sulfur for oxygen indicates that sulfur produces a stronger ligand field than does oxygen in the nickel(II) chelates. This finding is in accord with the ligand-exchange studies[22] and stability constant measurements of monothio-β-diketone–nickel(II) chelates[23] which reveal that nickel(II) in a square-planar environment binds sulfur more strongly than oxygen. The high-frequency values of the first spin-allowed band in the spectra of the N-phenyl- and N-methyl-2-thiocarbamoyldimedone complexes of nickel(II) place these ligands higher in the spectrochemical series than the monothio-β-diketones investigated by Chaston and Livingstone.[22] Conjugation with the second carbonyl group in the dimedone ring may contribute to the crystal field stabilization energy.

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(22) S. H. H. Chaston, S. E. Livingstone, and T. N. Lockyer, *Australian J. Chem.*, **19**, 1401 (1966).

(23) S. H. H. Chaston and S. E. Livingstone, *ibid.*, **19**, 2035 (1966).

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The Dipolar Ion Structure of Phosphoramidic Acid. Heats of Ionization[1]

By Dana Levine and Irwin B. Wilson

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Although there is conclusive evidence that amino acids exist in solution as dipolar ions, there is no completely definitive evidence concerning the structure of phosphoramidates in solution. The phosphoramidate anion, $[NH_3^+PO_3^{2-}]^-$, is believed to have a tripolar ion structure in crystals of the sodium salt.[2] This conclusion is based upon X-ray diffraction studies showing trigonal symmetry of the N–P axis with three N–O hydrogen bonds of equal length and with equal ONP angles. Although the positions of the hydrogen atoms were not determined, it is probable that if one of the hydrogen atoms were contributed by oxygen, the symmetry would be measurably disturbed.

This crystallographic study is the primary basis on which phosphoramidates are taken to have the tripolar ion structure in solution. The second ionization constant of phosphoramidic acid was found to increase in the presence of formaldehyde;[3] the first was not investigated. This result suggests and favors the dipolar ion structure of phosphoramidic acid, but it is also consistent with the "uncharged" formulation because it is quite reasonable to consider that the formaldehyde adduct of the uncharged form would be more acidic than the original structure.

The dipolar ion structure has been postulated in the treatment of hydrolysis and reactions of nucleophiles with phosphoramidic acid.[3–7]

In this work we have endeavored to obtain additional information bearing on the structure of phosphoramidic acid in solution by evaluating the heats of ionization of the first and second ionizations of phosphoramidic acid. The rationale of this method is based upon data presented in Table I. The heats of ionization of phosphoric acid at 25° involving the ionization of hydroxyl groups are numerically small for both the first and second ionizations. Therefore, we must expect that those ionizations of phosphoramidic acid which correspond to the ionization of hydroxyl groups will have numerically low heats of ionization. The heat of ionization of an ammonium function in phosphoramidic acid, on the other hand, might well be expected to have a large positive value if its ionization were the second ionization, *i.e.*, the one with $pK_a = 8.6$.

Results and Discussion

The ionizations of phosphoramidic acid can be represented as

$$\begin{tikzcd}
\begin{matrix}[NH_3{}^+PO_3H^-]^0, \\ NH_2PO_3H_2\end{matrix} \arrow[r, shift left, "-H^+,\ K_1"] & \arrow[l, shift left] \left\{\begin{matrix}[NH_3{}^+PO_3{}^{2-}]^-, \\ NH_2PO_3H^-\end{matrix}\right\} \arrow[d, shift left, "-H^+,\ K_2"] \\
& NH_2PO_3{}^{2-} \arrow[u, shift left]
\end{tikzcd}$$

Values for ionization constants K_1 and K_2 were obtained from titration curves at different temperatures and ionic strengths. The ionization constants at infinite dilution (Table I) were obtained for 25° by extrapolation of the lines obtained by plotting pK *vs.* $\sqrt{\mu}/(1 + 1.15\sqrt{\mu})$ (Figure 1) in accordance with eq 1, where A and B are functions of the temperature and have the values of 0.509 and 0.329 at 25°.[8] The sym-

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(1) This research was supported by National Institutes of Health Fellowship 5F1-GM-25,555-02 and Grant NB 00573-16 and by National Science Foundation Grant 18926.

(2) E. Hobbs, D. Corbridge, and B. Reistrick, *Acta Cryst.*, **6**, 621 (1953).

(3) W. Jencks and M. Gilchrist, *J. Amer. Chem. Soc.*, **86**, 1410 (1964).

(4) J. D. Chanley and E. Feageson, *ibid.*, **85**, 1181 (1963).

