compound of choice as it is a well-behaved compound, and its pK_a in acetonitrile has been determined with great care by Kolthoff et al.[42] using three different experimental methods and has been found to be 11.0 ± 0.1. Using **13** as the anchoring point has only the disadvantage that it stands far away from the stronger members of the scale. A reliable absolute pK_a value for a "well-behaved" strong acid would be very desirable. The other acids from Table 2 are not as suitable as references. **29** and **34** are cationic acids (see below), **19** is not a "well-behaved" compound, and **1** and **35** stand in far ends of the scale and the quality of their pK_a values is not as high as that of **13** (see refs 40 and 6). The agreement between the values from this work and those from the literature can be considered satisfactory. The most deviating result is that of **35**. As the experimental part of the original reference[6] does not describe the pK_a measurements, it is not possible to discuss this deviation thoroughly. Partially, it could be due to the "distance" between **35** and **13**.

The pK_a values for individual acids were found by minimizing the sum of squares of differences between directly measured ΔpK_a values and the assigned pK_a values.

$$u = \sum_{i=1}^{n_m} (\Delta pK_a^i - (pK_a(HA_2) - pK_a(HA_1)))^2 \quad (16)$$

The sum is taken over all the measurements whereby ΔpK_a^i is the result of a relative acidity measurement of acids HA_1 and HA_2 (HA_2 is the acid whose pK_a is higher). pK_a(HA_1) and pK_a(HA_2) are the absolute pK_a values for the two acids as found by the least-squares procedure. The precision and the consistency of the results can be assessed using a standard deviation as defined by eq 17,

$$s = \sqrt{\frac{u}{n_m - n_c}} \quad (17)$$

where $n_m = 74$ is the number of measurements and $n_c = 36 - 1 = 35$ is the number of pK_as determined. For our results, $s = 0.04$ pK_a units. This is a low enough value for the scale to be considered self-consistent. This value was also taken as the basis for giving the absolute pK_a values with two decimal digits in Table 1. This approach to estimate the precision of the results has a drawback however; the precision of the measurements is different for different classes of compounds, being dependent on the spectral properties of the compounds, namely the difference between the spectrum of the anion and the neutral and the wavelength(s) at which the measurements were carried out. The precision is highest for the phenylmalononitriles and phenols and lowest for acids **19** and **23**. The pK_a values for the latter two acids are given with lower precision in Table 1.

It is important to note that this estimate of precision must be interpreted as the precision of single measurements and not the precision of the absolute pK_a values. The reference compound **13** has an acidity that is about 7 pK_a units away from those for the strongest acids of the scale, and for obtaining their acidity, a large number (about 10) of ΔpK_a values has to be added, which significantly decreases the precision of the pK_a values for the compounds that stand far away from the reference compound. Still another point is the accuracy of the pK_a values, which is also significantly lower.

The spectra of the compounds (with the exception of the sulfonic acids **19** and **23**) in solutions of different acidity did not show any irregular behavior. The spectra of partially ionized acids could always be expressed as linear combinations of the spectra of the neutral and the anion. With most of the compounds, the spectra contained isosbestic points, and these were always sharp. These observations rule out the possibility that conjugation reactions take place to an appreciable extent under our experimental conditions as the homo- and heteroconjugation reactions are known to cause distortions in spectra.[44] The protonation−deprotonation process was reversible with all acids.

The spectra of **19** and **23** were slightly distorted, and the spectra of partially ionized acids could not be expressed as linear combinations of spectra of the anion and the neutral. The concentrations of **19** and **23** were sufficiently high to permit the homoconjugation reaction to occur. The homoconjugation constant K_{AHA} for **19** is 800 L/mol.[18] The same value was used for **23**, as the values of K_{AHA} do not differ very much within classes of acids as long as the acidities are not too different and steric factors are not involved. The resulting corrections to the ΔpK_a values were in the range of 0.05−0.2 pK_a units. The assumption in eq 9 is reasonable with **19** and **23** because the acidity center OH is not directly conjugated with the aromatic ring and no significant intramolecular charge transfer is involved in the anions on excitation. However, this assumption would probably not be valid for compounds with the acidity center conjugated

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directly to the aromatic ring (phenols, anilines) because the wavelengths and intensities of the spectral bands with intramolecular charge-transfer character of the deprotonated species are very sensitive to hydrogen bonding.[44,45]

Some of the measurements were made using $HClO_4$ instead of TfOH. Because of the procedure of preparing of solutions of $HClO_4$, substantial amounts of acetic acid (AcOH), a potential hydrogen bond donor, were introduced with this titrant. Two of the measurements were performed with both titrants, and the agreement is good (see Table 1). The rest of the measurements agree well with the overlapping measurements made with TfOH. Thus, we have reason to believe that AcOH does not interfere seriously. This is not unexpected because AcOH (pK_a in AN 22.3[18]) is a very weak acid compared to those studied in this work and the hydrogen bond donating ability of an acid in the conjugation reactions is directly related to its acidity.[46] Also, the basicity of AcOH (pK_a in AN 1.1[18]) as well as its concentration is too low to enable significant stabilization of the neutral acids by hydrogen bonding.

We have included two cationic acids (**29** and **34**) in the scale. The log *f* values were in the range of 0.02−0.05. However, as there is no general procedure for calculating activity coefficients for ions in nonaqueous media, these cationic acids cannot be rigorously regarded as belonging in the scale.

The pK_a values determined in this work together with those from other authors[18,20,23−25] set up a continuous acidity scale in AN ranging from pK_a 4 to 27.

Discussion

Method. The presence of the acids in the same solution eliminates many possible sources of error or reduces their influence. (1) The disturbing effects (traces of water in the medium, concentration errors) affect both acids with the same magnitude and are expected to partially cancel out. Thus, we expect that the effect of traces of water on the measurements will be less pronounced than in such measurements where the pK_a of a single acid is measured at a time. (2) The solutions can be very dilute and consequently very weakly buffered with this method. If a minor acidity change of the solution occurs, it will affect both of the acids and its effect will cancel out. (3) An important point is that the method eliminates the need for quantitative measurement of acidity of the medium.

A disadvantage of the method is that only the acidity of such acids that absorb in the UV−vis spectral region and for which the spectra of the acid and the anion are different can be measured. Also, the pK_as of the acids must not be very different from each other (preferably not more than 1.5−2 pK_a units).

Sulfonimides and Sulfonic Acids. pK_a values for aromatic sulfonimides are almost lacking from the literature. One reason is that sulfonimides are strong acids and their acidity cannot be measured in nondifferentiating solvents. On the other hand, aromatic sulfonimides totally lack volatility, and therefore no gas-phase data are available. To the best of our knowledge, there has been only one work[47] where pK_a data of some aromatic sulfonimides in water have been reported.

When the Hammett equation was applied to the aromatic sulfonimides **9**, **11**, **12**, **14**, **16**, **18**, and **20**, the following relationship was found:

$$pK_a = (11.25 \pm 0.05) - (1.97 \pm 0.07)(\sigma_1 + \sigma_2) \quad (18)$$

where $r^2 = 0.993$, $s = 0.11$, and σ_1 and σ_2 are the Hammett constants of the corresponding substituents (taken from ref 48).

The aromatic sulfonic acids **19**, **23**−**25** show a similar sensitivity toward substitution,

$$pK_a = (8.0 \pm 0.2) - (1.9 \pm 0.4)\sigma \quad (19)$$

where $r^2 = 0.92$ and $s = 0.3$, although the correlation is poorer. It is of interest to compare these data with those obtained for substituted benzoic acids. Applying the Hammett equation to a set of substituted benzoic acids (4-H, 4-Br, 4-NO_2, 4-COOH, pK_a values from ref 18) gives:

$$pK_a = (20.8 \pm 0.1) - (2.6 \pm 0.2)\sigma \quad (20)$$

where $r^2 = 0.987$ and $s = 0.12$. It can be seen that the acidity of benzoic acids is about 1.3 times more sensitive toward substitution in the aromatic ring than that of the acids with acidity center SO_2XH. The probable cause is that the $-SO_2-$ fragment in the anion is bigger, more polarizable, and more electronegative than the corresponding −CO− fragment in carboxylates and can therefore "hold" more charge and has less of a tendency to delocalize it into the aromatic ring.

Yagupolskii's Substituents. Compounds **30**, **33**, and **36** can be considered to be derivatives of **12**, **14**, and **18**, respectively, where an =O fragment of a sulfonyl group adjacent to the NH acidity center is replaced by =N−Tf. The acidifying effects of the substitution are 5.8, 5.7, and 5.4 pK_a units for **12**, **14**, and **18**, respectively. The following values of pK_a have been found for $C_6H_5SO_2NH_2$, $CH_3C_6H_4SO(=N-Tf)NH_2$, and $CH_3C_6H_4S(=N-Tf)_2NH_2$ in DMSO: 16.0,[49] 8.0,[50] and 3.4.[50] It can be seen that the acidity increase is not additive; the first substitution increases the acidity by 8 pK_a units, while the second substitution increases it by 4.6 pK_a units (the small effect of the 4-methyl group can be neglected here). It has not yet been possible to measure the acidities of these compounds or the sulfonimides in the gas phase,[4] but there is a value of gas-phase acidity available for a "superacidic" aniline 4-$(CF_3SO(=N-Tf))-C_6H_4NH_2$, $\Delta G_{acid} = 313.4$ kcal/mol.[4] This compound is 13 orders of magnitude (!) more acidic in the gas phase than the corresponding unmodified aniline 4-Tf−$C_6H_4NH_2$ ($\Delta G_{acid} = 331.3$ kcal/mol[11]).

The sulfonimides **26**, **28**, and **32** can be considered to be derivatives of sulfonic acids in which an =O fragment of the sulfonyl group is replaced with =N−Tf. These compounds can exist in two tautomeric forms:

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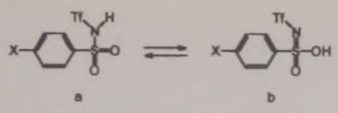

a b (21)

Evidence (NMR) has been presented by one of us that **a** is the dominating form[9] in acetone and chloroform. The differences in acidities between **26**, **28**, and **32** and the corresponding sulfonic acids **19**, **23**, and **25** are 2.3, 1.8, and 2.2 pK_a units, respectively. These results can be regarded as evidence in favor of the structure **a** in AN, as otherwise the differences should be similar to the ones obtained for sulfonimides.

Phenols. A pK_a value of −1.0 for **31** has been reported in aqueous H_2SO_4.[51] This is about 1.3 pK_a units lower than the pK_a of picric acid.[51] In AN, according to Table 1, **31** is about 6 pK_a units more acidic than picric acid. This qualitatively higher difference leads to a conclusion that some solvent effect is in operation here. 2-nitrophenols are known to give intramolecular hydrogen bonding[40] in AN. In water, on the other hand, this hydrogen bonding is absent due to the competition from water.[40] This intramolecular hydrogen bond causes considerable extrastabilization of the neutral in AN compared to water. There are no data in the literature on the hydrogen bond acceptor properties of the trifluoromethanesulfonyl group, but it is likely that these are weak in comparison with those of the nitro group because the hydrogen bond basicity of sulfones is generally very low. Another factor might be that the picrate anion, due to its nitro groups, is likely to be more solvated in water than the deprotonated **31**.

Phenylmalononitriles. An interesting result with this class of compounds is that the pK_as of **6** and **7** in AN are practically equal, or that of **7** is even slightly lower. The same behavior of these two compounds has been observed by one of us also in dimethoxyethane.[35] This is not completely unexpected. A similar effect is seen when comparing phenol to 4-fluorophenol; the latter is more acidic in water than the former by only 0.2 pK_a units.[17] The reason for this might be that although fluorine is an electronegative substituent, it is also a weak resonance donor.[48] The F is in the 4 position to the acidity center, which means that the inductive/field effect (but not the resonance effect) is weakened by the distance. The final factor is the strong electron deficiency of the ring that still weakens the inductive/field effect by saturation. These arguments have to be treated with caution however because the same is not observed in the gas phase, where **6** is more acidic by about 2 kcal/mol.

The correlation between the pK_as in AN and the gas-phase acidities[4] is poor,

$$pK_a = (-108 \pm 36) + (0.40 \pm 0.12)\Delta G_{acid} \quad (22)$$

where $r^2 = 0.69$ and $s = 0.91$. This is surprising because of all the compounds' classes in this scale, the phenylmalononitriles should be relatively weakly influenced by solvation, and they are very suitable for measurements both in AN and in the gas phase.

Conclusion

The present results together with those from the other authors furnish a unified acidity scale in AN ranging from pK_a 4 to 27 and help to establish a solid basis for the reliable pK_a determinations of strong acids in AN. We have demonstrated that the acidities of strong acids are readily measurable and that AN is a suitable medium for this work.

Acknowledgment. This work was supported by Grants 3060 and 3366 from the Estonian Science Foundation. We thank Mr. Viljar Pihl for helpful discussions, Dr. Vahur Mäemets for the NMR analyses, and Mrs. Lilja Lahe for the elemental analyses.

JO981124L

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II

Self-Consistent Spectrophotometric Basicity Scale in Acetonitrile Covering the Range between Pyridine and DBU

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Received May 24, 2000

A self-consistent spectrophotometric basicity scale in acetonitrile, including DBU, ten (arylimino)-tris(1-pyrrolidinyl)phosphoranes, two (arylimino)tris(dimethylamino)phosphoranes, 2-phenyl-1,1,3,3-tetramethylguanidine, 1-(2-tolyl)biguanide, benzylamine, two substituted benzimidazoles, pyridine, and ten substituted pyridines, has been created. The span of the scale is almost 12 pK_a units. Altogether, 29 different bases were studied and 53 independent equilibrium constant measurements were carried out, each describing the relative basicity of two bases. The scale is anchored to the pK_a value of pyridine of 12.33 that has been measured by Coetzee et al. Comparison of the basicity data of phenyliminophosphoranes and phenyltetramethylguanidines implies that the P=N bond in the (arylimino)tris(1-pyrrolidinyl)phosphoranes involves contribution from the ylidic (zwitterionic) structure analogous to that found in phosphorus ylides.

Introduction

Neutral organic bases have found a wide field of applications in organic synthesis as reagents in base-mediated transformations and are often irreplaceable.[1,2] Compared to ionic bases, they have many substantial advantages, such as milder reaction conditions, the enhanced reactivity of the more naked anions in the poorly associating ion pairs formed, and better solubility.[3−5] Several new and very promising families of strong neutral bases, "proton sponges",[6,7] guanidines,[8] amidines, phosphazenes,[3−5,9] and phosphorus ylides,[10,11] have emerged.

Exact quantitative basicity data are very important in applying bases in various fields of chemistry and in designing new bases with desired properties. Although a vast amount of acid−base data has been collected,[12−14] the situation is still far from ideal.

The intrinsic basicity can be measured in gas-phase experiments or calculated in simpler cases. Up to the present much gas-phase basicity data[12] have been collected. However, in this field there is enough space for improvement; basicities of only a few superbases (by definition the superbases have *GB* (gas-phase basicity) over 239 kcal/mol)[9] have been determined. The reasons are the lack of suitable reference bases,[15] low volatility of strong bases, long stabilization periods, and tendency of strong bases to undergo fragmentation in the experiments.

In condensed media investigation of basicities of strong bases has some limitations. The acid strength of solvent, various association processes, and decomposition of base or solvent are only some of the effects that have to be considered while measuring the basicities in condensed media. In relatively acidic media (such as water) strong bases are leveled up and only the basicity of moderately strong bases can be reliably measured. Alcohols have also been common solvents for such studies.[13] Stronger bases can be reliably investigated in solvents that have lower acidity and are relatively inert if exposed to the bases. In such solvents the basicities are not leveled and the intrinsic properties of bases are expressed to a higher

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10.1021/jo005521j CCC: $19.00
Published on Web 08/25/2000

extent. Several solvents, including dimethyl sulfoxide, acetonitrile (AN), THF, etc., have found wide application as media for studies of strong bases, and a vast number of papers and several compilations have been published.[13,14] Although AN has been probably the most popular solvent among the above-mentioned and a vast number of basicity measurements in AN have been carried out, the situation is still not good. Basicity data obtained by different authors even by the same method often lack consistency and are rather with gaps. Even for well studied pyridines included in this work the measured values deviate often by up to one or more pK_a units. That is more than stated with the experimental errors (see below for examples).

Our aim was to improve the situation in this field by building a reliable continuous self-consistent and sufficiently wide spectrophotometric basicity scale in AN covering the range from 12 to 24 pK_a units.

Basicity of a base B in solvent S is defined using eq 1

$$HB^+ + S \rightleftarrows B + HS^+ \quad (1)$$

$$K_a = \frac{a(HS^+)\cdot a(B)}{a(HB^+)} \quad (2)$$

and is expressed as dissociation constant K_a of the conjugate acid HB^+ of the base B or more commonly its negative logarithm pK_a.

AN has some advantages over other aprotic solvents as a medium for acid–base studies. It is a very weakly basic and acidic dipolar aprotic solvent with high dielectric constant (36.0[16]) and hence favors the dissociation of ion pairs into free ions. The autoprotolysis constant K_{auto} of AN is very low: p$K_{auto} \geq 33$,[17] (even values of pK_{auto} as high as 44 have been suggested[18,19]), and this makes it a good differentiating solvent.

Because AN solvates ions more weakly than water, various association processes have to be considered.[16,20] These processes may have an important role in explaining the fact that different results are obtained with different methods and conditions by different authors. AN is a weak electron-pair donor and hence solvates cations better than anions.[16,21] Therefore the main association processes, homoconjugation (eq 3) and heteroconjugation (eq 4) are not as extensive as in systems

$$HB^+ + B \rightleftarrows BHB^+ \quad (3)$$

$$HB^+ + B_1 \rightleftarrows BHB_1^+ \quad (4)$$

involving anions.[22] It is generally observed that homoconjugation increases with an increasing number of acidic hydrogen atoms, and also poor steric hindrance of the protonation center is a source of increased homoconjugation.[23] The easiest way to decrease the influence of the association processes to a negligible extent (they are always an additional source of poorer accuracy) is to use methods that allow the use of very dilute solutions. The UV–vis spectrophotometric method employed in the present work has the advantage over many others methods (potentiometry, ^{13}C NMR spectroscopy, etc.) that in certain cases (π-electron-rich systems conjugated with acidity center) sufficiently low concentrations of solutes may be used to minimize hetero- and homoconjugation processes.

To exclude the necessity for measuring the hydrogen ion activity (see eq 1) we studied the equilibrium between two bases B_1 and B_2:

$$B_2 + HB_1^+ \rightleftarrows HB_2^+ + B_1 \quad (5)$$

The relative basicity, ΔpK_a, of B_1 and B_2 was found as described elsewhere:[24,25]

$$\Delta pK_a = pK_a(HB_2^+) - pK_a(HB_1^+) = \log\frac{a(HB_2^+)\cdot a(B_1)}{a(HB_1^+)\cdot a(B_2)} \quad (6)$$

The method consists of UV–vis spectrophotometric titration of a solution, where both of the bases are present, with an optically transparent acid or base.

