Table I. Structural Data for $\{[M_3O_2(O_2CC_2H_5)_6(H_2O)_2]_2(H_3O_2)\}Br_3 \cdot 6H_2O$

		bond distances, Å[a]					
compd	M	M(1)–M(2)	M(1)–M(3)	M(2)–M(3)	M-O(H_2O)	M-O($H_3O_2^-$)	O-O($H_3O_2^-$)
1	W	2.743 (1)	2.765 (1)	2.760 (1)	2.11 (1)	1.99 (1)	2.48 (1)
2	Mo	2.750 (2)	2.777 (2)	2.773 (1)	2.11 (1)	2.01 (1)	2.52 (1)
3	Mo	2.743 (1)	2.780 (1)	2.770 (1)	2.11 (1)	2.03 (1)	2.44 (1)
		2.751 (1)	2.779 (1)	2.779 (1)	2.11 (1)	2.02 (1)	

[a] Numbers in parentheses are esd's occurring in the last significant figure. Crystallographically distinct but chemically equivalent distances have been averaged.

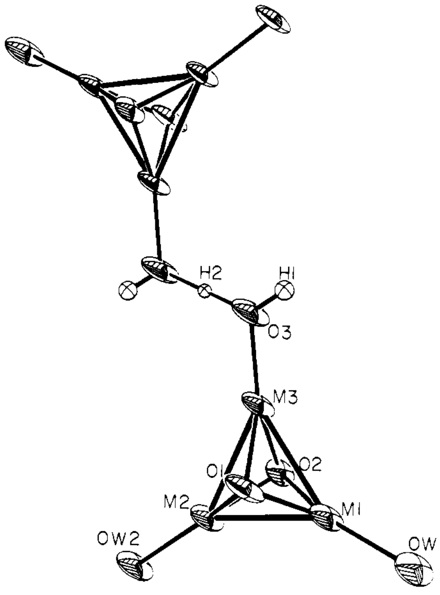

Figure 1. Skeletal structure of $\{[M_3O_2(O_2CC_2H_5)_6(H_2O)_2]_2(H_3O_2)\}^{3+}$ found in compounds **1** and **2**, M = W, Mo. The propionato groups were omitted for the sake of clarity. Hydrogen atoms H(1) and H(2) were located from the difference Fourier transform in compound **2**.

$Br_3 \cdot 6H_2O$ where M = Mo or W.[8] The principal structural results are given in Table I. Compounds **1** and **2** are isostructural and were obtained by elution with 0.5 M KBr. Compounds **2** and **3** are polymorphs; **3** was obtained by the use of 4 M HBr as eluant. The skeletal structure of the above compounds including the bridging $H_3O_2^-$ ligand is depicted in Figure 1. In **1** and **2** there is a crystallographic inversion center midway between the two oxygen atoms of the $H_3O_2^-$ unit while in **3** the asymmetric unit contains an entire $\{[Mo_3O_2(O_2CC_2H_5)_6(H_2O)_2]_2(H_3O_2)\}^{3+}$ ion. The short O–O distance in the bridging group and the presence of one hydroxide anion in the stoichiometry of the complex supports the assignment of $H_3O_2^-$ to this group. As expected the O–O distances in the coordinated $H_3O_2^-$ are longer (by ca. 0.15–0.23 Å) than in the previously reported free $H_3O_2^-$ ion.[4]

From the data in Table I it is obvious that the equilateral metal triangle found in the fluoroborate salt[6a] is subjected in these new bromide salts to considerable distortion. The resulting isosceles triangle is formed by displacement of the bridged metal atom, M(3), toward the $H_3O_2^-$ unit. As a result, the M(3)–M distance [M = M(1) or M(2)] increases by ca. 0.02 Å in **1** and by 0.03 Å in **2** and **3**. The M(3)–O($H_3O_2^-$) distance decreases by ca. 0.1 (compared to the average M–O(H_2O) distance of 2.11 Å) to ca. 2.00 Å. A similar decrease of the M–O distance was observed when other negatively charged ligands such as acetate ion replaced a water molecule coordinated to a W(IV) atom in a similar triangular cluster.[6a]

