Changes in the Relative Contribution of Specific and General Base Catalysis in Cationic Micelles. The Cyclization of Substituted Ethyl Hydantoates

Iva B. Blagoeva,*,† Maria M. Toteva,† Nadia Ouarti,‡ and Marie-Françoise Ruasse*,‡

Institute of Organic Chemistry, Bulgarian Academy of Sciences, ul. Acad. G. Bonchev bl.9, Sofia 1113, Bulgaria, and Institut de Topologie et de Dynamique des Systèmes de l'Université Paris 7-Denis Diderot, associé au CNRS, UPRESA 7086, 1, rue Guy de la Brosse, 75005 Paris, France

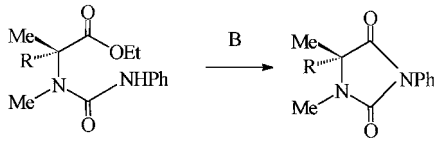
ruasse@paris7.jussieu.fr

Received December 19, 2000

The rate-surfactant profiles for the HO^-- and AcO^--catalyzed ring closure of two ethyl hydantoates, **E2** and **E3**, to hydantoins with three cetyltrimethylammonium salts (CTAX, X = Br^-, Cl^-, or AcO^-) are measured in 0.02 and 0.2 M acetate buffers 50% base with starting pH 4.65. Marked accelerations associated with large pH increases are found in 0.02 M buffered CTAOAc. Smaller accelerations and smaller pH changes are observed in 0.2 M buffered CTAOAc and CTACl. From these profiles, the micellar rate constants for the specific base- and general base-catalyzed reactions, $k_{2,m}^{HO^-}$ and $k_{2,m}^{AcO^-}$, respectively, of **E2** and **E3** are obtained separately. The resulting values of $k_{2,m}/k_w$, **E2**/**E3** rate constant ratios, and kinetic solvent isotope effects, KSIEs, are consistent with a strong predominance of the HO^- reaction in the dilute buffer, while in the more concentrated buffer, specific **and** general catalysis compete for the two substrates. This result is in sharp contrast with that observed in water in which the reaction of **E2** is almost exclusively specifically catalyzed. The increase in the general base-catalyzed pathway for **E2** is attributed not to an increase in the rate constant for this pathway in micelles but to a smaller decrease than that for the specific catalysis ($k_{2,m}/k_w$ = 0.2 and 0.4 for the specific and general catalysis, respectively). The different responses of the rate constants to the micellar media are interpreted as a larger effect of the interfacial polarity on the specific than on the general catalysis. The apparent contradiction between the rate constant decreases and the marked accelerations in micellar media is discussed in terms of pH changes, i.e., [HO^-] changes, and of acetate inclusion via ion exchanges at micellar interfaces.

Micellar systems provide models for understanding the ways by which changes in the microenvironment can affect rates of reactions of bioorganic interest.[1] Contrary to nucleophilic substitutions and specific acid−base-catalyzed reactions, general acid−base catalysis, which plays an important role in the accelerations of enzyme-catalyzed reactions,[2] has not been widely studied in micellar systems. This is due to numerous drawbacks of both conceptual and experimental origins.[1a,3,4] General catalysis has been most convincingly identified in cases of mixed micelles containing covalently bound catalyst or on the basis of structure−reactivity relationships and kinetic solvent isotope effects (KSIE).[5,6]

Scheme 1



E2 R=H

E3 R = Me

A suitable reaction for investigating general acid−base catalysis in micelles is the ring closure of sterically strained ethyl hydantoates to hydantoins (Scheme 1). This reaction has been extensively studied as a model reaction for the carboxylation of biotin.[7]

Some of the model compounds react at convenient rates by hydroxide and general base-catalyzed reactions at pH values below 7,[7d,e] which provides a unique opportunity to investigate catalysis by weak bases such as acetate ion in micellar media at low pH. We expected to maximize catalysis at the micellar interface by using a surfactant

* To whom correspondence should be addressed. (M.-F.R.) Tel.: +33 1 44 27 68 42. Fax: +33 1 44 27 68 14. (I.B.) E-mail: iva@orgchm.bas.bg.

† Bulgarian Academy of Sciences.

‡ Institut de Topologie et de Dynamique des Systèmes de l'Université Paris 7-Denis Diderot.

(1) (a) Bunton, C. A.; Savelli, G. *Adv. Phys. Org. Chem.* **1986**, *22*, 213. (b) Menger, F. M. *Angew. Chem., Int. Ed. Engl.* **1991**, *30*, 1086. (c) Tašcioğlu, S. *Tetrahedron* **1996**, *34*, 11113. (d) Dugas, H.; Penney, C. In *Bioorganic Chemistry*; Springer-Verlag: New York, Heidelberg, Berlin, 1981. (e) Perrin, C. L.; Chen, J.-H.; Ohta, B. K. *J. Am. Chem. Soc.* **1999**, *121*, 2448. (f) Bunton, C. A.; Yatsimirsky, A. K. *Langmuir* **2000**, *16*, 5921.

(2) Kirby, A. J. *Acc. Chem. Res.* **1997**, *30*, 290.

(3) Jencks, W. P. In *Catalysis in Chemistry and Enzymology*; McGraw-Hill: New York, 1969.

(4) Romsted, L. S.; Bunton, C. A.; Yao, J. *Curr. Opin. Colloid Interface Sci.* **1997**, *2*, 622.

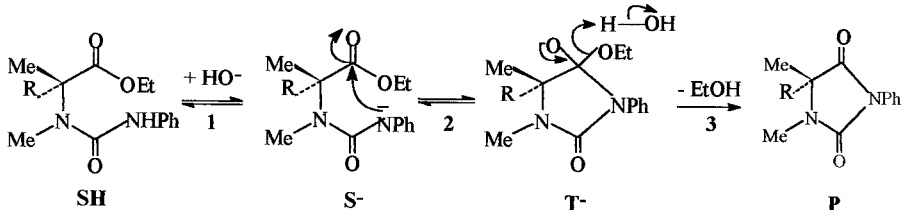
(5) Froehner, S. J.; Nome, F.; Zanette, D.; Bunton, C. A. *J. Chem. Soc., Perkin Trans. 2* **1996**, 673.

(6) Rav-Acha, Ch.; Ringel, I.; Sarel, S.; Katzhendler, J. *Tetrahedron* **1988**, *44*, 5879.

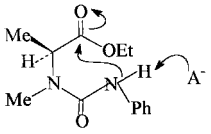
(7) (a) Blagoeva, I. B.; Pojarlieff, I. G.; Kirby, A. J. *J. Chem. Soc., Perkin Trans.* 2 **1984**, 745. (b) Knowles, J. R. *Annu. Rev. Biochem.* **1989**, *58*, 195. (c) Blagoeva, I. B. *J. Chem. Soc., Perkin Trans.* 2 **1987**, 127. (d) Blagoeva, I. B.; Tashev, D.; Kirby, A. J. *J. Chem. Soc., Perkin Trans. 1* **1989**, 1157. (e) Atay, E.; Blagoeva, I. B.; Kirby, A. J.; Pojarlieff, I. G. *J. Chem. Soc., Perkin Trans. 2* **1998**, 2289.