Experimental Section

Chemicals. The (arylimino)tris(1-pyrrolidinyl)phosphoranes **2**–**4**, **6**, **8**, **9**, **12**, **13**, **15**, and **22** were synthesized and purified as described elsewhere.[26] 2-Phenyl-1,1,3,3-tetramethylguanidine **7** was synthesized as described in ref 8. 2-Methylpyridine **28** (REAKHIM) was distilled fractionally from KOH in the atmosphere of dry argon, 2,4,6-trimethylpyridine **21** (REAKHIM) was distilled fractionally from KOH under reduced pressure, 4-methoxypyridine **25** (Aldrich, 97%) was distilled fractionally from $MgSO_4$ under reduced pressure, and 2,6-dimethylpyridine **27** (REAKHIM) was distilled fractionally from BaO under reduced pressure. Benzylamine **17** (REAKHIM, "pure") was kept on NaOH overnight and distilled from it fractionally under reduced pressure. 1-(2-Tolyl)biguanide **10** (Aldrich, 98%), 4-aminopyridine **16** (Aldrich, 98%), 2,3-diaminopyridine **20** (Aldrich, 98%), and 2-aminopyridine **24** (REAKHIM, "pure for analysis") were recrystallized from ethanol. 2-Amino-1-methylbenzimidazole **18** (Aldrich, 95%) and 2-aminobenzimidazole **19** (Aldrich, >97%) were recrystallized once from ethanol and then from water. DBU **1** (Aldrich, 98%), 4-(dimethylamino)pyridine **14** (TCI, >99%), 2,6-diaminopyridine **23** (Aldrich, 99+%), 3-aminopyridine **26** (Aldrich, 99%), and pyridine **29** (Fluka, >99.8%) were used without additional purification. (Phenylimino)tris(dimethylamino)-phosphazene **5** was synthesized according to the procedure described in ref 5. The same procedure was employed for synthesis of (2-chlorophenylimino)tris(dimethylamino)phosphorane **11** (starting compounds: 2-Cl-C_6H_4-N_3[27,28] and $(Me_2N)_3P$ (Fluka)). The crude product (a brown oil) was distilled to give a yellowish oil (yield 65%, bp 137 °C (0.15 Torr)). Anal. Calcd for $C_{12}H_{22}N_4P$: C, 49.93; H, 7.68; N, 19.40. Found: C, 49.52;

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H, 8.02; N, 18.88. ^{1}H NMR (200 MHz, CH_3CN) δ 2.66 (d, $^3J_{P-H}$ = 9.6 Hz, 18H), 6.48 (dddd, $^3J_{H-H}$ = 7.8 Hz, $^3J_{H-H}$ = 7.2 Hz, $^4J_{H-H}$ = 1.7 Hz, $^6J_{P-H}$ = 0.6 Hz, 1H), 6.73 (ddd, $^3J_{H-H}$ = 8.0 Hz, $^4J_{H-H}$ = 1.7 Hz, $^4J_{P-H}$ = 1.1 Hz, 1H), 6.94 (ddd, $^3J_{H-H}$ = 7.2 Hz, $^3J_{H-H}$ = 8.0 Hz, $^4J_{H-H}$ = 1.7 Hz, 1H), 7.17 (ddd, $^3J_{H-H}$ = 7.8 Hz, $^4J_{H-H}$ = 1.7 Hz, $^5J_{P-H}$ = 2.5 Hz, 1H); ^{13}C NMR (50 MHz, CH_3CN) δ 37.7 ($^2J_{P-C}$ = 3.3 Hz), 117.3, 123.7 ($^3J_{P-C}$ = 7.2 Hz), 127.9, 128.2 ($^3J_{P-C}$ = 21 Hz), 130.0, 149.1 ($^2J_{P-C}$ = 8.8 Hz).

Solutions of trifluoromethanesulfonic acid (TfOH) (Aldrich, 99+%) and (ethylimino)tris(1-pyrrolidinyl)phosphorane (EtP_1(pyrr)) were used as acidic and basic titrant, respectively. Synthesis and purification of the latter one has been described earlier.[4] EtP_1(pyrr) was fractionally distilled from BaO under reduced pressure and stored under dry argon.

Solvent. AN (>99.9%, Super Purity Solvent (far UV), water content < 0.005%) was purchased from Romil (Cambridge, U.K.) and was used without further purification. It was stored in dark bottles in a desiccator over P_2O_5. It has low absorbance in the UV region down to 200 nm, and its absorbance did not change upon addition of acidic or basic titrant.

Experimental Setup. A spectrophotometric titration method of previous works[24,25] was modified. As a result of the possible sensitivity of the bases to the moisture and oxygen, all weighing operations (except the weighing of TfOH for the standard acid solution), preparation of all solutions (all solutions were made daily), titration, and spectrophotometric measurements were carried out in a glovebox in an atmosphere of dry nitrogen. For continuous drying and purifying of the atmosphere in the glovebox from volatile basic and acidic contaminants, we used molecular sieves (Aldrich, 4 Å), powdered P_2O_5, and KOH pellets. A Perkin-Elmer Lambda 2S spectrophotometer was equipped with an external sample compartment (ESC). The 2-m long quartz-fiber light conductor cables of the ESC were guided through the wall of the glovebox and fastened airtightly. The spectrophotometer cell was closed with a hollow PTFE stopper with PTFE/silicone septa and open-top screw cap. The acidic and basic titrants were added into the cell through the septa using Hamilton gastight microliter syringes. The concentrations of both titrants were in the mM range. We found that the syringe needles were resistant to 10^{-2}–10^{-3} M TfOH solution. A reference cell with pure AN was placed in the spectrophotometer cell holder. The concentrations of individual bases were usually in the 10^{-5} M range, and their total concentration in our experiments never exceeded 2.2×10^{-4} M.

The water content of the titrated solutions was determined to be between 0.9 and 2.4 mM by means of Karl Fischer titration. Temperature was kept at 25 ± 1 °C during the measurements.

Calculation Methods. From each titration experiment, the ΔpK_a was determined as the mean of 5–30 values using the calculation methods described previously.[24,25] The essence of the general calculation method is the following. When two partially protonated bases B_1 and B_2 are in the same solution, then the following equation holds for absorbance A at wavelength λ (1 cm path length):

$$A^\lambda = [HB_1^+]\cdot\epsilon^\lambda_{HB_1^+} + [B_1]\cdot\epsilon^\lambda_{B_1} + [HB_2^+]\cdot\epsilon^\lambda_{HB_2^+} + [B_2]\cdot\epsilon^\lambda_{B_2} \quad (7)$$

The molar absorptivities ϵ can be found separately from the spectra of the free bases and fully protonated bases. If we use concentrations that are normalized to 1 (see ref 25) then we may write $[HB_1^+] = 1 - [B_1]$ and $[HB_2^+] = 1 - [B_2]$. After a mathematical transformation of eq 7 we get

$$\frac{A^\lambda - \epsilon^\lambda_{HB_f^+} - \epsilon^\lambda_{HB_f^+}}{(\epsilon^\lambda_{B_2} - \epsilon^\lambda_{HB_f^+})} = [B_1]\cdot\frac{(\epsilon^\lambda_{B_1} - \epsilon^\lambda_{HB_f^+})}{(\epsilon^\lambda_{B_2} - \epsilon^\lambda_{HB_f^+})} + [B_2] \quad (8)$$

If the spectra are recorded over a range of wavelengths then $[B_1]$ and $[B_2]$ can be found from eq 8 as the slope and intercept of a regression line. If $[B_1]$ and $[B_2]$ are known, the calculation of ΔpK_a of the bases is straightforward. In many cases (e.g., when the bases have absorption maxima in different wavelength ranges) it was possible to use various simpler calculation procedures (see refs 24 and 25). As a rule, the results obtained using different calculation procedures agreed well. It was found that ΔpK_a values over 2 ΔpK_a units could be reliably determined over different ranges of wavelengths only by methods that apply the least-squares of linear combinations method (see ref 25).

All bases used in this work and their conjugate acids have a significant difference in some region of their UV–vis spectra, and thus they can be treated as "visible" bases. Even DBU and benzylamine, which have maximum differences between absorbances of the neutral and protonated forms in a narrow region in UV, could be treated as "visible" bases.

Results

All in all, 53 individual relative acid–base equilibrium measurements between 29 bases were carried out to give the continuous basicity scale presented in Table 1. The relative basicity of any two bases in the scale can be obtained by combining at least two independent sets of measurements. Multiple overlapping measurements make the results more reliable and help to estimate their self-consistency. The entire basicity range covered involves at least two independent pathways of measurements, and the relative basicity of any two bases can be obtained by combining at least two independent sets of measurements. Reversibility of the protonation/deprotonation process of all bases was checked. All equilibria were reached in minutes and were stable.

In AN, homo- and heteroconjugation reactions (see eqs 3 and 4) of bases must be taken into account if the concentration of solutes is higher than 10^{-3} M or if the homoconjugation constant K_{BHB} is high.[29] Pyridines have a K_{BHB} in AN between 4 and 100;[30,22,23] benzylamine, 15;[23] and DBU, 35[31] L/mol. It is evident that phosphazene bases have lower homoconjugation constants as a result of their higher degree of charge delocalization and steric hindrance of the protonation center. In our experiments the concentration of the bases was usually 1.6×10^{-5} M and never exceeded 1.2×10^{-4} M. Therefore we assume that there was no need to consider homo- and heteroconjugation in the calculation procedures.

The method used in this work has the disadvantage that only relative basicities can be determined. To assign the absolute pK_a values for the conjugate acids of the bases, the scale has to be anchored to a reference compound or compounds for which the pK_a value(s) are known. Direct anchoring of the present scale to the "well-behaved" picric acid,[24,32] having a pK_a value that has been measured with great care, is probably not the best solution since it has a different charge type. In addition, the introduction of another anion (besides the TfO^-) to the solution may have some effect on the pK_a values of bases through the possible difference in ion-pairing as compared to the TfO^-. Minor influence of the counter-

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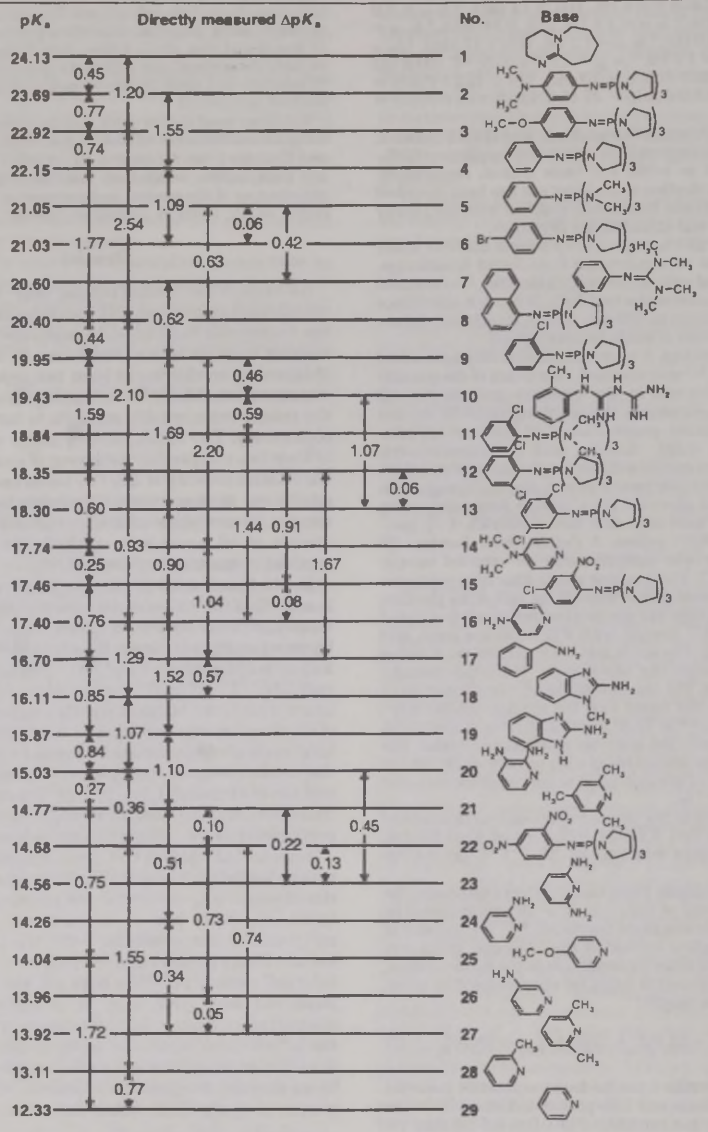
Table 1. Directly Measured Relative Basicity Values for Bases and Estimated Absolute pK_a Values for Their Conjugate Acids in AN



anion on the basicity of bases has been observed for potentiometric (and conductometric) measurements at higher concentrations.[23] The dissociation constant for the pyridinium picrate ion pair was measured in AN and was found to be 3.0×10^{-3} mol/L;[33] this is sufficiently low to have some influence upon the acid–base equilibria. Anchoring the scale to some other neutral acid with highly delocalized charge in the anionic form and with reliably measured pK_a in AN (e.g., with 2-(pentafluorophenyl)malono-1,3-dinitrile (pK_a in AN 13.01)[24] or 2-(2-perfluoronaphthyl)malono-1,3-dinitrile (pK_a in AN 12.23),[24] etc.) is also not preferred since these compounds are not very common and the ion-pairing reactions of these compounds have not been investigated. However, the majority of basicity values in AN given in the literature

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Table 2. Basicity Values in AN Determined in This Work for Which There Are Corresponding Values Reported in the Literature

	pK_a in AN		
base	this work	literature	difference
1	24.13	23.9[a]	+0.2
		24.33[b]	−0.20
2	23.69	23.9[c]	−0.2
3	22.92	23.4[c]	−0.5
4	22.15	22.6[c]	−0.45
5	21.05	20.9[c]	+0.15
6	21.03	21.0[c]	0
7	20.60	20.6[a]	0
8	20.40	20.7[c]	−0.3
9	19.95	19.8[c]	+0.15
12	18.35	18.0[c]	+0.35
13	18.30	17.9[c]	+0.4
14	17.74	18.18[d]	−0.44
15	17.46	17.5[c]	0
16	17.40	17.61[e]	−0.21
		18.38[d]	−0.98
		17.00[f]	+0.40
17	16.70	16.76[g]	−0.06
19	15.87	15.95[e]	−0.08
20	15.03		
21	14.77	14.38[h]	+0.39
22	14.68	14.5[c]	+0.2
23	14.56		
24	14.26	14.66[d]	−0.40
		14.43[e]	−0.17
25	14.04		
26	13.96	14.35[d]	−0.39
27	13.92	14.41[d]	−0.49
28	13.11	12.76[h]	+0.35
		13.88[d]	−0.77
29	taken as 12.33	12.33[g]	0
		12.60[d]	−0.27
		12.52[f]	−0.19

[a] Reference 8. [b] Reference 4, anchored to the pK_a value of 1,4-diaminobutane taken as 20.12. [c] Reference 26. [d] Reference 30a. [e] Reference 34b. [f] Reference 35c. [g] Reference 23. [h] Reference 36. [i] Reference 35d.

that have been obtained from potentiometric measurements have indirectly been measured relative to picric acid, because picrate buffers have commonly been used to calibrate the glass electrode.

Although the pK_a values in AN have been reported for several compounds (see Table 2) from the present scale there are often disagreements between the results from different authors that are higher than the stated experimental errors. For example, see compounds **16** and **28** in Table 2. Anchoring the scale to more than one point may distort the final results because some values are more influenced by possible erroneous values of anchoring points than others. Anchored to one point, all absolute values of the given scale are influenced to same extent, and relative values, calculated from the overlapping measurements, remain unaffected.

Criteria for the evaluation of the best anchoring point should be the correctness of the pK_a determination method, holding the principles of thermodynamics and reversibility and also considering the solutes activities and association processes. The number of parallel runs and consistency with results from other authors may also be criteria. Since in most works the potentiometric method was used, in which picrate buffers were used for the calibration of the glass-electrode, it would be predicted that the pK_a values near to the pK_a value of picric acid should be with the highest level of confidence. Analysis of the references concerning the basicity data in AN given in Table 2 shows good consistency between relative basicities of some compounds, namely, for **7**, **17**, **19** and **29** determined by other authors and in this work (see Table 2). pK_a values for various substituted pyridines given in ref 30a in comparison with results from this work show also consistency between relative basicities with two exceptions, **16** and **28**.

On the basis of these considerations we decided to anchor our scale to the pK_a value of pyridine (pK_a = 12.33) determined by Coetzee and Padmanabhan.[23]

The absolute pK_a values of the bases were calculated similarly as in a previous paper[24] by minimizing the sum of squares of differences between directly measured ΔpK_a values and the assigned pK_a values, while keeping the pK_a value of pyridine constant and equal to 12.33. However, it should be stressed that the absolute pK_a values of bases given in Table 1 are not as accurate as the relative pK_a's. One could anchor the scale to any other absolute pK_a value at one's own discretion; the relative basicities will remain the same. Precision s of the measurements was calculated as in ref 24; $n_m = 53$, $n_c = 29 - 1 = 28$. For our results, $s = 0.03$ pK_a units.

AN has the disadvantage that very strong bases tend to oligomerize this solvent.[16,18] The pK_a of the conjugated acid of EtP_1(pyrr) in AN, the strongest base involved in our experiment, is reported to be 28.89,[4] being sufficiently low not to decompose the solvent in a short time but still sufficiently high to be able to deprotonate the conjugate acids of the bases under study. However, if the standard solution of EtP_1(pyrr) was left to stand for several days, some discoloration of the solution was observed.

Discussion

Comparison of the Present pK_a Data with Those from the Literature. For a number of bases investigated in this work there are also pK_a values available from the literature (Table 2).

For the aryliminophosphoranes the results of this work can be compared with earlier results obtained using ^{13}C NMR spectroscopy.[26] In the present work for several compounds new pK_a values, somewhat different from the ones from ref 26, were found. The advantages of the UV–vis spectrophotometric method over ^{13}C NMR spectroscopy are the higher obtainable precision (due to the larger number of indicator ratios used for the calculations to obtain ΔpK_a of two compounds), larger measurable differences of pK_a values, and above all, the possibility to work with very dilute solutions that significantly minimizes the influence of various association processes on the results. Also, in the NMR study there were several cases when only one measurement of relative basicity was carried out for a given iminophosphorane base.

Analysis of the data concerning pyridines from Table 2 reveals that there are generally systematic differences between the pK_a values from different authors. This is not unexpected since potentiometry, which is the most exploited method for pK_a measurements, works well in aqueous media but has several problems (restrictions of electrode systems, association processes, variable activities of solutes, etc.) in nonaqueous media. The present work has seven coincident points with ref 30a. The mean systematic difference of −0.4 pK_a units was observed between our pK_a values and the values from that work (see Table 2), with two exceptions, 4-aminopyridine and 2-methylpyridine, which differ markedly more. The value

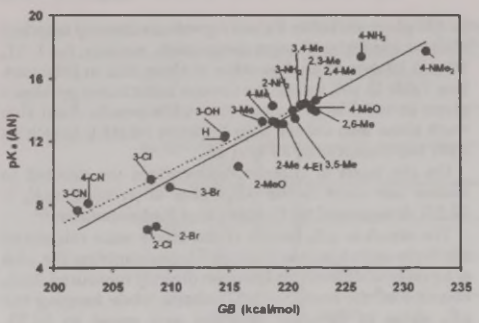


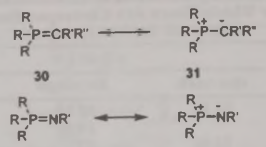
Figure 1. Correlation between gas-phase basicity (*GB*, 1 cal = 4.184 J) values and pK_a values in AN for pyridines. The solid line corresponds to the whole series of pyridines (pK_a(AN) = (−73.9 ± 6.7) + (0.397 ± 0.031) *GB*; n = 23; r^2 = 0.888; s = 1.07); the dotted line corresponds to the series of 3-substituted pyridines (pK_a(AN) = (−66.4 ± 8.4) + (0.365 ± 0.039) *GB*; n = 6; r^2 = 0.955; s = 0.60). The pK_a data for the pyridines not studied in this work have been taken from refs 30a, 35b, and 35d and have been corrected by adding −0.4, −0.19, and −0.19, respectively in order to take into account the systematic differences between those refs and this work (see the discussion). The *GB* values have been taken from ref 12.

for the basicity of 4-aminopyridine in ref 30a is surprising, since in the present work, in water[13] and in the gas phase,[12] 4-(dimethylamino)pyridine is a stronger base than 4-aminopyridine; there should not be any specific interactions in AN that could shift the pK_a value in such a way. The present work has two coincident points with the ref 36, namely, **21** and **28**. The mean difference is +0.37 pK_a units.

Comparison of the pK_a values of various substituted pyridines with their gas-phase basicities is presented in Figure 1. Similar trends of attenuation of substituent effects while going from solvent to the gas phase were observed also in previous studies.[7,38] These results confirm the previous findings[7] that the basicities of neutral bases are significantly less sensitive toward structural effects than the basicities of the anionic bases.

