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Rate of Hydrolysis and Transfer Free Energies of Aliphatic Alkyl Nitrites at Micellar Interfaces. A Kinetic Study

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The acid hydrolysis of methyl, ethyl, 2-methoxyethyl, 2-ethoxyethyl, 2-bromoethyl, 2-chloroethyl, n-propyl, isopropyl, n-butyl, 2-methyl-2-propyl, n-pentyl, n-hexyl, cyclohexyl, 2-phenylethyl, 1-phenylethyl, and isooctyl nitrites is kinetically studied in the presence of tetradecyltrimethylammonium bromide (TTABr) and sodium dodecyl sulfate (SDS) micelles. Cationic micelles inhibit the reaction through an effective association of alkyl nitrites (R–ONO) to the micellar pseudophase. The reactivity of the bound substrate becomes severely reduced due to the absence of protons in the micellar Stern layer. The kinetic results allow a quantitative estimation of the association constants. Kinetic data corresponding to the reaction with SDS are analyzed by using the simple pseudophase ion-exchange (PPIE) model, taking into account partitioning of the reactants between the micellar and aqueous pseudophases and competition between H^+ and Na^+ for binding to the micelles. Binding constants and rate constants for the reaction in the micellar pseudophase and in water, obtained from the kinetic analysis, are reported. Experimental values of the kinetic and thermodynamic parameters are explained in terms of changes both in the structures of alkyl nitrites and in the nature of the interface.

Introduction

The effect of micelles upon overall reaction rates and equilibria depends upon the incorporation of solutes into the micellar pseudophase, and the micellar effects therefore typically increase with increasing hydrophobicity of the reactants.[1-4]

Micellar effects on a number of bimolecular reactions have been interpreted largely in terms of the distribution of both reactants between the aqueous and micellar pseudophases.[5-9]

In the literature, there are a number of extremely important studies on the direct determination, by using different analytical methods, of interfacial partition coefficients for a large variety of molecules between water–organic solvents,[10] water–surfactant interfaces[11,12] or water–biological membranes or liposomes.[13] These studies have identified the driving forces for partitioning and have established quantitative general rules that determine the partitioning behavior.

As Diamond, Katz, and Wright[13,14] establish in their excellent reviews, the permeability of biological membranes to nonelectrolytes depends intimately upon molecular structure, so that a small change in the structure of a solute often causes a deep change in its permeating power. Two factors seem to determine the distribution of a solute between interfaces: (a) its membrane–water partition coefficient and (b) its diffusion coefficient in the membrane.

Partition coefficient, the most important parameter, is a true measure of interfacial association, and, as such, it may provide valuable insights into the driving forces for interfacial solubilization. General factors determine its magnitude: the hydrophobic effect, the magnitude of intermolecular forces solute–solvent and solute–membrane, and the fluidity of the membrane.

Although micelles are not good models of membranes, it can be said[15] that rules for solubilization in biological membranes apply to micelles as well. The present work is an attempt at connecting the micellar solubilization process and the literature on solubilization in biological membranes. The partitioning behavior of nonelectrolytes between water and gall bladder epithelial cells studied by Diamond and Wright follows the same general trends we observe in our present kinetic study on the acid hydrolysis of a great structural variety of alkyl nitrites in the presence of cationic micelles (tetradecyltrimethylammonium bromide, TTABr) and anionic micelles (sodium dodecyl sulfate, SDS).

Several facts add interest to the present work: Alkyl nitrites have proved to be highly interesting substrates from a biomedical point of view, because of their well-known vasodilatory and muscle relaxation properties.[16] However, no values of reactivity and partitioning of this important family of compounds at interfaces have hitherto

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been reported. In addition, they can act as nitrosating agents to give nitroso compounds,[17] some of which possess carcinogenic properties. Finally, the use of a great structural variety of alkyl nitrites sheds light on the nature of micellar interfaces.

Experimental Section

Alkyl nitrites were synthesized by treating the corresponding alcohols with sodium nitrite in aqueous–sulfuric acid,[18] purified by fractional distillation and stored at low temperature over molecular sieves to prevent their hydrolysis.