Dodson et al. proposed a mechanism for the isotopic Fe^{2+}/Fe^{3+} exchange reaction based on H-atom transfer via an $H_3O_2^-$ bridge between the iron atoms.[5]

Preliminary results[9] on the base-catalyzed decomposition of the trinuclear cluster $[Mo_3(CCH_3)_2(OAc)_6(H_2O)_3]^{2+}$ in aqueous solution showed that this process is accompanied by an electron transfer from one trimer to another. The axial water molecules in this bulky and inert cluster offer a plausible route for an electron transfer (or H atom transfer) via a bridging $H_3O_2^-$ ligand as found in the stable species reported here.

A series of future papers will provide full reports on the compounds described here and on others that have been discovered. Attempts are made to grow crystals of **1** suitable for neutron diffraction study.

Acknowledgment. We are grateful to the Israel Commission for Basic Research for financial support.

Supplementary Material Available: Tables of atomic positional parameters for all three compounds (7 pages). Ordering information is given on any current masthead page.

---

(8) Compound **1**: $\{[W_3O_2(O_2CC_2H_5)_6(H_2O)_2]_2H_3O_2\}Br_3 \cdot 6H_2O$; space group $P\bar{1}$; $a = 14.320$ (3), $b = 12.101$ (2), $c = 11.417$ (2) Å; $\alpha = 111.13$ (2), $\beta = 105.65$ (2), $\gamma = 66.51$ (2)°; $V = 1674$ (2) Å^3; $Z = 1$; d(calcd) = 2.48, d(obsd) = 2.49 ± 0.01 g cm^{-3}; $R = 0.068$, $R_w = 0.076$. Compound **2**: $\{[Mo_3O_2(O_2CC_2H_5)_6(H_2O)_2]_2H_3O_2\}Br_3 \cdot 6H_2O$; space group $P\bar{1}$; $a = 14.360$ (4), $b = 12.115$ (3), $c = 11.407$ (2) Å; $\alpha = 111.12$ (2), $\beta = 105.75$ (3), $\gamma = 66.65$ (2)°; $V = 1680$ (2) Å^3; $Z = 1$, d(calcd) = 1.95, d(obsd) = 1.96 ± 0.01 g cm^{-3}; $R = 0.055$, $R_w = 0.066$. Compound **3**: $\{[Mo_3O_2(O_2CC_2H_5)_6(H_2O)_2]_2H_3O_2\}Br_3 \cdot 6H_2O$; space group $P\bar{1}$; $a = 18.308$ (5), $b = 15.754$ (4), $c = 11.470$ (3) Å; $\alpha = 98.44$ (2), $\beta = 94.57$ (2), $\gamma = 96.67$ (2)°; $V = 3234$ (2) Å^3; $Z = 2$; d(calcd) = 2.02, d(obsd) = 2.01 ± 0.01 g cm^{-3}; $R = 0.056$, $R_w = 0.069$.

(9) Bino, A.; Gibson, D., unpublished results.

Relationship between Solution Entropies and Gas-Phase Entropies of Nonelectrolytes

Michael H. Abraham

Department of Chemistry, University of Surrey
Guildford, Surrey, United Kingdom

Received May 20, 1981

There has been considerable discussion as to the origin of entropies of hydration of gaseous nonelectrolytes. On Pierotti's version of scaled-particle theory (SPT) which yields for solution in water calculated $\Delta S_s°$ values in excellent agreement with experiment,[2] the entire $\Delta S_s°$ term results from the entropy of cavity formation in the solvent,[3] but other workers have suggested that restriction of rotation of flexible molecules is important.[4-6] Wertz[7] has recently put forward the proposition that on hydration all molecules lose a constant fraction of entropy and has argued from this that the $\Delta S_s°$ values arise from loss of internal and (especially) external degrees of freedom of the solute molecules and that there is but little contribution to $\Delta S_s°$ from the solvent itself. It was also suggested by Wertz,[7] though not demonstrated, that a similar proposition would hold for entropies of solution in solvents other than water. In view of the various comments previously made,

---

(1) R. A. Pierotti, *Chem. Rev.*, **76**, 717 (1976).

