**Figure 2.** Projection of the crystal structure of bis(pyridine)-osmate(VI) ester of adenosine along the $a$ axis. The osmium atom is shaded black.

the adenosine molecule, and the entire complex assumes an octahedral coordination with the two osmyl groups trans to one another; this was anticipated on spectral evidence by Griffith and Rosetti.[9] The nucleoside is in the syn conformation, rather than the usually preferred anti conformation which is found in the free nucleoside.[10-12] The glycosidic torsion angle ($\chi$ = 121°), while slightly out of the normal range for syn $\beta$-purine nucleosides, is similar to that observed for formycin[13] and virazole[14] and $\psi$[15] the C(4')–C(5') conformation angle is 55°. The furanose ring is in an envelope ($^2E$) conformation C(2') endo; the phase angle of pseudorotation (P) is 166° and the amplitude of pucker is 32°.[16] Essentially, these parameters agree, to within the limits of reliability of the data, with the profile considered typical of syn $\beta$-purine nucleosides.[17]

(9) W. P. Griffith and R. Rossetti, *J. Chem. Soc., Chem. Commun.* 1449 (1972).

(10) T. F. Lai and R. E. Marsh, *Acta Crystallogr., Sect. B,* **28,** 1982 (1972).

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The packing in the unit cell is partially determined by parallel stacking of the first pyridine ring and the adenine, separated by about 3.5 Å from molecules related by symmetry. A hydrogen bond of length 2.9 Å between O(5') and N(6) in symmetry-related molecules may stabilize the syn conformation. The constraints of parallelism between the adenine and first pyridine rings hold this first pyridine ring perpendicular to the plane of O(2'), O(3'), 1N1, and 2N1. The second pyridine ring stacks parallel to a symmetry-related copy of itself. This pyridine ring seems free to rotate slightly about its Os–2N1 axis since systematically shortened and elongated bonding distances characteristic[18] of such rotation are observed.

From a study of this structure, we infer that the osmium binding observed in yeast phenylalanine tRNA may occur primarily at the location of the only free cis diol, at the 3'OH end of the molecule.[4] However, two different hydroxyl groups in the correct position along the tRNA chain might also bind an osmium, provided that these hydroxyl groups are close enough. Subsequent work on the osmium heavy atom derivatives of tRNA make it seem likely that the 3'OH end of tRNA is indeed the primary osmium binding site, while other minor binding sites are observed which could correspond to other modes of ligand binding.[4]

(17) M. Sundaraligam, The Fifth Jerusalem Symposia on Quantum Chemistry and Biochemistry, Vol. V, 417 (1973).

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## Influence of Solvation Factors on Acidity. Thermodynamics of Hydration of a Series of Meta- and Para-Substituted Phenols[1,2]

*Sir:*

Volumes and entropies of ionization ($\Delta \bar{V}_1°$ and $\Delta \bar{S}_1°$, respectively) appear to be valuable quantities in understanding the nature and degree of solute–solvent interaction.[3-7] It has recently been reported that the differ-

(1) Taken in part from the thesis of P. T. Kasudia, Georgia State University, 1974.

(2) Part VI in the series, "Thermodynamics of Acid–Base Equilibria," Part V: C. L. Liotta, H. P. Hopkins, Jr., and E. M. Perdue, *J. Amer. Chem. Soc.,* in press.

(3) (a) J. W. Larson and L. G. Hepler, "Solute-Solvent Interactions," J. F. Coetzee and C. D. Ritchie, Ed., Marcel Dekker, New York, N. Y., 1969; (b) L. G. Hepler and E. M. Woolley, "Water-A Comprehensive Treatise," Vol. 3, F. Franks, Ed., Plenum Press, New York, N. Y., 1973, Chapter 3.

(4) (a) C. L. Liotta, A. Abidaud, and H. P. Hopkins, Jr., *J. Amer. Chem. Soc.,* **94,** 8624 (1972); (b) C. L. Liotta, E. M. Perdue, and H. P. Hopkins, Jr., *ibid.,* **95,** 2439 (1973); (c) C. L. Liotta, D. F. Smith, Jr., H. P. Hopkins, Jr., and K. Rhodes, *J. Phys. Chem.,* **76,** 1909 (1972); (d) C. L. Liotta, J. H. Leavell, and D. F. Smith, Jr., *ibid.,* **71,** 3091 (1967).

