a light yellow, mobile liquid of $n^{20}D$ 1.5452, with a characteristic phosphate ester smell. Capillary g.c. analysis indicated that it contains two major components in two to three ratio.

Saponification Value. Calcd. for the monoadduct, $C_{14}H_{21}O_2PS_2$: 316. Found: 314.

In another experiment, 9.3 g. (0.05 mole) of diethyldithiophosphoric acid was added to 3.3 g. (0.025 mole) of dicyclopentadiene in the same manner. Three hours after the addition 48% of the dithiophosphoric acid had reacted. This percentage remained essentially unchanged during 3 days standing of the reaction mixture at room temperature. Work-up of the reaction mixture in a manner described above yielded the monoadduct.

Dimethyl- and diisopropyldithiophosphoric acid reacted with dicyclopentadiene in a similar manner to yield liquid monoadducts.

Addition of Diethyldithiophosphoric Acid to Aldrin.—To a solution of 18.2 g. (0.05 mole) of Aldrin (m.p. 101–102°) in 50 ml. of *n*-heptane, 9.3 g. (0.05 mole) of diethyldithiophosphoric acid was added. The reaction mixture was allowed to stand for 16 hr. A subsequent titration of a sample with potassium hydroxide in the presence of neutral red indicator showed that only 12.5% of the diethyldithiophosphroic acid remained unchanged. The diethyldithiophosphoric acid monoadduct of Aldrin crystallized from the solution on cooling. The crude crystals were recrystallized from ethanol to yield 13.8 g. (50%) of *exo*-2-diethylthiophosphorylmercapto - *endo* - 5,6,7,8,9,9 - hexachloro-*exo–endo*-1,2,3,4,4a,5,8 - dimethanonaphthalene, II, as colorless crystals of m.p. 89.5–90.5°.

A similar procedure starting with 7.9 g. (0.05 mole) of dimethyl dithiophosphoric acid and 18.2 g. (0.05 mole) of Aldrin yielded 14.3 g. (55%) of *exo*-2-dimethylthiophosphorylmercapto-*endo*-5,6,7,8,9,9-hexachloro-*exo–endo*-1,2,3,4,4a,5,8,8a-octahydro-1,4,-5,8-dimethanonaphthalene, m.p. 112–113.5°.

Anal. Calcd. for $C_{14}H_{15}Cl_6O_2PS_2$: C, 32.08; H, 2.88; Cl, 40.59; O, 6.11; P, 6.10; S, 12.24. Found: C, 32.18; H, 2.89; Cl, 40.80; O, 6.10; P, 5.94; S, 12.60.

Addition of Diethyldithiophosphoric Acid to Norbornadiene.—To 9.2 g. (0.1 mole) of norbornadiene, 37.4 g. (0.02 mole) of diethyldithiophosphoric acid was added in the manner described in the previous example. The spontaneous exothermic reaction was 90% completed in 3 hr. On further standing, partial crystallization of the mixture occurred. The crystals were filtered off and recrystallized from ethanol to yield 23.7 g. (51%) of the colorless diadduct, m.p. 97.5–99°.

In another experiment equimolar amounts of norbornadiene and diethyldithiophosphoric acid were reacted in the same manner. Titration of the reaction mixture indicated that practically all the diethyldithiophosphoric acid reacted within an hour. A slightly yellow liquid adduct of $n^{20}D$ 1.5236 was obtained which according to g.c. consisted 70% of one and 30% of another isomer. G.c. analysis of the liquid part of the diaddition reaction showed the presence of the smaller component only.

Addition of Benzenethiol to Norbornadiene.—To 22 g. (0.2 mole) of benzenethiol, 9.2 g. (0.1 mole) of norbornadiene was added in the manner described in the previous example. However, the sequence of addition was reversed to repress the formation of 3-nortricyclyl 4-tolyl sulfide by the excess of thiol present. After the completion of the addition, the mixture was heated to 70° and kept at that temperature for 1 hr. A subsequent thiol determination indicated that only 28% of the original thiol was still present unreacted in the reaction mixture.

On scratching and cooling, crystals were formed in the reaction mixture. To complete the crystallization, the mixture was kept and recrystallized from *n*-heptane. In this manner, 10.8 g. (24%) of bisphenylmercaptonorbornadiene, m.p. 122–123° was obtained as colorless, large rhombic crystals.

Anal. Calcd. for $C_{19}H_{20}S_2$: C, 73.03; H, 6.45; S, 20.52 Found: C, 72.92; H, 6.53; S, 20.36.