Methyl and ethyl nitrites, which are gases at room temperature, were prepared in acetonitrile by treating known quantities of 2-methyl-2-propyl nitrite with 10^2-fold excess of the appropriate alcohol.[19] This yielded an equilibrium mixture of alkyl nitrite, its alcohol, and 2-methyl-2-propanol. (At reaction conditions there are, for example, in the reaction mixture ca. 5×10^{-4} mol·dm^{-3} of 2-methyl-2-propanol, 5×10^{-4} mol·dm^{-3} of methyl nitrite, and 10^{-2} mol·dm^{-3} of methanol.)

All other reagents used were of the maximum commercially available purity, supplied either by Merck or Sigma, and were used as supplied.

The kinetics of the reactions were studied using a Hi-Tech SF-61 stopped flow apparatus, noting the decreasing absorbance due to the consumption of alkyl nitrites, in the 240–250 nm region.

Stock solutions of alkyl nitrites were prepared in dioxane. A small amount of this solution was dissolved in weakly basic, 2×10^{-3} mol·dm^{-3}, carbonate buffer, in which alkyl nitrites are stable, and loaded into one syringe, and the appropriate quantity of hydrochloric acid into the other. The percentage of dioxane in the final reaction mixture never exceeded 1% by volume. Both syringes contained the same surfactant concentration, to prevent sudden changes in the viscosity of the medium in the mixing process. The concentration of alkyl nitrite used was $(2–5) \times 10^{-4}$ mol·dm^{-3}. All experiments were performed at 25 °C.

The absorbance–time data of all kinetic experiments were in perfect agreement with first-order integrated equations. The reported first-order rate constant, k_o, is the mean of at least five separate measurements, whose values never differed in more than 2% and were usually reproducible to within a similar range.
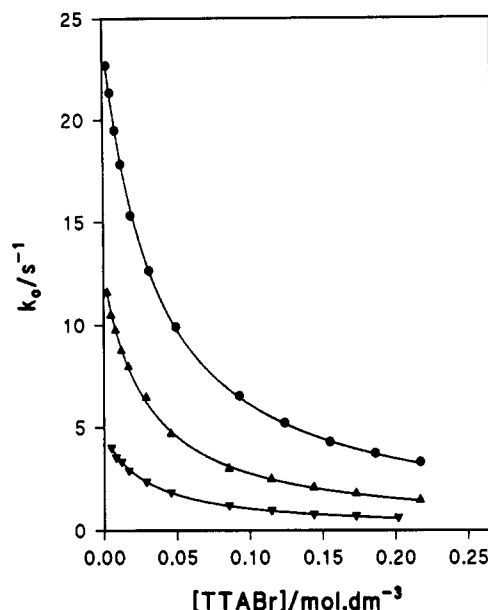
Results

Previous studies[20] from our laboratory have shown that the reaction of several alkyl nitrites with H$^+$ ions in aqueous medium proceeds via a concerted mechanism, in which protonation of the O atom and breaking of the O–N bond take place concurrently.

1. Reaction in the Presence of a Cationic Surfactant. The acid hydrolysis of several alkyl nitrites was studied in the presence of TTABr (0–0.25 mol·dm^{-3}) for different concentrations of surfactant and a fixed concentration of H$_3$O$^+$ (HCl). As expected, the addition of surfactant resulted in an inhibition of the reaction. Figure 1 shows typical results.

Kinetic results can be explained by means of the pseudophase model[21] which treats micelles and water as separate reaction media.

According to this model, the overall reaction rate constant, k_o, will be the sum of the contributions of rate constants in water, k_W', and in the micelles, k_M', and will therefore depend upon the distribution of reactants between each pseudophase and the appropriate rate constants in each pseudophase (eq 1)

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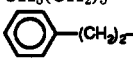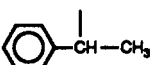(18) Noyes, W. A. *Organic Syntheses*; Wiley: New York, 1943; Collect. Vol. II.