(5) M. Halmann, A. Lapidot, and D. Samuel, *J. Chem. Soc.*, 1299 (1963).

(6) O. T. Quimby, A. Narath, and F. H. Lohman, *J. Amer. Chem. Soc.*, **82**, 1099 (1960).

(7) W. Jencks and M. Gilchrist, *ibid.*, **87**, 3199 (1965).

TABLE I

IONIZATION CONSTANTS, HEATS OF IONIZATION, AND ENTROPIES OF IONIZATION OF PHOSPHORAMIDIC ACID AND SOME COMPARISON COMPOUNDS AT 25°

Acid	Ref	pK_0	ΔH, kcal/mole	$\Delta S°$, eu
NH_3PO_3H	This work	3.08	~0	−14
$NH_3PO_3^-$	This work	8.63	+5.2 ± 0.1[a]	−22
H_3PO_4	*c*	2.12	−1.8[b]	−15.8
$H_2PO_4^-$	*c*	7.21	+0.8[b]	−30.5
NH_4^+	*d*	9.25	+12.5[b]	
$C_6H_5NH_3^+$	*d*	4.6	+7.5	
$C_5H_5NH^+$	*d*	5.2	+5.1	
Glycine	*d*	2.35	+1.2[b]	
Glycine	*d*	9.78	+10.8[b]	

[a] By "least squares" from data of Figure 2. [b] Bates[8] gives values of $\partial \log K/\partial T$ which yield ΔH values of −1.8 and +0.9 kcal/mole for the first and second ionizations of phosphoric acid, +11.9 kcal/mole for ammonium ion and +1.1 and +10.2 kcal/mole for glycine. [c] H. Harned and B. Owen, "Physical Chemistry of Electrolytic Solutions," Reinhold Publishing Corp., New York, N. Y., 1950. [d] E. Cohn and J. Edsall, "Proteins, Amino Acids and Peptides," Reinhold Publishing Corp., New York, N. Y., 1943.

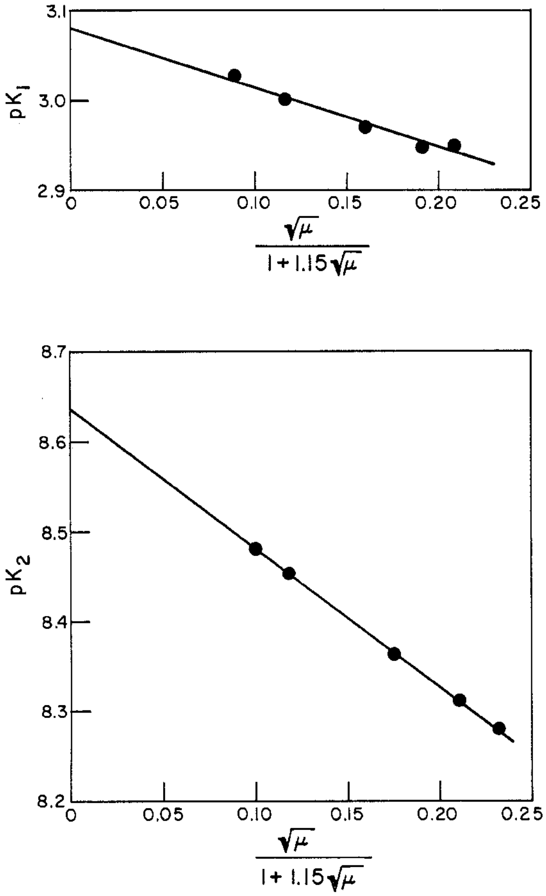

Figure 1.—The variation of apparent ionization constant with ionic strength at 25°.

$$pK = pK_0 + \frac{A\sqrt{\mu}(Z_i^2 - Z_j^2)}{1 + Ba\sqrt{\mu}} \qquad (1)$$

bol a refers to the average radius of the acid and conjugate base and was taken as 3.5 Å. It is difficult to assign an exact value for this constant, but the result for weak acids is not very sensitive to a[9] and a value is customarily taken which gives a straight line.[8] The symbols Z_i and Z_j refer to the net electrical charge of the acid and the conjugate base. For the first ionization $Z_i^2 - Z_j^2 = -1$ and for the second ionization $Z_i^2 - Z_j^2 = -3$. We obtain from the slope of the line −1.3 and −3.0, respectively, for $Z_i^2 - Z_j^2$.

The rate of change of pK with temperature was converted to apparent heats of ionization from the equation $\partial pK/\partial T = -\Delta H/2.3\,RT^2$.