In another experiment, 9.2 g. (0.1 mole) of norbornadiene was added to 11 g. (0.1 mole) of benzenethiol with cooling at room temperature. Then the mixture was allowed to stand for 5 days. By that time, all the thiol had reacted. Analysis of the crude product by capillary g.c. showed four peaks of 1.5%, 34.4%, 6.9% and 57.2% intensity. G.c. analysis of the liquid part from the diaddition experiment showed the absence of the 34.4% peak.

Addition of Diethyldithiophosphoric Acid to Conjugated Dienes.—A mixture of 0.25 mole of a thiol and 0.25 mole of a diene reacted in a closed 100-ml. round-bottom flask with magnetic stirring at room temperature. In the case of 2,5-dimethyl-2,4-hexadiene ultraviolet irradiation of the reaction mixture was necessary to increase the reaction rate. Water cooling of the irradiated reaction mixtures was necessary to avoid any rise of the temperature. Relative rates of the additions were determined on the basis of the decrease of the diethyldithiophosphoric acid content. Some of the comparative rate data are shown in Tables I and II.

After an arbitrary length of time the reaction was discontinued and the mixture was worked up. The unchanged diethyldithiophosphoric acid was removed by washing with 5% aqueous sodium carbonate solution. If separation was difficult, the organic phase was diluted with ether or benzene. After the drying, the separated organic phase was fractionally distilled. The adducts were usually obtained at about 1 mm. as mobile, colorless or slightly yellow liquid distillates having a characteristic phosphate ester odor. Separation of the isomeric monoadducts by superfractionation was not attempted. Some of the physical and analytical data of the isomeric adduct mixtures are given in Table II. The various isomers were identified and semiquantitatively determined by n.m.r. (Table III). Quantitative determination of the isomers was accomplished by capillary g.c. (Table II).

Acknowledgment.—The authors wish to thank T. G. Jermansen and A. M. Palmer for valuable technical help.

Reactivity of Thiophosphates. I. Hydrolysis of Phosphorothioic Acid[1–3]

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The hydrolysis of phosphorothioic acid (H_3PO_3S) in aqueous solution at constant ionic strength has rate maxima at pH ~3.0 and at pH ~8.0 and the rate minima at pH ~7.0 and at pH ~ −0.30. The reactivity of the dianion at pH 8 is unusual since the dianions of most simple organic phosphate esters are inert. In acid solutions stronger than 2 *M* a rate increase is observed which is interpreted as a salt effect on the rate of hydrolysis of the neutral acid.

Phosphorothioic acid, H_3PO_3S, is reported to decompose readily to hydrogen sulfide and phosphoric acid although its anhydrous trisodium salt is stable.[5] This investigation was undertaken in order to compare quantitatively the rate of hydrolysis of phosphorothioic

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[1] We are grateful to the National Institutes of Health for support of this work (grant A-1023).

[2] For experimental details, see O. B. Ramsay, Ph.D. thesis, University of Pennsylvania, 1960.

[3] Reported at the 139th National Meeting of the American Chemical Society, St. Louis, Mo., March, 1961.

[4] Department of Chemistry, Syracuse University, Syracuse 10, N. Y.

[5] A. Wurtz, *Compt. rend.*, **24**, 288 (1847); R. Klement, *Z. anorg. allgem. Chem.*, **253**, 237 (1947); J. V. Karabinos, R. A. Paulson, and W. H. Smith, *J. Res. Natl. Bur. Std.*, **48**, 322 (1952); S. K. Yasuda and J. L. Lambert, *J. Am. Chem. Soc.*, **76**, 5356 (1954); F. Binkley, *J. Biol. Chem.*, **181**, 317 (1949).

acid with its thioalkyl analogs.[6,7] The lack of any nonpolar hydrocarbon fragment in this thiophosphate made it possible that the compound would show behavior different from that of the alkyl esters as the medium was altered. This report discusses the effect of the acidity of the medium on the rate of hydrolysis since this is an effect which has been extensively studied with other esters of phosphoric acid.[2,8]

Ionization Constants.—Values for pK_2 and pK_3 of phosphorothioic acid have been given as 5.6 and 10.2 at 25°.[9] A titration of a dilute solution of the trisodium salt at *ca.* 25° gave $pK_1 = 2.05$, $pK_2 = 5.6$, $pK_3 = 10.3$. The p*K* values in solutions of 1 *M* ionic strength were $pK_1 = 1.35$, $pK_2 = 4.95$, $pK_3 = 9.8$. A pK_1 of 1.26 can be calculated from Hammett parameters given by Kabachnik.[10] The pK_1 determined from a titration curve of the trisodium salt may be in error because of decomposition of the salt and because of the low concentrations of substrate used. The second and third ionization constants are greater than the corresponding constants for orthophosphoric acid.