The nature of the phosphorus−nitrogen double bond in (arylimino)tris(1-pyrrolidinyl)phosphoranes and (arylimino)tris(dimethylamino)phosphoranes is an important question in rationalizing the basicity data of these compounds. The chemistry of iminophosphoranes is somewhat analogous to the chemistry of phosphorus ylides. The phosphorus−carbon bond in the phosphorus ylides is commonly described by two canonical structures (see Scheme 1), *ylenic* **30** and *ylidic* **31**.[37] Modern theoretical calculations and experimental physical methods have shown that the ylidic structure has higher contribution to the phosphorus ylides. Interestingly, the nature of the phosphorus−nitrogen double bond has received far less attention. It is most commonly represented as a formal double bond, and in the case simple of alkyl and aryl iminophosphoranes this approach is supported by X-ray crystallography and electron diffraction experiments.[10]

Scheme 1



When correlating the pK_a values in AN for ring-substituted (unsubstituted, 4-MeO, 4-Br, 2-Cl, 2,5-Cl_2) (arylimino)tris(1-pyrrolidinyl)phosphoranes and ring-substituted 2-phenyl-1,1,3,3-tetramethylguanidines from ref 8, then the following equation is obtained: pK_a(PhP1-(pyrr)) = (−5.7 ± 0.9) + (1.36 ± 0.05) pK_a(PhTMG); n = 5; r^2 = 0.996; s = 0.127. It appears that the basicities of the mentioned phosphoranes are 1.3 − 1.4 times more sensitive toward substitution in the phenyl group. The second interesting feature is that the pK_a difference (6.5 pK_a units) is far higher between MeN=P(NMe_2)$_3$ (pK_a = 27.55[4]) and **5** than between pentamethylguanidine (pK_a = 25.00[4]) and **7** (4.4 pK_a units). At first these findings seem surprising because the degree of delocalization of the positive charge in the protonated iminophosphoranes is expected to be higher as a result of a larger number of NMe_2 fragments (3 vs 2) and the more electropositive character of phosphorus atom in phenyliminophosphoranes as compared to carbon atom in PhTMGs. If one assumes that the phosphorus−nitrogen bond is a double bond, then no appreciable delocalization of the lone electron pair from the imino nitrogen takes place into the aromatic ring in either group of compounds.

If, however, we assume in these iminophosphoranes a certain contribution of the ylidic (zwitterionic) structure (see Scheme 1) analogous to the ylidic structure in phosphorus ylides, the situation seems far more logical. The ylidic structure is isoelectronic with phenols, and therefore delocalization of the electrons on the imino nitrogen into the aromatic ring can be expected.

Thus, our basicity measurements provide evidence about some contribution of the ylidic structure in the (arylimino)tris(1-pyrrolidinyl)phosphoranes and (arylimino)tris(dimethylamino)phosphoranes.

Tang et al.[5] have found the polycyclic phosphazene **32** (see Scheme 2) to be a weaker base than DBU and stronger than **5** in deuterated AN. In other words, its basicity is quite similar to the basicity of **4**, although the exact pK_a is not given. They have suggested the transannulation in the polycyclic cage as shown in **33** to significantly enhance the stability of the protonated **32**. However, this base strengthening effect cannot be considered very large (maybe around 2 pK_a units) because even substitution of three dimethylamino groups in **5** with three quite similar pyrrolidino groups already gives a base-strengthening effect of about one pK_a unit.

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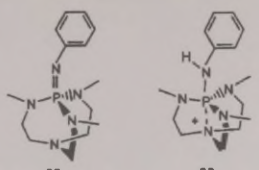
Scheme 2



32 **33**

Comparison of the two sets of compounds, **4**, **5**, **7**, **32** vs. phenyl-substituted (arylimino)tris(1-pyrrolidinyl)-phosphoranes, makes it clear that the latter series has a larger potential in differentiation of basicities, with pK_a values quite well predictable from analogues. This knowledge is very useful for creating new reference base series for basicity measurements in regions where good references are lacking.

Conclusion

A self-consistent spectrophotometric basicity scale in acetonitrile including DBU, ten (arylimino)tris(1-pyrrolidinyl)phosphoranes, two (arylimino)tris(dimethylamino)-phosphoranes, 2-phenyl-1,1,3,3-tetramethylguanidine, 1-(2-tolyl)biguanide, benzylamine, two substituted benzimidazoles, pyridine, and ten substituted pyridines has been created. All together 29 different bases were studied, and 53 independent equilibrium constant measurements were carried out, each describing the relative basicity of two bases. The scale is anchored to the formerly measured pK_a value of pyridine of 12.33. Our basicity measurements provide evidence for some contribution of the ylidic structure in the (arylimino)tris(1-pyrrolidinyl)phosphoranes and (arylimino)tris(dimethylamino)phosphoranes.

Acknowledgment. This work was supported by grants 4376, 3992, and 3366 from the Estonian Science Foundation. We are greatly indebted to Mr. Viljar Pihl for his assistance with various experimental aspects of this work and to Dr. Vahur Mäemets for the NMR measurements.

JO005521J

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Inorganica Chimica Acta 340 (2002) 87–96

Inorganica Chimica Acta

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Sitting-atop complex formation of 2,3,7,8,12,13,17,18-octaethylporphyrin with copper(II) ion in acetonitrile

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Received 23 January 2002; accepted 3 April 2002

Abstract

The reaction of 2,3,7,8,12,13,17,18-octaethylporphyrin (H_2OEP) with copper(II) triflate and copper(II) perchlorate in acetonitrile was studied using spectrophotometry. The reaction product is the so-called sitting-atop complex where two pyrrolenine nitrogen atoms of the porphyrin coordinate to the incoming metal ion and two protons on the pyrrole nitrogen atoms still remain. The composition of the sitting-atop complex was determined by the mole ratio method, and it was found that the H_2OEP molecule binds two copper(II) ions in the product. The mechanism of the reaction was confirmed to be a series of second-order reactions with the first and second step of the reactions being the outer sphere complex formation between the H_2OEP molecule and copper(II) ion and the rate determining sitting-atop complex formation reaction, respectively, based on the mole ratio method. The reaction is relatively fast, and the second-order rate constants for the reaction of H_2OEP with copper(II) ion was determined to be $k = (3.2 \pm 0.3) \times 10^6$ M^{-1} s^{-1} ($T = 25.0$ °C) for the copper(II) triflate and $k = (3.0 \pm 0.2) \times 10^6$ M^{-1} s^{-1} ($T = 25.0$ °C) for the copper(II) perchlorate under the second-order conditions. The pK_a values of the mono- and diprotonated forms of the conjugate acid of several porphyrins including H_2OEP were determined by spectrophotometric titration in acetonitrile. The higher reactivity of H_2OEP toward copper(II) ion as compared with other porphyrins such as 5,10,15,20-tetraphenylporphyrin was attributed to its higher basicity.

Keywords: Porphyrin metallation; Copper(II) complexes; Reaction intermediate; Sitting-atop complexes; pK_a of porphyrin in acetonitrile

1. Introduction

Metalloporphyrins and related compounds play important roles in biological systems. One of the most significant processes is the metalloporphyrin formation reaction with special reference to the heme biosynthesis [1–6]. The terminal step in the heme biosynthesis, the insertion of ferrous iron into protoporphyrin IX to form protoheme, is catalyzed by the enzyme ferrochelatase. Although the crystal structures of the ferrochelatases along with years of biophysical and kinetic studies have led to a better understanding of the catalytic mechanism of the ferrochelatase, the mechanism of the complex formation reaction of porphyrins has not been fully established not only in biological systems but in solution. The overall metalloporphyrin formation reaction consists of the coordination of pyrrole nitrogen atoms to metal ion and the release of two pyrrole protons. Concerning the pyrrole protons, a typical reaction intermediate was proposed by Fleischer and Wang in 1960 [7]. They studied the reaction of protoporphyrin IX derivatives with several metal ions in chloroform and observed the formation of the reaction intermediate. This intermediate is called a sitting-atop complex, in which two pyrrolenine nitrogen atoms of the porphyrin coordinate to the metal ion and two pyrrole protons still remain. Several sitting-atop complexes of naturally occurring and synthetic porphyrins including such ions as platinum(II), copper(II), and rhodium(I) were then reported [8–10]. Recently, more clear evidence for the sitting-atop complex was spectrophotometrically obtained during the reaction of 5,10,15,20-tetraphenylporphyrin (H_2TPP) with copper(II) ion in acetonitrile [11]. The Brønsted basicity of acetonitrile is very weak [12,13], and, therefore, the pyrrole protons of H_2TPP

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PII: S0020-1693(02)01044-7

are not easily released even after forming the intermediate with the copper(II) ion without additional proton acceptors. This acid-base property of the solvent is responsible for the stability of the sitting-atop intermediate. The rate of the sitting-atop complex formation reaction of H_2TPP with several 3d-block metal(II) cations in acetonitrile has also been determined, and the general mechanism of the reaction has been proposed on the basis of the parallel correlation between these values and the rate of the solvent exchange reaction of the corresponding cations [14]. The kinetics of the sitting-atop complex formation reaction have also been investigated for other *meso*-substituted porphyrins, i.e. 5,10,15,20-tetrakis(4-chlorophenyl)porphyrin ($H_2T(4\text{-}ClP)P$), and 5,10,15,20-tetramesitylporphyrin (H_2TMP), and the mechanism of the reaction was discussed in terms of the electronic and steric effects of the substituents of the porphyrin peripherals on the reaction rate [15]. By using acetonitrile, having a weak basicity and a coordination ability to metal ions, as a solvent [16], it was possible to characterize the sitting-atop complex of H_2TPP and other *meso*-substituted porphyrins.

In a previous paper, we also studied the sitting-atop complex formation of 2,3,7,8,12,13,17,18-octaethylporphyrin (H_2OEP) with copper(II) ion in acetonitrile and attempted to determine the rate constant of the reaction [15]. However, the rate is too fast to be followed under pseudo-first-order conditions where the copper(II) ion exists in large excess over H_2OEP. The higher reactivity of H_2OEP toward the complex formation reactions may be ascribed to its higher basicity and/or higher flexibility compared with other *meso*-substituted porphyrins. In the present study, we investigated the sitting-atop complex formation reaction of H_2OEP with copper(II) ion in acetonitrile and successfully determined the rate constant under the second-order conditions. We also determined the pK_a values of the porphyrins in acetonitrile using spectrophotometric titration in order to discuss the reactivity of the porphyrins towards metal ions in terms of the basicity of the porphyrins as well as the molecular structure and substituent effects of the porphyrins.

2. Experimental

2.1. Chemicals

Acetonitrile (Romil, >99.9%, Super Purity Solvent (far UV), water content <0.005%) was the same as used in a previous study [17]. Acetonitrile (Wako Pure Chemical Industries, Japan) was dried over activated 3 Å molecular sieves for several days and distilled under nitrogen atmosphere. H_2OEP (Tokyo Chemical Industry) was used without further purification. Pyridine (Fluka, >99.8%), 2,6-dimethoxypyridine (Aldrich, 98%), 2-chloropyridine and 3-chloropyridine (both TCI, >98%) were also used without additional purification. 8-Aminoquinaldine (REAKHIM, 'pure') was recrystallized from ethyl ether. 2-Methoxypyridine (Aldrich, 98%) was fractionally distilled from $MgSO_4$ under reduced pressure. Solutions of trifluoromethanesulfonic acid (Aldrich, 99+%) and triethylamine (Aldrich, 99%) were used as the acidic and basic titrants, respectively. Copper(II) perchlorate, $Cu(ClO_4)_2$, was prepared by treating copper(II) oxide (Wako Pure Chemicals, 99.9%) with perchloric acid (Wako Pure Chemicals, Super Special Grade), recrystallized from distilled water, and vacuum dried. Copper(II) triflate, $Cu(CF_3SO_3)_2$, was prepared according to the literature procedure [11].

2.2. Measurements

The UV–Vis absorption spectra were measured using a Hitachi U-3000 spectrophotometer. The spectral change for the reaction of the porphyrin with the copper(II) ion in acetonitrile was measured by a stopped-flow apparatus equipped with a rapid-scanning spectrophotometer (RSP-801, Unisoku, Japan). The temperature of all the measurements was controlled within ±0.1 °C using a circulating thermostated water bath. The water content of the solutions was determined by Karl Fischer titration to be between 1.3 and 3 mM with the one exception of 20 mM for the spectrophtometric titration study and less than 15 mM for the kinetic study.

2.3. Spectrophotometric titration

The spectrophotometric titration method in a glove box was very similar to the one used in a previous study [17]. The main difference from the previous study was that the saturated solutions of the porphyrins were made and these solutions were diluted for titration experiments due to the poor solubility of the porphyrins in acetonitrile. The dilution factor was from 1.1 to 5, depending on the porphyrin and the reference compound. The concentrations of the porphyrins in the titration solutions were from 0.8 to 2×10^{-6} M. The concentrations of the other bases were usually in the 10^{-5} M range and the total concentration of the bases in our experiments never exceeded 1.8×10^{-4} M.

2.4. Calculation methods

The basicity of a polybasic base, H_mB^{m+}, in solvent S is defined using Eq. (1) and is expressed as the dissociation constant, K_a, of the conjugate acid, $H_{m+1}B^{(m+1)+}$, of the base, H_mB^{m+}, or more commonly its negative logarithm, pK_a.

$$H_{m+1}B^{(m+1)+} + S \overset{K_a}{\rightleftharpoons} H_mB^{m+} + HS^+ \quad (1)$$

$$K_a = \frac{a(H_mB^{m+})a(HS^+)}{a(H_{m+1}B^{(m+1)+})} \quad (2)$$

$$pK_a = -\log\left(\frac{a(H_mB^{m+})a(HS^+)}{a(H_{m+1}B^{(m+1)+})}\right) \quad (3)$$

If $m=0$, then $H_{m+1}B^{(m+1)+}$ is the conjugate cationic acid of the neutral base B and if $m=+1$, then $H_{m+1}B^{(m+1)+}$ is the conjugate dicationic acid of cationic base HB^+.

To exclude the necessity for measuring the hydrogen ion activity (see Eq. (3)), we studied the equilibrium between two bases, $H_nB_1^{n+}$ and $H_mB_2^{m+}$.

$$H_mB_2^{m+} + H_{n+1}B_1^{(n+1)+} \rightleftharpoons H_{m+1}B_2^{(m+1)+} + H_nB_1^{n+} \quad (4)$$

The relative basicity, ΔpK_a, of bases $H_nB_1^{n+}$ and $H_mB_2^{m+}$ is expressed as:

$$\Delta pK_a = pK_a(H_{m+1}B_2^{(m+1)+}) - pK_a(H_{n+1}B_1^{(n+1)+}) = \log\left(\frac{a(H_{m+1}B_2^{(m+1)+})a(H_nB_1^{n+})}{a(H_mB_2^{m+})a(H_{n+1}B_1^{(n+1)+})}\right) \quad (5)$$

The mixture of bases as well as both bases separately were titrated with an optically transparent acid or base and the data for the calculations was obtained from the UV–Vis spectra. From each titration experiment, the ΔpK_a was determined as the mean of 5–20 values using mostly the same calculation methods as previously described [17–19]. For the complicated cases appearing in this study, where $pK_a(HB^+)$ and $pK_a(H_2B^{2+})$ of the two polybasic porphyrins are too close (less than 3 pK_a units) to observe the pure spectrum of the monoprotonated form (HB^+), a different approach was used. At wavelength λ, the absorbance of a solution (with a 1 cm path length) containing all forms of the dibasic compound B is expressed as follows:

$$A^\lambda = [B]\varepsilon_B^\lambda + [HB^+]\varepsilon_{HB^+}^\lambda + [H_2B^{2+}]\varepsilon_{H_2B^{2+}}^\lambda \quad (6)$$

The analytical concentration of the base expresses:

$$C = [B] + [HB^+] + [H_2B^{2+}] \quad (7)$$

To use the least-squares of the linear combinations method (see [18]), the spectra of B, HB^+ and H_2B^{2+} are needed. The absorbances of the solutions containing only B (Eq. (8)) or H_2B^{2+} (Eq. (10)) are experimentally obtained and the absorbance of HB^+ (Eq. (9)) is calculated as described below.

$$A_B^\lambda = C\varepsilon_B^\lambda \quad (8)$$

$$A_{HB^+}^\lambda = C\varepsilon_{HB^+}^\lambda \quad (9)$$

$$A_{H_2B^{2+}}^\lambda = C\varepsilon_{H_2B^{2+}}^\lambda \quad (10)$$

In the solutions of low acidity, the concentration of H_2B^{2+} is very low and the third member in Eqs. (6) and (7) can be ignored. Expressing the ε–s from Eqs. (8) and (9) and HB^+ from Eq. (7) and substituting these into Eq. (6) gives:

$$A^\lambda = \frac{[B]}{C} A_B^\lambda + \frac{C-[B]}{C} A_{HB^+}^\lambda \quad (11)$$

After rearrangement and substituting $x=[B]/C$ and $y=1/(1-[B]/C)$, we get for i-th spectrum.

$$A_{HB^+(i)}^\lambda = y_iA^\lambda - x_iy_iA_B^\lambda \quad (12)$$

On the other hand, in the solutions of high acidity, the concentration of B approaches zero and in a similar way, we can express the absorbance of HB^+:

$$A_{HB^+(i)}^\lambda = y_iA^\lambda - x_iy_iA_{H_2B^{2+}}^\lambda \quad (13)$$

If we have a set spectra for which either Eq. (12) or Eq. (13) can be applied, we then get the spectrum of HB^+ using the least-squares minimization. The minimized quantities are the sums S (Eq. (15)) of dispersions σ_λ^2 (Eq. (14)) of $A_{HB+(i)}^\lambda$ over the range of wavelengths λ_{start}–λ_{end}. This range of wavelengths was chosen so that it would incorporate the most characteristic absorption bands of the three forms of base B.

$$\sigma_\lambda^2 = \frac{\sum_{i=1}^{n}(\bar{A}_{HB^+}^\lambda - A_{HB^+(i)}^\lambda)^2}{n-1} \quad (14)$$

$$S = \sum_{\lambda=\lambda_{start}}^{\lambda_{end}} \sigma_\lambda^2 \quad (15)$$

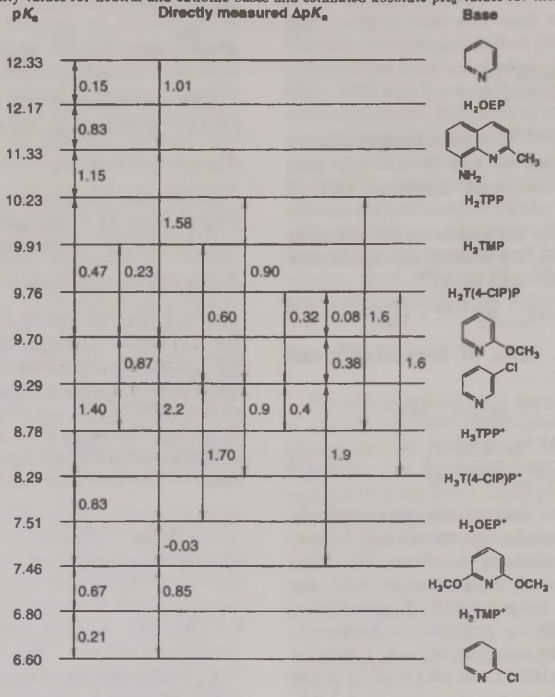
$\bar{A}_{HB^+}^\lambda$ is the average of the $A_{HB^+(i)}^\lambda$ values at wavelength λ. The adjusted quantities are x_i and y_i. In addition, if the isosbestic points were present in the wavelength range λ_{start}–λ_{end}, then the spectra were 'pushed' through the isosbestic points by setting higher weights (usually 100 times) to the dispersions $\sigma^2_{\lambda_{isosbestic}}$ in Eq. (15). The spectrum of HB^+ obtained this way is used in the ΔpK_a calculation procedure to calculate the difference between $pK_a(HB^+)$ and $pK_a(H_2B^{2+})$. If in Eq. (5), $m=n=0$, then we assumed that the ratio of the activity coefficients does not change during titration experiments. If $m=+1$ and $n=0$, then we calculated the activity coefficients using the Debye–Hückel equation as described in [20]. It appears that in very diluted solutions ($<10^{-4}$ M), the corrections arising from the accounting activity coefficients are very small (<0.04 pK_a units).