(5) H. L. Friedman and C. V. Krishnan, ref 3b, Chapter 1.

(6) F. J. Millero in "Water and Aqueous Solutions," R. A. Horne, Ed., Wiley, New York, N. Y., 1971.

(7) E. J. King, *J. Phys. Chem.,* 73, 1220 (1969).

**Table I.** Thermodynamics of Solution, Vaporization, and Hydration of Meta- and Para-Substituted Nitro-, Cyano-, and Formylphenols[a]

| | $\Delta\bar{G}_v^{\circ}$ [d] | $\Delta\bar{G}_s^{\circ}$ [b] | $-\Delta_h\bar{G}^{\circ}$ | $\Delta\bar{H}_v^{\circ}$ [e] | $-\Delta\bar{H}_s^{\circ}$ [c] | $\Delta\bar{H}_h^{\circ}$ | $\Delta\bar{S}_v^{\circ}$ | $\Delta\bar{S}_s^{\circ}$ | $-\Delta\bar{S}_h^{\circ}$ |
|---|---|---|---|---|---|---|---|---|---|
| 3-Cyanophenol | 8.70 | 0.93 | 7.77 | 22.19 | 5.34 | 16.9 | 45.3 | 14.8 | 30.5 |
| 4-Cyanophenol | 9.50 | 1.22 | 8.28 | 22.30 | 5.53 | 16.8 | 43.0 | 14.5 | 28.5 |
| 3-Formylphenol | 9.04 | 1.42 | 7.62 | 22.38 | 6.31 | 16.0 | 44.7 | 16.4 | 28.3 |
| 4-Formylphenol | 9.80 | 1.49 | 8.31 | 22.57 | 6.25 | 16.3 | 42.8 | 15.9 | 26.9 |
| | 9.63[g] | | | 21.8[g] | | | 40.8[g] | | |
| | 9.90[f] | | | 23.5[f] | | | 45.5[f] | | |
| 3-Nitrophenol | 9.11 | 1.36 | 7.74 | 21.95 | 5.71 | 16.2 | 43.0 | 14.6 | 28.5 |
| | | | | 21.9[h] | | | | | |
| 4-Nitrophenol | 9.66 | 1.29 | 8.38 | 21.0 | 5.58 | 16.4 | 41.3 | 29.9 | 26.4 |
| | | | | 21.8[h] | | | | | |
| | 10.05[f] | | | 23.6[f] | | | 45.5[f] | | |

[a] $\Delta\bar{G}^{\circ}$ and $\Delta\bar{H}^{\circ}$ values in kcal/mol; $\Delta\bar{S}^{\circ}$ values in cal/(deg mol). [b] The $\Delta\bar{G}_s^{\circ}$ ($\pm0.01$ or better) values reported here were derived from solubility data ($\Delta\bar{G}_s^{\circ} = -RT \ln S$) assuming that the activity coefficients of each of the neutral species is one. The solid phase was found to be in equilibrium with the aqueous solution in all cases. [c] The values of $\Delta\bar{H}_s^{\circ}$ ($\pm0.05$ or better) are the average of at least five determinations. [d] The vapor pressure of the solids were determined by the Knudsen cell method. Our data are in reasonable agreement with literature values. [e] The $\Delta\bar{H}_v^{\circ}$ values have been derived from a temperature dependence of the vapor pressures (30–55°) with a least-square uncertainty of $\pm0.15$ kcal/mol or better. [f] G. H. Parsons, C. H. Rochester, and C. E. C. Wood, *J. Chem. Soc. B*, 533 (1971). [g] A. Aihara, *Bull. Chem. Soc. Jap.*, **33**, 194 (1960). [h] K. L. Wolf and H. Weghofer, *Z. Phys. Chem., Abt. B*, **39**, 194 (1938); I. Nitta, S. Seki, M. Momotani, and K. Sato, *J. Chem. Soc. Jap.*, **71**, 328 (1950).