10.1021/jo001759w CCC: $20.00
Published on Web 02/20/2001

Scheme 2



Scheme 3



with the same reactive counterion as the studied general base. Cetyltrimethylammonium acetate is a good candidate for an efficient micellar general base catalysis on these substrates.

We report now results on the rates of ring closure of two ethyl hydantoates to hydantoins (**E2**, R = H and **E3**, R = Me) with three cetyltrimethylammonium salts (CTAX, X = Br^-, Cl^-, or AcO^-) in acetate buffers.

The model compounds **E2** and **E3** were chosen because their base-catalyzed reactions proceed by different mechanisms in water. The extra methyl group of **E3** changes the mechanism of the HO^--catalyzed reaction (Scheme 2) from specific base-catalyzed formation of the tetrahedral intermediate T^- (step 2 for **E2**) to its general acid-catalyzed breakdown by water (step 3 for **E3**) by slowing the proton transfer involved in the decomposition of T^-.[7e] General catalysis by buffer bases, A^-, also goes via different mechanisms for **E2** and **E3**.[8] The acetate-catalyzed reaction of **E2** ($A^- = AcO^-$, Scheme 3) does not involve any prior proton removal from SH but the direct, rate-determining formation of T^-. In other terms, the **E2** reaction with A^- consists of a true general base catalysis. In contrast, the general base-catalyzed reaction of **E3** (Scheme 4) goes through the kinetically equivalent specific base and general acid catalysis, with rapid deprotonation of SH and rate-limiting decomposition of T^- catalyzed by AH (acetic acid).

In water, the pseudo-first-order rate constants, k_w', are described by eq 1 in which all rate constants (k with superscripts referring to the several catalysts) have been measured previously[7e,9] (see footnotes of Table 1):

$$k_w' = k^{H_2O} + k_w^{HO^-}[HO^-] + k_w^{AcO^-}[AcO^-] + k_w^{AcOH}[AcOH] \quad (1)$$

The two substrates have the particularity of different reactivity ratios **E2**/**E3** (k_{E2}/k_{E3}) for the hydroxide- and acetate-catalyzed reactions.[7a,e] At pH values above 4 where these two catalysts compete and where the H_2O and AcOH reactions can be neglected, **E2** is more reactive than **E3**, with $k_w^{HO^-}$ for **E2** 4.4 times larger than $k_w^{HO^-}$ for **E3**, while the second-order rate constant $k_w^{AcO^-}$ is two times smaller for **E2** than for **E3**. Significant kinetic effect of CTAX micelles on these two reactions are expected because of favorable electrostatic interactions between the two anionic catalysts and the cationic surfactant headgroups. In this paper, we show that the relative contribution of the specific and general catalysis is different in micellar systems because of a different sensitivity of their rate constants to the change in the polarity of the reaction medium.

Experimental Section

Materials. Inorganic reagents and buffer components were of analytical grade and used without further purification. Buffer solutions were prepared with CO_2-free distilled water. D_2O, 99.99 atom % D, was from Aldrich. Cetyltrimethylammonium chloride, CTACl, 25 wt % solution in water, was from Aldrich and used without further purification. Cetyltrimethylammonium bromide, CTABr, from Merck was used after recrystallization from ethanol. Cetyltrimethylammonium acetate, CTAOAc, was prepared according to Sepulveda et al.[10] and purified by several recrystallizations from methanol/ether 1:20. The purity of the product was checked by titration with HCl of the acetate ions in 90:10 (v/v) ethanol/water. Purity was ca. 98%. The preparation of ethyl hydantoates **E2** and **E3** was described previously.[11]

Kinetic Measurements. Rate constants were determined at 25.0 ± 0.01 °C under pseudo-first-order conditions in the thermostated cell compartment of a Unicam SP-800 or Carl Zeiss Jena spectrophotometer. Reactions were initiated by injecting 20−50 μl of 1×10^{-2} M stock solution of the substrate in THF into 2.7 mL of preheated buffer solution containing the surfactant. The rates for cyclization of the substituted esters were followed at 238 nm by monitoring the decrease of the absorbance due to the phenylureido group. Pseudo-first-order rate constants, k_{obs}, were obtained by nonlinear regression fitting to the equation $A_t = A_0\, e^{-k_{obs}t} + A_\infty$ where A_t, A_0 and A_∞ are the absorbances at time t, zero, and infinity, respectively; k_{obs} values were found to be reproducible within 5%.

pH Measurements. pH values were measured directly after each kinetic run using a Radiometer pH M 84 research pH meter, with a GK 2401 C electrode standardized at pH 6.87 and 4.01.[7c] The electrode standardization was checked after each series of runs and found not affected by the surfactant solutions studied. Experiments in D_2O and H_2O solutions were run simultaneously in the multicell compartment of the spectrophotometer. pD values were obtained by adding 0.4 to the pH meter readings. To convert pH to $[OH^-]$ and pD to $[OD^-]$, pK_w of 14.0 and of 14.86 were taken, respectively.[12]

Results and Discussion

The extent of micellar catalysis was determined from experimental rate-surfactant profiles (Figures 1 and 2)

(8) Blagoeva, I. B.; Kirby, A. J.; Koedjikov, A. H.; Pojarlieff, I. G. *Can. J. Chem.* **1999**, *77*, 849.

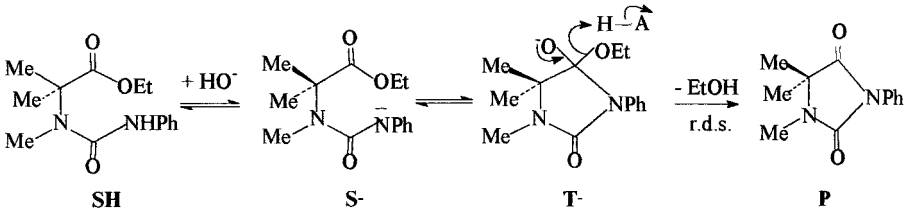
(9) Catalysis by acetic acid could not be detected for **E2**, and catalysis by acetate ions is weaker for this substrate (ref 7e). In any case, as a result of the favorable electrostatic effect, we expect only catalysis by anionic species to be significant at the interface of cationic micelles.

(10) Sepulveda, L.; Gabrera, W.; Gamboa, C.; Meyer, M. *J. Colloid Interface Sci.* **1987**, *117*, 460.

(11) Pojarlieff, I. G.; Blagoeva, I. B.; Kirby, A. J.; Mikhova, B. P.; Atay, E. *J. Chem. Res., Synop.* **1997**, 220.

(12) (a) Covington, A. K.; Robinson, R. A.; Bates, R. G. *J. Phys. Chem.* **1966**, *70*, 3820. (b) Fishbein, J. C.; Jencks, W. P. *J. Am. Chem. Soc.* **1988**, *110*, 5075.