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Figure 1. Pseudo-first-order rate constants for the acid hydrolysis of 2-methyl-2-propyl nitrite as a function of increasing [TTABr] and acidity of the medium: (●) [H$^+$] = 2.7 × 10^{-2} mol·dm^{-3}; (▲) [H$^+$] = 1.24 × 10^{-2} mol·dm^{-3}; (▼) [H$^+$] = 5 × 10^{-3} mol·dm^{-3}. Solid lines fit by eq 2, for parameters see Table I.

$$k_o = \frac{k_W' + K_S^N k_M'[D_n]}{1 + K_S^N[D_n]} \quad (1)$$

where K_S^N is the binding constant of the alkyl nitrite and [D$_n$] represents the micellized surfactant, that is, the total concentration of surfactant D less the concentration of monomeric surfactant given by the critical micelle concentration, cmc.

The observed inhibition arises because of the low concentration of H$^+$ near the cationic surface causing the reactivity of the associated substrate to be much less than that of the substrate in the aqueous phase. If $k_M'/K_S^N \ll k_W'$ (the electrical charge of the micellar surface almost excludes—although not completely[22]—the H$^+$) eq 1 becomes eq 2

$$k_o = \frac{k_{H_2O}}{1 + K_S^N([D] - cmc)} \quad (2)$$

where k_{H_2O} represents the observed rate constants in the absence of surfactant at the working acidity.

Equation 2 was found to satisfactorily fit the experimental data.

The theoretical fits—solid lines—to the experimental data are shown in Figure 1. (Only the experimental points obtained at surfactant concentrations above its cmc value are plotted in the figures.) Calculated binding constants, K_S^N, for the different alkyl nitrites are collected in Table I, in which the working overall acidities employed are also indicated. The table also includes the cmc values estimated as the lowest concentration of surfactant at which kinetic alteration was observed.

2. Reaction in the Presence of an Anionic Surfactant. As in the previous case, the influence of the concentration of SDS on the rate of hydrolysis of alkyl nitrites was studied in the presence of fixed concentrations of HCl.

Figure 2 shows typical results. First-order rate constants increase with increasing [SDS], go through maxima, and
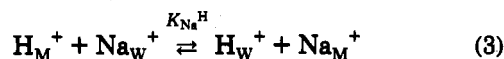
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Table I. Experimental Conditions and Values of Observed Rate Constants in Water, k_{H_2O}, and Calculated Values of K_S^N for the Partitioning of Alkyl Nitrites, R–ONO, between Water and TTABr Micelles

	R–	$10^2[H^+]^a$	k_{H_2O}/s^{-1}	$K_S^{N\ b}$	cmc^c
1.	CH$_3$–	1.27	4.80	1.6 ± 0.1	2.5
2.	CH$_3$CH$_2$–	1.27	6.87	4.90 ± 0.05	3.0
3.	CH$_3$O(CH$_2$)$_2$–	1.27	11.96	1.6 ± 0.1	2.5
4.	CH$_3$CH$_2$O(CH$_2$)$_2$–	0.74	6.30	3.6 ± 0.1	3.0
5.	ClCH$_2$CH$_2$–	1.27	15.90	11.0 ● 0.2	2.8
6.	BrCH$_2$CH$_2$–	0.75	9.55	15.5 ± 0.2	1.0
7.	(CH$_3$)$_2$CH–	0.65	4.04	10.9 ● 0.3	3.5
		1.29	8.26	11.3 ± 0.3	3.5
		variable		11.25^d	
8.	CH$_3$(CH$_2$)$_2$–	1.27	6.84	17.0 ± 0.4	3.2
9.	CH$_3$(CH$_2$)$_3$–	1.27	6.31	49 ± 1	3.0
10.	(CH$_3$)$_3$C–	0.50	4.42	32 ± 1	2.0
		1.24	11.60	32.5 ● 0.5	2.5
		2.70	23.25	28.5 ± 0.3	1.5
11.	CH$_3$(CH$_2$)$_4$–	1.27	6.10	168 ● 6	2.5
12.	C$_6$H$_{11}$–	1.24	8.59	150 ± 2	2.1
13.	CH$_3$(CH$_2$)$_5$–	1.27	5.80	584 ± 5	1.2
14.	⬡–(CH$_2$)$_2$–	1.27	9.61	230 ± 6	2.0
15.	⬡–CH—CH$_3$	1.24	8.20	309 ± 2	2.2
16.	(CH$_3$)$_2$CH(CH$_2$)$_5$–	1.27	0.648	540 ± 6	3.0