3. Results

3.1. Basicity of porphyrins

The relative basicity measurements between ten compounds were carried out and 26 relative acid–base

Table 1
Directly measured relative basicity values for neutral and cationic bases and estimated absolute pK_a values for their conjugate acids in acetonitrile



equilibrium constants were found. These values produce the self-consistent and continuous basicity scale presented in Table 1. Multiple overlapping measurements make the results more reliable and help to estimate their self-consistency. The entire basicity range involves at least two independent pathways of measurements and the relative basicity of any two bases can be obtained by combining at least two independent sets of measurements. The reversibility of the protonation/deprotonation process of all bases was checked. All equilibria were reached in minutes and were stable.

The method used in this study has the disadvantage that only the relative basicity can be determined. To assign the absolute pK_a values for the conjugate acids of the bases, the scale has to be anchored to a reference compound or compounds for which the pK_a value(s) are known. The absolute pK_a values of the bases were calculated in a way similar to the previous studies [17,19]. The scale was anchored to the pK_a value (12.33) of pyridine [21]. The precision s of the measurements was calculated as in [19]; $n_m = 26$, $n_c = 14-1 = 13$. For our results, $s = 0.09$ pK_a units.

3.2. Formation of the sitting-atop complex in acetonitrile

Fig. 1 shows the spectra of H_2OEP and its copper(II) complexes in acetonitrile. Upon the addition of copper(II) ion to the porphyrin solution, the spectrum rapidly changes to spectrum **B** in Fig. 1 with broader and weaker Soret band. The characteristic broader Soret band of the product indicates that the porphyrin ring should be distorted from the planar structure. The addition of nitrogen bases such as pyridine to the obtained solution causes a spectral change to that of the copper(II)–OEP complex, [Cu(OEP)]. These findings indicate that the intermediate species still binds the pyrrole protons and that the added pyridine acts as an acceptor of the pyrrole protons to accelerate the formation of [Cu(OEP)]. Without any bases, acetonitrile cannot easily accept these pyrrole protons due to its very weak basicity. A similar phenomenon was previously observed for the reaction of H_2TPP with copper(II) triflate in acetonitrile [11]. The reaction intermediate was ascribed to the so-called sitting-atop complex. Although the basicity of H_2OEP is much higher than that of H_2TPP, it is reasonably understood that the

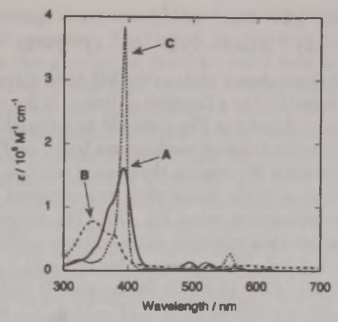


Fig. 1. UV–Vis absorption spectra of H_2OEP and copper(II) complexes in acetonitrile. H_2OEP (**A**), the product after the reaction of H_2OEP with copper(II) ion (**B**), [Cu(OEP)] (**C**). Wavelengths of the absorption maximum at the Soret band are 393, 348, and 394 nm for **A**, **B**, and **C**, respectively.

reaction intermediate was observed for both porphyrins in acetonitrile.

The mole ratio method was employed in order to clarify the composition of the sitting-atop complex. The spectra of solutions containing a constant amount of H_2OEP and different concentrations of copper(II) ion were measured using the rapid scanning spectrophotometer. As shown in Fig. 2, the plot of the absorbance of the solution at 340 and 393 nm as a function of the ratio of the total concentrations of copper(II) ion and H_2OEP clearly demonstrates that the introduction of the first equivalent amount of copper(II) ion to the H_2OEP solution does not cause any spectral change and that the produced sitting-atop complex contains two copper(II) ions based on the inflection point at $C_{Cu}/C_{OEP} = 2$. These findings indicates the two-step reaction for the sitting-atop complex formation reaction:

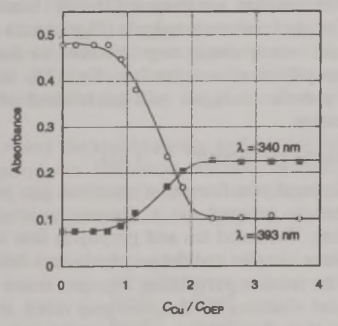


Fig. 2. Dependence of the absorbance on the ratio of the total concentrations of copper(II) ion and H_2OEP for the reaction of H_2OEP and copper(II) triflate at $T = 25.0$ °C. Total concentration of the porphyrin is 2.94×10^{-6} M.

$$H_2OEP + Cu^{2+} \rightleftharpoons [H_2OEP \cdots Cu^{2+}] \quad (16)$$

$$[H_2OEP \cdots Cu^{2+}] + Cu^{2+} \rightleftharpoons [H_2OEP \cdots (Cu^{2+})_2] \quad (17)$$

The mole ratio dependence of the absorbance A shown in Fig. 2 was analyzed by the least squares fitting program employing the following equation:

$$A = \frac{\varepsilon_1 + \varepsilon_2 K_1[Cu^{2+}] + \varepsilon_3 K_1 K_2 [Cu^{2+}]^2}{1 + K_1[Cu^{2+}] + K_1 K_2 [Cu^{2+}]^2} C_{OEP} \quad (18)$$

where ε_1, ε_2, and ε_3 represent the molar absorption coefficients of H_2OEP, $[H_2OEP \cdots Cu^{2+}]$, and $[H_2OEP \cdots (Cu^{2+})_2]$, respectively, and K_1 and K_2 are the equilibrium constants of reaction 16 and 17, respectively. The determined value of $\varepsilon_2 C_{OEP}$ is 0.071 ± 0.010 (340 nm) and 0.486 ± 0.021 (393 nm), which can be considered to be equal to that of $\varepsilon_1 C_{OEP}$ (0.074 ± 0.006 at 340 nm and 0.481 ± 0.006 at 393 nm) within experimental error. These findings indicate that the introduction of the first equivalent amount of the copper(II) ion to the H_2OEP solution is spectrophotometrically silent. The equilibrium constants are calculated to be $\log(K_1\ (M^{-1})) = 9.2 \pm 0.4$ and $\log(K_2\ (M^{-1})) = 7.8 \pm 0.4$ ($T = 25.0$ °C) for the copper(II) triflate system based on the assumption that ε_2 is equal to ε_1. The solid lines in Fig. 2 are calculated curves based on the obtained equilibrium constants and explain the results very well. For copper(II) perchlorate, similar results were obtained, and the equilibrium constants were determined to be $\log(K_1\ (M^{-1})) = 9.4 \pm 0.3$ and $\log(K_2\ (M^{-1})) = 8.2 \pm 0.3$ ($T = 25.0$ °C). These findings indicate that the counter anion of copper(II) has little effect on the sitting-atop complex formation reaction of H_2OEP.

3.3. Kinetics of the sitting-atop complex formation reaction

The reaction of H_2OEP with the copper(II) ion is too fast to be followed by a stopped-flow apparatus under the pseudo-first-order conditions where the copper(II) ion exists in large excess over the porphyrin in acetonitrile. The reaction rate was, therefore, measured under second-order conditions. The reaction was spectrophotometrically followed by monitoring the absorbance change associated with the formation of the sitting-atop complex under the conditions where the reaction quantitatively proceeds. Taking into account the results of the mole ratio method, several reaction mechanisms should be considered to analyze the reaction curve. The first candidate is the series of second-order reactions in which the first step (reaction 16) is the rate-determining step. In this case, the rate equation can be expressed by:

$$(b - 2a)^{-1} \ln\{a(b - 2x)b^{-1}(a - x)^{-1}\} = kt \quad (19)$$

where a and b represent the initial concentrations of H_2OEP and copper(II) ion, respectively, x is the

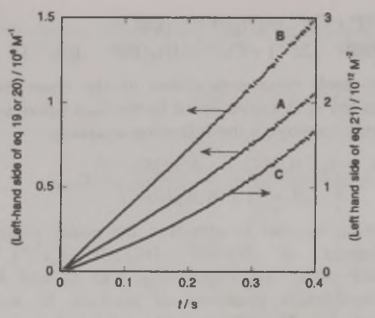


Fig. 3. Kinetic plots for the reaction of H_2OEP and copper(II) triflate at $T = 25.0$ °C. Each plot is based on Eq. (19) (**A**), Eq. (20) (**B**), and Eq. (21) (**C**). $C_{Cu} = 6.83 \times 10^{-6}$ M, $C_{OEP} = 2.94 \times 10^{-6}$ M.

concentration of the produced sitting-atop complex, and k is the second-order rate constant of reaction 16. The second possible mechanism is the series of second-order reactions in which the second step (reaction 17) is the rate-determining step, i.e. the mechanism including the fast pre-association equilibrium between H_2OEP and copper(II) ion. In this case the rate equation can be expressed by:

$$(b-a)^{-1} \ln\{a(b-x)b^{-1}(a-x)^{-1}\} = kt \qquad (20)$$

where k is the second-order rate constant of reaction 17. Because the intermediate species $[H_2OEP\cdots Cu^{2+}]$ is produced immediately after mixing the solution of H_2OEP and copper(II) ion, the value of b is set at $C_{Cu}-C_{OEP}$. Another reaction mechanism is the third-order reaction with two reactants, i.e. $-d[H_2OEP]/dt = k[Cu^{2+}]^2[H_2OEP]$. In this case the rate equation can be expressed as:

Table 2
Second-order rate constant of the sitting-atop complex formation of H_2OEP in acetonitrile at $T = 25.0$ °C

Copper(II) compound	C_{OEP} (10^{-6} M)	C_{Cu} (10^{-6} M)	k (10^6 M^{-1} s^{-1}) [a]
Copper(II) triflate	2.94	6.83	3.6
	2.94	8.14	3.5
	2.34	8.14	3.1
	2.94	9.54	3.5
	2.34	9.58	3.1
	2.94	10.8	3.0
	2.34	11.2	3.0
	2.39	12.9	3.0
copper(II) perchlorate	2.81	7.32	3.1
	2.81	8.68	3.1
	2.81	10.1	3.0
	2.81	11.6	2.9

$$2(2a-b)^{-1}\{(b-2x)^{-1} - b^{-1}\} + 2(2a-b)^{-2} \ln\{a(b-2x)b^{-1}(a-x)^{-1}\} = kt \qquad (21)$$

In Fig. 3 are shown plots of the left-hand sides of Eqs. (19), (20) and (21) as a function of time t. If the equation was correct, the plot in Fig. 3 should be a straight line. It is clear that the reaction mechanisms based on Eqs. (19) and (21) are not the case for the present reaction. On the other hand, a fairly linear plot was obtained for the reaction mechanism using Eq. (20). The second-order rate constant thus obtained under various conditions is listed in Table 2. The average value is $k = (3.2 \pm 0.3) \times 10^6$ M^{-1} s^{-1} ($T = 25.0$ °C) for copper(II) triflate and $k = (3.0 \pm 0.2) \times 10^6$ M^{-1} s^{-1} ($T = 25.0$ °C) for copper(II) perchlorate. The activation enthalpy and activation entropy were determined from the temperature dependence of the second-order rate constant in the temperature range between 15.0 and 35.0 °C for the reaction of copper(II) triflate: $\Delta H^{\ddagger} = 58 \pm 2$ kJ mol^{-1}, $\Delta S^{\ddagger} = 74 \pm 6$ J K^{-1} mol^{-1}. Almost the same rate constant was obtained for both copper(II) compounds indicating that the counter anions (triflate and perchlorate ions) of the copper(II) salts do not affect the rate of the sitting-atop complex formation reaction. Since both anions have a weak coordination ability toward metal ions, they at most exist as a solvent-separated ion pair with the copper(II) ion and do not directly coordinate to the copper(II) ion even in acetonitrile. These findings are consistent with the results for the solution chemistry of the copper(II) ion in acetonitrile, where the copper(II) ion exists as a hexasolvated ion $[Cu(AN)_6]^{2+}$ [14,22].

4. Discussion

The kinetics of the metalloporphyrin formation reaction has been extensively studied for many kinds of porphyrins and metal ions in various solvents in order to clarify the reaction mechanism [23–26]. Usually the reaction of the free-base porphyrin (H_2por) with a metal ion occurs via a single step reaction to form the metalloporphyrin as expressed in Eq. (22), and two pyrrole protons in H_2por will be released after the complexation.

$$M^{n+} + H_2por \rightleftharpoons [M(por)]^{(n-2)+} + 2H^+ \qquad (22)$$

Metalloporphyrin formation reactions are generally interpreted to proceed via a dissociative-interchange mechanism. The metal ion and porphyrin first form an outer sphere complex and the porphyrin core deforms to present the reactive pyrrolenine nitrogen atoms having lone-paired electrons to the incoming metal ion. The metal ion then dissociates a solvent molecule to form a bond between the nitrogen atom of the porphyrin and metal ion. A sitting-atop (or activated) complex is thus formed. In the solvents with a higher Brønsted basicity,

the following deprotonation should occur from the activated complex. The trend in the rate constant variation for a series of divalent metal ions including the 3d-block metal ions for the solvent exchange of DMF and for the metallation of H_2TPP and its *N*-methylated derivative in DMF is very similar, although the metallation rate is slower than the solvent exchange rate by several orders of magnitude [27–29]. Therefore, the rate-determining step for the metallation of the porphyrin is thought to be the dissociation of coordinated solvent molecules from the metal ion in the outer-sphere association complex. The overall rate constant for the metalloporphyrin formation, k_f, is given by the product of the equilibrium constant for the outer-sphere complex formation (K_{OS}), the equilibrium constant for the porphyrin ring deformation (K_D), and the rate constant for the dissociation of the coordinating solvent molecules in the inner-sphere of the metal ions (k): $k_f = K_{OS}K_Dk$. The slow metalation rate of the porphyrin can be explained by the small value of K_D.

The transient species, which includes both the incoming metal ion and two pyrrole protons, can exist as an intermediate of the complex formation reaction of the porphyrin, and this intermediate is assumed to be the sitting-atop complex first proposed by Fleischer and Wang [7]. However, in solvents such as water and DMF where the pyrrole protons of the intermediate may be easily released, the sitting-atop intermediate can hardly be observed and the metalloporphyrin is simply produced as the product during the metallation reaction. On the other hand, the existence of the sitting-atop complexes was evidenced for the protoporphyrin IX derivatives, H_2TPP, and other porphyrins including H_2OEP in the solvents such as chloroform and acetonitrile [7–11,14,15]. It is obvious that the low basicity of these solvents is ascribed to the stability of the sitting-atop intermediate. Although the sitting-atop complex was claimed to have a composition of M^{n+}:porphyrin = 1:1 for many cases, the results of the present study clearly suggest that two copper(II) ions are included in the sitting-atop complex of H_2OEP as evidenced by the mole ratio method. The proposed structure of this sitting-atop complex is shown in Chart 1.

Concerning the stoichiometry and mechanism of the metalloporphyrin formation reaction, Hambright has proposed that two metal ions are always required for the metal ion incorporation into porphyrins, one to distort the planar porphyrin core, and the second to incorporate from the opposite side of the porphyrin [30]. This proposal was supported by many kinetic studies. Tanaka et al. investigated the kinetics of the reaction of H_2TPP with copper(II), zinc(II), and cadmium(II) ions in DMF, and evidence for the involvement of the two metal ions in the reaction was obtained from the kinetics of the reaction [27]. The metalloporphyrin formation reaction is first-order with respect to the porphyrin, and the pseudo-first-order rate constant was obtained as:

$$k_{obsd} = (k_1K[M^{2+}] + k_2K[M^{2+}]^2)(1 + K[M^{2+}])^{-1} \quad (23)$$

On the basis of the rate law, the following reaction mechanism was proposed.

$$H_2TPP + Cu^{2+} \overset{K}{\rightleftharpoons} [H_2TPP \cdots Cu^{2+}] \quad (24)$$

$$[H_2TPP \cdots Cu^{2+}] \overset{k_1}{\rightarrow} [Cu(TPP)] + 2H^+ \quad (25)$$

$$[H_2TPP \cdots Cu^{2+}] + Cu^{2+} \overset{k_2}{\rightarrow} [Cu(TPP)] + Cu^{2+} + 2H^+ \quad (26)$$

The k_2 path involving two metal ions is consistent with Hambright's proposal. Similar kinetic features were observed for the other metalloporphyrin formation reaction [31,32]. On the other hand, it was demonstrated that the rate of the metalloporphyrin formation reaction of the one-face-hindered fenced porphyrin with copper(II), zinc(II), and cadmium(II) ions is about the same as for the sterically unhindered porphyrin in DMF or pyridine [33], and these findings are considered to be the evidence against the Hambright mechanism. Nevertheless, the Hambright mechanism should apply to the complex formation reaction of many sterically unhindered porphyrins.

The two-step reactions of the sitting-atop complex formation of H_2OEP observed in the present study corresponded to the mechanism in which two metal ions interact with the porphyrin molecule from the opposite side of the porphyrin core as proposed by Hambright. This type of sitting-atop complex should be regarded as the intermediate of the metalloporphyrin formation reaction in the solvents such as DMF and water. What is significant is the extremely large K value for the formation of the outer sphere association complex of H_2OEP with copper(II) ion in acetonitrile. The corresponding value of the pre-association constant K was reported to be 1.6×10^4 M^{-1} at 25.0 °C for the reaction of H_2TPP with copper(II) ion in DMF [27]. Formation of the outer sphere complex is usually postulated as the primary step for the ligand substitution reaction of the metal complexes. The formation constant of the outer sphere complex can be theoretically estimated to be 0.3 for the reaction of the noncharged ligand at the distance of closest approach of two species of 5 Å [34,35]. Although the pre-association

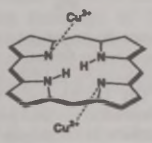


Chart 1. Proposed structure of the sitting-atop complex.

constant for the reaction of H_2TPP with Cu^{2+} in DMF is much greater than the predicted value for the outer-sphere complex, the first step of the metalation reaction (Eq. (24)) can be ascribed to the outer sphere complex formation due to the following spectral features. Under the conditions where the outer-sphere pre-associated complex [$H_2TPP\cdots Cu^{2+}$] is quantitatively formed, the spectral change for the metallation reaction was reported to be from the spectrum of H_2TPP to that of [Cu(TPP)], and no spectral evidence was obtained for the formation of the outer-sphere complex [27]. This fact indicates that the kinetically detected intermediate [$H_2TPP\cdots Cu^{2+}$] has the same absorption spectrum as that of H_2TPP. A similar spectral feature was observed for the sitting-atop complex formation reaction of H_2OEP in the present study. The pre-association constant of H_2OEP with copper(II) ion in acetonitrile is much greater than that of the H_2TPP case in DMF. Taking into account the low solubility of H_2OEP in acetonitrile being on the order of 10^{-5} M, the large pre-association constant can be rationalized by the stabilization of the H_2OEP molecule by forming the outer sphere complex with copper(II) ion in the solvent.

In the present study, we determined the second-order rate constant for the reaction of H_2OEP with the copper(II) ion. The rate of the sitting-atop complex formation of H_2OEP is faster than that of H_2TPP ($k = 3.6 \times 10^5$ M^{-1} s^{-1}) [11]. The activation enthalpy, $\Delta H^{\ddagger}$, for the H_2OEP and H_2TPP is almost the same, i.e. 58 ± 2 kJ mol^{-1} for the former and 56 ± 5 kJ mol^{-1} for the latter and the activation entropy for the H_2OEP (74 ± 6 J K^{-1} mol^{-1}) is greater than that for H_2TPP (46 ± 19 J K^{-1} mol^{-1}) [11]. The activation enthalpy for these reactions is much greater than the usual ligand substitution and solvent exchange reactions of the copper(II) ion. For example, $\Delta H^{\ddagger}$ for the reaction of the copper(II) ion with ammonia in aqueous solution is 18.8 kJ mol^{-1} [36]. The $\Delta H^{\ddagger}$ for the solvent exchange is 11.5 kJ mol^{-1} for water [37], 24.3 kJ mol^{-1} for DMF [38], and 17.2 kJ mol^{-1} for methanol [39,40]. The larger activation enthalpy for the reaction of the present sitting-atop complex formation can be attributed to the energetic requirement due to the distortion of the porphyrin core prior to the rate-determining step of the binding between the metal ion and porphyrin.