Scheme 4



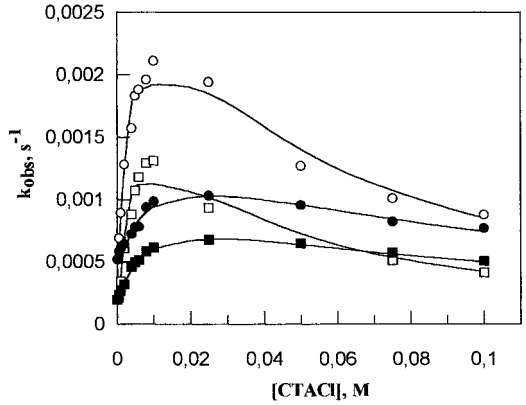


Figure 1. Dependence of the observed rate constants for ring closure of **E2** and **E3** at 25 °C in micelles of CTACl in acetate buffers on the concentration of surfactant. ○, **E2** in 0.02 M buffer; □, **E3** in 0 02 buffer; ●, **E2 i**n 0.2 M buffer; ■, **E3** in 0.2 M buffer. The curves are calculated with the rate constants from Tables 1 and 2 (see text).

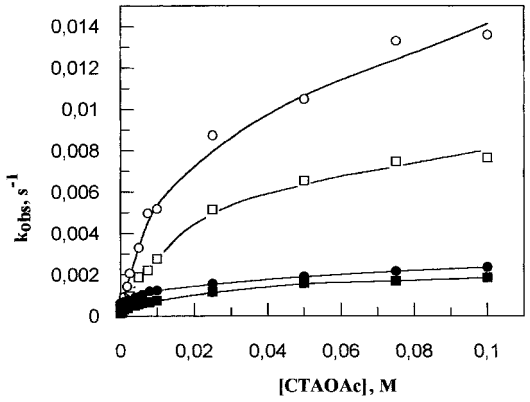
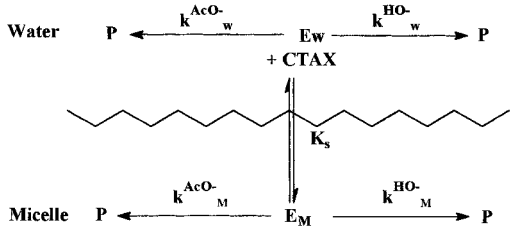


Figure 2. Dependence of the observed rate constants for ring closure of **E2** and **E3** at 25 °C in micelles of CTAOAc in acetate buffers on the concentration of surfactant. ○, **E2** in 0.02 M buffer; □, **E3** in 0.02 M buffer; ●, **E2** in 0.2 M buffer; ■, **E3** in 0.2 M buffer. The curves are calculated with the rate constants from Tables 1 and 2 (see text).

with cetyltrimethylammonium salts in the 0−0.1 M range. The standard aqueous medium for the rate-surfactant profiles is 0.02 M acetate buffer 50% base with pH 4.65 at [CTAX] = 0. The effect of increasing the concentration of the buffer to 0.2 M is also measured,[13] as well as the kinetic solvent isotope effects at both buffer concentrations.

Scheme 5



The ring closure of **E2** and **E3** is first order in [HO⁻] in the investigated pH region.[7d] In acetate buffers, as shown in Scheme 5, hydroxide- and acetate-catalyzed reactions are expected to take place in both micellar and aqueous pseudophases.

In the pseudophase model[1a] the observed rate constant, k_{obs}, is described by eq 2, in which k are second-order rate

$$k_{obs} = \frac{k_w^{HO^-}[HO^-]_w + k_w^{AcO^-}[AcO^-]_w + \{k_M^{HO^-}[HO^-]_M + k_M^{AcO^-}[AcO^-]_M\}K_S}{1 + K_S([D]_T - cmc)} \quad (2)$$

constants, with concentrations in square brackets referring to molarity in the total solution volume. The subscripts w and M refer to water and micellar reactions, respectively, and the superscripts HO^- and AcO^- to the catalysts. K_S is the binding constant of the substrate, $[D]_T$ is the surfactant concentration, and cmc is the critical micelle concentration. In eq 2, k_w are in units of M^{-1} s^{-1}, and k_M are in s^{-1} units. To obtain the second-order rate constants, $k_{2,m}$, for the reaction in the micellar phase, k_M is multiplied by the micellar volume, V_m ($k_{2,m} = k_M V_m$).[1a]

In the usual pseudophase ion exchange (PIE) model,[14] the ion concentrations in both phases are obtained from the equilibrium constants for exchange of anions Y^- from the aqueous phase for surfactant counterions X^-:

$$K_{ex} = K_{X^-}^{Y^-} = \frac{[Y^-]_M[X^-]_w}{[Y^-]_w[X^-]_M} \quad (3)$$

In the case of parallel base-catalyzed reactions, $[HO^-]_M$ and $[AcO^-]_M$ cannot be determined separately since these two concentrations are interdependent. Therefore, $[AcO^-]_M$ is first obtained from the pH dependence on the surfactant concentration (vide infra). $[HO^-]_M$ is then calculated from the previously measured $K_{AcO^-}^{HO^-}$ exchange constant. Consequently, the pH changes of the buffer solutions provoked by the surfactant addition are analyzed to obtain the micellar catalyst concentrations *at every*

(13) In accordance with the procedure usually adopted in work on micellar reactivity, the ionic strength of the water phase is not controlled by addition of any salt to avoid additional ion exchanges. The underlying assumptions of this procedure (negligible changes in pH, pK, and k in the water pseudophase as compared to aqueous buffer solutions) have been widely discussed and its reliability fairly well established (ref 1f and references therein).

(14) (a) Bunton, C. A.; Nome, F.; Quina, F. H.; Romsted, L. S. *Acc. Chem. Res.* **1991**, *24*, 357. (b) Romsted, L. S. *Surfactants in Solution*; Mittal, K. L., Lindman, B. Eds; Plenum Press: New York, 1984; Vol. 1, p 1015.

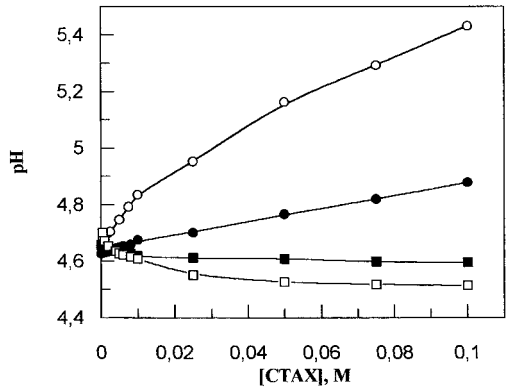


Figure 3. Dependence of pH values of acetate buffers 50% base on the concentration of CTAOAc and CTACl. Open symbols, 0.02 M buffer; closed symbols, 0.2 M buffer; circles, CTAOAc; squares, CTACl.

surfactant concentration, before applying eq 2 to the rate data for calculating k_M.

pH Changes with Surfactant Concentration. When the surfactant counterion is identical to that of the buffer base (CTAOAc in acetate buffer), the $[AcO^-]_w/[AcOH]_w$ buffer ratio changes as a result of the dissociation of the surfactant.[1a,14] The pH changes of acetate buffers with increasing surfactant concentrations are shown in Figure 3. In the case of 0.02 M buffer, the pH increases up to 0.8 pH units with the addition of 0.1 M surfactant, while in the more concentrated (0.2 M) buffer, where the amount of acetate ions released by the surfactant is less significant as compared to the buffer acetate ions, the pH changes only by 0.3 pH units.

In CTACl and CTABr the pH values decrease with addition of surfactant as a result of opposite changes in the buffer ratio.[15,16] $[AcO^-]_w$ decreases since chloride and bromide surfactant counterions exchange for acetate ions from the buffer. The observed decreases of pH are in these cases not larger than 0.2 pH units.