pH-Rate Profile.[11]—A plot of the observed first-order rate constants against pH is shown in Fig. 1. The rate data are given in Table I. The rate constants for hydrolysis at 52.8° and at 1 *M* ionic strength for the neutral acid ($k_N = 2.43 \times 10^{-5}$ sec.$^{-1}$), monoanion ($k_M = 16.7 \times 10^{-5}$ sec.$^{-1}$), dianion ($k_D = 3.81 \times 10^{-5}$ sec.$^{-1}$), and trianion ($k_T = 1.56 \times 10^{-5}$ sec. $^{-1}$) were calculated from the observed rate constants and the ionization constants (at 25°) by the method of least squares from

Table I

Rate Constants for Hydrolysis of Phosphorothioic Acid at 52.8° from pH 1–10.6 at 1 *M* Ionic Strength

pH	$10^5 k_{obsd}$, sec.$^{-1}$	pH	$10^5 k_{obsd}$, sec.$^{-1}$
1.00	7.45	5.00	9.51
1.05	7.75	6.00	4.29
1.50	11.4	7.20	3.91
1.55	11.9	7.50	3.87
2.05	16.0	8.00	3.94
2.10	14.4	8.50	3.92
2.65	15.9	9.07	3.79
2.70	16.5	9.55	3.21
3.50	15.7	10.25	2.41
4.00	14.8	10.60	1.71

$$k_{obsd} = k_N \frac{C_N}{C_P} + k_M \frac{C_M}{C_P} + k_D \frac{C_D}{C_P} + k_T \frac{C_T}{C_P}$$

where C_N, C_M, C_D, and C_T are concentrations of neutral acid, monoanion, dianion, and trianion, respectively, and $C_P = C_N + C_M + C_D + C_T$. From the calculated rate constants, a pH-rate profile can be constructed which agrees well with the observed profile. An equally good fit of the calculated curve to the observed can be obtained if $k_T = 0$ and $pK_3 = 10.3$. It is probable that k_T is very small and that the pK_3 determination was in error.

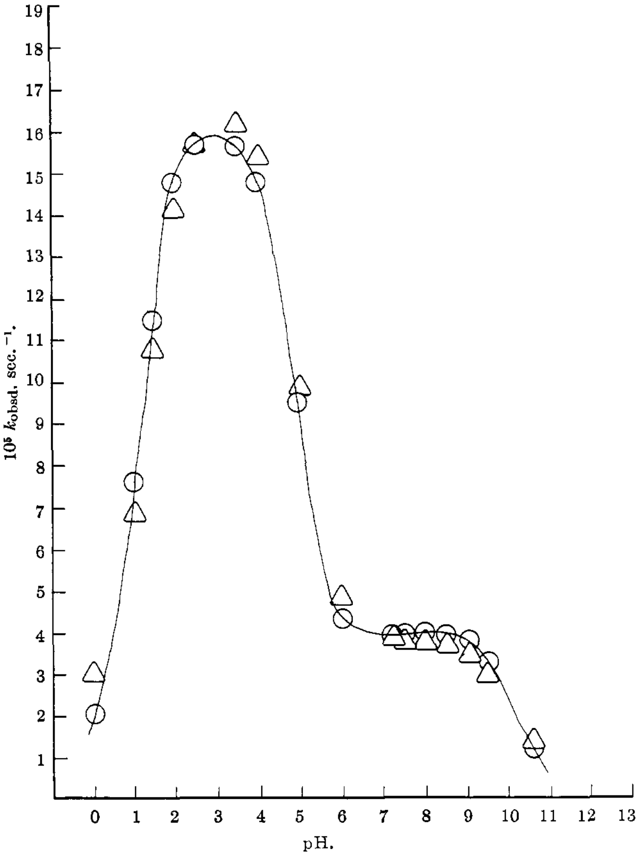

Fig. 1.—pH-Rate profile for the hydrolysis of phosphorothioic acid at 1 *M* ionic strength: ○, experimental points; △, calculated points.

