

# Micellar effects on aromatic esters hydrolysis

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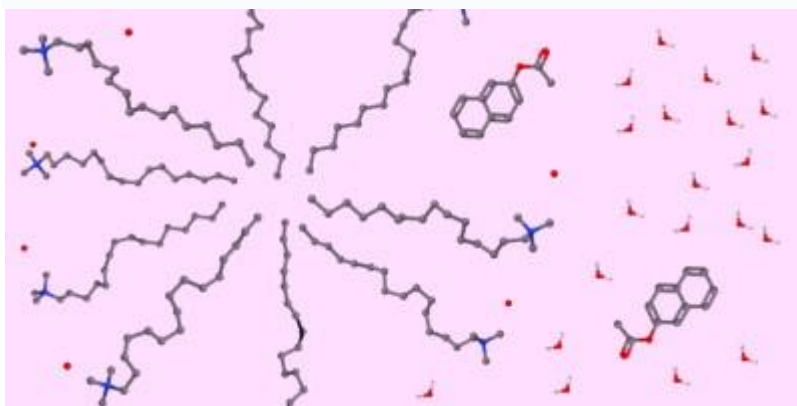
## Highlights

- Cationic surfactants catalyze the alkaline hydrolysis of aromatic esters.
- Variation of counter-ions association degree to micelles explains experimental data.
- Adsorption equilibrium constants of esters do not depend on micelle counter-ion.

## Abstract

The alkaline hydrolysis of two aromatic esters, 2-naphthyl acetate (2NA) and phenyl acetate (PhA) has been tackled in this work. The reaction has been followed in water and in the presence of cationic surfactants with different chain lengths: dodecyltrimethylammonium bromide (DTABr), tetradecyltrimethylammonium bromide (TTABr), and hexadecyltrimethylammonium bromide (CTABr) and the corresponding hydroxides (DTAOH, TTAOH and CTAOH) at 25 °C. The pseudo-first order rate constants increase with surfactant concentration in the presence of surfactants with reactive counter-ions, while a maximum is reached for bromide surfactants with a constant concentration of hydroxyl ion. Both, the association equilibrium constants of 2NA and PhA with micelles and the rate constant in this medium were determined. The degree of counter-ion association to micelles was considered as a variable parameter, using the adsorption equilibrium of bromide and hydroxide ions to the micelle.

## Graphical abstract



## Keywords

Hydrolysis; Ion-exchange equilibrium; Micelles; Mass action model; Pseudo-phase model

## 1. Introduction

Phenyl acetate is a really interesting compound: it is present in plasma because is a metabolite of phenylalanine, and it is useful in cancer chemoprevention because of its antitumor properties [1] and [2]. It has been used in humans, and in spite of the high concentrations needed to show antitumor activity, no toxicity has been observed. Besides, naphthyl acetates are used as substrates in the detection of carboxylesterases, the enzymes that catalyze the hydrolysis of different esters and are quite important in the metabolism [3].

On the other hand, micelles are supramolecular assemblies originated by association of surfactant molecules [4], [5] and [6]: after the critical micellar concentration the amphiphilic monomers associate to form micelles, with an inner core integrated by the hydrophobic part of the surfactant, while the hydrophilic part is oriented toward the aqueous solvent, stabilizing the aggregates. They are mimetic entities that resemble cell membranes and have been used as drug delivery systems.

The association of both compounds used in this study with micelles has been investigated by a kinetic method, following the variation in the rate of the alkaline hydrolysis in the presence of surfactant. It is a well-known fact that micelles can modify reaction rates, because they provide micro-environments different from bulk water, and besides they may exert a concentration effect because they can shift equilibrium position [7]. In the frame of the pseudo-phase model, micelles are considered a phase

different from aqueous medium, where reaction rates and solubilities of the substrates can vary considerably. In reactions with ionic species, an important fact is the charge of the surfactant head groups and counter-ions. In this way, it is expected that the rate of hydrolysis of hydrophobic esters by  $\text{OH}^-$  ions will be enhanced by cationic micelles, that can include the ester in their core and attract reactive ions of opposite sign, that is, hydroxyl ions. When the counter-ions of the surfactant are non reactive, for instance, bromide ions, there is a competition between  $\text{OH}^-$  and  $\text{Br}^-$  attracted toward the surface of the micelle. In this case the equilibrium ion-exchange constant can be used to quantify the concentration of hydroxyl ions in the micellar surface, together with the assumption that the degree of association of counter ions to the micelle,  $\beta$ , can be considered constant; this is essentially the pseudo-phase ion-exchange (PIE) model. However, when counter ions are very hydrophilic, e.g. when the counter ion of the micelle is the hydroxyl ion, cannot longer be considered constant and the adsorption equilibrium of  $\text{OH}^-$  to the micelles has to be used to determine the concentration of this reactive ion in the micellar surface; this is known as the mass action model.

In this paper the reaction of two aromatic esters, 2-naphthyl acetate (2NA) and phenyl acetate (PhA) has been followed in aqueous alkaline medium and in the presence of cationic surfactants with different lengths of the alkyl chains, and different counter ions ( $\text{Br}^-$  and  $\text{OH}^-$ ). The mass action model has been applied to explain the experimental behavior, and different approximations have been compared, as it will be discussed in more detail in the next sections.

## 2. Experimental

### 2.1. Materials

All reagents were of high degree of purity and used as received. Esters, 2-naphthyl acetate (2NA) and phenyl acetate (PhA), were supplied by Sigma–Aldrich, the same that the surfactants: dodecyltrimethylammonium bromide (DTABr, 99%), tetradecyltrimethylammonium bromide (TTABr, 99%), and hexadecyltrimethylammonium bromide (CTABr, 99%). Hydroxides of alkyltrimethylammonium (DTAOH, TTAOH and CTAOH) were prepared using Amberlite IRA-400 ion exchange resin to substitute halogen ions by  $\text{OH}^-$ . A weighted amount of surfactant bromide was dissolved in water and passed slowly through the resin column saturated with hydroxide ion. The column was rinsed with water several times to avoid lost of reagent, and the filtered solution was brought to the desired volume. The absence of bromide ion was confirmed with silver nitrate. Solutions were freshly prepared before use and stored refrigerated for no more than two days.

All other reagents were supplied by Merck and the solutions were prepared with doubly-distilled water obtained from a permanganate solution.

Esters were dissolved in acetonitrile, and the working solutions were prepared daily by dilution of the stock solution. The percentage of organic solvent in the working solution was less than 2% by volume.

## 2.2. Methods

Hydrolysis of the esters was carried out under pseudo-first order conditions in excess of hydroxide ion, being the concentration  $<0.23$  mM for 2-NA, and 0.41 mM for PhA. These concentrations were selected to follow the reaction by recording the absorbance of the product of the hydrolysis, 2-naphtholate or phenolate ions. The integrated method was used, fitting the experimental absorbance–time data to the first-order integrated rate equation:  $A = A_{\infty} + (A_0 - A_{\infty}) \exp(-k_0 t)$  by non-linear regression analysis, to obtain  $A_{\infty}$ ,  $A_0$  and  $k_0$  as optimizable parameters (with  $A$ ,  $A_0$  and  $A_{\infty}$  being the absorbance readings at times  $t$ , zero and infinity, respectively; while  $k_0$  represents the pseudo-first order rate constant).