^{a} In mol·dm^{-3}. ^{b} In mol^{-1}·dm^3. ^{c} In mmol·dm^{-3}. ^{d} From experiments at fixed [TTABr] and varying [HCl].

Figure 2. Pseudo-first-order rate constants for the acid hydrolysis of (▲) 2-phenylethyl nitrite at [H$^+$] = 1.27 × 10^{-2} mol·dm^{-3}, (O) cyclohexyl nitrite at [H$^+$] = 1.24 × 10^{-2} mol·dm^{-3}, and of (□) n-butyl nitrite at [H$^+$] = 1.27 × 10^{-2} mol·dm^{-3} as a function of increasing [SDS]. Full lines were drawn fitting eq 4 to the experimental points obtained at [SDS] above its cmc value by a simulation program. For parameters see Table II.

decrease at still higher concentrations of SDS. This behavior is normal for micellar-catalyzed bimolecular processes[4,7,23] of counterions with hydrophobic substrates and are explained by the simple pseudophase ion-exchange (PPIE) model developed by Romsted.[24] The concentration of H$^+$ at the micellar surface of SDS depends upon competition with the inert counterion, Na$^+$, and this competition can be treated quantitatively by the ion-exchange equilibrium 3

$$H_M^+ + Na_W^+ \overset{K_{Na}^H}{\rightleftarrows} H_W^+ + Na_M^+ \qquad (3)$$

Therefore the pseudo-first-order rate constant for disappearance of R–ONO, k_o, will follow eq 4

$$k_o = \frac{k_W[H^+]_{total} + \left(\dfrac{k_2^m}{V}K_S^N - k_W\right)[H_m^+]}{1 + K_S^N[D_n]} \qquad (4)$$

where k_W and k_2^m are the second-order reaction rate constants in the aqueous and micellar phases, respectively, V is the molar reaction volume per mole of micellized surfactant for reaction in the micellar phase, and [H$_m^+$], the concentration of H$^+$ bound to the micelle, can be calculated from the ion-exchange equilibrium (3) as described in the literature.[25] (All concentrations are relative to the total volume of solution).

In the simple ion-exchange model the fraction of micellar charge neutralized, β, is kept constant. This is only an approximation. As determined by Pérez-Benito and Rodenas[26] the ion exchange constant between Na$^+$ and H$^+$ increases with HCl concentration until reaching a
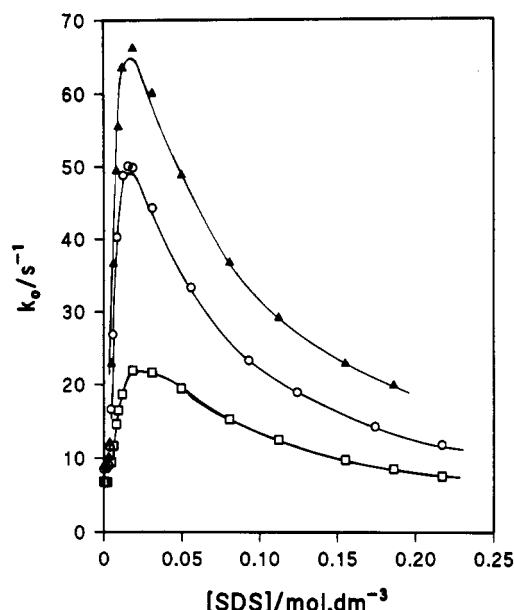
maximum value that depends upon SDS concentration. K_{Na}^H typically ranges from 0.6 to 1.