(2) M. H. Abraham and A. Nasehzadeh, *J. Chem. Soc., Faraday Trans. 1*, **77**, 321 (1981).

(3) There is, of course, a correction term to take into account the difference in standard states in the gas phase and in solution.

(4) M. Aranow and L. Witten, *J. Phys. Chem.*, **64**, 1643 (1960).

(5) O. W. Howarth, *J. Chem. Soc., Faraday Trans. 1*, **71**, 2303 (1975).

(6) M. Osinga, *J. Am. Chem. Soc.*, **101**, 1621 (1979).

(7) D. H. Wertz, *J. Am. Chem. Soc.*, **102**, 5316 (1980).

Table I. Ratio of $-\Delta S_s°$ to $\bar{S}°(g)$[a]

solute	solvent				
	water[b]		methanol[c]	benzene[c]	cyclohexane[c]
He	0.62[d]	0.80[e]	0.48[e]	0.35[e]	0.33[e]
Ar	0.73	0.83	0.43	0.35	0.37
Rn	0.72	0.82	0.53	0.40	0.37
H_2	0.66	0.81		0.36	0.36
O_2	0.48	0.63	0.33	0.26	0.27
CF_4	0.47	0.59		0.21	0.21
methane	0.59	0.72	0.38	0.30	0.31
n-octane	0.43	0.50	0.25	0.19	0.21
H_2O	0.44[f]	0.63[f]			
n-octanol	0.43[g]	0.51			
NH_3	0.46	0.61			
n-hexylamine	0.45[h]	0.51[h]			

[a] All values of $\bar{S}°$ taken from D. R. Stull, E. F. Westrum, and G. C. Sinke, "The Chemical Thermodynamics of Organic Compounds", Wiley, New York, 1969, or from CODATA recommended values in *J. Chem. Thermodyn.*, **8**, 603 (1976) and **10**, 903 (1978), unless shown otherwise. [b] $\Delta S_s°$ values from ref 9 unless shown otherwise. [c] $\Delta S_s°$ values from E. Wilhelm and R. Battino, *Chem. Rev.*, **73**, 1, (1973), and ref 2. [d] Wertz' standard states of 55.5 mol dm^{-3} in each phase. [e] Standard states of 1 atm gas phase and unit mole fraction solution. [f] From the entropy of vaporization. [g] K. Bocek, *J. Chromatogr.*, **162**, 209 (1979). [h] $\Delta S_s°$ from F. M. Jones III and E. M. Arnett, *Prog. Phys. Org. Chem.*, **11**, 263 (1974). $\bar{S}°$ (gas, 1 atm) estimated as 105.5 cal K^{-1} mol^{-1} by the method of group contributions (see footnote *a*).

Table II. Correlations of $\Delta S_s°$ through Equation 3[a]

solvent	*m*(calcd)	solutes	no.	*m*	*c*	*ρ*
water	−0.48	rare gases	6	−0.95	5.1	0.971
		C_1-C_8 alkanes[b]	9	−0.33	−17.2	0.994
		C_1-C_8 *n*-alcohols[c]	8	−0.43	−9.8	0.993
		C_1-C_6 *n*-alkylamines[d]	6	−0.35	−16.8	0.977
		inorganic gases[e]	4	−0.33	−15.3	0.987
methanol	−0.32	rare gases	5	−0.60	5.1	0.971
		C_1-C_8 alkanes[f]	9	−0.17	−8.8	0.990
ethanol	−0.31	rare gases	4	−0.44	1.0	0.940
		C_1-C_8 alkanes[f]	9	−0.19	−7.3	0.993
		inorganic gases	4	−0.09	−11.0	0.918
benzene	−0.22	rare gases	6	−0.53	6.7	0.860
		C_1-C_8 alkanes[f]	9	−0.12	−8.1	0.997
		inorganic gases	4	−0.07	−9.0	0.903
hexane	−0.14	rare gases	6	−0.50	5.2	0.944
		C_1-C_8 alkanes[f]	9	−0.20	−3.0	0.995
		inorganic gases	4	−0.10	−8.1	0.986