When the rate constant was low, the ultraviolet absorption spectra were recorded with a Uvikon-XS (Bio-Tek Instruments) double-beam spectrophotometer, but the vast majority of the measurements were made in a Bio-Logic SFM-20 stopped-flow system interfaced with a computer and operated by a Bio-Kine32 software (V4.51, 2009). In all the experiments the temperature was kept constant at 25 °C.

## 3. Results and discussion

### 3.1. Reaction in basic media

Esters hydrolysis yields the corresponding alcohol and acid as the reaction products, which in alkaline media can be ionized. In the case of the aromatic esters used in this study, 2-naphthyl acetate yields 2-naphtholate ion and acetate, while phenyl acetate yields phenolate ion and acetate. In both cases, the ionic form of the aromatic compound absorbs in the UV region of the spectra. The absorption spectra of 2NA in basic media presents a wide band centered at 346 nm that increases with time due to the formation of 2-naphtholate. The position of the maximum is shifted to 356 nm in the presence of surfactant as can be seen in Fig. 1a, where spectra were recorded every minute in the presence of CTAOH. The same is true for PhA where the maximum of the band of phenolate ion is located at 289 nm in the absence of surfactant and at 293 nm in CTAOH, as it can be seen in Fig. 1b.

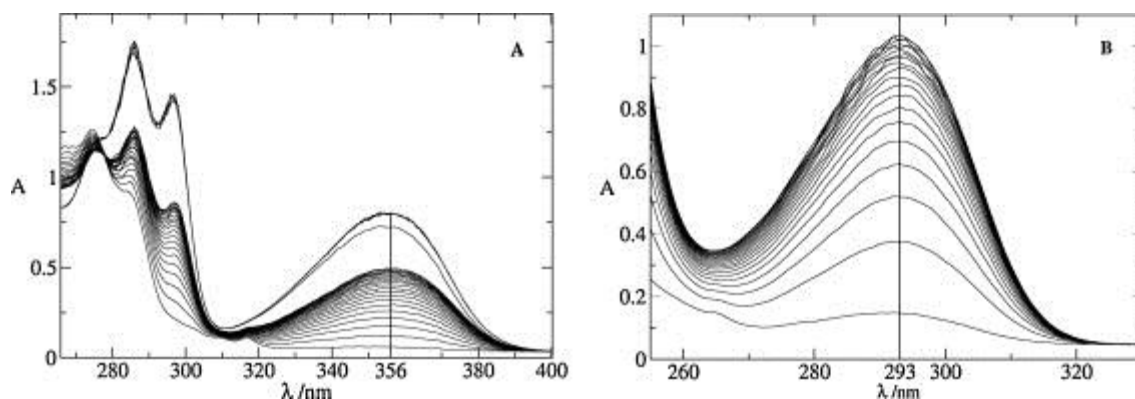


Fig. 1. (A) UV-visible spectra of 2NA (0.23 mM) showing the alkaline hydrolysis in [CTAOH] = 0.75 mM (lower curves) and [CTAOH] = 4.2 mM (upper curves), scan repeated every minute. (B) UV-visible spectra showing the alkaline hydrolysis of PhA (0.41 mM) in [CTAOH] = 2.1 mM, scan repeated every 1.5 min ( $M = \text{mol dm}^{-3}$ ).

The increasing absorbance at the wavelength of the maximum was recorded with time and used to calculate the pseudo-first order rate constant,  $k'_w$ , that is related to the

second order constant,  $k_w$ , by the equation  $k'_w = k_w [\text{OH}^-]$ ; fitting of experimental data to this expression yields the following values:

$$k_w = (1.465 \pm 0.009) \text{ M}^{-1} \text{ s}^{-1} \text{ (PhA)} \quad k_w = (1.77 \pm 0.04) \text{ M}^{-1} \text{ s}^{-1} \text{ (2NA)}$$

where  $[\text{OH}^-]$  was varied from 0.02 M to 0.20 M, in absence of surfactant.

### 3.2. Reaction in micelles functionalized with $\text{OH}^-$

The hydrolysis of both esters was studied in the presence of functionalized cationic surfactants. Hydroxides of alkyltrimethylammonium with different size of the alkyl chain were used, dodecyl-, tetradecyl- and hexadecyltrimethylammonium hydroxides, denoted DTAOH, TTAOH and CTAOH, respectively. Rate constants were obtained varying the concentration of surfactant without added  $\text{OH}^-$  and adding a constant amount of  $\text{OH}^-$ , as it can be seen for 2NA in Fig. 2 and Table 1S (supplementary information), and for PhA in Fig. 3 and Table 2S (supplementary information). On the other hand, the critical micelle concentration, for each series of data is given in Table 1.

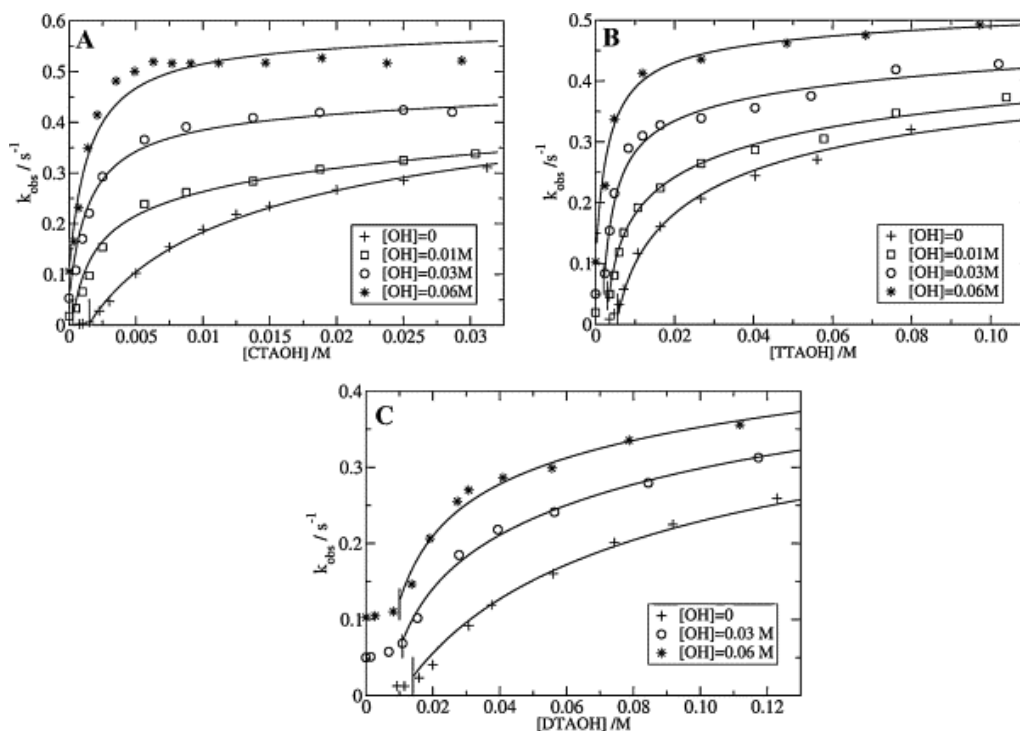


Fig. 2. Variation of pseudo-first order rate constant for hydrolysis of 2NA with concentration of surfactant (CTAOH, TTAOH and DTAOH), at fixed added  $[OH]$  and  $25\text{ }^{\circ}\text{C}$ . Symbols represent experimental data and lines correspond to model of Eq. (3) with the parameters of Table 2. CMC is indicated with a small vertical line ( $M = \text{mol dm}^{-3}$ ).