Experimental kinetic values can be adapted to this model by simulation techniques.[7,8] The experimental data were fitted to the model to estimate the "best" values of K_S^N, K_{Na}^H, and k_2^m by using constant values of β (=0.8), cmc (3–3.5) × 10^{-3} mol·dm^{-3}, and V, the volume element of the reaction in the micellar phase, estimated for SDS[27] as 0.14 dm^3 mol^{-1}. We found that values of K_{Na}^H in the range 0.7–0.8 all gave very good fits to the experimental points, yielding quite similar values of the optimized parameters. We therefore chose a value of 0.75 for K_{Na}^H in every case. This value is moreover the value found by us[7a,c] in other studies of acid catalyzed reactions in SDS micelles. The kinetic cmc is significantly lower than that in pure water (ca. 8 × 10^{-3} mol dm^{-3}). The difference is consistent with the effect of HCl and organic substrates upon cmc.[8,26]

Table II shows the calculated parameters which best explain the experimental results. The root mean square deviation of the experimental points to the model was always below 8% (typically between 1 and 5%). Figure 2 shows the theoretical fits to the experimental points. (Only observed rate constants measured at [SDS] higher than its cmc value are considered in the fitting process.)

Discussion

1. Reactivities in Aqueous and Micellar Pseudophases. Our kinetic procedure allows estimation of the bimolecular rate constants in the micellar phase of SDS (with units dm^3 mol^{-1} s^{-1}), k_2^m, that are listed in Table II.

Several conclusions can be drawn from results in this table.

(a) The second-order rate constants, k_W, in water are substantially higher (4–25 times) than those in the micellar pseudophase, k_2^m. This lower reactivity in micelles seems

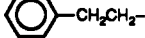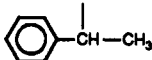(23) van de Langkruis G. B. and Engberts J. B. F. N. *J. Org. Chem.* **1984**, *49*, 4152.

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Table II. Values of Rate and Equilibrium Constants That Best Fit the Experimental Results for the Acid Hydrolysis of Alkyl Nitrites, R–ONO, in the Presence of SDS Micelles, Using the Pseudophase Ion-Exchange Model (These Values Correspond to $K_{Na}^H = 0.75$ and $\beta = 0.80$)

	R–	$10^2[H^+]^a$	$K_S^{N \, b}$	k_W^c	$k_2^{m \, c}$	k_W/k_2^m	k_{rel}
1.	CH_3–	1.27	too low	378			
2.	CH_3CH_2–	1.27	3.3	541	51	10.5	1.6
3.	$CH_3O(CH_2)_2$–	1.27	2.3	942	134	7.05	1.7
4.	$CH_3CH_2O(CH_2)_2$–	0.75	2	851	188	4.5	2.8
		1.27	2.2	772	168	4.6	2.5
5.	$ClCH_2CH_2$–	1.27	4.9	1252	131	9.5	2.3
6.	$BrCH_2CH_2$–	0.75	4	1273	188	6.8	2.9
		1.27	6.1	1283	135	9.5	2.7
7.	$(CH_3)_2CH$–	0.85	8	575	42	13.7	2.3
		1.35	8.9	575	39	14.7	2.2
8.	$CH_3CH_2CH_2$–	1.27	10.6	538	29	18.5	2.1
9.	$CH_3(CH_2)_3$–	1.27	24	497	27	18.7	3.2
10.	$(CH_3)_3C$–	0.495	9	893	93	9.7	4.5
		1.24	15	935	73	12.8	3.7
		1.90	13.0	921	78	11.8	3.2
11.	$CH_3(CH_2)_4$–	1.27	52	480	24	20	5.2
12.	C_6H_{11}–	1.24	62	693	38	18	5.8
13.	$CH_3(CH_2)_5$–	1.27	140	457	19	24.3	6.5
14.	⬡–CH_2CH_2–	1.27	46	756	56	13.5	7.4
15.	⬡–CH—CH_3	1.24	79	661	45	14.7	8.1
16.	$(CH_3)_2CH(CH_2)_5$–	1.27	700		51	10	5.3	42.2

^{a} mol·dm^{-3}. ^{b} dm^3·mol^{-1}. ^{c} dm^3·mol^{-1}·s^{-1}.

to be typical of reactions of hydrophilic ions and is consistent with the decrease of rate of our reaction in dioxane–water mixtures where reduction of the water content of the solvent causes a strong decrease in reaction rate.[20] Water activity at a micellar surface is lower than in bulk water.[28]

Nevertheless, the PPIE analysis confirms the dominant influence of the concentration effect, which is completely responsible for the observed rate acceleration in the presence of SDS micelles.