[a] In all cases, $\Delta S_s°$ refers to standard states 1 atm (gas) and unit mole fraction (solution) and $\bar{S}°$ to 1 atm (gas); all values in cal K^{-1} mol^{-1}. For references see footnotes to Table I except where shown. [b] C_1-C_8 *n*-alkanes and isobutane; values of $\Delta S_s°$ from ref 2 and 9. [c] From $\Delta G_s°$ values by J. H. Rytting, L. P. Huston, and T. Higuchi, *J. Pharm. Sci.*, 67, 615 (1978) and $\Delta H_s°$ values by J. Konicek and I. Wadso, *Acta Chem. Scand.*, **25**, 1541 (1971). See also footnote *g* in Table I. [d] See footnote *h*, Table I. [e] H_2, N_2, O_2, and CO. [f] Values either from ref 2 or from unpublished data by M. H. Abraham.

it seemed important to investigate the generality of Wertz' proposition and the deductions that might be made from it.

Wertz[7] suggested that if standard states are chosen to be 55.5 mol dm^{-3} in both gas phase and aqueous solution, then all molecules lose 46% of their gas-phase entropy on solution in water. In Table I is given the ratio $-\Delta S_s°/\bar{S}°$ for solution in water, on the basis of these standard states, for various solutes. Although the ratio is close to 0.46 for a number of molecules (H_2O, NH_3, *n*-hexylamine, CF_4, and O_2),[8,9] it cannot be considered as a constant, the spread being from 0.73 (argon) to 0.43 (*n*-octane and *n*-octanol). If standard states are chosen to be the ideal gas at 1 atm and the ideal solution at unit mole fraction, the spread in $-\Delta S_s°/\bar{S}°$ for solution in water is now from 0.83 (argon) to 0.50 (*n*-octane). When the latter standard states are used, it is clear also that for representative nonaqueous solvents $-\Delta S_s°/\bar{S}°$ is not very constant either (Table I).

An alternative formulation was given by Wertz[7] in the form of eq 1 where standard states were taken as 1 atm for the gas and 0.0408 mol dm^{-3} for solution in water. If standard states for the

$$\Delta S_s° \text{ (in water)} = -0.46\bar{S}° + 6.6 \qquad (1)$$

aqueous phase are chosen differently, only the constant term in eq 1 is affected. Thus for standard states of 1 atm gas and unit mole fraction in aqueous solution, eq 1 is transformed into eq 2.

$$\Delta S_s° \text{ (in water)} = -0.46\bar{S}° - 7.8 \qquad (2)$$

Because the slope of these equations is independent of standard state, it is probably more sound to analyze the data in this way, at least for a series of solutes.

Results of such an analysis using the general eq 3 are in Table

$$\Delta S_s° \text{ (in solvent)} = m\bar{S}° + c \qquad (3)$$

II for various series of solutes in water and in some representative nonaqueous solvents; also given are the calculated values of slopes in eq 3 deduced by the method given by Wertz.[7] Quite good correlations are found in most cases so that eq 3 represents a useful method of predicting values of $\Delta S_s°$ from known values of $\bar{S}°$. However, for a given solvent, the slopes and intercepts differ from one set of solutes to another so that it is doubtful if correlations through eq 1–3 represent any fundamental general property. Thus for solution in water *m* varies from −0.33 to −0.95 (calcd −0.46) and *c* varies from −17.2 to +5.1 cal K^{-1} mol^{-1} (calcd −7.8, eq 2).