The apparent heats (kcal/mole) of ionization (Table I) were obtained at ionic strength 0.04 for the first ionization (Table II) and at 0.05 and 1.05 for the second ionization from a plot of pK_a *vs.* $1/T$ (Figure 2). The values of heats of ionization can often be used to identify the ionizing group.[9] The values of the apparent heats of ionization are very different for the first and second ionizations of phosphoramidic acid and a comparison with the values for phosphoric acid and for various kinds of ammonium ions enables a clean-cut assignment of the groups undergoing ionization. The two ionization constants of phosphoramidic acid, broadly speaking, have similar values as those of phosphoric acid. If both ionizations were derived from hydroxyl groups, we should expect fairly similar apparent heats of ionization for phosphoramidic acid and phosphoric acid. Data for the first ionization of phosphoramidic acid are similar to the data for phosphoric acid and fit for a hydroxyl group, but data for the second ionization are not similar to the data for phosphoric acid and do not fit for a hydroxyl group. The second heat of ionization is reasonable for an ammonium group. The data enable an assignment of the first ionization to a hydroxyl group and the second to an ammonium group. Therefore, the prevailing species of phosphoramidic acid in solution is the dipolar ion $[NH_3^+ \text{-} PO_3H^-]^0$, and the prevailing species of the monoanion is the tripolar ion $[NH_3^+PO_3^{2-}]^-$. These conclusions are consistent with the cystallographic structural assignment for the monoanion.

TABLE II

THE TEMPERATURE VARIATION OF THE FIRST IONIZATION CONSTANT OF PHOSPHORAMIDIC ACID[a]

Temp, °C	pK_1	ΔH, kcal/mole
18.2	2.93	~0
25.3	2.92	
32.0	2.93	
39.0	2.94	

[a] The temperature variation of pK_1 is very small. If the very slight indication of a trend is real, *i.e.*, an increase of 0.01 in 20°, ΔH is −0.2 kcal/mole.

Experimental Section

Disodium Phosphoramidate Hexahydrate.—O,O′-Diphenylphosphoramidate (mp 149.5–151°, lit. 149°) was prepared according to the procedure described by Chambers and Khorana[10]

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(8) R. G. Bates, "Determination of pH, Theory and Practice," John Wiley and Sons, Inc., New York, N. Y., 1964, pp 52–53.

(9) E. J. King, "Acid-Base Equilibria," The Macmillan Co., New York, N. Y., 1965.

(10) R. Chambers and H. Khorana, *J. Amer. Chem. Soc.*, **80**, 3749 (1958).

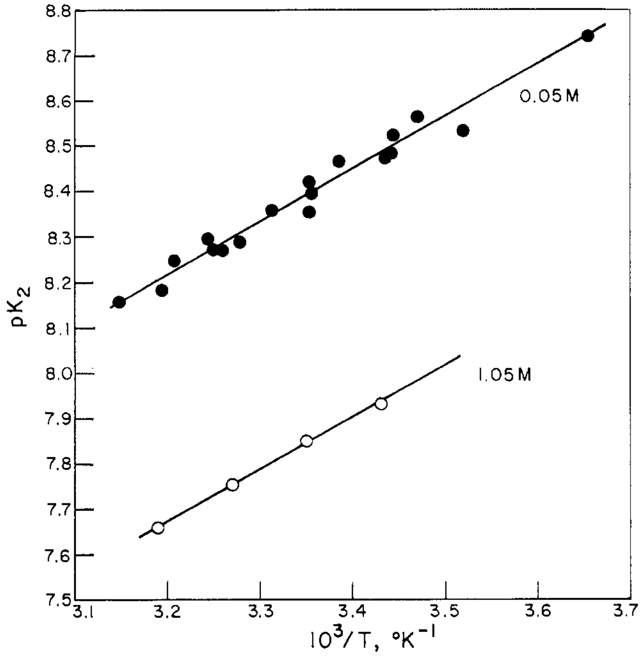

Figure 2.—The variation of apparent ionization constant at ionic strengths 0.05 and 1.05 M with temperature. The data are plotted in accordance with the equation $pK_2 = (\Delta H/2.3RT) +$ constant.

and hydrolyzed in strong sodium hydroxide by a modification of the procedure described by Klement[11] to yield disodium phosphoramidate hexahydrate. O,O′-Diphenylphosphoramidate (6.25 g, 0.025 mole) was boiled for 12 min in 4 g of NaOH dissolved in 15 ml of water. Addition of 5 ml of water and slow cooling to 0° yielded a fine, needlelike precipitate which was filtered and washed several times with 95% ethanol; yield 4.5 g (72%); mp 64–66°. Disodium phosphoramidate hexahydrate was recrystallized from a small volume of water containing one pellet of NaOH; addition of ethanol precipitated the product (mp 64–66°). Titration with standard 0.05 M HCl gave a molecular weight of 249 (calcd 249). This compound was prepared as needed a few days before use. It was protected from the light and stored in a desiccator in a freezer.