(b) For the same reaction, micellar rate constants, k_2^m, decrease with increasing substrate hydrophobicity. We can clearly see this effect in comparing values of ratios k_W/k_2^m for reactions of the series $CH_3(CH_2)_nONO$ ($n =$ 0–6). This behavior can be explained in terms of the two-site model for substrate solubilization in micelles.[29] By increase of the substrate hydrophobicity, a greater fraction of the substrate will be located closer to the micellar core away from the reaction site. Further arguments for this conclusion can be drawn by observing the ratio k_W/k_2^m for 2-phenylethyl and 1-phenylethyl nitrites; hydrogen bonds between aromatic π-electrons and hydroxyl groups might cause molecules containing aryl moieties to reside on the average in a more aqueous region of the micelle than that occupied by compounds with aliphatic chains. On the whole, our results point to a nonuniform micellar structure.[29]

(c) The relative catalysis, k_{rel} ($=k_o^{max}/k_W'$) usually increases with the hydrophobicity of the substrate. This generalization holds in our case as we can see from Table II (entry 7).

(d) Finally, the results obtained for isooctyl nitrite do not follow the same trends described by the other alkyl nitrites. The high hydrophobicity of this substrate probably causes a strong perturbation of micellar structure.
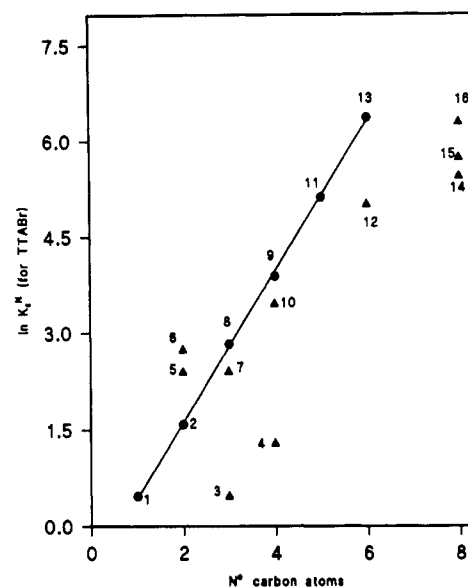
Figure 3. Correlation between partition coefficients of alkyl nitrites in tetradecyltrimethylammonium bromide + water and the number of carbon atoms in the side chain. The numbering of the solutes is the same that for Table I. Solutes not used in the correlation (▲).

In particular, the possible existence of submicellar species cannot be ruled out.

2. Relations between the Chemical Structure of Alkyl Nitrites and Their Ability To Permeate Micellar Interface. A change in the chemical structure of the solute affects the interfacial partition coefficients or binding constants by altering interface–solute and water–solute intermolecular forces (enthalpic part of the free energy of transfer) and also the "degree of order" in water and in the interface (entropic part of the free energy of transfer).

Following the rules set out by Diamond and Wright, we have used data from Tables I and II to assess the effect of various molecular groups on the partitioning behavior.

(a) Effect of Methylene Moiety. As we can see in Figures 3 and 4 introduction of each successive –CH$_2$– unit multiplies the binding constants of alkyl nitrites by a roughly constant quantity (varies between 3–3.4 with TTABr micelles and 2.2–3.2 with SDS micelles). A similar behavior was found by other authors working with different interfaces and substrates.[11,30,31] The longer the hydrocarbon chain in the substrate, the stronger the van der Waals interactions with the hydrocarbon micellar chain, an effect which "pulls into" the lipid phase longer chain substrates by stronger intermolecular forces. In addition, nonpolar solutes in aqueous solution cause the structure of the surrounding water to be more organized than that of bulk water, therefore creating "iceberg"[32] regions, so that long chain substrates are "pushed out of" the aqueous phase by the resulting gain in entropy.