In any case, the mere observation of a mathematical relationship between two variables does not by itself imply that there is necessarily any causal relationship. Thus for the alkanes in water, eq 3 holds very well (ρ = 0.994, confidence limit = 99.9999%). But for these solutes $\bar{S}°$ is linearly related to the solute hard-sphere

---

(8) The value for methane in Table I is not the same as that calculated by Wertz[7] because of quite different experimental entropies of solution. I have used the value recently suggested.[9]

(9) E. Wilhelm, R. Battino, and R. J. Wilcock, *Chem. Rev.*, **77**, 219 (1977).

diameter[2] ($\rho = 0.956$), the solute molar volume[2] ($\rho = 0.988$), and the solute polarizability[2] ($\rho = 0.988$) so that values of $\Delta S_s°$ are also linearly related to these solute properties.[10] If the relationship between $\Delta S_s°$ and $\bar{S}°$ is fundamental, then the observed $\Delta S_s°$ values might be attributed to loss of degrees of freedom of the solute. However, if, for example, the fundamental correlation is that between $\Delta S_s°$ and solute size, the origin of the $\Delta S_s°$ values might then be interpreted in terms of SPT—the larger the solute, the larger the required cavity in the solvent and the more negative the entropy of cavity formation. This illustrates the difficulty (if not the impossibility) of obtaining information on the molecular level from purely macroscopic quantitities (i.e., $\Delta S_s°$ and $\bar{S}°$).

Wertz[7] also suggested that partition coefficients, or $\Delta G_t°$ values, of hydrocarbons between water and 1-octanol could be calculated through eq 4, where $\Delta S_s°$ is the entropy of solution of the hydrocarbon in water. In the deduction of eq 4, Wertz assumed that the corresponding value of $\Delta H_t°$ was zero. Using eq 4, Wertz was able to reproduce the $\Delta G_t°$ values for the alkanes methane

$$\Delta G_t° \text{ (water} \rightarrow \text{1-octanol)} = -0.63\text{T}\Delta S_s° \text{ (in water)} \quad (4)$$

to *n*-butane. There is a difficulty over standard states, since $\Delta G_t°$ seems to refer to standard states of unit mole fraction in water and 1-octanol whereas $\Delta S_s°$ in eq 4 corresponds to standard states of 0.0408 mol dm^{-3} in water and the gas phase. But this aside, the fundamental assumption that $\Delta H_t° = 0$ is just not correct. Thus for methane, $\Delta G_t° = 2.8$ kcal mol^{-1} (mole fraction standard states) but $\Delta H_t°$ is -2.4 kcal mol^{-1},[9,11] so the apparent success of eq 4 is somewhat fortuitous.

---

(10) Correlation constants for plots of $\Delta S_s°$ in water against these solute properties are ($n = 9$ in all cases) 0.938 (hard-sphere diameter), 0.977 (molar volume), and 0.991 (solute polarizability).

(11) R. J. Wilcock, R. Battino, W. F. Danforth, and E. Wilhelm, *J. Chem. Thermodyn.*, **10**, 817 (1978).

Free Radical Epoxidation of 7,8-Dihydroxy-7,8-Dihydrobenzo[*a*]pyrene by Hematin and Polyunsaturated Fatty Acid Hydroperoxides