ence in the $\Delta\bar{V}_i^{\circ}$ of isomeric meta- and para-substituted phenols has its origin exclusively in the partial molar volumes of the neutral (un-ionized) species.[4a] Quite surprisingly, the partial molar volumes of the isomeric charged (ionized) species were found to be nearly identical. These results contrasted sharply with the postulated origin of the corresponding differences in $\Delta\bar{S}_i^{\circ}$. Hepler[3a,8] has suggested that it is the extent of charge delocalization in the isomeric anions which influences the differences in $\Delta\bar{S}_i^{\circ}$. The more localized the charge, the greater the degree of solvent orientation and the more negative the $\Delta\bar{S}_i^{\circ}$. Using this anion solvation model, Hepler has successfully explained the relative entropies of ionization of isomeric m- and p-nitrophenols, cyanophenols, and formylphenols. While the $\Delta\bar{V}_i^{\circ}$ values are determined by measuring the initial and final states independently, the reported $\Delta\bar{S}_i^{\circ}$ values are measures of a difference between initial and final states. Consequently, in the absence of hydration data for each of the species in the equilibrium, the origin of $\Delta\bar{S}_i^{\circ}$ differences is largely conjecture. In order to determine if there is indeed a disparity in the origin of $\Delta\bar{V}_i^{\circ}$ and $\Delta\bar{S}_i^{\circ}$ differences, the hydration thermodynamics for a series of isomeric phenols have been determined. Within the framework of a complete thermodynamic cycle (Scheme 1), the relative contribu-

**Scheme I**

tions of the isomeric anions to $\Delta\bar{S}_i^{\circ}$ have been estimated.

The changes in the standard Gibbs free energy, enthalpy, and entropy for the transfer of a substituted phenol from the gas phase to aqueous solution (hydration) were calculated from vapor pressure and solution thermochemical data as determined by previously described methods.[9,4c] The derived thermodynamic parameters for the transfer process are summarized in Table I.

The difference in the entropies of hydration of two isomeric anions ($\delta$) may be calculated from eq 3. From

$$(\Delta\bar{S}_h^{\circ})_p^{-} - (\Delta\bar{S}_h^{\circ})_m^{-} = \delta =$$
$$\Delta\bar{S}^{\circ}(aq) + (\Delta\bar{S}_h^{\circ})_p - (\Delta\bar{S}_h^{\circ})_m - \Delta\bar{S}^{\circ}(g) \quad (3)$$

the data in Table I and the published $\Delta\bar{S}_i^{\circ}$ values for the isomeric m- and p-nitro-, formyl-, and cyanophenols[3a] it may be shown that

$$\delta_{NO_2} = [+6.3 - \Delta\bar{S}^{\circ}\,(g)]\ \text{cal/(deg mol)} \quad (4)$$

$$\delta_{CHO} = [+4.8 - \Delta\bar{S}^{\circ}\,(g)]\ \text{cal/(deg mol)} \quad (5)$$
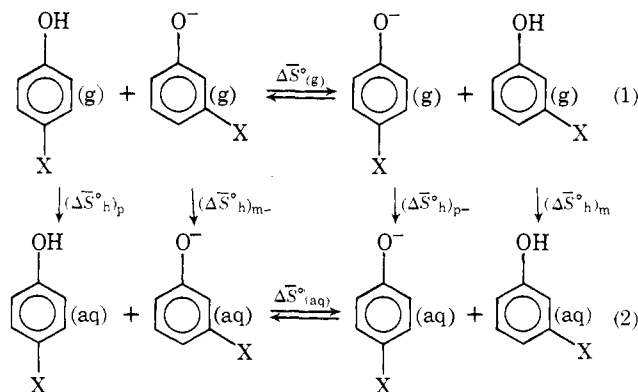
$$\delta_{CN} = [+3.9 - \Delta\bar{S}^{\circ}\,(g)]\ \text{cal/(deg mol)} \quad (6)$$

In order to estimate $\Delta\bar{S}^{\circ}(g)$ for each of the isomeric series, the rotational and vibrational contributions to $\Delta\bar{S}^{\circ}(g)$ must be calculated. The rotational component is approximately zero, except for a $-R \ln 2$ term due to the inclusion of a symmetry factor for the p-NO₂ and p-CN phenoxides. Since the vibrational frequencies for the two anions and the two neutral phenols in eq 1 are not expected to differ appreciably, the vibrational contributions to $\Delta\bar{S}^{\circ}(g)$ are assumed to be small.[2,10,11] As

(8) L. P. Fernandez and L. G. Hepler, *J. Amer. Chem. Soc.*, **81**, 1783 (1959).