α, Degree of Micelle Dissociation, in CTAOAc Buffered Solutions. These pH data (Figure 3) allow for the calculation of the degree of CTAOAc dissociation. Micelle dissociation is characterized by the parameter α, while β represents the fraction of counterions bound to the micelle[1a]:

$$\alpha = (1 - \beta) = \frac{[AcO^-]_w}{([D]_T - cmc)} \qquad (4)$$

When it is assumed that the pK_a of acetic acid[13,17] and the concentration of acetic acid in the water phase of the diluted buffer remain unchanged (eq 5 with $[AcOH]_w = [AcOH]_0$, index 0 indicating initial concentrations of buffer species), α is calculated from the pH change via eq 6, pH_0 being the value in the absence of surfactant.

$$pH = pH_0 + \log[AcO^-]_w - \log[AcOH]_w \qquad (5)$$

$$pH = pH_0 + \log\{[AcO^-]_0 + cmc + \alpha([D]_T - cmc)\} - \log[AcOH]_0 \qquad (6)$$

We obtain $\alpha = 0.48 \pm 0.01$ and cmc = (0.0004 ± 0.0002) M using least-squares nonlinear regression fit of the data for 0.02 M buffer. The corresponding β value of 0.52 is in the range of published values for CTAOAc.[18] The cmc values of CTAX (X = AcO^-, Cl^-, and Br^-) in pure water[18] are in the range of $1-1.5 \times 10^{-3}$ M, but the cmc values of CTACl in the presence of 0.1 and 0.01 M buffers are in the range of 2×10^{-4} to 6×10^{-4} M, whatever the buffer concentration.[19]

If the α value of 0.48, obtained in 0.02 M buffer, is applied to the pH data in 0.2 M buffer, the calculated pH of a 0.1 M solution of CTAOAc is only 0.17 pH units higher than that of the pure aqueous buffer. The experimental increase is, however, 0.30 pH units and would correspond to $\alpha = 0.8$, which value is highly unlikely. Most likely, this additional pH increase arises, in contrast to an earlier assumption,[20] from an intake of acetic acid in the micelles from the buffer, since the concentration of the acid in the concentrated buffer is very large (0.1 M).

The association constant, K_S^{AcOH}, for incorporation of acetic acid in cetyltrimethylammonium micelles can be estimated as 1.7 M^{-1} from the linear solvation free energy relationship proposed by Quina et al.[21] With this value we calculate that up to 14.5% of the acetic acid of the buffer is incorporated at the maximum concentration of CTAOAc ($[AcOH]_M = 1.45 \times 10^{-2}$ M for [CTAOAc] = 0.1 M). Accordingly, eq 6 is transformed in eq 7 in which the

$$pH = pH_0 + \log\{[AcO^-]_0 + cmc + \alpha([D]_T - cmc)\} - \log\left\{\frac{[AcOH]_0}{1 + K_S^{AcOH}([D]_T - cmc)}\right\} \qquad (7)$$

concentration term for the acetic acid in the water phase is calculated by means of K_S^{AcOH}. The value of α in 0.2 M acetate buffer from eq 7 is now 0.48 ± 0.02, in agreement with that found from 0.02 M buffer[22] in which the micellar inclusion of acetic acid can be neglected ($[AcOH]_M = 1.5 \times 10^{-3}$ M only for $[CTAOAc] = 10^{-1}$ M).

Rate-Surfactant Profiles. The dependences of the rate constants for ring closure of **E3** and **E2** in 0.02 and 0.2 M acetate buffers on the concentration of CTACl are shown in Figure 1. The maxima in the rate profiles are consistent with second-order micelle-mediated ion−molecule reactions.[1a] The experimental first-order rate constants, k_{obs}, increase with increasing [CTACl] and then decrease. The largest accelerations in 0.02 M buffer for **E3** and **E2** (expressed as k_{max}/k_w) are 10 and 4, respectively. In 0.2 M buffer, the rate constants are surprisingly smaller (vide infra), but the decrease at high surfactant concentrations due to dilution of the reagents in the micellar phase is very small.

(15) Ouarti, N.; Marques, A.; Blagoeva, I.; Ruasse, M.-F. *Langmuir* ***2000***, *16*, 2157.

(16) (a) Quina, F. H.; Politi, M. J.; Cuccovia, I. M.; Baumgarten, E.; Martins-Franchetti, S. M.; Chaimovich, H. *J. Phys. Chem.* **1980**, *84*, 361. (b) Romsted, L. S. *J. Phys. Chem.* **1985**, *89*, 5107, 5113.

(17) pK_a in the micellar phase is about the same as that in bulk water. (El Seoud, O. A. *Adv. Colloid Interface Sci.* **1989**, *30*, 1. Berthold, A.; Saliba, C. *Analusis* **1985**, *13*, 437.)

(18) Gallion, L.; Hamidi, M.; Lelievre, J.; Gaboriaud, R. *J. Chim. Phys.* **1997**, *94*, 707.

(19) Unpublished measurements by tensiometry and spectrofluorometry.

(20) Toullec, J.; Couderc, S. *Langmuir* **1997**, *13*, 1918.

(21) Quina, F. H.; Alonso, E. O.; Farah, J. P. S. *J. Phys. Chem.* **1995**, *99*, 11708.

(22) Soldi, V.; Keiper, J.; Romsted, L. S.; Cuccovia, I. M.; Chaimovich, H. *Langmuir* **2000**, *16*, 59.

Table 1. Equilibrium and Rate Constants for Ring Closure of E2 and E3 at 25 °C in Micelles of CTAX in 0.02 M Acetate Buffers 50% Base[a]

substrate	surfactant	K_s, M^{-1} [f]	$k \times {}^{HO^-}_M\ 10^{-5}$, s^{-1} [f]	k_{2m}/k_w [g]	KSIE	**E2**/**E3** [h]	$k_M^{AcO^-}$, s^{-1} [i]	k_{2m}/k_w	**E2**/**E3** [j]
E2[b]	CTAOAc	50 ± 6	13.7 ± 0.5	0.2		1.9	0.0304	12	1.7
E2[b]	CTACl	47 ± 5	9.71 ± 0.67	0.14		2.1	0.0248	10	2.1
E2[b]	CTABr	55 ± 7	11.2 ± 1.0	0.16		1.2	0.0274	11	1.4
E2[c]	CTAOAc/D_2O	57 ± 7	17.8 ± 0.7	0.17	0.78 ± 0.06				
E3[d]	CTAOAc	58 ± 5	7.06 ± 0.19	0.45			0.0181	3.7	
E3[d]	CTACl	74 ± 14	4.68 ± 0.56	0.30			0.0118	2.4	
E3[d]	CTABr	39 ± 4	9.17 ± 0.6	0.59			0.0198	4.1	
E3[e]	CTAOAc/D_2O	36 ± 6	4.58 ± 0.3	0.53	1.5 ± 0.1				