Thomas A. Dix and Lawrence J. Marnett*

Department of Chemistry, Wayne State University
Detroit, Michigan 48202

Received June 1, 1981

The hydroperoxidase activity of the heme protein prostaglandin endoperoxide synthetase catalyzes the hydroperoxide-dependent oxidation of polycyclic hydrocarbons, aromatic amines, and nitrofurans.[1] The natural hydroperoxide substrate is the hydroperoxy intermediate of prostaglandin biosynthesis, PGG_2, but polyunsaturated fatty acid hydroperoxides will also trigger oxidation.[2] One of the compounds oxidized is 7,8-dihydroxy-7,8-dihydrobenzo[*a*]pyrene (BP-7,8-diol) which is a metabolite of the ubiquitous chemical carcinogen, benzo[*a*]pyrene.[3] BP-7,8-diol is oxygenated during prostaglandin biosynthesis to strongly mutagenic derivatives which have been identified as diol epoxides.[4] Since these diol epoxides are generally regarded as the ultimate carcinogenic forms of benzo[*a*]pyrene, it is possible that prostaglandin endoperoxide synthetase plays a role in polycyclic hydrocarbon carcinogenesis.[5] We have found that other heme proteins with peroxidase activity, such as methemoglobin, will catalyze the hydroperoxide-dependent epoxidation of BP-7,8-diol.[6] We now report that hematin, in the absence of protein, will catalyze the epoxidation of BP-7,8-diol by fatty acid hydroperoxides and that peroxy radicals derived from the carbon skeleton of the hydroperoxide appear to be the epoxidizing agents.

Table I. Hydroperoxide Specificity of O_2 Uptake and BP-7,8-diol Oxidation

hydroperoxide	O_2 Uptake, μM^a	diol,[b] μM/min
13-OOH-18:2[c]	160 ± 10	12.3 ± 0.9
9-OOH-18:2[d]	160 ± 5	12 ± 1
15-OOH-20:4[e]	160 ± 10	12 ± 1
cumene hydroperoxide	0	1.9 ± 0.1
n-butyl hydroperoxide	0	0
hydrogen peroxide	0	0

[a] [Hydroperoxide] = 500 μM, [hematin] = 5 μM, [Tween 20] = 200 μM, 0.1 M $NaPO_4$ (pH 7.8), 25 °C. [b] [Hydroperoxide] = 50 μM, [hematin] = 0.5 μM, [BP-7,8-diol] = 18 μM, [Tween 20] = 200 μM, 0.1 M $NaPO_4$(pH 7.8), 25 °C. [c] 13-hydroperoxy-9,11-octadecadienoic acid. [d] 9-hydroperoxy-10,12-octadecadienoic acid. [e] 15-hydroperoxy-5,8,11,13-eicosatetraenoic acid.

The initial velocity of BP-7,8-diol oxidation was determined by monitoring the decrease in absorbance at 370 nm. Saturation kinetics were observed for hematin and 13-hydroperoxy-9,11-octadecadienoic acid with half-maximal velocities at 0.25 and 20 μM, respectively. Maximal velocities were observed at 0.5 μM hematin and 50 μM hydroperoxide. A linear concentration dependence was observed for BP-7,8-diol in the range 0.1–50 μM; higher concentrations could not be experimentally attained.

No oxidation of BP-7,8-diol is observed at Tween 20 concentrations below 50 μM, its critical micellar concentration.[7] A sharp increase to maximal initial velocities occurs between 70 and 100 μM Tween. There is an increase in solubility accompanied by a shift in the absorbance maximum of BP-7,8-diol to longer wavelengths as the Tween 20 concentration is increased from 50 to 100 μM. This indicates that the detergent solubilizes the hydrophobic substrate. In addition, it appears that detergents stimulate the peroxidase activity of the hematin. Oxidation of the water-soluble aromatic hydrocarbon, 9,10-bis(carboxyethyl)anthracene,[8] is increased sixfold by raising the Tween 20 concentration from 25 to 100 μM. Thus, it appears that the basis for the detergent requirement is more complex than simple substrate solubilization.

The major products of BP-7,8-diol oxidation are the isomeric tetraols **3**–**6** which are the hydrolysis products of epoxides **1** and **2** (Scheme I).[9] In a typical experiment 13-hydroperoxy-9,11-octadecadienoic acid (5 μmol) is added to a solution of [7-^{14}C]-BP-7,8-diol (1.8 μmol) and hematin (0.05 μmol) in 100 mL 0.1 M sodium phosphate (pH 7.8) containing Tween 20 (20 μmol). After 5 min, 2-*tert*-butyl-4-methoxyphenol is added to terminate the reaction, and the solution is extracted with ethyl acetate. The extract is concentrated and analyzed by HPLC on a reversed-phase column (Radial Pak B) eluted with water–methanol gradients. Tetraols **3**–**6** account for 33% of the total radioactivity eluting