(9) W. Jackson, T. S. Hung, and H. P. Hopkins, Jr., *J. Chem. Thermodyn.*, **3**, 347 (1971).

(10) E. J. King, "Acid-Base Equilibria," MacMillan, New York, N. Y., 1965.

(11) A small vibrational contribution might exist if the barriers to free rotation for the hydroxyl, nitro, and formyl groups were substantially different for the meta and para isomers. Applying simple resonance logic, it appears that the barrier to the out-of-plane motion (bending and torsional) of the p-NO₂ and p-CHO groups with respect to the ring should be greater in the ionized than in the un-ionized forms. If it is assumed that there is no difference in the out-of-plane motion frequencies of the substituents when located in the meta positions of the neutral and charged species, then the overall vibrational contribution to $\Delta S^{\circ}(g)$ should make $\delta$ more positive. The torsional problem is not encountered with the CN substituent, but the arguments with regard to the bending contribution should be similar to the p-NO₂ and p-CHO cases.

a consequence of this, they will be neglected. Therefore, the $\delta$ values for the nitro, formyl and cyanophenoxides are estimated to be $+7.7$, 4.8 and 5.2 cal/(deg mol), respectively.

These results clearly indicate that the origin of the $\Delta \bar{S}_i{}°$(aq) for the proton transfer reaction, eq 2, is almost exclusively due to the difference in entropy of hydration of the anions. This is in agreement with the interpretation of Hepler for the variations found for the entropies of ionization of isomeric substituted phenols. It has been stated in a previous publication that, based upon the corresponding $\Delta \bar{V}_i{}°$ differences and the partial molar volumes of the species involved in the equilibria, the solvation contribution to differences in acidity of isomeric phenols may have their primary origin in solvation differences in the un-ionized forms.[4a] It now appears that $\Delta \bar{S}_i{}°$ and $\Delta \bar{V}_i{}°$ measure different aspects of the solvation phenomenon and that the relationship between these parameters suggested by simple electrostatic[10] theory and from one of the Maxwell relationships[12] deserves reconsideration.

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## [1,3]Sigmatropic Shifts of Carbon–Carbon Bonds in Acid Catalyzed Rearrangements of Cyclohexadienones[1]

*Sir:*

Suprafacial [1,3]sigmatropic shifts of carbon–carbon bonds are classified as "forbidden" reactions.[2,3] The known examples of [1,3] rearrangements, whether they proceed by symmetry allowed antarafacial paths[4,5] or by other mechanisms,[6,7] proceed at temperatures high enough to provide most or all of the energy necessary for homolytic cleavage of the migrating bonds.[4-8]

(1) Reactions of Cyclohexadienones, XXXIII. Part XXXII, B. Miller and L. Lewis, *J. Org. Chem.*, **39**, 2605 (1974).

(2) R. B. Woodward and R. Hoffmann, "The Conservation of Orbital Symmetry," Academic Press, New York, N. Y., 1970.

(3) M. J. S. Dewar, *Angew. Chem., Int. Ed. Engl.*, **10**, 761 (1971).

(4) (a) J. A. Berson and G. L. Nelson, *J. Amer. Chem. Soc.*, **89**, 5503 (1967); **92**, 1096 (1970); (b) W. R. Roth and A. Friedrich, *Tetrahedron Lett.*, 2607 (1969); (c) S. Masamune, S. Takada, N. Nakatsuka, R. Vukov, and E. N. Cain, *J. Amer. Chem. Soc.*, **91**, 4422 (1969).

(5) J. A. Berson, *Accounts Chem. Res.*, **5**, 406 (1972).

(6) See J. S. Swenton and A. Wexler, *J. Amer. Chem. Soc.*, **93**, 3066 (1971), for references to [1,3] shifts in rearrangements of vinylcyclopropanes.