The difference in the reactivity of the porphyrins towards a metal ion has been often discussed in terms of the acid–base properties of the porphyrin as well as the structural properties of the molecule, especially concerning the planarity of the porphyrin core. Before discussing the relationship between the reactivity of the porphyrin with the copper(II) ion and the porphyrin basicity, we will first consider the acid–base properties of H_2OEP and some *meso*-substituted porphyrins. The pK_a value of the conjugate acid of the porphyrins was spectrophotometrically determined in acetonitrile. In Table 1 are shown the pK_a values of the mono- and diprotonated porphyrins. It seems reasonable to assume that the higher the degree of delocalization of the charge introduced by the first proton from the center of the porphine moiety, the closer the second pK_a will be to the first. This explains why H_2OEP has the greatest ΔpK_a of the porphyrins, because the delocalizing power of the phenyl rings of TPP is greater than that of the ethyl groups of OEP. Although the phenyl groups in the free base of TPP are almost perpendicular to the porphyrin core, this does not hold for the mono- and diprotonated forms, for which the porphyrin core is also quite distorted. Such a situation can be seen in the molecular structure of the porphyrin diacid species [41–48]. The crystal structure of the porphyrin diacid species, $[H_4OEP](ClO_4)_2$, $[H_4TPP](ClO_4)_2$, $[H_4TMP](ClO_4)_2$, indicates that the distortion of the porphyrin core from planarity should result from the severe transannular crowding between the four (or three for the monoacid form) hydrogen atoms of the pyrrole nitrogen atoms [48]. The crystal structures of the D_{2d}-symmetry distortion of the porphyrin core depend on the nature of the peripheral substituents. The phenyl and 4-Cl-phenyl substituents can rotate to some extent under the D_{2d}-symmetry distortion, so that they will be more coplanar with the porphyrin core. The mean porphyrin-phenyl group dihedral angles in $[H_4TPP](ClO_4)_2$ are 27(2)°. This enables the resonance interaction that is very efficient in charge delocalization. As a consequence, the ΔpK_a of TPP and T(4-ClP)P is far less than the OEP. TMP is in a different situation from TPP, because it has two *o*-methyl substituents on each aryl substituent. The eight *o*-methyl groups of $[H_4TMP](ClO_4)_2$ are positioned above and below the plane of the porphyrin ring, approximately over the *meso*-carbons. Such a steric hindrance significantly hinders the rotation of the aryl groups about the C_m–C_p bond. The porphyrin-aryl

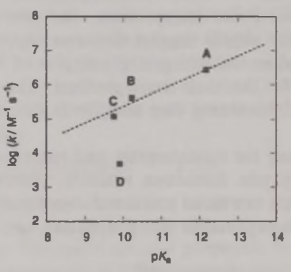


Fig. 4. Relationship between the second-order rate constant k of the reaction of the sitting-atop complex formation with copper(II) triflate and the basicity of the porphyrin for H_2OEP (**A**), H_2TPP (**B**), $H_2T(4\text{-}ClP)P$ (**C**), and H_2TMP (**D**). The pK_a value represents that of the conjugate acid of the corresponding free base porphyrin in acetonitrile. The values of the rate constants for the porphyrins other than H_2OEP are from [15].

group dihedral angles of $[H_4TMP](ClO_4)_2$ are less acute than those of $[H_4TPP](ClO_4)_2$, averaging 63(13) and 75(15)° in the two forms of molecules in the crystal. Consequently, $[H_4TMP](ClO_4)_2$ show substantially smaller core distortions than $[H_4TPP](ClO_4)_2$, though both are considerably more saddled than $[H_4OEP](ClO_4)_2$. These steric effects cause a smaller charge delocalization which is given by the resonance effect than in the case of TPP. However, even without the resonance effect, the mesityl groups are more efficient in charge delocalization than the ethyl groups. As a result, the ΔpK_a of TMP is intermediate between OEP on one hand and TPP on the other hand.

The dependence of the second-order rate constant of the sitting-atop complex formation reaction on the basicity of the porphyrins is shown in Fig. 4. The rate constant decreases in the order of $H_2OEP \gg H_2TPP > H_2T(4\text{-}ClP)P \gg H_2TMP$. The difference in the reactivity can be interpreted by the electronic and steric factors. It has been demonstrated how the reactivity of the porphyrins with metal ions is influenced by the porphyrin basicity. A linear relationship between the log k (k is the second-order rate constant of the metalloporphyrin formation reaction) and the pK_a of the conjugate acid of the monoprotonated porphyrin has been demonstrated, for example, for the reaction of the water-soluble porphyrins with the copper(II) and zinc(II) ions [49,50]. Also, the correlation between the porphyrin reactivity and the reduction potential of the porphyrins has been explored [26,51]. The faster metalloporphyrin formation reaction was observed for the porphyrin with a stronger basicity and a more negative reduction potential for the metal ion incorporation reaction into the water-insoluble porphyrins. The reduction potentials parallel the basicity scale of the porphyrins, and thus the porphyrin bearing the stronger basicity reacts with the zinc(II) ion faster than the less basic porphyrins [51]. As shown in Fig. 4, a similar correlation was observed in the present study except for H_2TMP, which exhibits the additional steric effect on the reaction. The interaction between the metal ion and the free base porphyrin should play an important role in the present sitting-atop complex formation reaction. The electrostatic attraction between the local negative charge on the pyrrolenine nitrogen atoms and the positive charge of the metal ion should drive the outer-sphere association between these two species. The electron density of the nitrogen atom also affects the interaction between the metal ion and the pyrrolenine nitrogen atom of the porphyrin during the rate-determining exchange of the bound solvent molecule by the incoming porphyrin ligand around the metal ion. For H_2TMP, as has already been discussed in a previous paper, the slow rate can be interpreted by the steric hindrance due to the α-methyl groups of the *meso*-aryl substituents of H_2TMP [15]. The *meso*-aryl rings of H_2TMP become more perpendicular to the porphyrin skeleton due to the steric interaction between the methyl groups of the aryl rings and the hydrogen atoms on the pyrrole rings. Due to this steric hindrance, the electrostatic approach between the copper(II) ion and H_2TMP is also inhibited and the more distant rupture in the bonding between the copper(II) ion and the dissociating acetonitrile molecule is also required since it is hard for the pyrrolenine nitrogen to approach closer to the copper(II) ion in the transition state. The correlation between the porphyrin reactivity and the acid-base properties of the porphyrins shown in Fig. 4 can thus be interpreted by the electronic and steric factors of the porphyrins.

5. Supplementary material

Figure showing the temperature dependence of the second-order rate constant for the sitting-atop complex formation reaction of H_2OEP with copper(II) triflate in acetonitrile (Fig. S1). The supplementary material is available from the authors on request.

Acknowledgements

M.I. gratefully acknowledges the helpful discussion of Professor Shigenobu Funahashi of Nagoya University, Japan. I.L. and I.K. gratefully acknowledge the support from grant 4376 from the Estonian Science Foundation.

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IV

Acid−Base Equilibria in Nonpolar Media. 2.[1] Self-Consistent Basicity Scale in THF Solution Ranging from 2-Methoxypyridine to EtP_1(pyrr) Phosphazene

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Received October 9, 2001

Relative ion-pair basicities ΔpK_{ip} of 25 substituted aryl and alkyl iminophosphoranes (phosphazenes) and 20 other N-bases (various pyridines, amines, amidines) have been measured in THF medium using the UV−Vis and/or ^{13}C NMR methods. The ΔpK_{ip} values were corrected for ion pairing using the Fuoss equation to obtain relative ionic basicities ΔpK_a. Based on the measurements, a basicity scale ranging from 2-methoxypyridine to EtP_1(pyrr) and having a total span over 18 pK units has been created. The scale has been anchored to the pK_a value of triethylamine (pK_a = 12.5). The results are compared to pK_a values in various other solvents and in the gas phase. The pK_a values give better correlations than the pK_{ip} values, thus indirectly validating the procedure of correction for ion pairing. The predictability of the basicity together with suitable spectral properties in the UV range make the phenylphosphazenes convenient neutral indicators in the high basicity range where the choice of neutral indicators is very limited.

Introduction

Numerous acidity studies, mostly focused on CH-acids, have been carried out in THF.[2,3,4] Alkali metal amides or carbanions with alkali metal counterions have mostly been used as deprotonating agents. On the basis of these measurements, ion-pair acidity scales have been developed relative to 9-phenylfluorene or fluorene.[2,3,4] Those scales have been anchored to the pK_a values of these reference compounds in aqueous sulfolane and DMSO, respectively. This approach has also been used in constructing acidity scales in various other nonpolar media—cyclohexylamine,[5] benzene,[6,7] dimethoxyethane,[8] diethyl ether,[9] etc.

Contrary to acidity measurements, studies on basicity in THF are scarce. Recently, the Morris group compiled an acidity scale in THF based on NMR measurements including numerous metal hydrides, phosphines, etc.[10] The observed pK_{ip} values were corrected for ion-pairing using the Fuoss equation.[11] Besides the phosphines the pK_{ip} values for a number of other organic compounds were established and the pK_a (here and henceforth the term "pK_a of the base" is understood as the acidity of conjugate acid of the base, and its definition is given in eq 1) value 12.5 of triethylamine was suggested as the secondary standard for anchoring acidity and basicity scales in THF.[10]

In nonpolar media the measured equilibrium constants do not generally reflect the free ion acidity but rather refer to ion pairs. An attempt to suppress the interactions between cations and anions of the CH-acids in nonpolar media was made by the Konovalov group.[12,2] They have used lithium [2.1.1]cryptate as the counterion for the anions of CH-acids. In this cryptate the [2.1.1]cryptand acts as a layer of solvent molecules separating the ions and so eliminating the specific interactions. In a previous work by some of us phosphazene t-BuP_4(dma) (here and henceforth "dma" denotes dimethylamino ($N(CH_3)_2$) and "pyrr" denotes 1-pyrrolidinyl ($N(CH_2\,CH_2)_2$) radical) was used as the deprotonating agent to create an acidity scale in n-heptane.[1] Protonated t-BuP_4(dma) is a large cation with a delocalized charge and has very weak interaction with anions.

We have found earlier that aryl-substituted P_1 phosphazenes $ArN{=}P(R)_3$, where R = dma or pyrr, are suitable indicators together with amines to compile a basicity scale in acetonitrile (AN).[13,14]

In this paper we report a basicity scale in THF medium that incorporates aryl and alkyl P_1 phosphazenes

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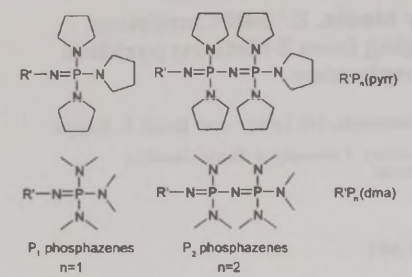
10.1021/jo016185p CCC: $22.00
Published on Web 02/22/2002

Scheme 1



R'N=P(R)$_3$ (see Scheme 1), some aryl P$_2$ phosphazenes R'N=P(R)$_2$N=P(R)$_3$, various substituted pyridines, and several other bases.

For a better comparison with the THF data and as a further extension of the earlier established basicity scale in AN, some additional pK_a values for several bases were also measured in the latter solvent.

In a polar solvent S (water, AN, etc.) at low or moderate concentrations the basicity of base B is defined using eq 1 and is expressed as dissociation constant K_a (eq 2) of the respective conjugate acid HB$^+$ of the base B or, more commonly, its negative logarithm pK_a

$$HB^+ + S \overset{K_a}{\rightleftharpoons} B + HS^+ \quad (1)$$

$$K_a = \frac{a(B)\,a(HS^+)}{a(HB^+)} \quad (2)$$

In media of low polarity ($D \leq 15-20$)[15] there is an extensive ion-pairing and formation of aggregates between ions and neutral molecules (homo- and heteroconjugation) that has to be considered. The extent of ion-pairing depends on the solvent, the size of the ions, and the charge distribution in ions. The general trend is that small ions tend to form solvent-separated ion-pairs (SSIP) (eq 3) while large ions with delocalized charge tend to form contact ion-pairs (CIP) (eq 4).

$$HB^+ + A^- \rightleftarrows HB^+_s \cdot A^-_s \quad (3)$$

$$HB^+ + A^- \rightleftarrows [HBA]_s \quad (4)$$

In media of low polarity the homo- and heteroconjugation processes (eqs 5 and 6, respectively) occur to a much lower extent as compared to ion-pairing. At a low concentration these conjugation processes between ions and neutrals can be neglected because most of the ions present in the solution are ion-paired.

$$HB^+ + B \rightleftarrows BHB^+ \quad (5)$$

$$HB^+ + B_1 \rightleftarrows BHB_1^+ \quad (6)$$

To exclude the necessity for measuring the hydrogen ion activity (see eq 2), the equilibrium between two (ion-paired) bases B$_1$ and B$_2$ was studied:

$$B_2 + HB_1^+A^- \overset{K_d^{HB_1^+A^-}}{\rightleftharpoons} B_2 + HB_1^+ + A^- \overset{K}{\rightleftharpoons}$$

$$HB_2^+ + B_1 + A^- \overset{1/K_d^{HB_2^+A^-}}{\rightleftharpoons} HB_2^+A^- + B_1 \quad (7)$$

The K_d are the dissociation constants of the respective ion pairs. The directly measured quantity is the relative ion-pair basicity, ΔpK_{ip}, of bases B$_1$ and B$_2$. It is expressed as follows:

$$\Delta pK_{ip} = pK_{ip}(HB_2^+A^-) - pK_{ip}(HB_1^+A^-) = \log \frac{KK_d^{HB_1^+A^-}}{K_d^{HB_2^+A^-}} = \log \frac{a(HB_2^+A^-)\,a(B_1)}{a(HB_1^+A^-)\,a(B_2)} \quad (8)$$

If the K_d values can be measured or estimated then the pK_a (an estimate of the pK_a) can be found as follows:

$$\Delta pK_a = pK_a(HB_2^+) - pK_a(HB_1^+) = \Delta pK_{ip} - \log \frac{K_d^{HB_1^+A^-}}{K_d^{HB_2^+A^-}} \quad (9)$$

Experimental Section

Chemicals. Most of the compounds were the same as used earlier.[13,14] The following compounds were of commercial origin and were used without further purification: TMG (Aldrich, 99%), 1,8-bisdimethylaminonaphthalene (DMAN) (Aldrich, >99%), *tert*-butylimino-tris(dimethylamino)phosphorane (*t*-BuP$_1$(dma)) base (Fluka, >98%), *tert*-butylimino-tris(pyrrolidino)phosphorane (*t*-BuP$_1$(pyrr)) base (Fluka, ≥98%), and all substituted anilines (Aldrich) used for synthesis of new phosphazenes. Aniline (Reakhim) was purified by refluxing with acetone and following recrystallization of HCl salt.[16] *N,N*-Dimethylaniline (Reakhim) was purified by refluxing with acetic anhydride and following recrystallization of HCl salt.[16] *o*-Toluidine (Reakhim) was distilled twice, the second time from CaH_2. Pyrrolidine (Merck or Fluka, >98%) was distilled from NaOH and was kept over NaOH. 2-Methoxypyridine (Aldrich, 98%) was distilled fractionally from $MgSO_4$ under reduced pressure.

Solutions of trifluoromethanesulfonic acid (TfOH) (Aldrich, 99+%) or methanesulfonic acid ($MeSO_3H$) (Fluka, >99%) were used as acidic titrants. Phosphazene bases EtP$_1$(pyrr)[17] or EtP$_2$-(dma) (Fluka, >98%) were used as basic titrants.

Solvents. THF was used as purchased (Romil, >99.9%, Super Purity Solvent, water content <0.005%) or purified (REAKHIM, pure) as follows:[16] THF was kept on KOH pellets and then boiled 2−2.5 h on CaH_2 in the flow of dried Ar. After that, THF was distilled through the 60 cm long column filled with steel rings, and the fraction with bp 66.2−66.5 °C was collected. To this distillate was added $LiAlH_4$, and then it was boiled with same column for 1 h under Ar and then distilled fractionally, fraction with bp 66.2−66.3 °C was collected. For syntheses was used THF stored over KOH and distilled from $LiAlH_4$. CCl_4 was distilled from P_2O_5, benzene from $LiAlH_4$ (all Reakhim).

We observed that ~0.2 M TfOH polymerizes THF within a few hours, and in dilute solutions the active concentration of this acid was lower than analytical concentration. Also we observed that with some batches of THF the absorbance below 280 nm decreased upon addition of TfOH.

AN (Romil, >99.9%, Super Purity Solvent (far UV), water content <0.005%) was the same used in previous work[14] and was used without further purification.

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General Method for the Synthesis of the P_1 Phosphazenes. New phenyl-substituted P_1 phosphazenes (s-PhP_1(pyrr)) were synthesized by the Kirsanov reaction[18] according to the scheme:

$$s\text{-}PhNH_2 + PCl_5 \xrightarrow[CCl_4]{} s\text{-}PhN{=}PCl_3 \xrightarrow[Bz\ or\ THF]{PyrrH} s\text{-}PhP_1(pyrr) \quad (10)$$

An amount of 20 mmol of the corresponding aniline was dissolved in 20 mL of CCl_4, and 20 mmol of PCl_5 was added (in some cases it was necessary to cool the mixture in the beginning of the reaction). The mixture was warmed and refluxed until the evolution of HCl finished. Insoluble in CCl_4, trichlorophosphazenes were filtered off and washed successively with CCl_4, benzene, and diethyl ether. When the trichlorophosphazene was soluble in CCl_4, the solvent was removed on a rotavapor, and the residue was dried in vacuo.

An amount of 15 mmol of obtained phenylimino trichlorophosphazene was dissolved in 40 mL of dried THF (in the case of 2-NO_2-4-CF_3 and 5-Cl-2-NO_2 substituents, 50 mL of dry benzene was used instead), and 90 mmol of pyrrolidine (solution in 15 mL of THF) was added by means of a dropping funnel. The mixture was heated to 50 °C and stirred ca. 1 h. Then the mixture was cooled to +5 °C, the HCl salt of pyrrolidine was filtered off, and the solvent was removed at reduced pressure (60 °C/5 Torr). The brown residue was washed with 70% aqueous solution of $EtNH_2$. The crystals of the product were collected, washed with 40% aqueous $EtNH_2$ solution, dried in vacuo, and recrystallized (2-NO_2-4-CF_3-$C_6H_3P_1$(pyrr) was isolated as HBF_4 salt).

4-CF_3-$C_6H_4P_1$(pyrr). According to the general method, after recrystallization from 70% $EtNH_2$ aq soln, 1.9 g of colorless crystals was obtained, yield 32%, mp 102.4–103.3 °C. Anal. Calcd for $C_{19}H_{28}F_3N_4P$: C, 56.99%; H, 7.07%; N 13.99%. Found: C, 57.18%; H, 7.12%; N, 14.01%. ^{1}H NMR (200 MHz, CD_3Cl) δ 1.83 (m, 12H), 3.19 (dt, 12H, J_{H-H} = 6.6, J_{P-H} = 3.4), 6.78 (d, 2H, J_{H-H} = 8.5), 7.27 (d, 2H, J_{H-H} = 8.5). ^{13}C NMR (50 MHz, CD_3Cl) δ 26.4 (d, J_{C-P} = 8.0), 46.9 (d, J_{C-P} = 4.0), 117.3 (q, J_{C-F} = 32.3), 122.2 (d, J_{C-P} = 18.2), 125.72 (d, J_{C-P} = 2.7), 125.74 (q, J_{C-F} = 270.3), 155.7.