---

(1) (a) Marnett, L. J. *Life Sci.* **1981**, *29*, 531–546. (b) Marnett, L. J.; Reed, G. A.; Johnson, J. T., *Biochem. Biophys. Res. Commun.* **1977**, *79*, 569–576. (c) Sivarajah, K.; Anderson, M. W.; Eling, T. E. *Life Sci.* **1978**, *23*, 2571–2578. (d) Sivarajah, K.; Lasker, J. M.; Abou-Donia, M. B.; Eling, T. E. *J. Biol. Chem.*, in press. (e) Davis, B. B.; Mattammal, M. B.; Zenser, T. V. *Nephron* **1981**, *27*, 187–196.

(2) (a) Marnett, L. J.; Wlodawer, P.; Samuelsson, B. *J. Biol. Chem.* **1975**, *250*, 8510–8517. (b) Marnett, L. J.; Reed, G. A. *Biochemistry* 1979, *18*, 2923–2929. (c) Zenser, T. V.; Mattammal, M. B.; Davis, B. B. *Cancer Res.* **1980**, *40*, 114–118.

(3) (a) Yang, S. K.; McCourt, D. W.; Roller, P. P.; Gelboin, H. V. *Proc. Natl. Acad. Sci. U.S.A.* **1976**, *73*, 2594–2598. (b) Thakker, D.; Yagi, H.; Lu, A. Y.; Levin, W.; Conney, A. H.; Jerina, D. M. *Ibid.* **1976**, *73*, 3381–3385.

(4) (a) Marnett, L. J.; Reed, G. A.; Denison, D. J. *Biochem. Biophys. Res. Comm.* **1978**, *82*, 210–216. (b) Marnett, L. J.; Johnson, J. T.; Bienkowski, M. J. *FEBS Lett.* **1979**, *106*, 13–16. (c) Sivarajah, K.; Mukhtar, H.; Eling. T. *Ibid.* **1979**, *106*, 17–20. (d) Marnett, L. J.; Bienkowski, M. J. *Biochem. Biophys. Res. Commun.* **1980**, *96*, 639–647.

(5) (a) Kapitulnik, J.; Wislocki, P. G.; Levin, W. Yagi, H.; Jerina, D. M.; Conney, A. H. *Cancer Res.* **1978**, *38*, 354–358. (b) Sims, P.; Grover, P. L.; Swaisland, A.; Pal, K.; Hewer, A., *Nature (London)* **1974**, *252*, 326–328. (c) Huberman, E.; Sachs, L.; Yang, S. K.; Gelboin, H. V. *Proc. Natl. Acad. Sci. U.S.A.* **1976**, *73*, 607–611. (d) Wislocki, P. G.; Wood, A. W.; Chang, R. L.; Levin, W.; Yagi, H.; Hernandez, O.; Jerina, D. M.; Conney, A. H. *Biochem. Biophys. Res. Commun.* **1976**, *68*, 1006–1012.

(6) Marnett, L. J.; Johnson, J. T.; Reed, G. A. *Proc. Am. Assoc. Cancer Res.* **1979**, *20*, 60.

(7) Helenius, A.; Simons, K. *Biochim. Biophys. Acta* **1975**, *415*, 29–79.

(8) Lindig, B. A.; Rodgers, M. A. J.; Schaap, A. P. *J. Am. Chem. Soc.* **1980**, *102*, 5590–5593.

(9) (a) Yagi, H.; Thakker, D. R.; Hernandez, O.; Koreeda, M.; Jerina, D. M. *Ibid.* **1977**, *99*, 1604–1611. (b) Keller, J. W.; Heidelberger, C.; Beland, F. A.; Harvey, R. G. *Ibid.* **1976**, *98*, 8276–8277.

0002-7863/81/1503-6744$01.25/0