(7) (a) G. S. Hammond and C. D. De Boer, *J. Amer. Chem. Soc.*, **86**, 899 (1964); (b) J. A. Berson and R. W. Holder, *ibid.*, **95**, 2037 (1973); (c) R. C. Cookson and J. E. Kemp, *Chem. Commun.*, 385 (1971).

(8) For arguments that the activation energy for a concerted rearrangement may *exceed* the dissociation energy of the migrating bond, see J. A. Berson, T. Miyashi, and G. Jones, II, *J. Amer. Chem. Soc.*, **96**, 3468 (1974).

Distinguishing between concerted and diradical processes is therefore difficult, and attempts to establish the rearrangement mechanisms must rely on estimates of the degree of stereospecificity to be expected of homolytic dissociation–recombination processes.

I have observed that addition of small amounts of sulfuric acid to solutions of cyclohexadienone 1 in methanol or acetic acid results in rapid disappearance of 1 to give the [1,3] rearrangement products 3 and 4, in addition to the [1,2] rearrangement product 2. The solvolysis products, 2,6-dimethylphenol (2,6-DMP) and 4-methylbenzyl methyl ether or 4-methylbenzyl acetate, are also obtained. Phenols 2 and 3 are obtained as an inseparable mixture, whose composition may be determined by vpc[9] or (preferably) by nmr[10] analysis. The structures
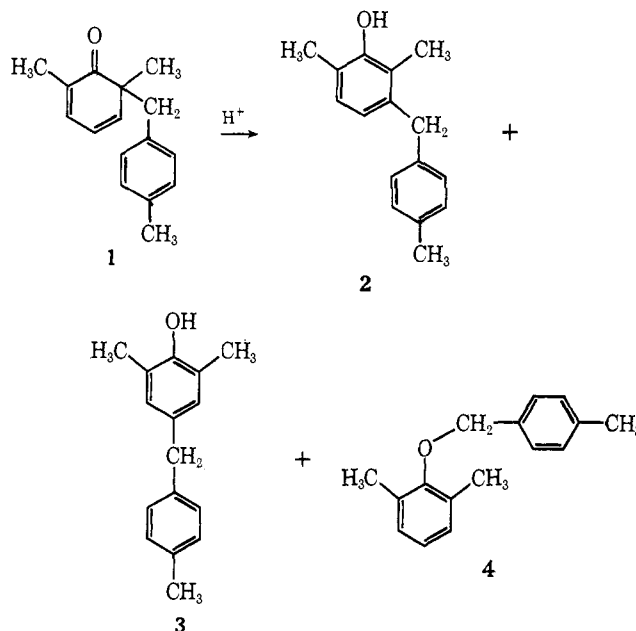


of these phenols were proved by their independent synthesis and by comparison of the spectra and vpc retention times of appropriate synthetic mixtures with those of the mixture of 2 and 3 obtained from the rearrangement of 1. Product yields are listed in Table I.

The evidence outlined below suffices to eliminate most of the possible mechanisms for formation of the [1,3] rearrangement products 3 and 4. (a) All of the reaction products, either individually or in combination, are stable for prolonged periods under the conditions of the rearrangements. (b) Addition of 10- to 20-fold molar excess of 2,6-DMP to solutions of 1 in small volumes of methanol or acetic acid (although not in more dilute solutions) resulted in an increase in the yield of 3. (See lines 4–6 in the table.) However, the ratio of 2 to 4, or 2 to 4-methylbenzyl acetate, did not significantly change when excess 2,6-DMP was added to the rearrangement mixtures. Rearrangement in the more nucleophilic solvent methanol, although it gave cleavage products as the major components of the reaction mixture, actually gave somewhat higher ratios of 3 and 4 to 2 than did rearrangement in acetic acid. Finally, rearrangement in an acetic acid–ethanethiol

(9) Partially overlapping peaks for 2 and 3 are obtained by vpc on a 5 ft × 1/8 in., 1.5% OV-101 column.

(10) The ratio of 2 to 3 in a mixture is best determined from relative areas of the singlets in the nmr spectrum of the mixture at $\delta$ 3.92 and at 3.80, due to the methylene groups in 2 and 3, respectively.