Hydrolysis in Neutral Acid Region.—Phosphorothioic acid is reported to yield sulfur in the presence of oxidizing acids or air.[5] Yields of hydrogen sulfide were 79–86% in up to 5.4 *M* perchloric acid. No visible sulfur was produced in the perchloric acid used, but considerable sulfur was produced in 9 *M* sulfuric acid. Aqueous solutions of perchloric acid are reported not to be reduced by hydrogen sulfide.[12]

The rate of hydrolysis reaches a minimum around 2 *M* hydrochloric or perchloric acid. At higher concentrations of acid the rate increases; increasing the concentration of neutral salt (lithium chloride) also increases the rate. The rate constant for the hydrolysis of the neutral acid is not proportional to the concentration of acid. The data on the hydrolysis in strong acid are given in Table II.

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(6) See Paper II, D. C. Dittmer, O. B. Ramsay, and R. E. Spalding, *J. Org. Chem.*, **28**, 1273 (1963).

(7) D. E. Koshland and E. B. Herr, *Biochem. Biophys. Acta*, **25**, 219 (1957).

(8) The hydrolysis of phosphate esters has been reviewed in "Phosphoric Esters and Related Compounds," Special Publication no. 8, The Chemical Society, London, 1957.

(9) P. D. Bartlett and C. G. Swain, *J. Am. Chem. Soc.*, **71**, 1406 (1949).

(10) M. I. Kabachnik, *Proc. Acad. Sci. USSR, Chem. Sect.* (*Eng. Transl.*), **110**, 577 (1956).

(11) Kinetic data cannot be used to distinguish between a reaction of water with a monoanion of phosphorothioic acid or of hydroxide ion with the neutral acid, nor can a reaction of water with a dianion be distinguished kinetically from a reaction of hydroxide ion with the monoanion. It may be noted, however, that hydroxide ion shows much greater nucleophilicity than does water toward phosphorus in triphenyl phosphate and trimethyl phosphate [R. F. Hudson and D. C. Harper, *J. Chem. Soc.*, 1356 (1958); P. W. C. Barnard, C. A. Bunton, D. R. Llewellyn, C. A. Vernon, and V. A. Welch, *ibid.*, 2670 (1961)]. For the sake of organization, the discussion is arranged according to the major species present in each pH range.

(12) "Gmelins Handbuch der anorganischen Chemie," 8th ed., Vol. 3. Chlorine, Verlag Chemie, Berlin, 1927, p. 387.

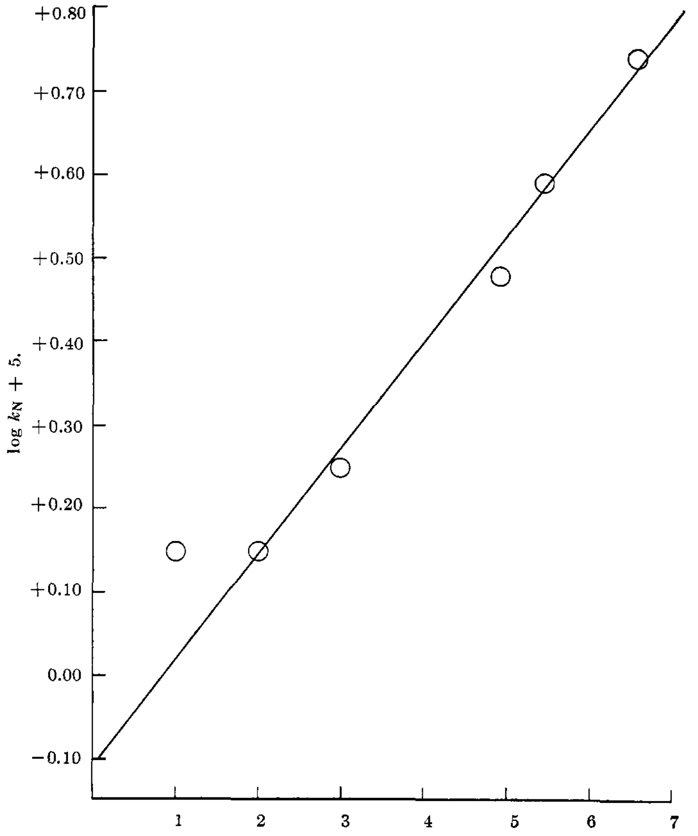


Fig. 2.—Relationship of log k_N to ionic strength.