[a] Calculated from eq 2 (see text), with $K_{AcO^-}^{HO^-} = 0.49$, $K_{Cl^-}^{HO^-} = 0.25$, $K_{Br^-}^{HO^-} = 0.048$, $K_{Cl^-}^{AcO^-} = 0.50$, $K_{Br^-}^{AcO^-} = 0.098$ from ref 10, cmc = 4×10^{-4} M, $\beta_{CTAOAc} = 0.52$, $\beta_{CTACl} = 0.6$, $\beta_{CTABr} = 0.7$. [b] $k_w^{HO^-} = 9.51 \times 10^5\ M^{-1}\ s^{-1}$; $k_w^{AcO^-} = 3.53 \times 10^{-4}\ M^{-1}\ s^{-1}$; $k_{H_2O} = 1.43 \times 10^{-5}\ s^{-1}$; from ref 7e. [c] $k_w^{DO^-} = 1.42 \times 10^6\ M^{-1}\ s^{-1}$; $k_w^{AcO^-} = 1.18 \times 10^{-4}\ M^{-1}\ s^{-1}$; $k_{D_2O} = 7 \times 10^{-6}\ s^{-1}$; from ref 7e. [d] $k_w^{HO^-} = 2.17 \times 10^5\ M^{-1}\ s^{-1}$; $k_w^{AcO^-} = 6.77 \times 10^{-4}\ M^{-1}\ s^{-1}$; $k_{H_2O} = 1.68 \times 10^{-5}\ s^{-1}$; $k_w^{AcOH} = 2.27 \times 10^{-4}\ M^{-1}\ s^{-1}$; from ref 7e. [e] $k_w^{DO^-} = 1.21 \times 10^5\ M^{-1}\ s^{-1}$; $k_w^{AcO^-} = 3.3 \times 10^{-3}\ M^{-1}\ s^{-1}$; $k_{D_2O} = 8.5 \times 10^{-6}\ s^{-1}$; $k_w^{AcOH} = 1.1 \times 10^{-4}\ M^{-1}\ s^{-1}$ from ref 7e. [f] From eq 2 with the term $k_M^{AcO^-}$ omitted. [g] With $V_m = 0.14\ M^{-1}$, ref 1a. [h] In water **E2**/**E3** = 4.4. [i] From eq 2 with the term $k_M^{HO^-}$ omitted. [j] In water **E2**/**E3** = 0.52.

The rate profiles in CTABr for both substrates (Figure S1, Supporting Information) parallel those in CTACl. The reaction in CTABr micelles is not markedely slower than that in CTACl micelles, the accelerations, k_{max}/k_w, being about 6 and 3 at small and large buffer concentrations, respectively.

Figure 2 shows the rate-surfactant profiles for ring closure of **E3** and **E2** in CTAOAc. As opposed to the profiles in CTACl, the reaction in CTAOAc micelles exhibits no maximum but a continuous increase in k_{obs} with [CTAOAc]. The largest accelerations, k_{max}/k_w, in 0.02 M buffers are now very large, 60 and 25 for **E3** and **E2**, respectively. As in the case of the two other surfactants, the accelerations are significantly larger for **E3** than for **E2**. This is in accordance with the expectations for micellar catalysis by acetate ions, since general base catalysis by acetate ions in water is more important for ester **E3**. However, these accelerations in CTAOAc micelles cannot be straightforwardly attributed to an increase in the total concentration of micellized acetate ions, since the associated increases in pH values indicate large increases in micellized hydroxide ions also. The rate-surfactant profiles in the more concentrated 0.2 M acetate buffer show again a decrease in the reaction rates, as compared to the less concentrated buffer. However, the reaction takes place at much lower pH values, and this effect probably overcomes the effect of increasing concentrations of micellized acetate ions.

Micellar Rate Constants of the Hydroxide-Catalyzed Reaction. Table 1 lists the rate and association constants calculated (eq 2) by nonlinear fits of the rate-surfactant profiles in CTAX micellar solutions **in 0.02 M acetate buffer**. The concentrations of HO^- and AcO^- in both phases of CTAOAc solutions are calculated by using eqs 8−10 with β and cmc values obtained from eq 6. A value of 0.49 is used[10] for $K_{AcO^-}^{HO^-}$.

$$[HO^-]_w = 10^{(pH-14)} \quad (8)$$

$$[AcO^-]_M = \beta([D_T] - cmc) = [AcO^-]_{tot} - [AcO^-]_w \quad (9)$$

$$[HO^-]_M = K_{AcO^-}^{OH^-}\frac{[AcO^-]_M[HO^-]_w}{[AcO]_w} \quad (10)$$

The calculations for CTACl and CTABr need, in addition, to take into account the exchanges of surfactant counterions for AcO^-. The concentration of acetate ions in water, $[AcO^-]_w$, at each surfactant concentration, is obtained from the quadratic eq 11, which describes the

$$[AcO^-]_w^2(K_{ex} - 1) + [AcO^-]_w\{[AcO^-]_0(2 - K_{ex}) + \beta([D]_T - cmc)(K_{ex} - 1)\} - [AcO^-]_0\{[AcO^-]_0 - \beta([D]_T - cmc)\} = 0 \quad (11)$$

change in $[AcO^-]_w$ with the surfactant concentration and its dependence on K_{ex}. $[AcO^-]_0$ is the initial concentration of acetate ions in the buffer, and K_{ex} is the constant for the exchange of Cl^- or Br^- for acetate ions. β and K_{ex} are optimized until the sum of the differences between experimental and calculated values of $[AcO^-]_w$ obtained from the pH data and eq 11, respectively, reaches a minimum (footnotes in Table 1).

The concentration of micellized hydroxide ions, $[HO^-]_M$, at pH 4−5 is orders of magnitude lower than the concentrations of the surfactant counterions (X = Cl^- or Br^- and AcO^-), and it is estimated in the mixed micelles by using $K_{X^-}^{HO^-}$ and $K_{AcO^-}^{HO^-}$ (eq 10).

Because of the $[HO^-]_m/[AcO^-]_m$ interdependence mentioned before, it is not possible to obtain at the same time K_s and the two micellar rate constants, $k_M^{HO^-}$ and $k_M^{AcO^-}$, by fitting the experimental data to the total eq 2. Almost equally good fits of the data are obtained when one or the other micellar rate term is omitted.

The K_s values (Table 1) obtained from either of these calculations vary between about 40 and 70 M^{-1}. The large uncertainties of their calculation preclude any quantitative discussion of the respective values for **E2** and **E3**. These values correspond to a weak binding of the relatively polar ureido esters. As both substrates differ only by one methyl group, the similarity of these K_s values is not surprising.

As regards the rate constants, $k_M^{HO^-}$ and/or $k_M^{AcO^-}$, the predominant catalysis is determined by comparing the values, for either pathway, of the $k_{2,m}/k_w$ ratios and of the relative reactivities of **E2** and **E3** (**E2**/**E3** ratios in Table 1). On one hand, $k_{2,m}/k_w$ values are expected to be smaller than unity as for most ion−molecule reactions and, in particular, hydroxide-catalyzed reactions in micelles.[4] On the other hand, the hydroxide-catalyzed reaction should be faster for **E2** compared to **E3** (**E2**/**E3** > 1), while the acetate catalysis is expected to be stronger for **E3** (**E2**/**E3** < 1)[7d] in micelles as in water.