Following a treatment for K_S^N similar to that reported by Bunton and Sepúlveda,[30] the free energy of transfer on crossing the interface is given by eq 5, where ΔG_{mc} corresponds to a model compound, H–ONO in our case, and N^o is the number of carbon atoms in the chain. Figures 3 and 4 show the excellent straight line resulting from a plot of $\Delta G_{tr}^{w \rightarrow M}$ against N^o. From the slope of the line we obtain values of $\Delta G_C = -696$ and -640 cal/mol for the free

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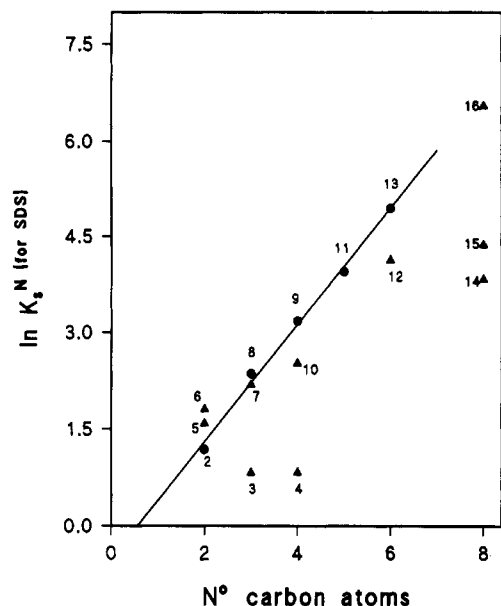
(29) (a) Al-Lohedan, H. A.; Bunton, C. A. *J. Org. Chem.* **1982**, *47*, 1160. (b) Al-Lohedan, H. A. *J. Chem. Soc., Perkin Trans. 2* **1989**, *1181*.

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(32) Nemethy, G.; Scheraga, H. A. *J. Chem. Phys.* **1962**, *36*, 3401.

Figure 4. Correlation between partition coefficients of alkyl nitrites in sodium dodecyl sulfate + water and the number of carbon atoms in the side chain. The numbering of the solutes is the same that for Table II. Solutes not used in the correlation (▲).

$$\Delta G_{tr}^{w \rightarrow M} \text{ (cal/mol)} = -RT \ln K_S^N - RT \ln 55.5$$

and

$$\Delta G_{tr}^{w \rightarrow M} \text{ (cal/mol)} = N^\circ \Delta G_C + \Delta G_{mc} \qquad (5)$$
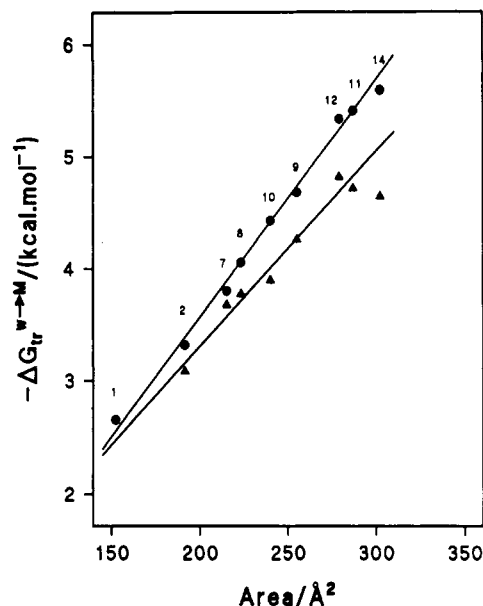
energy of transfer per methylene group in the case of TTABr micelles and SDS micelles, respectively.