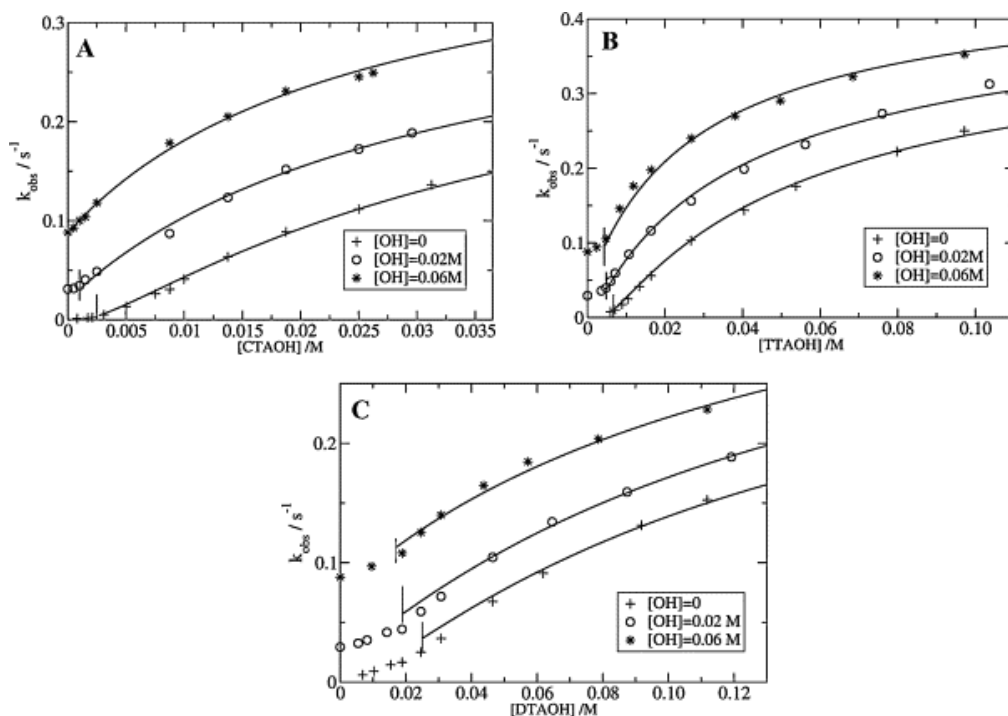


Fig. 3. Variation of pseudo-first order rate constant for hydrolysis of PhA with concentration of surfactant (CTAOH, TTAOH and DTAOH), at fixed added  $[OH]$  and  $25\text{ }^{\circ}\text{C}$ . Symbols represent experimental data and lines correspond to model of Eq. (3) with the parameters of Table 2. CMC is indicated with a small vertical line ( $M = \text{mol dm}^{-3}$ ).

Table 1.

Critical micelle concentration, CMC, in mol dm<sup>-3</sup> determined at 25 °C, for the indicated surfactants, as the minimal surfactant concentration required to observe kinetic effects in the alkaline hydrolysis of 2NA and PhA.

[OH] (M)	[CTAOH]	[TTAOH]	[DTAOH]	[CTABr]	[TTABr]
hydrolysis of 2NA					
0	0.0015	0.0056	0.014		
	0.0018 <sup>a</sup>	0.005–0.007 <sup>a</sup>	0.030–0.031 <sup>a</sup>		
0.01	0.0003	0.0030			
0.03		0.0020	0.011		
0.06			0.010	0.00020	0.00091
0.12				0.00018	
[OH] (M)	[CTAOH]	[TTAOH]	[DTAOH]		
hydrolysis of PhA					
0	0.0025	0.0067	0.025		
0.02	0.0010	0.0048	0.019		
0.06		0.0043	0.017		

<sup>a</sup> Taken from Ref. [10].

At surfactant concentrations below the CMC, the monomers of the surfactant increase a little the rate of reaction because the counter ion is OH<sup>-</sup> that reacts with the esters; then a small linear increase is expected; above the CMC, micelles are formed and the reaction rate increases sharply. The concentration at which this abrupt change of rate takes place was determined graphically as the CMC of the solution [8]. Of course the CMC decreases with increasing the alkyl chain length, and besides it is smaller for 2NA than for PhA solutions, because naphthyl moiety of 2NA is more hydrophobic than phenyl group of PhA, which enhances the hydrophobicity and promotes micellization. The data in the literature are in good agreement with that obtained in this study (see Table 1); according to Ref. [9] CMC at 25 °C varies in the following ranges 1.8–3.4 mM for CTAOH, 4.5 mM for TTAOH, 0.8–1.0 mM for CTABr and 3.3–3.8 mM for TTABr, for pure surfactants. Data from literature [10], obtained from conductivity and pH measurements, have been included in Table 1 for purpose of comparison. Coincidence is quite good for CTAOH and TTAOH, while a slightly bigger discrepancy is found for DTAOH. This can be attributed to the fact that the rate enhancement is lower in DTAOH than in the other surfactants, making more difficult the exact determination of the concentration of DTAOH that speeds up the reaction. On the other

hand, the addition of electrolyte decreases the CMC, for all the surfactants, as it can be seen in Table 1, because the repulsion between the head groups is reduced by the electrolyte ions [11].