5-Cl-2-NO_2-$C_6H_3P_1$(pyrr). According to the general method, after recrystallization from 70% $EtNH_2$ aq soln, 4.5 g of yellowish crystals was obtained, yield 73%, mp 82.7–83.2 °C. Anal. Calcd for $C_{18}H_{27}ClN_5O_2P$: C, 52.49%; H, 6.61%; N, 17.0%. Found: C, 52.47%; H, 6.59%; N, 16.80%. ^{1}H NMR (200 MHz, THF) δ 1.8 (overlapped by solvent, 12H), 3.19 (dt, 12H, J_{H-H} = 6.7, J_{P-H} = 4.2), 6.42 (dd, 1H, J_{H-H} = 8.6, 2.2), 6.77 (dd, 1H, J_{H-H} = 2.2, J_{P-H} = 1.0), 7.42 (dd, 1H, J_{H-H} = 8.6, J_{P-H} = 2.4). ^{13}C NMR (50 MHz, THF) δ 27.0 (d, J_{C-P} = 7.7), 47.6 (d, J_{C-P} = 3.7), 114.34, 124.2 (d, J_{C-P} = 9.4), 126.0, 137.4, 144.2 (d, J_{C-P} = 24.3), 147.9 (d, J_{C-P} = 7.4).

2-NO_2-4-CF_3-$C_6H_3P_1$(pyrr). The raw product (4.9 g) synthesized according to the general method was dissolved in 40 mL of 5% aq HCl, and the solution of 1.3 g of $NaBF_4$ in water was added. The precipitate was filtered, dried, and recrystallized from ethyl acetate. A 3.3 g (6.2 mmol) amount of light yellow crystals of the HBF_4 salt of the phosphazene (mp 186–187 °C) was dissolved in the mixture of 8 mL MeOH and 12 mL AN. A 6.2 mmol amount of MeOK as 25% solution in MeOH was added. The solvent was removed, and the residue was refluxed with hexane for 0.5 h. The mixture was filtered, and the hexane was removed under the reduced pressure. The brownish solid residue was recrystallized from 70% $EtNH_2$ aq soln to give 2.0 g (yield 72.5%, mp 70.5–71.2 °C) of light yellow crystals of the product. Anal. Calcd for $C_{19}H_{21}F_3N_5O_2P$: C, 51.23%; H, 6.11%; N, 15.72%. Found: C, 51.42%; H, 6.12%; N, 15.66%. ^{1}H NMR (200 MHz, THF) δ 1.8 (overlapped by solvent, 12H), 3.20 (dt, 12H, J_{H-H} = 6.7, J_{P-H} = 4.2), 6.90 (d, 1H, J_{H-H} = 8.8), 7.29 (dd, 1H, J_{H-H} = 8.8, 2.5, J_{P-H} = 0.6), 7.72 (br m, 1H, J_{H-H} = 2.5). ^{13}C NMR (50 MHz, THF) δ 27.1 (d, J_{C-P} = 7.7), 47.6 (d, J_{C-P} = 4.0), 115.5 (q, J_{C-F} = 33.8), 122.4 (br m), 125.0 (d, J_{C-P} = 9.2), 125.7 (q, J_{C-F} = 270.2), 128.2 (m), 144.9 (d, J_{C-P} = 26.3), 150.0 (d, J_{C-P} = 7.8).

2,6-Cl_2-4-NO_2-$C_6H_2P_1$(pyrr). According to the general method, 4.2 g (yield 63.2%, mp 95.3–96.9 °C, recrystallized from 70% $EtNH_2$ aqueous solution or from hexane) of yellow crystals was obtained. Anal. Calcd for $C_{18}H_{26}Cl_2N_5O_2P$: C, 48.77%; H, 5.91%; N, 15.8%. Found: C, 48.37%; H, 5.87%; N, 15.80%. ^{1}H NMR (200 MHz, $CDCl_3$) δ 1.82 (m, 12H), 3.21 (dt, 12H, J_{H-H} = 6.6, J_{P-H} = 3.4), 8.13 (d, J_{P-H} = 1.2). ^{13}C NMR (50 MHz, $CDCl_3$) δ 26.4 (d, J_{C-P} = 8.6), 47.0 (d, J_{C-P} = 5.0), 124.0, 128.2 (d, J_{C-P} = 9.0), 136.1, 151.4 (d, J_{C-P} = 11.7).

2,6-$(NO_2)_2$-$C_6H_3P_1$(pyrr). According to the general method, 3.6 g (yield 57%, mp 132.5–132.8 °C, recrystallized from 3:1 mixture MeOH–$CHCl_3$) of orange crystals was obtained. Anal. Calcd for $C_{18}H_{27}N_6O_4P$: C, 51.17%; H, 6.44%; N, 19.89%. Found: C, 50.77%; H, 6.36%; N, 19.89%. ^{1}H NMR (200 MHz, CD_3Cl) δ 1.81 (m, 12H), 3.09 (dt, 12H, J_{H-H} = 6.6, J_{P-H} = 3.0), 6.59 (dt, 1H, J_{H-H} = 8.0, J_{P-H} = 1.3), 7.56 (dd, 2H, J_{H-H} = 8.0, J_{P-H} = 1.2). ^{13}C NMR (50 MHz, CD_3Cl) δ 26.4 (d, J_{C-P} = 8.7), 46.6 (d, J_{C-P} = 5.1), 113.4, 126.6, 137.6 (d, J_{C-P} = 12.8), 148.1 (d, J_{C-P} = 7.7).

2,5-Cl_2-$C_6H_3P_1$(pyrr).[13] ^{1}H NMR (200 MHz, THF) δ 1.8 (overlapped by solvent, 12H), 3.05 (dt, 12H, J_{H-H} = 6.7, J_{P-H} = 4.3), 6.37 (ddd, 1H, J_{H-H} = 8.4, 2.5, J_{P-H} = 0.5), 6.68 (dd, 1H, J_{H-H} = 2.5, J_{P-H} = 1.2), 7.05 (dd, 1H, J_{H-H} = 8.4, J_{P-H} = 2.5). ^{13}C NMR (50 MHz, THF) δ 27.0 (d, J_{C-P} = 7.5), 47.6 (d, J_{C-P} = 3.7), 115.9, 122.3 (d, J_{C-P} = 8.3), 127.3 (d, J_{C-P} = 26.4), 130.1 (d, J_{C-P} = 1.5), 132.2, 150.5 (d, J_{C-P} = 5.8).

Synthesis of $HBPh_4$ Salts of the P_2 Phosphazenes. 2-Cl-$C_6H_4P_2$(pyrr)·$HBPh_4$. In three-necked flask, equipped with magnetic stirrer, powder adding system, and rubber gas balloon with Ar, of 15 mL (140 mmol) of 2-chloroaniline was added 0.28 g (12 mmol) of NaH. The mixture was heated slowly to 60 °C. After the evolution of H_2 was ceased, the mixture was stirred and cooled to room temperature, and 2.6 g (4.7 mmol) Cl-P(pyrr)$_2$=N-P$^+$(pyrr)$_3$·BF_4^- was added gradually by means of powder adding system.[19] The mixture was heated again to 60 °C, held at that for 6 h, and then left for 2 days. The excess of chloroaniline was removed under reduced pressure (110 °C/6 mmHg). To the residue dissolved in CH_2Cl_2 was added some water, and the CH_2Cl_2 extract of product HBF_4 salt was washed with acidified water to remove the rest of chloroaniline. The CH_2Cl_2 was removed in vacuo, the residue, 0.8 g of dark oil, was dissolved in 10 mL MeOH, treated with charcoal, and product was precipitated as $HBPh_4$ salt by adding 0.43 g of $NaBPh_4$ solution in 1.5 mL of MeOH. It was filtered and recrystallized from a 1:1 mixture of EtOH and AN. Yield 0.2 g (18%, mp 155.0–156.1 °C) of nearly colorless crystals. Anal. Calcd for $C_{50}H_{65}BClN_7P_2$: C, 68.84%; H, 7.51%; N, 11.24%. Found: C, 68.82%; H, 7.71%; N, 11.20%. ^{1}H NMR (200 MHz, THF) δ 1.8 (overlapped by solvent, 20H), 2.95 (dt, 12H, J_{H-H} = 6.6, J_{P-H} = 3.6), 3.22 (m, 8H), 6.34 (br d, 1H, J_{P-H} = 11.4), 6.69 (t, 4H, J_{H-H} = 7.2), 6.84 (t, 8H, $J_{H-H(av)}$ = 7.4), 7.0 (m, 2H), 7.19 (d, 1H, J_{H-H} = 4.5), 7.26 (m, 8H), 7.39 (d, 1H, J_{H-H} = 4.5). ^{13}C NMR (50 MHz, THF) δ 26.9 (d, J_{C-P} = 9.0), 27.1 (d, J_{C-P} = 9.4), 47.5 (d, J_{C-P} = 5.2), 47.9 (d, J_{C-P} = 5.7), 121.7, 121.8 (d, J_{C-P} = 2.9), 124.8, 125.6 (m, J_{C-B} = 2.9), 128.5, 130.5, 130.6 (d, J_{C-P} = 5.0) 137.2 (m, J_{C-B} = 1.5), 137.5 (d, J_{C-P} = 2.7), 165.3 (m, J_{C-B} = 49.9).

PhP_2(dma)·$HBPh_4$. The same procedure was used as for 2-Cl–$C_6H_4P_2$(pyrr)·$HBPh_4$. To 15 mL of aniline (distilled from Zn-dust) was added 15.2 mmol of NaH, and the mixture was heated for a short time. At room temperature 7.6 mmol (5.2 g) of ClP(dma)$_2$=N–P$^+$(dma)$_3$·BPh_4^- (made from BF_4^- salt[19]) was added. The excess of aniline was washed out with acidified water and the residue extracted with warm 70% $EtNH_2$ aq soln. By diluting this extract with water, a brownish precipitate of desired (raw) product as a $HBPh_4$ salt was collected. The recrystallization from 70% $EtNH_2$ aq soln and finally from

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ethyl acetate gave 1.0 g of crystals with gentle violet tint (yield 18.8%, mp 135.2–136.3 °C). Anal. Calcd for $C_{40}H_{56}BN_7P_2$: C, 67.89%; H, 7.99%; N, 13.85%. Found: C, 67.30%; H, 8.07%; N, 13.87%. ^{1}H NMR (200 MHz, THF) δ 2.41 (d, 18H, J_{P-H} = 10.2), 2.63 (d, 12H, J_{P-H} = 10.8), 6.70 (t, 4H, J_{H-H} = 7.2), 6.85 (t, 8H, $J_{H-H(av)}$ = 7.4), 6.9–7.0 (m, 3H), 7.2 (m, 2H), 7.28 (m, 8H). ^{13}C NMR (50 MHz, THF) δ 36.9 (d, J_{C-P} = 5.2), 119.9 (d, J_{C-P} = 7.1), 121.7, 123.2, 125.6 (m, J_{C-B} = 2.8), 130.0, 137.2, 140.8, 165.2 (m, J_{C-B} = 49.5).

PhP_2(pyrr)·$HBPh_4$. The same procedure was used as for 2-Cl–$C_6H_4P_2$(pyrr)·$HBPh_4$, except that NaH was not used. To 5 mL of aniline 5 mmol (2.4 g) of ClP(pyrr)$_2$=N–P^+(pyrr)$_3$·BF_4^- was added,[19] and the mixture was stirred and heated at 150 °C for 20 h. Then aniline was removed under the reduced pressure. The residue was extracted with 10 mL of CH_2Cl_2, and the extract was washed twice with water and acidified water, respectively. The extract was concentrated, and a dark sticky residue (1.9 g) was dissolved in 8 mL of MeOH. The solution was filtered, and a solution of 1.22 g (3.5 mmol) of $NaBPh_4$ in 3.5 mL MeOH was added. Collected violet crystals of product were recrystallized from 70% $EtNH_2$ aq soln as described in ref 19 or from 4:1 mixture of MeOH/$CHCl_3$. A 1.8 g (yield 44%, mp 179.2–180.2 °C) amount of crystals with beige tint was obtained. Anal. Calcd for $C_{50}H_{56}BN_7P_2$: C, 71.67%; H, 7.94%; N, 11.70%. Found: C, 71.64%; H, 7.90%; N, 11.64%. ^{1}H NMR (200 MHz, THF) δ 1.8 (overlapped by solvent, 20H), 2.99 (dt, 12H, J_{H-H} = 6.6, J_{P-H} = 3.6), 6.68 (t, 4H, J_{H-H} = 7.2), 6.83 (t, 8H, $J_{H-H(av)}$ = 7.5), 6.9–7.0 (m, 3H), 7.19 (d, 2H, J_{H-H} = 7.6), 7.25 (m, 8H). ^{13}C NMR (50 MHz, THF) δ 27.0 (d, J_{C-P} = 8.6), 27.1 (d, J_{C-P} = 9.2), 47.5 (d, J_{C-P} = 5.2), 47.7 (d, J_{C-P} = 5.8) 119.8 (d, J_{C-P} = 7.2), 121.6, 123.0, 125.6 (m, J_{C-B} = 2.8), 129.9, 137.2 (br m), 141.2, 165.3 (m, J_{C-B} = 49.5).

4-MeO–$C_6H_4P_2$(pyrr)·$HBPh_4$. 5 mmol (0.61 g) of anisidine, 5 mmol (2.75 g) of Cl–P(pyrr)$_2$=N–P^+(pyrr)$_3$·BF_4^- and 5 mmol (0.5 g) of Et_3N were mixed in a 20 mL solution of THF/AN (1:2) at room temperature. The mixture was refluxed for 20 h. The solvent was removed under reduced pressure, and the residue, 2.2 g of viscous oil, was dissolved in 10 mL of AN. A 1.7 g amount of $NaBPh_4$ in 5 mL of MeOH was added. An oily substance separated. The solution was decanted from the oily substance, and the substance was twice extracted with 40 mL of hot MeOH. The crystals collected from the methanolic solutions were recrystallized from 1:4:9 mixture of water–AN–MeOH. Yield 0.7 g (16%, mp 165.0–166.4 °C) of colorless crystals. ^{1}H NMR (200 MHz, THF) δ 1.8 (overlapped by solvent, 20H), 2.98 (dt, 12H, J_{H-H} = 6.5, J_{P-H} = 3.6), 3.7 (overlapped by solvent, 3H), 6.54 (d, 4H, J_{H-H} = 12.0), 6.68 (t, 4H, J_{H-H} = 7.2), 6.8 (br m, 12H,), 7.25 (m, 8H). ^{13}C NMR (50 MHz, THF) δ 27.0 (d, J_{C-P} = 8.6), 27.1 (d, J_{C-P} = 9.2), 47.5 (d, J_{C-P} = 5.0), 47.7 (d, J_{C-P} = 5.7), 55.7, 115.2, 121.7, 122.0 (d, J_{C-P} = 6.7), 125.6 (m, J_{C-B} = 2.8), 133.6, 137.2, 156.8, 165.3 (m, J_{C-B} = 49.5).

Liberation of P_2 Phosphazene Bases from Their $HBPh_4$ Salts with KOMe. The corresponding $HBPh_4$ salt was dissolved in a possibly small amount of dried MeOH, and a calculated (with light excess) amount of 25% KOMe solution in MeOH was added. The precipitated $KBPh_4$ was filtered off in a glovebox, and MeOH was removed under reduced pressure. The residue was extracted with hexane, the extract was filtered, and hexane was removed in vacuo. The free bases were used for spectrometric measurements.

2-Cl–$C_6H_4P_2$(pyrr): colorless crystals. ^{1}H NMR (200 MHz, THF) δ 1.8 (overlapped by solvent, 20H), 3.12 (dt, 12H, J_{H-H} = 6.6, J_{P-H} = 4.2), 3.23 (m, 8H), 6.17 (ddd, 1H, J_{H-H} = 7.7, 6.8, 1.9), 6.71 (ddd, 1H, J_{H-H} = 8.1, 6.8, 1.7), 6.79 (ddd, 1H, J_{H-H} = 8.1, 1.9, J_{P-H} = 1.2), 7.01 (ddd, 1H, J_{H-H} = 7.7, 1.7, J_{P-H} = 2.6). ^{13}C NMR (50 MHz, THF) δ 27.1 (d, J_{C-P} = 8.2), 27.4 (d, J_{C-P} = 8.4), 47.4 (d, J_{C-P} = 4.6), 47.9 (d, J_{C-P} = 4.0), 113.8, 122.1 (d, J_{C-P} = 11.9), 126.5, 127.9 (d, J_{C-P} = 30.7), 129.4 (d, J_{C-P} = 2.7), 151.3 (d, J_{C-P} = 5.1).

PhP_2(dma): colorless crystals. ^{1}H NMR (200 MHz, THF) δ 2.55 (d, 18H, J_{P-H} = 10.1), 2.65 (d, 12H, J_{P-H} = 10.0), 6.27 (ddt, 1H, J_{H-H} = 7.1, 1.5, J_{P-H} = 1.2), 6.60 (br d, 2H,, J_{H-H} = 7.9), 6.82 (br t, 2H,, $J_{H-H(av)}$ = 7.5). ^{13}C NMR (50 MHz, THF) δ 37.3 (d, J_{C-P} = 4.4), 38.1 (d, J_{C-P} = 2.6), 114.3, 123.2 (d, J_{C-P} = 20.1), 128.4 (d, J_{C-P} = 1.7), 154.9.

PhP_2(pyrr): colorless needles. ^{1}H NMR (200 MHz, THF) δ 1.8 (overlapped by solvent, 20H), 3.12 (dt, 20H, J_{H-H} = 6.7, J_{P-H} = 4.3), 6.58 (m, 2H, J_{H-H} = 8.3), 6.78 (m, 2H). ^{13}C NMR (50 MHz, THF) δ 27.1 (d, J_{C-P} = 8.1), 27.3 (d, J_{C-P} = 8.4), 47.4 (d, J_{C-P} = 4.7), 47.9 (d, J_{C-P} = 3.8), 113.9, 122.9 (d, J_{C-P} = 20.8), 128.3 (d, J_{C-P} = 1.9), 155.3.

Synthesis of $HBPh_4$ Salts of P_1 Phosphazenes and Amines. A methanolic solution of bases (10 mmol in 10 mL MeOH) was acidified with 15% HCl aq soln, and a slight excess of $NaBPh_4$ solution in small quantity of MeOH was added. Precipitate of the salt was filtered, washed several times with MeOH, recrystallized (except TBD) from a 4:1 mixture of MeOH and $CHCl_3$, and dried in vacuo.

Et_3N·$HBPh_4$: mp 184–187 °C (dec); ^{1}H NMR (200 MHz, THF) δ 0.88 (t, 3H, J_{H-H} = 7.4) 2.53 (q, 2H, J_{H-H} = 7.4), 6.75 (t, 4H, J_{H-H} = 7.2), 6.89 (t, 8H, $J_{H-H(av)}$ = 7.4), 7.31 (m, 8H). ^{13}C NMR (50 MHz, THF) δ 9.3, 47.8, 122.0, 125.9 (m, J_{C-P} = 2.7), 137.1, 165.1 (m, J_{C-P} = 49.5).

2-Cl–$C_6H_4P_1$(dma)·$HBPh_4$: mp 162.0–163.3 °C. ^{1}H NMR (200 MHz, THF) δ 2.47 (d, 18H, J_{P-H} = 10.2), 6.72 (t, 4H, J_{H-H} = 7.2), 6.86 (t, 8H, $J_{H-H(av)}$ = 7.6), 7.10 (m, 1H), 7.2 (m, 2H) 7.28 (m, 8H), 7.47 (m, 1H). ^{13}C NMR (50 MHz, THF) δ 37.3 (d, J_{C-P} = 4.4), 121.9, 125.7 (m, J_{C-B} = 2.8), 128.7 (d, J_{C-P} = 2.0), 129.2 (d, J_{C-P} = 5.7), 131.3, 132.1 (d, J_{C-P} = 6.6), 134.3, 137.2, 165.1 (m, J_{C-B} = 49.5).