When the data for **E2** are calculated as if the only reaction at the micellar interface was catalysis by hy-

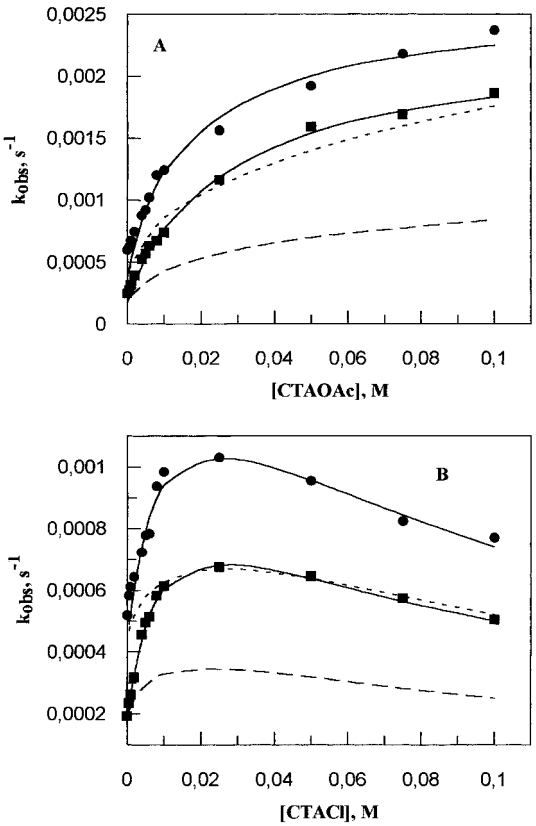


Figure 4. Rate constants for ring closure of **E2** and **E3** in 0.2 M acetate buffers at 25 °C in micelles of (A) CTAOAc and (B) CTACl. ●, **E2**; ■, **E3**. The full lines are calculated by complete eq 2 with the rate constants $k_M^{AcO^-}$ from Table 2. The dashed lines (- - - - - - for **E2** and — — — — for **E3**) represent the rate surfactant profiles calculated with the respective rate constants for $k_M^{HO^-}$ from Table 1 (see text).

droxide ions, a "normal" $k_{2,m}/k_w$ ratio of 0.2 is obtained, whatever the surfactant counterion. When the same data are analyzed as if only catalysis by acetate ions was present, unreasonably large values for catalysis by acetate, $k_{2,m}/k_w \sim 10$, are obtained. The calculations for **E3**, the substrate with more pronounced acetate catalyzed reaction in water, give similar results. The $k_{2,m}/k_w$ ratios for $k_M^{HO^-}$ are higher than the those calculated for **E2** but still less than unity, whereas the ratios $k_{2,m}/k_w$, for $k_M^{AcO^-}$, of ~3 are larger than unity as in the case of **E2**. Therefore, the hydroxide-catalyzed pathway is also predominant for **E3**, although some acetate catalysis cannot be totally excluded.

The micellar rate constant ratios, **E2**/**E3**, are larger than unity with both assumptions ($k_M^{HO^-}$ and $k_M^{AcO^-}$ in Table 1). They are consistent only with a predominant HO^- pathway since they agree with what is expected for the hydroxide-catalyzed reaction but not for the acetate catalysis. This result can be readily understood when the micellization of the two catalysts in the 0.02 M buffer is considered. In CTACl, the concentrations of micellized acetate ions are too small to overwhelm catalysis by HO^-. In CTAOAc, the large changes in the pH values make the variation in the rates of the HO^--catalyzed reaction predominant, masking the contribution of AcO^-.

Finally, the kinetic solvent isotope effects (KSIE), $k_w^{HO^-}/k_w^{DO^-}$, measured in water[7e] (0.67 for **E2**, typical of a specific base-catalyzed reaction and 1.79 of the general base-catalyzed HO^- reaction of **E3**) were decisive in assigning the reaction mechanisms for these substrates. The KSIEs ($k_M^{HO^-}/k_M^{DO^-}$) measured in micellar solutions of 0.78 ± 0.06 for **E2** and of 1.5 ± 0.1 for **E3** (data from Table 1) are very similar to the values in water and support the predominance of the hydroxide-catalyzed reaction of the two substrates in 0.02 M acetate buffers.

In conclusion, only the calculations related to the $k_M^{HO^-}$ pathway (Table 1) are chemically meaningful because they are consistent with the KSIEs, $k_{2,m}/k_w$, and **E2**/**E3** ratios, whereas those with the $k_M^{AcO^-}$ assumption are not.

Table 2. Equilibrium and Rate Constants for Ring Closure of E2 and E3 at 25 °C in Micelles of CTAX in 0.2 M Acetate Buffers 50% Base[a]

substrate	surfactant	K_s, M^{-1}	$k_M^{AcO^-} \times 10^3$, s^{-1} [f]	k_{2m}/k_w [g]	**E2**/**E3** [h]
E2[b]	CTAOAc	90 ± 16	1.03 ± 0.18	0.44	0.4
E2[b]	CTACl	76 ± 11	1.13 ± 0.16	0.45	0.9
E2[c]	CTAOAc/D_2O	96 ± 27	0.716 ± 0.153		0.8
E3[d]	CTAOAc	47 ± 5	2.38 ± 0.13	0.49	
E3[d]	CTACl	83 ± 4	1.22 ± 0.04	0.25	
E2[e]	CTAOAc/D_2O	64 ± 4	0.936 ± 0.025		

[a] Calculated from eq 2 (see text), with $K_{AcO^-}^{HO^-} = 0.49$, $K_{Cl^-}^{HO^-} = 0.25$, $K_{Br^-}^{HO^-} = 0.048$, $K_{Cl^-}^{AcO^-} = 0.50$, $K_{Br^-}^{AcO^-} = 0.098$ from ref 10, cmc = 4×10^{-4} M, $\beta_{CTAOAc} = 0.52$, $\beta_{CTACl} = 0.6$, $\beta_{CTABr} = 0.7$. [b] $k_w^{HO^-} = 9.51 \times 10^5$ M^{-1} s^{-1}; $k_w^{AcO^-} = 3.53 \times 10^{-4}$ M^{-1} s^{-1}; $k_{H_2O} = 1.43 \times 10^{-5}$ s^{-1}; from ref 7e. [c] $k_w^{DO^-} = 1.42 \times 10^6$ M^{-1} s^{-1}; $k_w^{AcO^-} = 1.18 \times 10^{-4}$ M^{-1} s^{-1}; $k_{D_2O} = 7 \times 10^{-6}$ s^{-1}; from ref 7e. [d] $k_w^{HO^-} = 2.17 \times 10^5$ M^{-1} s^{-1}; $k_w^{AcO^-} = 6.77 \times 10^{-4}$ M^{-1} s^{-1}; $k_{H_2O} = 1.68 \times 10^{-5}$ s^{-1}; $k_w^{AcOH} = 2.27 \times 10^{-4}$ M^{-1} s^{-1}; from ref 7e. [e] $k_w^{DO^-} = 1.21 \times 10^5$ M^{-1} s^{-1}; $k_w^{AcO^-} = 3.3 \times 10^{-3}$ M^{-1} s^{-1}; $k_{D_2O} = 8.5 \times 10^{-6}$ s^{-1}; $k_w^{AcOH} = 1.1 \times 10^{-4}$ M^{-1} s^{-1} from ref 7e. [f] From eq 2 with the term $k_M^{HO^-}$ from Table 1. [g] With $V_m = 0.14$ M^{-1}, ref 1a. [h] In water **E2**/**E3** = 0.52.