Ionization Constant of Phosphoramidic Acid and Its Anion as a Function of Ionic Strength at 25°.—Disodium phosphoramidate hexahydrate (25–200 mg) was dissolved in 20 ml of CO_2-free water in a water-jacketed beaker containing a thermometer and connected to a constant-temperature bath maintained at 25 ± 0.01°. One or two equivalents of standard 1 N HCl, dispensed from a 1-ml Gilmont micropipet in small increments, was added and the solution was rapidly stirred. The pH was measured after each addition of acid on a Corning pH meter, Model 12. The pH meter was previously standardized with two standard buffers. The pH 7.00 ± 0.01 standard buffer was purchased from Coleman. The pH 4.00 ± 0.01 standard buffer was 0.050 M potassium acid o-phthalate. The meter was checked at pH 10.00 ± 0.01 with a standard buffer purchased from Coleman. The pH meter can be read to ±0.002 unit or better on expanded scale, but repeat measurements on the same solution were reproducible to only ±0.003.

Concentrations are in moles per liter.

The second equivalent of acid was added rapidly in fewer increments so that no correction for hydrolysis of the sample was necessary. Corrections were applied below pH 4 for the amount of acid required to titrate the water.

The acid was standardized by potentiometric titration against 1 ml of standardized 1 M NaOH. The sodium hydroxide was standardized by potentiometric titration against dry reagent grade potassium acid o-phthalate.

The pK was defined as the pH of the solution at $^1/_2$ and $^3/_2$ neutralization. The ionic strength was varied by taking different weights of disodium phosphoramidate and the molar ionic strength was calculated from the concentration of all ionic species at $^1/_2$ and $^3/_2$ neutralization. The experimental end points agreed within 1% with the calculated end points.

Ionization Constant of Phosphoramidic Acid and Its Anion as a Function of Temperature.—A solution of the disodium salt in CO_2-free water, or 1 M NaCl in CO_2-free water, was titrated with 1 or 2 equiv of standard 1 N HCl and the pH was measured as before. The temperature of the solution was maintained constant to ±0.02° by means of a loosely covered water-jacketed beaker connected to a constant-temperature bath. The pH meter was previously standardized with two standard buffers at each temperature as described above. The ionic strength of the solution was 0.05 and 1.05 M for the addition of 0.5 equiv of acid and 0.04 M for the addition of 1.5 equiv of acid. The pK was defined as the pH of the solution at $^1/_2$ and $^3/_2$ neutralization. The experimental end points agreed within 1% with the calculated end points.

At higher ionic strength (1 M NaCl) only four temperatures were used and only the second ionization constant was studied. Three titrations were done at each temperature and the pK_a values were reproducible to ±0.02.

The first ionization constant was measured at only four temperatures, three titrations at each temperature and only at the lower ionic strength. Reproducibility was 0.01.

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[11] R. Klement, *Inorg. Syn.*, **6**, 100 (1960).

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Reactions of Coordinated Ligands. XV. Demonstration of the Kinetic Coordination Template Effect

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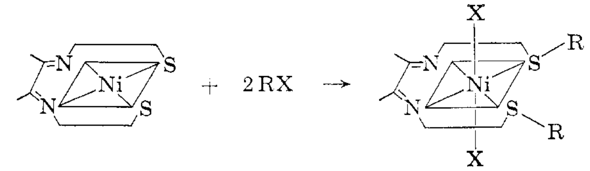
2,3-Pentanedionebis(mercaptoethylimino)nickel(II) (Ni(PE)) is a planar diamagnetic complex which contains two terminal *cis*-mercaptide groups.[1,2] The integrity of this *cis* configuration is assured by the presence of three fused five-membered rings, and the crystal structure[3] reported for the related 2,3-butanedione derivative confirms these conclusions. When this complex reacts with any monofunctional alkylating agent, both mercapto groups are alkylated and the resulting product contains coordinated thioether[4]



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[1] M. C. Thompson and D. H. Busch, *J. Am. Chem. Soc.*, **84**, 1762 (1962).

[2] M. C. Thompson and D. H. Busch, *ibid.*, **86**, 213 (1964).

[3] Q. Fernando and P. Wheatley, *Inorg. Chem.*, **4**, 1726 (1965).

[4] D. H. Busch, D. C. Jicha, M. C. Thompson, J. W. Wrathall, and E. Blinn, *J. Am. Chem. Soc.*, **86**, 3642 (1964).