The more evident feature of Fig. 2 and Fig. 3 is the enhancement in the reaction rates originated by the presence of the cationic surfactant. This is a reasonable behavior taking into account that aromatic substrates are hydrophobic in nature and tend to associate to the micelles by hydrophobic effect, mainly. Besides, the reaction is with an anion,  $\text{OH}^-$ , the counter ion of the micelle, this supramolecular assembly acts as a microreactor increasing the concentration of the two reactants in a medium, in any case, different from bulk water. In this situation, a really important fact is the degree of association,  $\beta$  (or dissociation),  $\alpha$  of the counter ion of the micelles,  $\text{OH}^-$ ; an increase in the  $\beta$  value implies a higher concentration of one of the reactants, that will enhance the reaction rate. Hydroxyl ions are very hydrophilic and as it has been demonstrated in the literature, in a micelle with hydrophilic counter ions,  $\beta$  is not constant and its variation has to be considered to fit the kinetic data [7] and [12]. The adsorption equilibria of the different counter-ions present in the system has to be taken into account to estimate the concentration of hydroxide in the micellar pseudophase,  $[\text{OH}]_m$ . In Ref. [13], the hydrolysis of an ester was carried out in CTAOH without and with added  $\text{OH}^-$ ; a variation of  $\beta$  from 0.37 to 0.82 was estimated when CTAOH was changed from 2 to 60 mM, without added hydroxide; while for  $[\text{OH}] = 0.03 \text{ M}$ , it was found that  $\beta$  was virtually constant with values:  $\beta \equiv 0.92$  for 2 mM CTAOH and  $\beta \equiv 0.93$  for 60 mM CTAOH. So it is expected that  $\beta$  will vary when hydroxides of alkyltrimethylammonium are used without, or with a small amount of added  $\text{OH}^-$ , although it may be constant at higher  $\text{OH}^-$  concentration. Therefore, in the next equations variation of  $\beta$  will be considered, besides the same formalism applies when it is constant, it is a more restrictive case that can be managed with the same equations, but with the advantage that no supposition is made about  $\beta$ .

The hydrolysis reaction of the substrate,  $S$ , takes place according to Scheme 1, where  $k'_w$  and  $k'_m$  are the first order rate constants in aqueous and micellar pseudophases, respectively, while  $K_s$  is the association constant of the ester to the micelle, and  $D_n$  is the concentration of micellized surfactant ( $[\text{Surfactant}] - \text{CMC}$ ) [14], [15], [16] and [17]. According to this mechanism, the pseudo-first order rate constant  $k_{\text{obs}}$  is given by:

$$k_{\text{obs}} = \frac{k'_w + k'_m K_s [D_n]}{1 + K_s [D_n]} \quad (1)$$

and the first order constants are related to the second order ones, by:

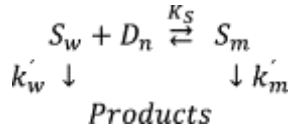
$$k'_w = k_w [\text{OH}]_w \quad (2a)$$



and

$$k'_m = \frac{k_m[\text{OH}]_m}{[\text{D}_n]} \quad (2b)$$

where  $k'_m$  is defined in terms of the ratio of hydroxyl ions concentration bound to the micelle.



Scheme 1.

Equilibrium and reaction step for hydrolysis of an ester (S) in cationic micelles.

Using Eq. (2a) and (2b),  $k_{\text{obs}}$  can be rewritten as:

$$k_{\text{obs}} = \frac{k_w[\text{OH}]_T + (k_m K_S - k_w)[\text{D}_n]m_{\text{OH}}}{1 + K_S[\text{D}_n]} \quad (3)$$

The concentration ratio of hydroxyl ion in the micellar pseudo-phase,  $m_{\text{OH}} = [\text{OH}]_m/[\text{D}_n]$  is the solution of the equation:

$$m_{\text{OH}}^2 - \left\{ 1 + \frac{[\text{OH}]_T}{[\text{D}_n]} + \frac{1}{K_{\text{OH}}[\text{D}_n]} \right\} m_{\text{OH}} + \frac{[\text{OH}]_T}{[\text{D}_n]} = 0 \quad (4)$$

where  $[\text{OH}]_T$  is the total amount of hydroxide ion in solution, that is coming from the functionalized surfactant (XTAOH) and from constant amount of sodium hydroxide added to each set of data  $[\text{OH}]_T = [\text{NaOH}]_{\text{added}} + [\text{XTAOH}]$

The aim of this treatment is to explain the kinetic data and calculate the constants of the system  $k_m$  and  $K_S$ , because in Eq. (3),  $k_w$ ,  $[\text{OH}]_T$ ,  $[\text{D}_n]$  are known parameters, while  $m_{\text{OH}}$  can be calculated with Eq. (4) with a certain value of  $K_{\text{OH}}$ . In fact, the fitting equation can be constructed by substitution of  $m_{\text{OH}}$ , given by solution of Eq. (4), in Eq. (3). For the hydrolysis of 2NA in CTAOH four series of data of  $k_{\text{obs}}$  were obtained varying surfactant concentration: one in solutions of pure CTAOH and three with a fixed added concentration of NaOH (0.01, 0.03 and 0.06 M). Data are plotted in Fig. 2. To determine the unknown constants, the values of  $K_S$  and  $K_{\text{OH}}$  were considered the same for the four series, while the value of  $k_m$  was varied for each set of data, as it can be seen in Table 2. This kind of fit, a *multi-branch* fit, is relatively complicated, there are multiple data sets with one or more common constants ( $K_S$  and  $K_{\text{OH}}$ ), but with parameters that are different for each set ( $k_m$ ) [18]. There are programs that are able to perform this kind of fit, one of them is Gnuplot[19], a very valuable free software

package for data fitting and graphical representation. The fitting equation for each series of data is given to the program,  $[\text{OH}]_{\text{T}}$  and  $k_{\text{m}}$  are different in each series, while the equilibrium constants are the same; the set of data corresponding to each equation are provided to the program, that finds the best parameters minimizing the total sum of squared residuals. Parameters in Table 2 were obtained in this way and were used together with Eq. (3) to calculate the lines that are plotted with the experimental data in Fig. 2, and as it can be seen, the concordance of experimental data and theoretical curves is good.

Table 2. Parameters for 2NA and PhA in CTAOH, TTAOH and DTAOH at 25 °C, calculated in the fit of experimental data to Eq. (3) (values without error are fixed parameters and  $\text{M} = \text{mol dm}^{-3}$ ).