2,5-Cl_2-$C_6H_3P_1$(pyrr)·$HBPh_4$: mp 153.4–153.6 °C. ^{1}H NMR (200 MHz, THF) δ 1.8 (overlapped by solvent, 12H), 3.06 (dt, 12H, J_{H-H} = 6.7, J_{P-H} = 3.7), 6.71 (t, 4H, J_{H-H} = 7.2), 6.86 (t, 8H, $J_{H-H(av)}$ = 7.4), 7.13 (dd, 1H, J_{H-H} = 2.4, J_{P-H} = 1.1), 7.22 (ddd, 1H, J_{H-H} = 8.6, 2.4, J_{P-H} = 0.9), 7.26 (m, 8H), 7.44 (dd, 1H, J_{H-H} = 8.6, J_{P-H} = 1.2). ^{13}C NMR (50 MHz, THF) δ 26.8 (d, J_{C-P} = 8.5), 48.4 (d, J_{C-P} = 4.7), 121.8, 125.6 (m, J_{C-B} = 2.9), 125.7 (overlapped by anions peak), 128.0, 128.7 (d, J_{C-P} = 8.0), 132.4, 134.1 (d, J_{C-P} = 1.7), 136.2, 137.1 (m, J_{C-B} = 1.4), 165.2 (m, J_{C-B} = 49.5).

***t*-BuP_1(dma)·$HBPh_4$**: mp 266–267 °C (dec). ^{1}H NMR (200 MHz, THF) δ 1.22 (d, 9H, J_{P-H} = 1.0), 2.50 (d, 18H, J_{P-H} = 10.0), 4.5 (br d, 1H, J_{P-H} = 9.7), 6.71 (t, 4H, J_{H-H} = 7.2), 6.86 (t, 8H, $J_{H-H(av)}$ = 7.4), 7.28 (m, 8H). ^{13}C NMR (50 MHz, THF) δ 31.5 (d, J_{C-P} = 4.6), 37.8 (d, J_{C-P} = 4.7), 53.2, 121.8, 125.7 (m, J_{C-B} = 2.8), 137.2, 165.2 (m, J_{C-B} = 49.5).

TBD·$HBPh_4$ (1,5,7-triazabicyclo[4.4.0]dec-5-enyltetraphenylborate): mp 240–241 °C (dec); ^{1}H NMR (200 MHz, DMSO-d_6) δ 1.81 (q, 4H, $J_{H-H(av)}$ = 5.9), 3.13 (t, 4H, J_{H-H} = 5.8), 3.18 (t, 4H, J_{H-H} = 6.0). ^{13}C NMR (50 MHz, DMSO-d_6) δ 20.2, 37.5, 46.2, 121.5, 125.3 (m, J_{C-B} = 2.7), 135.5 (m, J_{C-B} = 1.2), 150.6, 163.3 (m, J_{C-B} = 49.5).

***t*-BuP_1(pyrr)·$HBPh_4$**: mp 233–235 °C. ^{1}H NMR (200 MHz, THF) δ 1.24 (d, 9H, J_{P-H} = 0.8), δ 1.8 (overlapped by solvent, 12H), 3.05 (dt, 12H, J_{H-H} = 6.7, J_{P-H} = 3.7), 6.71 (t, 4H, J_{H-H} = 7.2), 6.85 (t, 8H, $J_{H-H(av)}$ = 7.4), 7.26 (m, 8H).^{13}C NMR (50 MHz, THF) δ 26.8 (d, J_{C-P} = 8.2), 31.7 (d, J_{C-P} = 4.5), 48.4 (d, J_{C-P} = 4.7), 53.2 (d, J_{C-P} = 0.8) 121.8, 125.7 (m, J_{C-B} = 2.8), 137.2, 165.2 (m, J_{C-B} = 49.5).

H_2NP_1(pyrr)·$HBPh_4$: mp 183.0–184.3 °C. ^{1}H NMR (200 MHz, THF) δ 1.8 (overlapped by solvent, 12H), 3.05 (dt, 12H, J_{H-H} = 6.6, J_{P-H} = 3.8), 3.35 (d, 2H, J_{P-H} = 10.4), 5.50 (d, 1H, J_{P-H} = 35.6), 6.73 (t, 4H, J_{H-H} = 7.2), 6.87 (t, 8H, J_{H-H} = 7.2), 7.28 (m, 8H). ^{13}C NMR (50 MHz, THF) δ 26.9 (d, J_{C-P} = 8.1), 47.9 (d, J_{C-P} = 4.0), 121.9, 125.8 (m, J_{C-B} = 2.8), 137.1, 165.2 (m, J_{C-B} = 49.7).

EtP$_1$(pyrr)·$HBPh_4$: ^{1}H NMR (200 MHz, THF) δ 1.08 (dt, 3H, J_{H-H} = 7.4, J_{P-H} = 1.2), δ 1.8 (overlapped by solvent, 12H), 2.82 (dq, 2H, J_{H-H} = 7.4, J_{P-H} = 9.0), 3.03 (dt, 12H, J_{H-H} = 6.7, J_{P-H} = 3.8), 6.71 (t, 4H, J_{H-H} = 7.2), 6.86 (t, 8H, $J_{H-H(av)}$ = 7.4), 7.26 (m, 8H). ^{13}C NMR (50 MHz, THF) δ 17.3 (d, J_{C-P} = 7.1), 26.9 (d, J_{C-P} = 8.0), 36.8, 48.0 (d, J_{C-P} = 4.6), 121.8, 125.7 (m, J_{C-B} = 2.8), 137.2, 165.2 (m, J_{C-B} = 49.5).

UV–Vis Spectrophotometric Determination of pK_a in AN and pK_{ip} in THF. The spectrophotometric titration method in THF and in AN media in the glovebox was similar to that used in the previous work.[14] Perkin-Elmer Lambda 2S or Lambda 40 spectrophotometer equipped with quartz fiber

Table 1. Ionic Radii Used for Correction for Ion Pairing

ion	ion-pair radius, Å[a]	ion	ion-pair radius, Å[a]
$RNH^+{=}PR'_3$; $RNH^+{=}PR''_3$	4	pyrrolidineH$^+$	2
$RNH^+{=}PR'_2N{=}PR'_3$	4.8	DMANH$^+$	3.2
$RNH^+{=}PR''_2N{=}PR''_3$	5.4	X-anilineH$^+$; X-pyridineH$^+$	2
TBDH$^+$	3	N,N-Me$_2$-anilineH$^+$	2.2
DBUH$^+$; DABH^{+b}	2.5	2,6-X$_2$-pyridineH$^+$	2.2
TMGH$^+$; TEAH$^+$	2.2	$CF_3SO_3^-$; $CH_3SO_3^-$	2.5
PhTMGH$^+$	2.7	BPh_4^-	4.4

[a] Ionic radii from literature (ref 10) were used when available. In cases when no literature data were available, the radii were estimated by PM3 calculations; R = alkyl or aryl; R′ = dma; R″ = pyrr; X = H or substituent. [b] DAB = 1,4-diaminobutane.

optic system, and an external sample compartment positioned into the glovebox was used. Two different gloveboxes were used. Earlier measurements were carried out in a Mecaplex glovebox in the atmosphere of nitrogen, continuously purified from water vapor, volatile acidic and basic impurities with molecular sieves, powdered P_2O_5, and KOH pellets. Later measurements were carried out in MBraun glovebox in the atmosphere of argon that was constantly circulated through a purification system containing molecular sieves and activated copper for removal of water vapor and oxygen, respectively. The residual concentrations of water and oxygen in the atmosphere of the glovebox during the measurements were constantly monitored and were generally below 1 ppm.

All solutions were prepared in glovebox and were made fresh daily. The concentrations of acidic titrants were from 1 to 3 mM, and basic titrants were from 0.5 to 5.5 mM. Higher concentrations (up to 0.36 and 0.025 M for methanesulfonic acid and EtP_1(pyrr), respectively) were made for studying relative ion-pair basicity dependence from concentrations in THF. The concentrations of studied bases were generally in mM range; higher concentrations (up to 0.23 M) were used for studying relative ion-pair basicity dependence from concentrations.

The phosphazene bases are more stable and convenient to handle if they are used as salts. Several counterions (BPh_4^-, PF_6^-, ClO_4^-) were used in the salts. From these, only salts with tetraphenylborate anion were soluble enough in THF.

Various disturbances were frequently observed with the tetraphenylborate anion both in UV–Vis (strange behavior of the spectra upon titration and the nonreversibility of the spectra upon back-titration, change of spectrum in time) and NMR measurements (change of spectrum in time, appearance of alien peaks). This could be due to the somewhat unstable nature of this anion. Therefore, we avoided this anion as counterion in UV–Vis spectrophotometric measurements and used free bases instead.

Calculation Methods for UV–Vis Spectrophotometric Measurements. The ΔpK_a calculation methods in AN are similar to those of the previous works.[1,20] The essence of the general calculation method is the following. When two partially protonated bases B_1 and B_2 are in the same solution, then the following equation holds for absorbance A at wavelength λ (1 cm path length):

$$A^\lambda = [HB_1^+]\epsilon^\lambda_{HB_1^+} + [B_1]\epsilon^\lambda_{B_1} + [HB_2^+]\epsilon^\lambda_{HB_2^+} + [B_2]\epsilon^\lambda_{B_2} \quad (11)$$

The molar absorptivities ϵ can be found separately from the spectra of the free bases and fully protonated bases. If we use concentrations, that are normalized to 1 then we may write: $[HB_1^+] = 1 - [B_1]$ and $[HB_2^+] = 1 - [B_2]$. After a mathematical transformation of eq 11 we get

$$\frac{A^\lambda - \epsilon^\lambda_{HB_1^+} - \epsilon^\lambda_{HB_2^+}}{(\epsilon^\lambda_{B_2} - \epsilon^\lambda_{HB_2^+})} = [B_1]\frac{(\epsilon^\lambda_{B_1} - \epsilon^\lambda_{HB_1^+})}{(\epsilon^\lambda_{B_2} - \epsilon^\lambda_{HB_2^+})} + [B_2] \quad (12)$$

If the spectra are recorded over a range of wavelengths then $[B_1]$ and $[B_2]$ can be found from eq 12 as the slope and intercept of a regression line. From the values of $[B_1]$ and $[B_2]$, the calculation of ΔpK_a of the bases is straightforward. In many cases (for example, when the bases have absorption maxima in different wavelength ranges) it was possible to use various simpler calculation procedures (see refs 1 and 20). The mixture of bases as well as both bases separately was titrated with an optically transparent acid and/or base, and the data for ΔpK_a calculations was obtained from UV–Vis spectra. From each titration experiment, the ΔpK_a was determined as the mean of 5–20 values.

In THF the general principle of the ΔpK calculations is the same. As UV–Vis spectrophotometry does not make any difference between free ions, solvent separated ion-pairs, and loosely bonded contact ion-pairs (these last two are the main forms of monocharged ions at concentration below 0.01 M in THF),[10] we get ΔpK_{ip}s instead of ΔpK_as. The ΔpK_{ip} values given in Table 2 are a mean of 5–23 measurements. The correction for ion pairing was calculated using the Fuoss equation as described in ref 10, and a ΔpK_a is then obtained. The ionic radii that were used are given in Table 1.

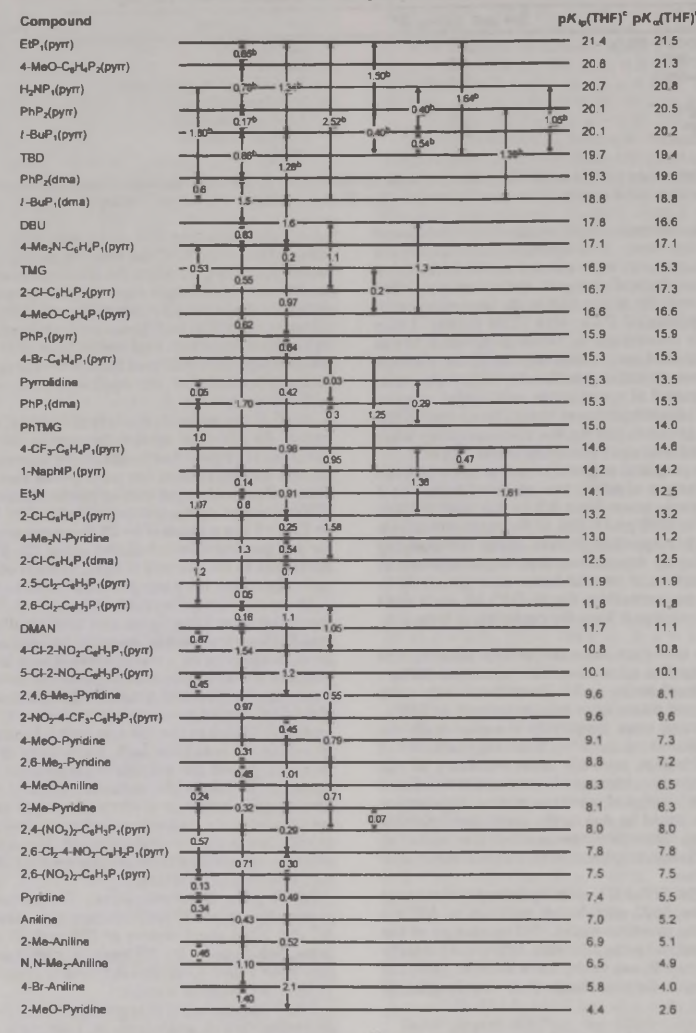
In some cases ("invisible bases", e.g., aliphatic amines (pyrrolidine and triethylamine) and t-BuP_1(dma), also DBU and TMG versus "visible" aromatic bases) the calculations have been carried out on a molar basis. The solution containing a mixture of known amounts (in moles) of "invisible" and visible base was titrated with titrant of known concentration. From the added titrant mass and its concentration, the amount (in moles) of titrant in the cell was found. Combining the spectra of solutions containing both bases fully deprotonated, fully protonated, and the mixture of protonated and deprotonated forms, we calculated the indicator ratio of the visible base, and knowing the amounts of the visible base and titrant added, we calculated the indicator ratio for "invisible" base. The ΔpK_{ip} calculation is then straightforward. The agreement between relative ion-pair basicities obtained with different calculation methods was satisfactory.

NMR pK_{ip} Determination. The standard 1D ^{1}H and proton-decoupled ^{13}C NMR spectra were recorded on a Bruker AC-200 NMR spectrometer at 200 and 50 MHz, correspondingly. Solutions (~ 0.1 M) were prepared and sealed off in 5 mm NMR tubes. Chemical shifts were determined relative to TMS as an internal standard.

The NMR spectra of phosphazenes and the method used for determination of their ΔpK_{ip} in THF were analogous to the corresponding NMR spectra and ΔpK_a calculations applied for phosphazenes in the AN.[13] There is no large difference of the ^{13}C and ^{1}H chemical shifts for phosphazenes in THF as compared to the corresponding shifts in AN. Usually, the differences were below 1 ppm for ^{13}C and 0.1 ppm for ^{1}H. The ΔpK_{ip}s of phosphazenes were determined using approximately equimolar mixture of phosphazene and indicator in THF. As it was in the case of AN solutions, there is the fast (in NMR time scale) exchange between phosphazene and indicator base and acid forms leading to the coalescence of NMR lines in the ^{13}C and ^{1}H spectra. Correspondingly, the chemical shifts of these forms were determined separately from the single component THF solutions of these species. The ΔpK_{ip} values were calculated as it was done previously with ΔpK_a values.[13] The indicator ratios for eq 8 were calculated (eqs 13 and 14) from chemical shifts of the individual species (both neutral and protonated forms of the bases) and their averaged values in the mixtures containing both forms.

(20) Leito, I.; Kaljurand, I.; Koppel, I. A.; Yagupolskii, L. M.; Vlasov, V. M. *J. Org. Chem.* **1998**, *63*, 7868–7874.

Table 2. Continuous Self-consistent Basicity Scale of Neutral Bases in THF Solution[a]



[a] The numbers on the arrows are the direct experimental ΔpK_{ip} values (uncorrected for ion pairing) obtained from UV–Vis spectrophotometric measurements if not indicated otherwise. [b] NMR measurements. [c] Absolute pK_{ip}(THF) and pK_{α}(THF) estimated values for conjugate acids of the respective bases.

$$\frac{a(HB_2^+A^-)}{a(B_2)} = \frac{\delta_{B_2} - \delta}{\delta - \delta_{HB_2^+A^-}} \quad (13)$$

$$\frac{a(B_1)}{a(HB_1^+A^-)} = \frac{\delta - \delta_{B_1}}{\delta_{HB_1^+A^-} - \delta} \quad (14)$$

In the case of alkylphosphazenes, the differences of the ^{13}C chemical shifts for alkyl substituent between phosphazene base and acid forms are markedly lower as compared to the corresponding differences of the arylic carbons of arylphosphazenes[13] and were in the range of 3–6 ppm for the ones measured.

Results

Altogether 96 individual acid–base equilibrium measurements in THF involving 45 bases were carried out using the UV–Vis spectrophotometric or ^{13}C NMR method. These measurements give a continuous basicity scale in THF as presented in Table 2, and some ^{13}C NMR results are given in Table 3.

Multiple overlapping measurements make the results more reliable and help to estimate their self-consistency. The entire basicity range covered involves at least two independent pathways of measurements and the relative basicity of any two bases can be obtained by combining

Table 3. ^{13}C NMR ΔpK_{ip}[a] Results Not Included in Table 2

base	reference base	ΔpK_{ip}
t-BuP$_1$(dma)	PhP$_2$(dma)	0.39
DBU	*t*-BuP$_1$(dma)	0.77
	2-Cl–C$_6$H$_4$P$_2$(pyrr)	0.00
	4-Me$_2$N-C$_6$H$_4$P$_1$(pyrr)	0.48
TMG	PhP$_1$(pyrr)	−1.11
DAB	2-Cl-C$_6$H$_4$P$_1$(pyrr)	−0.53
DMAP	2-Cl-C$_6$H$_4$P$_1$(dma)	0.38
	2,5-Cl$_2$-C$_6$H$_3$P$_1$(pyrr)	−0.45
Et$_3$N	2-Cl-C$_6$H$_4$P$_1$(dma)	0.87
	2,5-Cl$_2$-C$_6$H$_3$P$_1$(pyrr)	0.08
	4-Cl-2-NO$_2$-C$_6$H$_3$P$_1$(pyrr)	−0.94
DMAN	2,5-Cl$_2$-C$_6$H$_3$P$_1$(pyrr)	0.31
	Et$_3$N	0.25

[a] $\Delta pK_{ip} = pK_{ip(\text{conjugate acid of the reference base})} - pK_{ip(\text{conjugate acid of the base})}$.

at least two independent sets of measurements. Reversibility of protonation/deprotonation process of all bases was checked. All equilibria presented in Table 2 were reached in minutes and were stable. Both ion-pair (pK_{ip}) values and values corrected for ion pairing (pK_a) are given in Table 2. Although somewhat arbitrary, the correction for ion-pairing is useful because it makes our data comparable to the data by the Morris group.[10]

The absolute pK_a values have been obtained by anchoring the scale to the pK_a value of triethylamine in THF ($pK_a = 12.5$),[10] a secondary standard proposed by the Morris' group. This is not a perfect choice, but is the most suitable anchoring point available for our data. See Discussion for additional comments.

It is not easy to find a suitable anchoring point for the ion-pair pK_{ip} values. The amount of available absolute pK_{ip} values of bases in THF is scarce. The data on acids is abundant but not directly comparable to pK_{ip} data of bases (see ref 21 for further discussion). Therefore the pK_{ip} values have been anchored to the pK_a value of PhP$_1$-(pyrr). This anchorage is arbitrary, but this way the core of the scale—the P$_1$ phosphazenes—have all practically the same values in both scales, which facilitates the discussion.

The absolute pK_a values were calculated as in the previous papers[14,20] by minimizing the sum of squares u of differences between directly measured ΔpK_a^I values and the assigned pK_a values while keeping the pK_a value of triethylamine constant and equal to 12.5.

$$u = \sum_{i=1}^{n_m} (\Delta pK_a^I - (pK_a(HB_2^+A^-) - pK_a(HB_1^+A^-))) \quad (15)$$

It should be stressed that the absolute pK_a values of bases given in Table 2 are not as accurate as the relative pK_as. One could anchor the scale to any other absolute pK_a value, and the relative basicities will remain the same. Precision s of the measurements was calculated as in the previous papers:[14,20]

$$s = \sqrt{\frac{u}{n_m - n_c}} \quad (16)$$

The number of measurements is $n_m = 83$; the number of pK_as determined $n_c = 43$. For our results, $s = 0.10$ (for pK_{ip}, $s = 0.09$). Some previously published pK_a values in AN[14] (Table 4) have minor corrections (up to 0.03 pK_a units) due to the new measurements.