TABLE II

RATE CONSTANTS FOR THE HYDROLYSIS OF PHOSPHOROTHIOIC ACID AT 52.8° IN 1–6.86 M ACID

Acid concn., M	$10^5 k_{obsd}$, sec.$^{-1}$
1.00 $HClO_4$	2.05
1.98 $HClO_4$	1.71
2.97 $HClO_4$	1.95
4.90 $HClO_4$	3.06
5.43 $HClO_4$	3.93
6.56 $HClO_4$	5.44
2.45 HCl	1.84
3.93 HCl	1.92
6.86 HCl	2.76
2.31 HCl + 4.86 M LiCl[a]	5.97
2.31 HCl + 3.69 M LiCl	5.10
2.31 HCl + 2.69 M LiCl	4.33
3.33 HCl + 3.80 M LiCl[a]	5.06
4.76 HCl + 2.43 M LiCl[a]	4.00

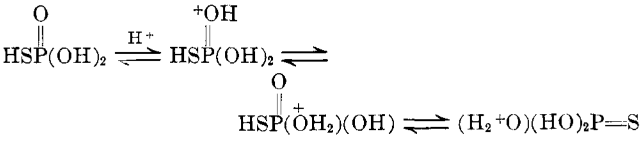
[a] Ionic strength is constant at 7.1–7.2 M.

A plot of log k_N[13] against concentration (ionic strength) of perchloric acid (Fig. 2) is linear from 2–6.6 M perchloric acid. An H_0 plot is linear also, but the slope is +0.24. The Hammett acidity function is related to the ionic strength.[14] A Bunnett plot[15] of (log $k_N + H_0$) *vs.* log a_{H_2O} is linear only from 5 M upwards (slope = +4.1). Below 5 M the slope increases greatly. A plot of H_0 *vs.* log a_{H_2O} for these points also is linear above 5 M; and the linearity observed in the Bunnett plot is attributed to this, since H_0 is changing more rapidly with increasing acidity than is log k_N. A similar behavior in the plot of (log $k_N + H_0$) *vs.* log a_{H_2O} for the S-*n*-butyl ester of phosphorothioic acid has been observed,[6] although there was still some curvature at high acidities (a slope of about +4.5 could be obtained). A plot of (log k_N − [$HClO_4$]) *vs.* log a_{H_2O} has a maximum between 3 and 5 M. The plot of (log k_N − [$HClO_4$]) for the S-*n*-butyl ester has no maximum and was similar in shape but shallower than the plot of (log $k_N + H_0$); a crude slope of about +0.5 could be obtained from points beyond 5 M.

It may be thought that the increase in rate with increasing acidity is caused by the formation and hydrolysis of the conjugate acid of phosphorothioic acid. The conjugate acid of phosphoric acid was invoked to explain the increasing rate of its exchange with $H_2{}^{18}O$ as the acidity of the medium was increased.[16] However, it is difficult to understand why S-*n*-butylphosphorothioic acid fails to show acid catalysis of its rate of hydrolysis.[2,6] Data in Table II show that hydrochloric acid affects the rate but slightly, less than does lithium chloride.

The conjugate acids of phosphoric acid[16] and its methyl ester[17] are reactive because on protonation the molecule obtains a better leaving group—water in the case of phosphoric acid, methanol in the case of methyl phosphate. Protonation of sulfur is less favored than protonation of oxygen, and the reactivity of the conjugate acids in the sulfur case and in the oxygen case would not be comparable since in the latter the leaving group has been provided by protonation. The conjugate acid of the sulfur compound (protonated on oxygen) may be reactive, but the results reported in this paper and in the paper on the hydrolysis of the S-*n*-butyl ester[6] indicate that it is not greatly reactive. The behavior of phosphorothioic acid and its S-*n*-butyl ester in acid medium may be attributed to salt effects on the rate of hydrolysis of the neutral acid. The lithium chloride data in Table II tend to support this interpretation.

It is conceivable that acid might catalyze the tautomerism of phosphorothioic acid; the conjugate acid may exist to a large extent in the thiono form which is less reactive than the neutral acid. The rate of hydrolysis of trialkylthionophosphates is less than the rate of hydrolysis of trialkylphosphates.[18] The addition of acid to aqueous solutions of phosphorothioic acid would have two effects: one to increase the rate of hydrolysis by a salt effect on the neutral acid, and two, to decrease the rate by formation of a less reac-



$$HS\overset{\overset{O}{\|}}{P}(OH)_2 \overset{H^+}{\rightleftharpoons} HS\overset{\overset{^+OH}{\|}}{P}(OH)_2 \rightleftharpoons HS\overset{\overset{O}{\|}}{P}(\overset{+}{O}H_2)(OH) \rightleftharpoons (H_2{}^+O)(HO)_2P{=}S$$

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(13) k_N is defined as

$$\frac{k_{obsd} - k_M a_w \dfrac{C_M}{C_P}}{C_N/C_P}$$

where a_w is the activity (relative humidity) of water and the other quantities are as defined earlier. At acidities greater than about 3.5 M, the term in k_M becomes negligible. If it is assumed that the transition state of the monoanion contains a molecule of water, transition state theory requires that the a_w term be included. The plot in Fig. 2 is not changed much by omitting a_w, but the points at low acidities show greater deviations from the line.