$[\text{OH}]$ (M)	$K_{\text{S}}$ ( $\text{M}^{-1}$ )	$k_{\text{m}}$ ( $\text{s}^{-1}$ )	$K_{\text{OH}}$ ( $\text{M}^{-1}$ )	$K_{\text{S}}$ ( $\text{M}^{-1}$ )	$k_{\text{m}}$ ( $\text{s}^{-1}$ )	$K_{\text{OH}}$ ( $\text{M}^{-1}$ )
2NA in CTAOH						
0	$678 \pm 70$	$0.62 \pm 0.07$	$74 \pm 22$	$716 \pm 64$	$0.70 \pm 0.02$	$55^{\text{a}}$
0.01		$0.57 \pm 0.06$			$0.63 \pm 0.02$	
0.03		$0.61 \pm 0.04$			$0.66 \pm 0.02$	
0.06		$0.70 \pm 0.03$			$0.74 \pm 0.01$	
$[\text{OH}]$ (M)	$K_{\text{S}}$ ( $\text{M}^{-1}$ )	$k_{\text{m}}$ ( $\text{s}^{-1}$ )	$K_{\text{OH}}$ ( $\text{M}^{-1}$ )			
2NA in TTAOH						
0	$319 \pm 35$	$0.51 \pm 0.03$	$53 \pm 10$			
0.01		$0.53 \pm 0.03$				
0.03		$0.57 \pm 0.03$				
0.06		$0.62 \pm 0.02$				
2NA in DTAOH						
0	$69 \pm 24$	$0.56 \pm 0.09$	$13 \pm 6$			
0.03		$0.62 \pm 0.09$				
0.06		$0.65 \pm 0.09$				
PhA in CTAOH						
0	$41 \pm 4$	$0.42 \pm 0.02$	$74^{\text{b}}$			
0.02		$0.45 \pm 0.02$				
0.06		$0.49 \pm 0.02$				
PhA in TTAOH						
0	$34 \pm 2$	$0.46 \pm 0.01$	$53^{\text{c}}$			
0.02		$0.50 \pm 0.01$				
0.06		$0.53 \pm 0.01$				
PhA in DTAOH						
0	$9 \pm 3$	$0.44 \pm 0.04$	$13^{\text{b}}$			
0.02		$0.49 \pm 0.05$				
0.06		$0.51 \pm 0.05$				

a  $K_{\text{OH}} = 55$ . Taken from Ref. [7].

b Value taken from 2NA fit.

The model seems to be adequate to explain the experimental behavior. The rate constants at different hydroxide concentration,  $k_m$ , are found to be the same within the experimental error. The value of  $K_{OH} = 74$ , obtained in CTAOH, is quite similar to that used by Bunton et al. for a great number of system,  $K_{OH} = 55$  (see Tables in Ref. [7]). Different approaches can be found in the literature regarding to  $K_{OH}$ . Some authors consider that the substrate can alter the interface in such a way that binding constant,  $K_{OH}$ , and ion-exchange constants depend on the compound under study [14] and [20]. On the contrary, other authors [7],[21], [22] and [23] use values of these constant obtained previously in the literature, assuming that they should remain the same with different substrates. If a value of  $K_{OH} = 55$  is used as a fixed parameter in the fit of the experimental data, the difference with  $k_m$  and  $K_S$  obtained when  $K_{OH} = 74$  is really small, as it can be seen in Table 2. When a value is considered equal for all the sets, the cells in the table are jointed and a unique value is found there, on the other hand, when a parameter is taken from the literature the reference is indicated and no error is given. The obtained value of  $K_{OH} = 74$ , can be used to calculate  $m_{OH}$  as a function of the CTAOH concentration, that is plotted in Fig. 4. As it can be seen, if  $[OH]$  is large enough  $m_{OH}$  is almost constant, but this is not true at low  $[OH]$ .

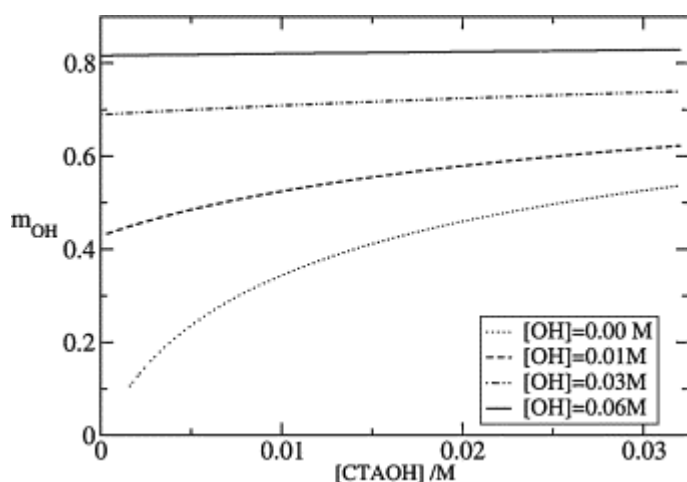


Fig. 4.  
 $m_{OH}$  in CTAOH calculated using Eq. (4) with  $K_{OH} = 74$  at different hydroxyl concentrations ( $M = \text{mol dm}^{-3}$ ).

A single value of the association constant has been considered because although it is expected to vary with  $[OH]$ , due to the salting out, the variation in  $[OH]$  is too small to observe this effect. In Ref. [24], the authors studied the variation of  $K_S$  with  $[OH]$ , but  $[OH]$  were increased up to 1 M. For small values of  $[OH]$ , up to 0.12 M, a single value of  $K_S = 100$  was considered, for the substrate used in that paper, while  $K_S = 110$  was found at  $[OH] = 0.25$  M, a small difference, in any case. Of course, an important point is that good fits are obtained using a single value of  $K_S$  for the esters under study.

The parameters in TTAOH and DTAOH for 2NA and PhA are given together with those for CTAOH in Table 2, and similar behavior is found. In relation with the data of

PhA, the value of  $K_{OH}$  obtained for 2NA has been used in the fit to avoid a great number of parameter to be calculated with a smaller number of data: there are three, instead of four sets of data in each surfactant because the variation of  $k_{obs}$  is smaller in this compound, a reasonable fact for a less hydrophobic substance, with one instead of two aromatic rings. The calculated parameters for PhA are given in Table 2, and the corresponding data are plotted in Fig. 3.

An outstanding feature of this model is that the value of  $k_m$  is almost constant for the same substrate in different surfactants and hydroxyl concentrations. It may be considered that there is a very slow increase in  $k_m$  when  $[OH^-]$  increases or it may be considered constant within the experimental error. On the contrary, the value of  $K_S$  increases sharply with the number of carbon atoms in the alkyl chain of the surfactant, what it is perfectly reasonable because a longer chain supposes a more hydrophobic character of the micelle, that favors the inclusion of the aromatic part of the molecule into the aggregate. However the reaction with the ester moiety takes place in the interface, where it is possible the contact between the  $OH^-$  ions, that is quite hydrophilic, and the substrate, partially immersed in the micelle. In this case the interface may be relatively similar for the three surfactants, that have the same head group, yielding a similar value of the rate constant,  $k_m$ ; that, in any case, may be different from that in water. It is interesting compare the rate of reaction in both media. To accomplish this task, a conversion from  $k_m$ , first order ( $s^{-1}$ ), to  $k_{m2}$ , second order rate constant ( $s^{-1} M^{-1}$ ) is needed. An estimation of the volume of that part of micelle where the reaction takes place has to be made. Assuming that the reaction occurs in the Stern layer and that its volume is the half of the total volume of the micelle, values of  $V_M = 0.14-0.30 M^{-1}$  are used in the literature [7]. Then,  $k_{m2} = k_m V_M$ , and values of this parameter are  $k_{m2} = 0.084 M^{-1} s^{-1}$  for 2NA and  $k_{m2} = 0.067 M^{-1} s^{-1}$  for PhA (using  $V_M = 0.14 M^{-1}$  and mean values of the rate constants,  $k_m = 0.60 s^{-1}$  for 2NA and  $k_m = 0.48 s^{-1}$  for PhA). The micellar catalysis can be originated by two effects: a change in the concentration of reactants or a change in the rate constant in the micellar media [7]. As second order rate constants in micellar media are much smaller than that in water, it is obvious that the dominant factor in the micellar catalysis is the concentration of reactants in the micelles, for the systems under study.