Micellar Rate Constants of the Acetate-Catalyzed Reaction. The pH values of the **concentrated (0.2 M) acetate buffer** in micellar solutions vary much less than those of the dilute buffer (Figure 3). In CTACl, the pH is almost constant and in CTAOAc, the pH increase is not larger than 0.3 pH units. Analogously, the micellar accelerations are markedly smaller (Figures 1 and 2). The contributions of the hydroxide reaction to the overall rates are estimated using the pH values in the concentrated buffer and the rate constants $k_M^{HO^-}$ obtained in the dilute buffer solutions (Table 1). The calculated rates for the hydroxide-catalyzed reaction are maximum evaluations since a $k_M^{HO^-}$ decrease can result from AcOH incorporation decreasing the interface polarity (vide infra) and since the minor contribution of acetate catalysis in the dilute buffer was neglected.

Figure 4 (dashed lines) shows that the contributions from a HO^--catalyzed reaction are insufficient to account for reaction rates in the more concentrated buffer. The differences between the calculated and experimental rate constants are attributed to the acetate catalysis. Table 2 lists the rate constants for acetate catalysis calculated from complete eq 2 with the $k_M^{HO^-}$ values of Table 1.[23]

(23) The data for CTABr in 0.2 M buffers did not give reliable results for acetate catalysis. This is probably due to the much less favorable exchange of the reactive anions from the solution for Br^- compared to Cl^-, the respective $K_{X^-}^{Y^-}$ being 0.048 and 0.48 for HO^- and 0.098 and 0.50 for AcO^- (ref 10).

Table 3. Ratios of Second-Order Rate Constants of Acetate- and Hydroxide-Catalyzed Reactions for Ring Closure of E2 and E3 in Micellar Media and in Water

substrate	medium	$k^{AcO^-}/k^{HO^-} \times 10^9$
E2	CTAOAc	0.75[a]
E2	CTACl	1.16[a]
E2	**H_2O**	**0.37**[b]
E3	CTAOAc	3.4[a]
E3	CTACl	2.6[a]
E3	**H_2O**	**3.1**[b]

[a] Calculated from data in Tables 1 and 2. [b] Calculated from data in ref 7e.

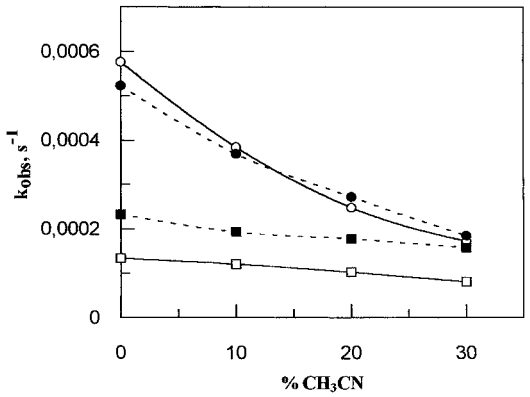


Figure 5. Effect of added acetonitrile (vol %) to acetate buffers on the observed rate constants for ring closure of **E2** and **E3** at 25 °C without addition of surfactant. ○, **E2** in 0.02 M buffer; □, **E3** in 0.02 M buffer; ●, **E2** in 0.2 M buffer; ■, **E3** in 0.2 M buffer.

The calculated $k_M^{AcO^-}$ values provide $k_{2,m}/k_w$ ratios smaller than unity, in the range expected for reactions taking place in the less polar environment of the micellar surface. The **E2**/**E3** ratios are also smaller than unity, in agreement with an acetate-catalyzed reaction for **E3** and for **E2**, as found in water. The KSIEs for both substrates (1.4 for **E2** and 2.5 for **E3**) are consistent with a slow proton transfer in the rate-determining step.

All of these results are strong evidence for a large acetate-catalyzed pathway at the micellar interface **for both E2** and **E3**, resulting from significant changes in the relative rate constants of the hydroxide and acetate reactions going from water to micelles. Table 3 shows that the ratios of the rate constants for the two base-catalyzed reactions of **E2** in micelles is larger than that measured in water, while for **E3** the ratio in water is maintained in the micellar media. The evidence for a significant acetate catalysis of the **E2** reaction is an important conclusion since in pure water this general catalysis is almost negligible[7d] and, therefore, difficult to be measured. The effect of the medium polarity on the rate constants of the reactions of the two substrates deserves, therefore, to be considered.

Medium Effects. To understand the origins of the micellar effects on these reactions, the sensitivity of the model reaction to the medium polarity needs to be known. Figure 5 shows the effect of increasing the concentration of acetonitrile in water on the rates of the ring closure reactions **in the absence of surfactant**. The difference in the solvent effects on **E2** and **E3** reactions is additional evidence for the different mechanisms by which both substrates react. For **E3** the solvent effect is small in both 0.02 and 0.2 M acetate buffers, the reactivity increase

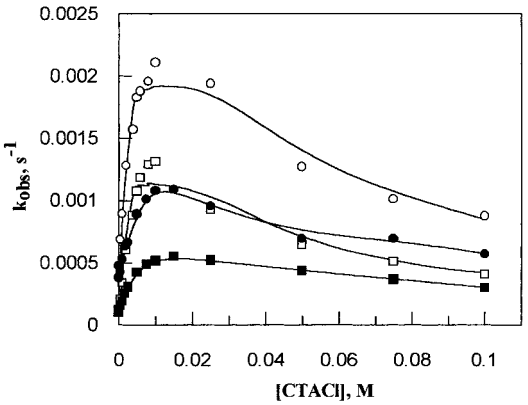


Figure 6. Effect of added 10 vol % ethanol on the observed rate constants for ring closure of **E2** and **E3** at 25 °C in 0.02 M buffer with increasing concentrations of CTACl. ○, **E2** in the absence of EtOH; □, **E3** in the absence of EtOH; ●, **E2** with added EtOH; ■, **E3** with added EtOH.

with the buffer concentration resulting primarily from the increased buffer catalysis for this substrate. The low sensitivity to the medium is consistent with a general base-catalyzed hydroxide reaction involving significant charge delocalization in the transition state.[2] By contrast, the marked rate decrease for **E2** with the decrease in solvent polarity supports specific catalysis. The data of Table 3 in micellar media agree with these medium effects. On going from water to the less polar micellar interface, the specific base-catalyzed reaction of **E2** is decelerated so that the unchanged general base catalysis becomes significant. The increase in $k_M^{AcO^-}/k_M^{HO^-}$ ratio for **E2** is, therefore, attributable to a decrease in $k_M^{HO^-}$ rather than to an increase in $k_M^{AcO^-}$.