(14) I. I. Moiseev and R. M. Flid, *J. Appl. Chem. USSR* (*Eng. Transl.*), **27**, 1047 (1954).

(15) J. F. Bunnett, *J. Am. Chem. Soc.*, **82**, 499 (1960); **83**, 4956, 4968, 4973, 4978 (1961).

(16) C. A. Bunton, D. R. Llewellyn, C. A. Vernon, and V. A. Welch, *J. Chem. Soc.*, 1636 (1961).

(17) C. A. Bunton, D. R. Llewellyn, K. O. Oldham, and C. A. Vernon *ibid.*, 3574 (1958).

(18) D. F. Heath, *ibid.*, 3804 (1956).

tive conjugate acid. From this viewpoint, one can understand the greater effect of lithium chloride over hydrochloric acid in increasing the rate of hydrolysis.

If, despite these arguments, it is assumed that the rate increase with phosphorothioic acid is caused solely by the formation of the conjugate acid, a negative value of the equilibrium constant for the conversion of the conjugate acid to the neutral acid is obtained. The rate constant for the hydrolysis of the neutral acid used in this calculation was obtained from the pH-rate profile where its use gives good agreement with experiment.

Attempts to study the hydrolysis in perchloric acid–sodium perchlorate solutions of constant ionic strength were not successful because of the production of considerable sulfur even at low acidities. This could be caused by some impurity such as chlorate in the sodium perchlorate or to the increase in the oxidation potential of perchloric acid on addition of perchlorate ions.

The hydrolysis of the neutral acid may be complicated by the following equilibrium.

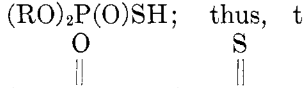
$$(HS)(HO)_2PO \rightleftharpoons (HO)_3PS$$

The reduced rate of hydrolysis of phosphorothioic acid as compared with its S-*n*-butyl ester[6] may be caused by such an equilibrium. A study of the phosphorus n.m.r. spectrum might reveal the extent to which the two species are present.

Hydrolysis may proceed (1) by a displacement (S_N2) on phosphorus by water, (2) by way of metaphosphoric acid as an intermediate, or (3) through a pentacovalent intermediate which decomposes by rupture of the S–P bond. An investigation of the exchange of $H_2{}^{18}O$ and $H_2{}^{35}S$ with phosphorothioic acid may give an indication of the mechanism.

Hydrolysis in Monoanion Region.—The most reactive species in water appears to be the monoanion,[19] a not unexpected finding since the monoanions of monoesters of phosphoric acid are the most reactive of the possible species.[8]

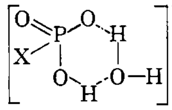
The similarity of the first ionization constants of phosphorothioic acid and phosphoric acid may indicate that the monoanion of the former is produced principally by an ionization of the O–H bond rather than an S–H bond. The ionization constants of $(RO)_2PO_2H$ are greater than the constants of $(RO)_2PS_2H$, and the ionization of O–H in $(RO)_2P(S)OH$ is reported to be greater than the ionization of S–H in $(RO)_2P(O)SH$; thus, the equilibrium (in water)



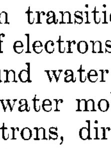
$(RO)_2\overset{O}{\overset{\|}{P}}SH \rightleftharpoons (RO)_2\overset{S}{\overset{\|}{P}}OH$ lies to the left.[20]

It is probable that both monoanions, $(HO)_2P(O)S^-$ and $H_2PSO_2O^-$, are present in solution and are interconvertible.