Table 4. Basicity Data of the Bases in AN, H_2O, and the Gas Phase

base	pK_a values in AN[a]	GB (kcal/mol)[b]	pK_a values in H_2O[c]
EtP$_1$(pyrr)	28.89[d]		
t-BuP$_1$(pyrr)	28.35[d]		
PhP$_2$(dma)	26.28		
TBD	25.96[e]	244.3	
t-BuP$_1$(dma)	26.88[d]		
2-Cl-C$_6$H$_4$P$_2$(pyrr)	25.24		
4-Me$_2$N-C$_6$H$_4$P$_1$(pyrr)	23.71		
DBU	24.16	239.6	
4-MeO-C$_6$H$_4$P$_1$(pyrr)	22.95		
PhP$_1$(pyrr)	22.17		
4-Br-C$_6$H$_4$P$_1$(pyrr)	21.05		
TMG	23.3[f]	238.4	13.6
PhP$_1$(dma)	21.07		
4-CF$_3$-C$_6$H$_4$P$_1$(pyrr)	19.93		
1-NaphtP$_1$(pyrr)	20.42		
PhTMG	20.63	236.9	12.18[g]
pyrrolidine	19.34	218.8	11.1
2-Cl-C$_6$H$_4$P$_1$(pyrr)	19.97		
2-Cl-C$_6$H$_4$P$_1$(dma)	18.87		
Et$_3$N	18.63	227	10.7
2,5-Cl$_2$-C$_6$H$_3$P$_1$(pyrr)	18.32		
2,6-Cl$_2$-C$_6$H$_3$P$_1$(pyrr)	18.36		
DMAP	17.75	232.1	9.53
DMAN	18.42	238.0	12.1[h]
4-Cl-2-NO$_2$-C$_6$H$_3$P$_1$(pyrr)	17.48		
5-Cl-2-NO$_2$-C$_6$H$_3$P$_1$(pyrr)	17.07		
2-NO$_2$-4-CF$_3$-C$_6$H$_3$P$_1$(pyrr)	16.33		
2,4,6-Me$_3$-pyridine	14.78		
2,4-(NO$_2$)$_2$-C$_6$H$_3$P$_1$(pyrr)	14.68		
2,6-Cl$_2$-4-NO$_2$-C$_6$H$_2$P$_1$(pyrr)	14.25		
2,6-(NO$_2$)$_2$-C$_6$H$_3$P$_1$(pyrr)	13.91		
4-MeO-pyridine	14.04	222.2	6.5
2,6-Me$_2$-pyridine	13.92	222.5	6.70
4-MeO-aniline	11.66	207.6	5.3
2-Me-pyridine	13.11	219.2	5.94
pyridine	12.33[f]	214.7	5.25
aniline	10.42	203.3	4.6
2-Me-aniline			4.4
N,*N*-Me$_2$-aniline	11.23	217.3	5.1
4-Br-aniline	9.25		3.9
2-MeO-pyridine			3.1

[a] Slightly revised pK_a values of conjugate acids of corresponding bases obtained in previous work (ref 14) or in this work if not noted otherwise. [b] Gas-phase basicities from ref 25. [c] pK_a values of conjugate acids of corresponding bases from refs 24 or 23 if not noted otherwise. [d] Reference 17. [e] Reference 26. [f] Reference 27. [g] Reference 28. [h] Reference 29. [i] Anchor of AN scale; value taken from ref 22.

Discussion

pK_a Values of Iminophosphoranes. Unsubstituted PhP$_1$(pyrr) is a strong base with basicity ($pK_a = 15.9$) between those of TMG and DBU. By substitution of the phenyl ring, the basicity can be varied over a wide pK_a range. In this work the pK_a values of substituted PhP$_1$-(pyrr) range from 7.5 (2,6-dinitro-) to 17.1 (4-(dimethylamino)-), that is — by almost 10 orders of magnitude. The influence of substituents in the phenyl ring on the basicity of the phosphorane is easily predictable giving the possibility to conveniently "tune" the basicity of the phosphorane.

Alkyliminophosphoranes are significantly stronger bases than aryliminophosphoranes. Thus, EtP$_1$(pyrr) is by ca. 5.5 and *t*-BuP$_1$(pyrr) by ca. 4 orders of magnitude stronger than PhP$_1$(pyrr). The inductive effect and some delocalization of the lone electron pair of the imino

(21) Streitwieser, A.; Kim, Y.-J. *J. Am. Chem. Soc.* **2000**, *122*, 11783–11786.

nitrogen into the aromatic ring (see ref 14 for discussion on this topic) are most probably the reasons.

As can be expected, the P_2 phenyliminophosphoranes are stronger bases than the corresponding P_1 phenyliminophosphoranes. The difference is ca. 4−5 pK units. For alkyliminophosphoranes the same difference is ca. 6 pK units in acetonitrile.[19] We do not have data on P_2 alkyliminophosphoranes in THF, so direct comparison is not possible.

The relatively good predictability of the basicity together with suitable spectral properties in the UV range make the phenyliminophosphoranes convenient neutral indicators in the medium to high basicity range. The choice of neutral indicators in the high basicity range is currently very limited.

Comparison of Basicities in THF with Those in Other Media. Correlation of pK_{ip} and pK_a values in THF with pK_a values in acetonitrile yields the following equations: pK_{ip}(THF) = (−2.68 ± 0.55) + (0.83 ± 0.03)·pK_a(AN); n = 39; r^2 = 0.959; s = 0.89. pK_α(THF) = (−5.08 ± 0.39) + (0.92 ± 0.02)·pK_a(AN); n = 39; r^2 = 0.983; s = 0.63. Correlation of pK_{ip} and pK_a values in THF with pK_a values in water yields the following equations: pK_{ip}(THF) = (1.78 ± 0.64) + (1.08 ± 0.08)·pK_a(H_2O); n = 17; r^2 = 0.926; s = 1.05. pK_α(THF) = (−0.31 ± 0.49) + (1.14 ± 0.06)·pK_a(H_2O); n = 17; r^2 = 0.960; s = 0.80. Correlation of pK_{ip} and pK_α values in THF with gas-phase basicity pK_a values (pK_a(GB) = GB (kcal·mol^{-1})/1.364 (kcal·mol^{-1})) is poor and yields the following equations: pK_{ip}(THF) = (−53.29 ± 12.52) + (0.39 ± 0.08)·pK_a(GB); n = 15; r^2 = 0.676; s = 2.57. pK_α(THF) = (−61.34 ± 12.48) + (0.43 ± 0.08)·pK_a(GB); n = 15; r^2 = 0.718; s = 2.56.

From these correlations it appears that the differentiating ability of THF for basicities is between water and AN. One can see that the transfer of this reaction series from THF to AN increases slightly its sensitivity toward substituent effects both in case of pK_{ip} and pK_α whereas the transfer from THF into water leads to the opposite result.

It is interesting to note that in all cases the correlation is better with the pK_α than with the pK_{ip} values. This result indirectly validates the method of correction for ion-pairing.

Concentration Dependence of pK_{ip} Values. If we assume that no larger associates than 1:1 ion pairs exist in the solution then the pK_{ip} values should not show any concentration dependence. The concentrations in the NMR measurements are intrinsically higher than in UV−Vis measurements, and the agreement between these two methods serves as a good indicator. According to the data in Tables 2 and 3, the results of the two methods agree well for the ΔpK_{ip} values of arylphosphazenes. With bases of smaller size, however, disagreements are observed. When comparing the results of the measurement of the same equilibrium carried out by different methods, the following is observed: the ΔpK_{ip} between DBU and 4-Me_2N−$C_6H_4P_1$(pyrr) according to the UV−Vis measurements is 0.83, according to NMR it is 0.48; according to the UV−Vis measurements DMAP is by 0.54 pK_{ip} units stronger base than 2-Cl−$C_6H_4P_1$(dma), whereas according to the NMR measurements DMAP is by 0.38 units weaker. The situation is even more serious with triethylamine: according to the UV−Vis measurements it is by 1.58 pK_{ip} units stronger base than 2-Cl−$C_6H_4P_1$(dma), according to the NMR it is by 0.87 pK_{ip} units weaker. These discrepancies are larger than the possible uncertainties of these measurements. On the UV−Vis measurements we did not observe noticeable ΔpK_{ip} dependence (see Table S1 in Supporting Information) on concentrations while changing the concentration of Et_3N over the wide range (from 4.5 × 10^{-5} M to 2.3 × 10^{-2} M) and keeping the 2-Cl−$C_6H_4P_1$(pyrr) concentration 10^{-4} to 10^{-5} M.

The disagreements can be due to the formation of aggregates of 1:1 ion pairs at higher concentrations, especially at those used in the NMR method. Due to this concentration dependence, the NMR measurements involving bases of small size were not included in the scale (Table 2).

We used the Streitwieser method[30] to estimate the mean aggregation numbers of the ion pairs. For the UV−Vis data these were around 1, indicating that no significant aggregation of the ion pairs was taking place during the measurements. In principle the method also permits to find the mean aggregation constants; however, due to the very low extent of aggregation we could not get reliable estimates for the aggregation constants. Also, it was not possible to apply that method to the NMR data, because it is necessary that one of the protonated bases would be in the solution only in the form of a monomeric ion pair. This condition is not met at the concentrations used for the NMR measurements.

Anchoring of the Scale. These disagreements cast some doubt on the suitability of triethylamine as anchoring point for our data: the pK_α value of triethylamine was determined by the Morris group using NMR measurements (concentrations were in the range of 0.02−0.07 M).[10] The other two compounds common in this work and ref 10, *N,N*-dimethylaniline (pK_α = 6.0)[10] and TMG (estimated pK_α = 15),[10] are not as suitable because they are, respectively, either at the very bottom of the scale or have a pK_α value that has only been estimated, not measured. They are of similar size to triethylamine, so that similar concentration dependence problems can be anticipated. The pK_α values of *N,N*-dimethylaniline and TMG from this work are 4.9 and 15.3, respectively. Thus, in the case of *N,N*-dimethylaniline there is a disagreement, but the general picture would remain the same if one of these two compounds would be used as the anchoring point. To the best of our knowledge, besides the work of the Morris group,[10] there is no other absolute basicity data in THF available in the literature that could be used to anchor our scale.

Conclusions

Relative ion-pair basicities ΔpK_{ip} of 25 substituted aryl and alkyl iminophosphoranes and 20 other N-bases

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(various pyridines, amines, amidines) have been measured in THF medium using the UV–Vis and/or ^{13}C NMR methods. The ΔpK_{ip} values were corrected for ion pairing using the Fuoss equation to obtain relative ionic basicities ΔpK_a. Based on these measurements, a basicity scale ranging from 2-methoxypyridine to EtP_1(pyrr) and having a total span over 18 pK units has been created. The scale has been anchored to the pK_a value of triethylamine ($pK_a = 12.5$). It is practically impossible to give any truly absolute pK_{ip} values at this time, because absolute pK_{ip} data of neutral bases in THF is almost nonexistent in the literature.

The results are compared to pK_a values in various other solvents and in the gas phase. The pK_a values give better correlations than the pK_{ip} values, thus indirectly validating the procedure of correction for ion pairing.

The predictability of the basicity together with suitable spectral properties in the UV range make the phenyliminophosphoranes convenient neutral indicators in the high basicity range where the choice of neutral indicators is very limited.

Acknowledgment. This work was supported by grants 4376, 3992, 3366, and 3370 from the Estonian Science Foundation.

Supporting Information Available: Table (Table S1) of detailed experimental data (includes data for all the UV–vis equilibrium measurements: concentrations, acid used, calculation method) and ^{13}C NMR spectra of selected compounds (Figures S2). This information is available free of charge via the Internet at http://pubs.acs.org.

JO016185P

Acid-Base Equilibria in Nonpolar Media 2. Self-Consistent Basicity Scale in THF Solution Ranging from 2-methoxypyridine to EtP_1(pyrr) Phosphazene

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Supporting Information

Table S1. Experimental results for UV-vis spectrophotometric measurements in THF.

A(stronger base)	B	C(A)·10^5 M	C(B)·10^5 M	ΔpK_{ip}[a]	Acid[b]	Calc[c]
PhP_2(dma)	*t*-BuP_1(dma)	3.4	3.5	0.6	T	NV
	DBU	3.9	4.47	1.5	M	NV
t-BuP_1(dma)	4-Me_2N-C_6H_4P1(pyrr)	3.28	3.5	1.6	M	NV
DBU	4-Me_2N-$C_6H_4P_1$(pyrr)	4.08	4.0	0.83	M	NV
	2-Cl-$C_6H_4P_2$(pyrr)	3.43	2.8	1.1	M	NV
	4-MeO-$C_6H_4P_1$(pyrr)	5.3	5.8	1.3	M	NV
4-Me_2N-$C_6H_4P_1$(pyrr)	TMG	4.2	3.7	0.2	M	NV
	2-Cl-$C_6H_4P_2$(pyrr)	2.9	1.5	0.53	M	S
	4-MeO-$C_6H_4P_1$(pyrr)	1.9	3.2	0.55	M	S
TMG	4-MeO-$C_6H_4P_1$(pyrr)	7.7	5.0	0.2	M	NV

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	PhP_1(pyrr)	5	4.9	0.97	M	NV
4-MeO-$C_6H_4P_1$(pyrr)	PhP_1(pyrr)	1.9	4.3	0.62	M	S
		3.2	2.3	0.62	M	S
PhP_1(pyrr)	4-Br-$C_6H_4P_1$(pyrr)	3.1	1.9	0.64	T	S
	1-NaphtP_1(pyrr)	2.2	2.1	1.70	T	S
4-Br-$C_6H_4P_1$(pyrr)	PhP_1(dma)	2.0	3.6	0.20	T	S
	1-NaphtP_1(pyrr)	2.8	2.2	1.25	T	S
	PhTMG	1.9	3.5	0.42	M	S
Pyrrolidine	PhTMG	8.2	5.0	0.29	M	NV
	PhP_1(dma)	6.0	7.4	0.05	M	NV
PhP_1(dma)	PhTMG	3.2	4.0	0.3	T	S
	1-NaphtP_1(pyrr)	2.4	1.8	1.0	T	S
PhTMG	Et_3N	5.5	4.9	0.95	M	NV
	1-NaphtP_1(pyrr)	4.3	1.4	0.98	T	S
4-CF_3-$C_6H_4P_1$(pyrr)	2-Cl-$C_6H_4P_1$(pyrr)	1.46	3.49	1.36	M	S
	1-NaphtP_1(pyrr)	1.64	2.2	0.47	M	S
	4-Me_2N-Pyridine	1.86	4.2	1.61	M	S
1-NaphtP_1(pyrr)	Et_3N	3.9	5.8	0.14	M	NV
	4-Me_2N-Pyridine	9.7	3.2	1.07	M	S
	2-Cl-$C_6H_4P_1$(pyrr)	2.0	3.2	0.91	T	S
Et_3N	2-Cl-$C_6H_4P_1$(pyrr)	3.68	4.93	0.63	M	NV
		20.8	20.0	0.73	M	NV
		198	19.8	0.69	M	NV
		4.47	4.13	0.78	M	NV

		23.2	19.3	0.86	M	NV
		230	21.8	0.73	M	NV
		2310	20.6	0.81	M	NV
	2-Cl-$C_6H_4P_1$(dma)	6.4	3.8	1.58	M	NV
2-Cl-$C_6H_4P_1$(pyrr)	2,5-Cl_2-$C_6H_3P_1$(pyrr)	1.7	1.9	1.3	M	S
	4-Me_2N-Pyridine	1.3	2.4	0.25	M	S
4-Me_2N-Pyridine	2-Cl-$C_6H_4P_1$(dma)	3.6	3.8	0.54	M	S
	2,6-Cl_2-$C_6H_3P_1$(pyrr)	3.7	3.9	1.2	M	S
2-Cl-$C_6H_4P_1$(dma)	2,5-Cl_2-$C_6H_3P_1$(pyrr)	3.4	1.3	0.7	M	S
2,5-Cl_2-$C_6H_3P_1$(pyrr)	2,6-Cl_2-$C_6H_3P_1$(pyrr)	1.9	2.8	0.05	M	S
	2-NO_2-4-Cl-$C_6H_3P_1$(pyrr)	1.6	2	1.1	M	S
2,6-Cl_2-$C_6H_3P_1$(pyrr)	DMAN	2.8	2.0	0.16	T	S
	2-NO_2-4-Cl-$C_6H_3P_1$(pyrr)	2.5	2.4	1.05	T	S
DMAN	2-NO_2-4-Cl-$C_6H_3P_1$(pyrr)	2.01	5.58	0.87	T	S
	2-NO_2-5-Cl-$C_6H_3P_1$(pyrr)	1.30	3.7	1.54	T	S
2-NO_2-4-Cl-$C_6H_3P_1$(pyrr)	2,4,6-Me_3-Pyridine	3.8	6.1	1.2	T	S
2-NO_2-5-Cl-$C_6H_3P_1$(pyrr)	2-NO_2-4-CF_3-$C_6H_3P_1$(pyrr)	2.2	2.5	0.55	T	S
	4-MeO-Pyridine	1.6	7.8	0.95	T	S
		2.2	5.5	1.00	T	S
	2,4,6-Me_3-Pyridine	3.3	8.3	0.45	T	S
2-NO_2-4-CF_3-$C_6H_3P_1$(pyrr)	4-MeO-Pyridine	4.1	6.8	0.45	T	S
	2,6-Me_2-Pyridine	3.2	8.8	0.79	T	S
4-MeO-Pyridine	2,6-Me_2-Pyridine	8.5	10	0.31	T	S

	2-Me-Pyridine	3.8	7.7	1.01	T	S
2,6-Me_2-Pyridine	4-MeO-Aniline	5.7	8.2	0.45	T	S
	2,4-$(NO_2)_2$-$C_6H_3P_1$(pyrr)	7.8	2.3	0.71	T	S
4-MeO-Aniline	2-Me-Pyridine	8	7.4	0.24	T	S
		16	6.5	0.2	T	S
	2,4-$(NO_2)_2$-$C_6H_3P_1$(pyrr)	1.6	8.6	0.32	T	S
2-Me-Pyridine	2,6-$(NO_2)_2$-$C_6H_3P_1$(pyrr)	7.0	2.3	0.57	T	S
	2,6-Cl_2-4-NO_2-$C_6H_3P_1$(pyrr)	3.3	7.9	0.29	T	S
	2,4-$(NO_2)_2$-$C_6H_3P_1$(pyrr)	8.7	2.5	0.07	T	S
2,4-$(NO_2)_2$-$C_6H_3P_1$(pyrr)	Pyridine	3.0	6.5	0.71	T	S
2,6-Cl_2-4-NO_2-$C_6H_2P_1$(pyrr)	2,6-$(NO_2)_2$-$C_6H_3P_1$(pyrr)	3.7	3.7	0.30	T	S
2,6-$(NO_2)_2$-$C_6H_3P_1$(pyrr)	Aniline	2.6	3.6	0.49	T	S
	Pyridine	4	5.8	0.13	T	S
Pyridine	Aniline	10	8.5	0.38	T	S
		8.9	6.0	0.30	T	S
	2-Me-Aniline	8.3	6.6	0.43	T	S
Aniline	N,N-Me_2-Aniline	3.8	2.9	0.52	T	S
2-Me-Aniline	4-Br-Aniline	4.2	3.0	1.10	T	S
	N,N-Me_2-Aniline	3.8	2.5	0.46	T	S
N,N-Me_2-Aniline	2-MeO-Pyridine	8.1	9.0	2.1	T	S
4-Br-Aniline	2-MeO-Pyridine	4.2	7.9	1.40	T	S

[a] $\Delta pK_{ip} = pK_{ip}(A) - pK_{ip}(B)$. [b] Abbreviation of the acid titrated with: T = CF_3SO_3H, M = CH_3SO_3H. [c] Calculation method: NV - one base is "non-visible", ΔpK_{ip} calculated on molar basis, S - calculated from UV-vis spectra.

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Tähtsamad teaduspublikatsioonid

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