The marked micellar rate decrease observed when the buffer concentration increases (Figures 1 and 2) is indicative of an additional reduction of the interface polarity, which can arise from the substantial micellar incorporation of the acetic acid (vide supra). Since it was not possible to investigate this acetic acid effect without changing markedly the relative contributions of the two catalysts, we considered the kinetic influence of the ethanol addition to the buffered micellar solutions. Significant decelerations of the reaction in the dilute buffer, approaching those observed on going from the dilute to the concentrated buffer, are found (Figure 6) only with the addition of very large EtOH concentration (1.7 M^{-1}), about 17 times larger than those of acetic acid (0.1 M in the concentrated buffer), whereas K_S^{EtOH} (0.6 M^{-1})[21] is only 3 times smaller than K_S^{AcOH} (1.7 M^{-1}). Therefore, the incorporation of acetic acid at the micellar interface cannot be the only factor responsible for the observed decrease in the interfacial polarity in the presence of the more concentrated buffer.

Other related factors and, in particular, a change in the surface potential can be invoked. It is well established[1e,f] that salt addition to the water phase decreases the surface potential of ionic micelles and, therefore, the electrostatic interactions between the micellar interface and the charged reagents and/or transition states. Consequently, a rate decrease in the overall reaction on going from dilute to concentrated acetate buffer can also result from a surface potential decrease. The specific-base-catalyzed reaction with a poorly delocalized charge at the transition state would respond to these electrostatic

interactions more strongly than the general base catalysis with a more delocalized transition state. This effect leads also to a larger decrease in $k_M^{HO^-}$ than in $k_M^{AcO^-}$.

In conclusion, the different sensitivities of the two substrates to medium supports the differences in the mechanisms of their catalysis in water and the significant contribution of the general catalyzed pathway for **E2** when it reacts in the less polar micellar environment.

Conclusion

The most striking result of this investigation is the change in the relative contributions of the two catalyzed pathways for the **E2** reaction. Whereas in pure water the specific catalysis is largely predominant, hydroxy and acetate catalysis compete at the micellar interface. This is, to the best of our knowledge, one of the first examples of micelle-mediated general base catalysis by a basic surfactant counterion. In the present case, the increased contribution of the acetate-catalyzed reaction arises not from an increase but from a greater decrease in the micellar rate constant of the specific base-catalyzed reaction than the general base-catalyzed one. This is quite obvious in Tables 1 and 2; $k_{2,m}/k_w$ is 0.2 for the specific reaction of **E2** and about 0.45 for the general hydroxy- and acetate-catalyzed reactions of **E2** and **E3** (Schemes 2−4).

Another finding is that all of the rate constants at the micellar interfaces are smaller than those in water, whatever the catalysis mechanism. This is consistent with the micellar interface being less polar than water, as supported by the kinetic results in acetonitrile−water mixtures in which the two mechanisms respond quite differently to the medium effect. Whereas the specific catalysis depends markedly on the medium polarity, the general catalysis does not. This is evidenced by the relative reactivity ratios **E2**/**E3** and their sensitivity to the reaction medium. In water, the rate constant for the HO^--catalyzed reaction is 4.4 times larger for **E2** than for **E3**, but in micelles this factor is only 2. In contrast, the **E2**/**E3** ratios for the general acetate-catalyzed reactions are quite similar in the two media (0.5 in water and about 0.7 in micelles). The small medium sensitivity of the general base catalysis is probably related to a significant charge delocalization in the corresponding transition states.[1a,24,25]

Finally, the rate constant decrease raises the question of the origin of the micellar accelerations. This contradiction is generally solved in terms of increased reagent concentrations in the small micellar volume. In our study, the largest accelerations are found (Figure 2) in CTAOAc micellar solutions with the more dilute buffer for both substrates. These rate enhancements are unambiguously attributed to large pH increase (Figure 3), i.e., to an increase in $[HO^-]$ in CTAOAc. In the more concentrated buffer, the small pH changes show that the micellized hydroxide ion concentrations do not change much with the surfactant concentration. With CTAOAc, the micellized acetate ion concentration does not depend significantly on that of the buffer since it is controlled mainly by the β-parameter of the surfactant. In this case, the rate decrease on going from the dilute to the concentrated buffer is to be related to the decrease in interfacial polarity. In contrast with CTACl, the increase in buffer concentration results in an increase in $|AcO^-]_M$ because of the exchange equilibrium 12:

$$CTACl + AcO^-_w \rightleftharpoons CTAOAc + Cl^-_w \qquad (12)$$

However, $|AcO^-]_M$ are quite similar in CTAOAc and CTACl solutions in the concentrated buffer, as shown by calculations using eq 11 (Table 4). For example, $[AcO^-]_M/([D]_T - cmc) = 0.52$ and 0.50 in CTAOAc and 0.01 M CTACl, respectively, so that the actual surfactant is CTAOAc in both cases. In other words, the use of a concentrated buffer favors the acetate catalysis by increasing the concentration of micellized acetate ions without altering markedly the hydroxide concentration. This agrees with a recent finding[15] that, in the presence of buffers, the actual counterion is the buffer base at small surfactant concentrations because of a significant shift of eq 12 arising from high buffer and small surfactant concentrations.

Insofar as micelles are reasonable models of the microenvironment of enzyme reactions, the significant increase in the general base catalysis resulting from the micellar effect found in this work opens new insights in bioorganic reactivity. More work is in progress to understand the interplay between the medium polarity dependence of the micellar rate constants and the surfactant-induced pH variations, i.e., the ion-exchange controlled concentrations of the catalysts, on the efficiency of micellar catalysis in acid−base-catalyzed reactions.

Table 4. Micellization of Acetate Ions in Solutions of CTACl in 0.02 and 0.2 M Acetate Buffers 50% Base

[CTACl], M	θ_{AcO^-} in 0.02 M acetate buffer[a,b]	θ_{AcO^-} in 0.2 M acetate buffer[a,b]	$[AcO^-]_M$ 0.2 M / $[AcO^-]_M$ 0.02 M
0.0005	0.55	0.59	1.09
0.001	0.50	0.59	1.17
0.002	0.44	0.58	1.32
0.004	0.35	0.56	1.59
0.005	0.32	0.55	1.71
0.006	0.30	0.54	1.81
0.008	0.26	0.52	2.03
0.01	0.23	0.50	2.22
0.025	0.12	0.41	3.33
0.05	0.07	0.32	4.51
0.075	0.05	0.27	5.29
0.1	0.04	0.23	5.85

[a] $\theta_{AcO^-} = [CTAOAc]_M/[CTACl]_M + [CTAOAc]_M$; calculated from eq 11 with $\beta = 0.60$ and $K_{Cl^-}^{AcO^-} = 0.50$. [b] θ_{AcO^-} in CTAOAc = β = 0.52, whatever the buffer concentration.

Acknowledgment. We are indebted to the BAS and CNRS for support of I.B. and M.F.R. collaboration. I.B. gratefully acknowledges the University of Paris 7-Denis Diderot for a visiting professorship. We thank Professor O. A. El Seoud for fruitful discussion.

Supporting Information Available: Rate-surfactant profiles for ring closure of **E2** and **E3** in micelles of CTABr in acetate buffers and in CTACl in acetate buffers with addition of 0.1 M ethanol. This material is available free of charge via the Internet at http://pubs.acs.org.

JO001759W

(24) Buurma, J.; Herranz, A. M.; Engberts, J. B. F. N. *J. Chem. Soc., Perkin Trans. 2* **1999**, 113.

(25) Bunton, C. A.; Gillitt, N. D.; Mhala, M. M.; Moffatt, J. R.; Yatsimirsky, A. K. *Langmuir* **2000**, *16*, 8595.