In the monoanion and neutral acid regions, phosphorothioic acid is considerably less reactive than the S-*n*-butyl ester.[2,6,7] This may be caused by the tautomeric equilibrium discussed above, or by an increased stability of the initial state of the phosphorothioic acid species (if it is assumed that the transition states for the ester and the free acid are of the same stability) or by different mechanisms for the two compounds. Phosphorothioic acid may fit into the water structure better than the butyl ester. The additional hydrogen bonding from the sulfhydryl group to water might stabilize the initial state of phosphorothioic acid relative to that of the butyl ester whose butyl groups is not likely to be greatly solvated. Both compounds presumably are solvated by water at their oxygen atoms; it is this kind of specific solvation, an example of which is shown below, which has been suggested to facilitate the hydrolysis of the monoanions of phosphate esters.[8]



Hydrolysis in the Dianion and Trianion Regions.—In the dianion region, phosphorothioic acid shows considerable reactivity in contrast to S-*n*-butylphosphorothioate[2,6,7] and other phosphate esters.[8] This appears to be the first phosphate which shows any reactivity toward hydrolysis in the dianion region.[21] This hydrolytic reactivity of the dianion of phosphorothioic acid contrasts with the apparent lack of reactivity of the dianion of phosphoric acid to exchange with $H_2{}^{18}O$. The leaving group in the case of the phosphorothioate is SH^- (or S^{-2}) which is expected to be a considerably better leaving group than OH^- (or O^{-2}). Rates of alkaline hydrolysis of S-alkyl esters of phosphorothioic acid are considerably greater than their oxygen analogs.[22] A rate constant of 0.1×10^{-6} sec.$^{-1}$ at 100° for exchange of the dianion of phosphoric acid with $H_2{}^{18}O$ would give a pH-rate profile nearly the same as the one obtained from the experimental data.[16] In the dianion of phosphorothioic acid, there still exists an hydroxyl group (or -SH) which, with a negatively charged oxygen, is held accountable for the great reactivity of monoanions of phosphate monoesters; an example of specific solvation by water follows.



Such specific hydration[8] positions a water molecule for attack on phosphorus, either to the rear of the semipolar P–O bond (sulfur equatorial in the transition state) or on the face of the phosphorus atom opposite sulfur (sulfur axial in transition state). The attack by an unshared pair of electrons in a *p* or hybrid orbital on oxygen in the bound water could be accomplished by a rotation of the water molecule which brings the unshared pair of electrons, directed away from phosphorus, into position for attack. This rotation would stretch or weaken the hydrogen bonds and would rupture any hydrogen bond from external water to the

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(19) Note the comment under ref. 11. The reaction of the neutral acid with hydroxide ion seemed unreasonable in the case of dibenzyl phosphate hydrolysis [J. Kumamoto and F. H. Westheimer, *J. Am. Chem. Soc.*, **77**, 2515 (1955)].

(20) M. I. Kabachnik, S. T. Joffe, and T. A. Mastryukova, *J. Gen. Chem. USSR (Eng. Transl.)*, **25**, 653 (1955).

(21) However, the monofluorophosphate ion is reported to be hydrolyzed rapidly in strongly alkaline solutions although it is stable in neutral or moderately alkaline solution [L. N. Devonshire and H. H. Rowley, *Inorg. Chem.*, **1**, 680 (1962)].

(22) R. F. Hudson and L. Keay, *J. Chem. Soc.*, 3269 (1956); E. M. Thain, *ibid.*, 4694 (1957).

unshared pair of electrons involved in the assault on phosphorus. The transition state for the attack of water may be either trigonal bipyramidal or tetragonal pyramidal, geometries which have been suggested for transition states for displacement reactions on phosphorus.[23]

Alternatively, the hydrogen bonding of water may be to sulfur to give an intermediate similar to one suggested for the hydrolysis of phosphate ester monoanions.[24]

Hydrolysis of the monoanion by hydroxide ion may be less likely because of repulsion between OH^- and $H_2PSO_3^-$. In the dianion region, the S-*n*-butyl ester is quite unreactive.[2,6]

The trianion of phosphorothioic acid seems very stable since it is prepared by refluxing thiophosphoryl chloride with aqueous sodium hydroxide.[5] However, until a quantitative study of the hydrolysis of the trianion is made, its reactivity, although clearly less than the other anionic species, must remain in doubt.

Experimental

Trisodium phosphorothioate was prepared as described elsewhere.[25] Analysis for inorganic phosphate, which is an impurity, indicated the salt was 97% pure.

Determination of Products of Hydrolysis.—To 40 ml. of 5.4 *M* perchloric acid at 52.8° was added 0.886 g. (0.00492 mole) of trisodium phosphorothioate. The hydrogen sulfide produced was swept from the reaction flask by a stream of nitrogen and passed through a condenser and into a solution of 5% mercuric cyanide. The gas stream then was passed through a solution of 10% sodium hydroxide. After 1293 min., the mercuric sulfide was collected, washed with distilled water, and dried at 110° for 15 min. The yield of mercuric sulfide was 0.868 g. (0.00373 mole) corresponding, after correction for the amount of unhydrolyzed substrate, to 86%. Yields of mercuric sulfide from hydrolyses in 6.5 *M* and 2.5 *M* perchloric acid were 79% and 84%, respectively. These yields are undoubtedly lower limits. There was no visible evidence of sulfur.