Similar situations are found in the literature, with  $k_{m2}$  values much smaller than  $k_w$ . The alkaline hydrolysis of acetylsalicylic acid in cationic micelles was studied by different authors and values of  $k_{m2}$  approximately 100–140 times smaller than  $k_w$  were found [14] and [25]. The hydrolysis of 1,3-benzoxazine-2,4-dione and its derivatives yielded  $k_{m2}$  values 300–600 times smaller than  $k_w$  [26].

Although in these cases  $k_{m2}$  is the smallest value, in general, it can be greater or smaller than  $k_w$  [7]. When the negative charge is localized on the oxygen in the transition state, this can be stabilized by forming hydrogen bonds with water molecules, making the reaction in water more favorable than in the micelle interface where water molecules are scarce. When the charge is delocalized in the transition state, as in aromatic nucleophilic substitution,  $k_{m2}$  is greater than  $k_w$  [27]. In this situation the negative charge density of

the transition state is low and it is stabilized by the favorable interaction with the cationic micellar head groups.

As it has been shown above, the two esters under study exhibit quite similar rates of hydrolysis, but  $k_m$  and  $k_w$  are slightly greater for 2NA than for PhA. The reason is that the leaving group for 2NA, that is the 2-naphtholate ion, with two aromatic rings, is more stabilized by resonance than the phenolate ion, with only one ring, in the case of PhA hydrolysis.

### 3.3. Reaction in surfactant bromides

2NA hydrolysis has been studied in the presence of a variable amount of bromide surfactant at constant  $[\text{OH}^-]$ . As it can be seen in Fig. 5, the pseudo-first order rate constant increases in the presence of a small amount of surfactant until it reaches a maximum; after this point an increase in surfactant concentration diminishes the observed rate constant. This behavior is explained by the fact that at low surfactant concentration, the amount of substrate increases in the micelles speeding up the reaction, but at higher concentrations of surfactant, the competition between  $\text{OH}^-$  and the increasing amount of  $\text{Br}^-$  and the dilution of the substrate in the micelles causes the slow down of the reaction rate [28]. Hydrolysis of PhA in bromide surfactants has not been studied due to the low values of  $K_S$  for this substrate, as it can be seen in Table 2. In Fig. 5c, data for 2NA in DTABr are plotted and a really small catalytic effect is appreciated, for PhA with lower association constants an even smaller effect would be expected.

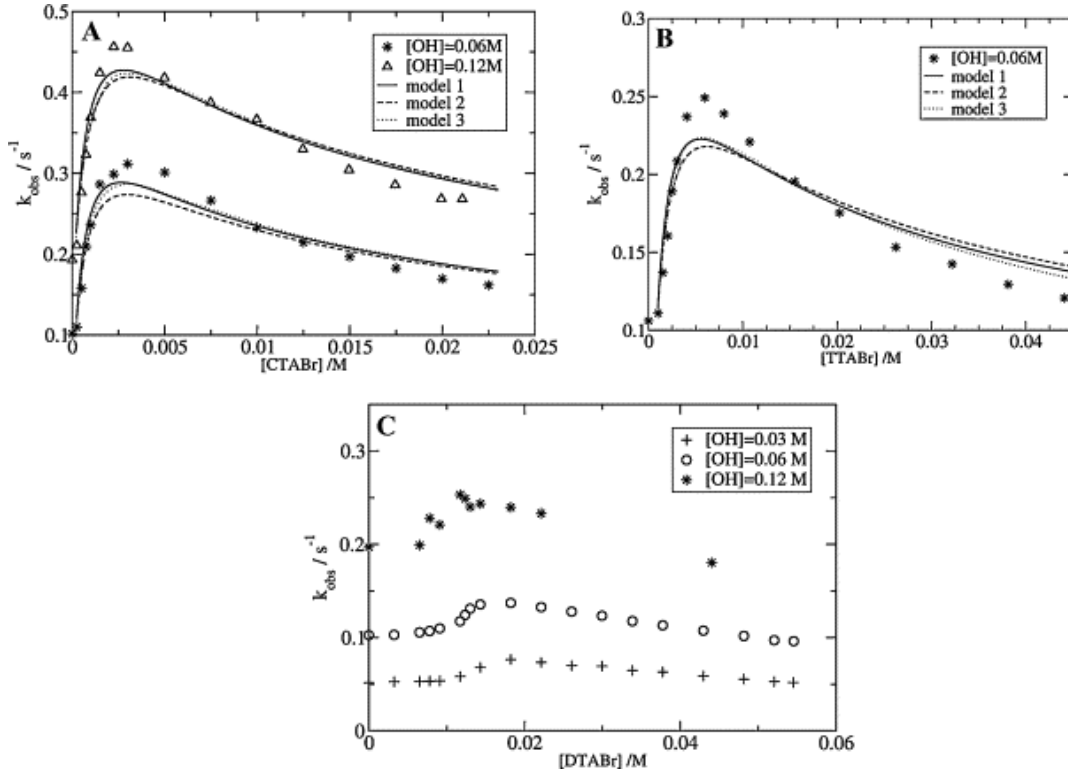
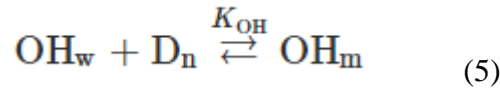


Fig. 5. Variation of pseudo-first order rate constant for hydrolysis of 2NA with concentration of surfactant (CTABr, TTABr and DTABr), at fixed added [OH] and 25 °C. Symbols represent experimental data and lines correspond to several models in Table 3 ( $M = \text{mol dm}^{-3}$ ).

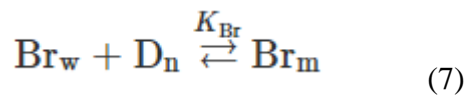
Eq. (3) for  $k_{\text{obs}}$  is still valid, but the way of calculation of  $m_{\text{OH}}$  or  $[\text{OH}]_{\text{m}}$  varies now, because two ion adsorption equilibria in the micelle have to be considered:



with constant:

$$K_{\text{OH}} = \frac{[\text{OH}]_{\text{m}}}{[\text{OH}]_{\text{w}}([\text{D}_{\text{n}}] - [\text{OH}]_{\text{m}} - [\text{Br}]_{\text{m}})} \quad (6)$$

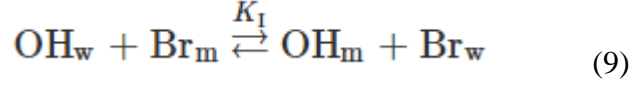
and



with constant

$$K_{\text{Br}} = \frac{[\text{Br}]_{\text{m}}}{[\text{Br}]_{\text{w}}([\text{D}_n] - [\text{OH}]_{\text{m}} - [\text{Br}]_{\text{m}})} \quad (8)$$

It is obvious that the combination of the two last equilibria yields ion-exchange equilibrium between  $\text{Br}^-$  and  $\text{OH}^-$  ions:



with constant

$$K_{\text{I}} = \frac{K_{\text{OH}}}{K_{\text{Br}}} \quad (10)$$

Therefore, if two of these parameters are known, the third can be calculated. This is important because values of  $K_{\text{I}}$  can be found in the literature:  $K_{\text{I}} = 0.048$  in CTA [29] and  $K_{\text{I}} = 0.07$  TTA [30].