These values are consistent with previously reported results on micelles,[31,33] liposomes,[14] microemulsions,[11] and on transfer free energies between monomers and their micelles.[30]

(b) Chain Branching Effect. Branched hydrocarbons typically[13] show partition coefficients lower than those of the analogous linear hydrocarbons, probably due to a reduction in van der Waals interactions in the lipid phase which are approximately proportional to the molecular surface area (branching in the chain results in less surface area than for the corresponding straight chain). Figure 5 corroborates this conclusion. The plot of free energy of transfer against molecular surface area yields a good straight line.[34]

Our data of ΔG_{tr} also correlate well (see Figure 6) with the transfer free energy of amino acids from water to the surfactant interfaces of water-in-oil microemulsions determined by Leodidis and Hatton,[11] and good correlation is also found with octanol–water[35] or hydrophobicity[36] scales.

(c) Halogen Atom and Ether Link Effects. The introduction of halogens generally increases permeability of a solute, an effect for which no satisfactory explanation has been put forward: In the absence of other substituents and with the exception of fluorine, it seems likely that a

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(34) Values of A_{area} refer to the R groups of alkyl nitrites (taken from Hermann R. B. *J. Phys. Chem.* (a) **1972**, *76*, 2754; (b) **1975**, *79*, 163). Two of them, isopropyl nitrite and 2-methyl-2-propyl nitrite, were estimated by comparison with analogous compounds.

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(36) Nozaki, Y.; Tanford, C. *J. Biol. Chem.* **1971**, *246*, 2211.

Figure 5. Free energy of transfer of alkyl nitrites from water to (●) TTABr interface and (▲) SDS interface as a function of the alkyl nitrites van der Waals molecular area.
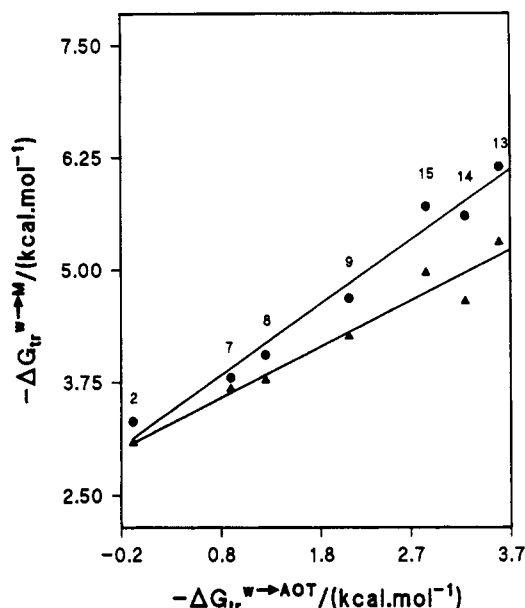
Figure 6. Correlation between free energy of transfer of alkyl nitrites from water to (●) TTABr micelles and (▲) SDS micelles and free energy of transfer of amino acids, with the same side chain, from water to AOT (Aerosol-OT, sodium bis(2-ethylhexyl) sulfosuccinate) interface. (Last data obtained from ref 11).

halogen is simply equivalent to a hydrocarbon residue of the same size in its direct effect on permeability.[11,12,37] This argument seems to hold in our case, where 2-chloroethyl nitrite and 2-bromoethyl nitrite have higher partition coefficient than ethyl nitrite and similar, at least in the case of bromo substituent, to the partitioning coefficient of *n*-propyl nitrite. On the other hand, we observed a considerable decrease in partition coefficients upon introduction of an ether linkage (compare data for 2-methoxy- and 2-ethoxyethyl nitrites with those for *n*-propyl and *n*-butyl nitrites), in agreement with the expectations based on the possibility of hydrogen bonding of the ether oxygen and a hydrogen of water.

(d) Aromatic Ring Effect. Comparison between permeabilities of aliphatic and aromatic compounds is not

(37) Roger, J. A.; Davis, S. S. *Biochim. Biophys. Acta* **1980**, *598*, 392.

a straightforward matter, since several opposing factors can play a role in determining the partitioning behavior. Hydrogen bonds between hydroxyl group and aromatic π electrons[37] will result in a reduced permeability of aromatic compounds. However, the existence of charge-transfer complexes between π electrons and TTABr head groups will favor micellar incorporation for aromatic compounds.