When 1.042 g. (0.00579 mole) of trisodium phosphorothioate was heated with 50 ml. of 9 *M* sulfuric acid on a steam-bath for 1 hr., 0.121 g. of mercuric sulfide and 0.157 g. of sulfur were produced. The total yield of mercuric sulfide plus sulfur corresponded to 94% hydrolysis.

Determination of Dissociation Constants.—A titration of an aqueous solution of trisodium phosphorothioate with perchloric acid gave the following pK values: pK_1 = 2.05, pK_2 = 5.6, pK_3 = 10.3. The pH was determined by means of glass and calomel electrodes with a Beckman Model H-2 pH meter. The pH of a solution of equal concentrations of the neutral acid and monoanion was 1.35 at an ionic strength of 1 *M* ($NaCl-HClO_4$). A sleeve-type calomel electrode was used in this determination. The pH values of solutions of 1 *M* ionic strength (KCl-HCl) of equal concentrations of monoanion and dianion and of dianion and trianion were, respectively, 4.95 and 9.8. All determinations were at approximately 25°.

Kinetic Procedure.—The rates of hydrolysis were followed by the rate of formation of inorganic phosphate which was determined by a colorimetric procedure similar to one described for the determination of phosphate in the presence of adenosine triphosphate.[26] Concentration of substrate in all kinetic runs was 1.27×10^{-3} *M*.

All kinetic runs from pH 1–10.6 were adjusted to an ionic strength of 1 *M* with potassium chloride. From pH 1–6, a phthalate buffer was used, from pH 7.2–8.5, a tris[27] buffer, and from pH 9.07–10.6, a borate buffer. In this work, no attempt was made to investigate general acid or base catalysis or specific buffer effects.

The molybdate reagent used in the phosphate determination was prepared by addition of 125 ml. of 10 *N* sulfuric acid to a solution of 12.5 g. of ammonium molybdate and dilution to 500 ml. An acid–ethanol solution was prepared by addition of 10 ml. of concentrated sulfuric acid to 490 ml. of absolute ethanol. The 1-amino-2-naphthol-4-sulfonic acid reagent was prepared by addition of 0.5 g. of the sulfonic acid to 195 ml. of a 15% solution of sodium bisulfite solution followed by addition of 5 ml. of a 20% solution of sodium sulfite. If a homogeneous solution was not obtained after vigorous shaking, more sodium sulfite solution could be added; but excess sodium sulfite should be avoided. The pale yellow solution is stable for about a week in a refrigerator.

Aliquots (1.0 ml.) from the kinetic runs were added to 10 ml. of the molybdate solution in a 60-ml. separatory funnel, and 1 ml. of the aminonaphtholsulfonic acid reagent and 10 ml. of isobutyl alcohol were added. The aqueous layer was discarded after the contents of the funnel were shaken vigorously for 15 sec. The isobutyl alcohol layer was removed as completely as possible with a pipet, and 2 drops of the acid–ethanol solution were added. After 30 min. the optical density of the blue solution was read at 660 mμ on a Bausch and Lomb "Spectronic 20" spectrophotometer. An isobutyl alcohol extract of a molybdate solution which contains no phosphate can be used as a blank, but it was sufficient in most cases to use isobutyl alcohol itself since the difference in blanks was only 0.02–0.03 units of optical density. Infinity optical density readings were determined for each run since this reading was sensitive to the amount of sodium perchlorate and perchloric acid present. These readings were taken either after 10 half-lives or after oxidation of a sample with bromine water. Too acidic molybdate solutions (as, for example, would occur in the runs in strong acid) must be avoided or the analytical results will be erratic. The acidity may be adjusted by addition of sodium hydroxide solution. The concentration of phosphate in the aliquot is read from a standard plot of optical density versus concentration.

Calculation of Rate Constants.—The concentration of phosphate is linearly related to the optical density of the reduced phosphomolybdate complex. The rate constant is

$$k_{\text{obs}} = (2.303/t) \log \frac{D - D_0}{D - D_t}$$

where D_0, D_t, and D are optical densities at zero time, time t, and at infinite time.

The values of a_w, the relative humidity of water, used in calculating the rate constants for the neutral acid used in Fig. 2, were the data of Robinson and Baker.[28] These data were obtained at 25° so that a small error was introduced on its use at 52.8°.

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