It is a well known fact that  $K_{\text{Br}} \gg K_{\text{OH}}$ , therefore bromide ion will be associated to the micelles in a greater extension than hydroxyl ion, then the approximation  $[\text{OH}]_{\text{T}} \gg [\text{OH}]_{\text{m}}$  or  $[\text{OH}]_{\text{T}} \approx [\text{OH}]_{\text{w}}$  can be employed and using this relation with Eqs. (6) and (8) together with the mass balances, the concentration of hydroxyl ion in the micellar pseudo-phase is the solution of the equation [15] and [16]:

$$K_{\text{Br}} \left\{ 1 + \frac{1}{K_{\text{OH}}[\text{OH}]_{\text{T}}} \right\} m_{\text{OH}}^2 + \{1 + K_{\text{OH}}[\text{OH}]_{\text{T}} + K_{\text{Br}}\text{CMC}\} \frac{m_{\text{OH}}}{[\text{D}_n]} - K_{\text{OH}} \frac{[\text{OH}]_{\text{T}}}{[\text{D}_n]} = 0 \quad (11)$$

Data in Fig. 5 (supplementary Table 1S) have been fitted to Eq.(3), with  $m_{\text{OH}}$  substituted by the solution of Eq.(11). The value of  $K_{\text{OH}}$  was taken from Table 2, while several possibilities have been considered related to the other parameters. In the case of CTABr two series of data were obtained at 0.06 and 0.12 M of hydroxyl ion. Gnuplot program was used to fit these two sets of data with a common value of  $K_{\text{S}}$  and  $K_{\text{I}}$  and letting  $k_{\text{m}}$  to vary for both series.

A first fit, model 1, was performed using  $K_{\text{I}} = 0.048$  in CTA as a fixed value and the other parameters were calculated (fixed values has no error in Table 3). In this case, at  $[\text{OH}^-] = 0.06$  M, it was found  $k_{\text{m}} = 0.70 \text{ s}^{-1}$  that it is the same value that was obtained in CTAOH with the same hydroxyl concentration, a reasonable result, taking into account that the interface should be quite similar. There is a difference in  $K_{\text{S}}$  that increases from

678 to 852  $M^{-1}$ . Although some characteristics of the interface could remain almost the same when the counter-ion is changed, the number of aggregation varies from 46 in CTAOH (23 °C, 50 mM) to 104 in CTABr (23 °C, 31 mM) [9]. Consequently, the roughness of the surface and the capacity to accommodate the substrate could vary yielding a different binding constant for the ester.

Table 3.

Parameters for 2NA in CTABr and TTABr at 25 °C, calculated with Eq. (3) (values without error are fixed parameters and  $M = \text{mol dm}^{-3}$ ).  $K_{OH} = 74$  in CTA and  $K_{OH} = 53$  in TTA, from Table 2 were used.

2NA in CTABr	[OH] (M)	$K_S (M^{-1})$	$k_m (s^{-1})$	$K_I (M^{-1})$
Model 1	0.06	852 ± 86	0.70 ± 0.02	0.048 <sup>a</sup>
	0.12		0.78 ± 0.02	
Model 2	0.06	678	0.70	0.047 ± 0.004
	0.12		0.80 ± 0.02	
Model 3	0.06	678	0.74 ± 0.02	0.048 <sup>a</sup>
	0.12		0.81 ± 0.02	
2NA in TTABr	[OH] (M)	$K_S (M^{-1})$	$k_m (s^{-1})$	$K_I (M^{-1})$
Model 1	0.06	375 ± 72	0.64 ± 0.03	0.07 <sup>b</sup>
Model 2	0.06	319	0.62	0.081 ± 0.007
Model 3	0.06	319	0.68 ± 0.02	0.07 <sup>b</sup>

a Taken from Ref. [29].

b Taken from Ref. [30].

On the contrary, if it is assumed that  $K_S = 678 M^{-1}$ , and  $k_m = 0.70 s^{-1}$  at  $[OH^-] = 0.06 M$ , that is, they remain the same that in CTAOH, model 2, then the other parameters obtained are  $k_m = 0.80 s^{-1}$  at  $[OH^-] = 0.12 M$  and  $K_I = 0.047$ , a really good value compared to that in the literature [29]. In Fig. 5, the curves corresponding to both fits are plotted and it can be seen that they are almost equally good to explain the experimental behavior, it could be said that model 1 is slightly better than 2. Although the difference is small, it seems interesting to check if the approximation used to obtain Eq. (11), that is has some effect in this result. In this way, as reasonable data for the equilibrium constants are available,  $K_I = 0.048$  (from Ref.[29]) and  $K_{OH} = 74$  (from Table 2), these values were used to calculate  $m_{OH}$ , without using any simplification. The concentrations of all the present species were determined with HYSS program [31], a software designed for speciation calculation: if the equilibrium constants and the initial concentrations of the reactants are provided to the program, the concentrations of all the species in the equilibrium can be obtained, among them



$[\text{OH}]_m$  and consequently,  $m_{\text{OH}}$ . In this situation, Gnuplot program was used to fit  $k_{\text{obs}}(m_{\text{OH}}, [\text{D}_n])$  given by Eq. (3), with  $K_S = 678 \text{ M}^{-1}$  from Table 1, and the values of  $k_m$  were obtained, the result is given in Table 3, model 3, and the lines are plotted in Fig. 5. There is no significant difference between curves of models 1 and 3.

Consistently, in this case it seems not be necessary to assume a change in the binding constant to explain data. However this option has been used in the literature for other systems [32], where  $K_S$  is different in CTACl and CTAOH for the substrate used in that study. In other references the same value of  $K_S$  in CTABr and CTAOH is assumed [24]. It is interesting to mention that in Ref. [20] hydrolysis of 2-naphthyl acetate in basic media was performed in the presence of CTABr (at three different hydroxyl concentrations, up to 0.03 M) and in CTAOH (only without added  $\text{OH}^-$ , only up to 15 mM). A value of  $K_S = 680 \text{ M}^{-1}$  and  $k_m = 0.53 \text{ s}^{-1}$  was found in CTABr, while  $K_S = 800 \text{ M}^{-1}$  was estimated in CTAOH, the authors used  $K_{\text{OH}}$  and  $K_{\text{I}}$  as adjustable parameters that depend on the substrate. These two constants seem not to have a great influence in the values of both the substrate binding and rate constants, similar result was found in Ref. [33] where  $K_{\text{I}}$  was varied orders of magnitude with a small variation in the least-square values of the fit. In this situation it is not surprised that good fits can be found with fixed or varying values of these constants, that has to be taken into account, they cannot be omitted, but they influence is not determinant. Therefore, in spite of the differences in  $K_{\text{OH}}$  and  $K_{\text{I}}$ , the value of  $K_S$  in CTABr in Ref. [20] is equal to the one obtained in this paper in CTAOH. Only few data in CTAOH in that study were determined, therefore although the value of  $K_S$  is higher than in CTABr the difference may be ascribed to the experimental error (errors are not given in the original reference) more than to real difference in the binding constant.

In view of these data it could be assumed that, sometimes, the difference in values of the constants in the same surfactant with different counter-ion may be due to the experimental error, approximations and differences in the assumed values of the known constants.

Similar models have been used to fit  $k_{\text{obs}}$  in TTABr and the obtained parameters are given in Table 3 and the corresponding curves in Fig. 5, but data obtained in the presence of DTTABr have not been fitted as due to the small variation observed in the pseudo-first order rate constant. In any case, curves in this figure deviate from experimental data in a systematic way, the lines predicted by the models go below experimental data at concentrations of surfactant corresponding to the maximum observed values of  $k_{\text{obs}}$  and go above them at higher surfactant concentrations, as it can be seen in Fig. 5. This fact may be due to unavoidable systematic errors introduced in the values of the *known* constants and it may be influenced by the fact that the points in the upward part of the curve, where the slope is really important, will have experimental errors bigger than other points and will be affected in a greater extent by the experimental error in the CMC. The catalytic effect takes place in a quite narrow range of surfactant concentrations, that besides lies near the CMC, and unavoidably, experimental data will be conditioned by this situation. This sticking point is inherent to

the kinetic profiles, but in spite of this problem, fits of the kinetic model to experimental data are acceptably good.

On the other hand, another possibility that has been widely used in the literature is consider that the association degree of ions to the micelle,  $\beta$ , is constant at a certain  $[\text{OH}^-]$  concentration [20], [32],[34] and [35]. In this case with  $K_I$  and using  $\beta = m_{\text{OH}} + m_{\text{Br}}$  as a constant that has to be estimated or taken from the literature, the value of  $m_{\text{OH}}$  can be calculated. To check the validity of this approximation for data in this paper, the values of  $m_{\text{OH}}$  and  $\beta = m_{\text{OH}} + m_{\text{Br}}$  for model 2 are plotted in Fig. 6, and as it can be appreciated  $\beta$  is almost constant at both hydroxyl concentration, while  $m_{\text{OH}}$  varies considerably. The model is essentially the same that the one used in this paper, because  $\beta$  is, in fact, practically constant in the presence of bromide ion. On the contrary, in the case of the functionalized micelles, where the counter-ion is just hydroxyl ion, only at high  $[\text{OH}^-]$  the value of  $m_{\text{OH}}$ , (that is equal to  $\beta$  because there is no bromide ion), the association degree, can be considered constant (see Fig. 4).

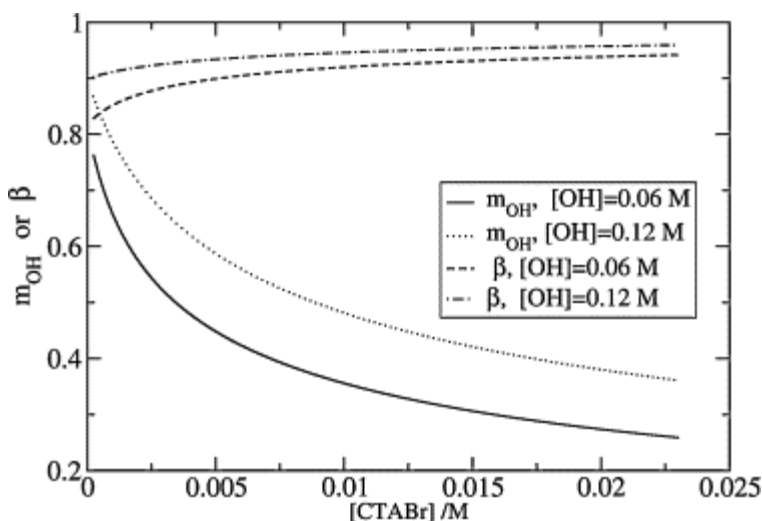


Fig. 6.  
 $m_{\text{OH}}$  and  $\beta = m_{\text{OH}} + m_{\text{Br}}$  for model 2 ( Table 3) in CTABr, at 25 °C ( $M = \text{mol dm}^{-3}$ ).

#### 4. Conclusions

The pseudo-first order rate constants of the alkaline hydrolysis of 2NA and PhA in water and in the presence of cationic surfactants, with inert and reactive counter-ions, can be explained with the mass action model, that considers the variation of the association degree of counter-ions to the micelles. This formalism yields as fitting parameters: (1) the adsorption equilibrium constants of both esters with micelles, confirming that can be considered equal with different counter-ion, for the same substrate, and (2) the rate constant in micellar medium, demonstrating that the dominant factor in the micellar catalysis is the concentration of reactants in the micelles.

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## Appendix A. Supplementary data

The following is supplementary data to this article:

**Table 1S.** Data of  $k_{\text{obs}}$ ,  $\text{s}^{-1}$ , for 2NA in CTAOH, TTAOH, DTAOH, CTABr and TTABr at 25°C

2NA			
[OH]=0.00 M	[OH]=0.01 M	[OH]=0.03 M	[OH]=0.06 M
[CTAOH] vs $k_{\text{obs}}$ , $\text{s}^{-1}$	[CTAOH] vs $k_{\text{obs}}$ , $\text{s}^{-1}$	[CTAOH] vs $k_{\text{obs}}$ , $\text{s}^{-1}$	[CTAOH] vs $k_{\text{obs}}$ , $\text{s}^{-1}$
0.00075 0.00218	0.0005 0.0332	0.00050 0.108	0.00035 0.1648
0.001 0.00307	0.00100 0.0657	0.00100 0.17005	0.00070 0.2311
0.0015 0.00577	0.00150 0.0976	0.00150 0.2205	0.00140 0.3493
0.00225 0.0281	0.00250 0.1538	0.00250 0.2926	0.00210 0.4144
0.00300 0.0471	0.00563 0.2387	0.00563 0.3655	0.00350 0.4813
0.00500 0.1025	0.00875 0.2618	0.00875 0.3907	0.00490 0.5002
0.00750 0.1538	0.01375 0.2839	0.01375 0.4088	0.00630 0.5191
0.01000 0.1883	0.01875 0.30745	0.01875 0.4191	0.00769 0.5162
0.01250 0.2183	0.02500 0.3247	0.02500 0.424	0.00909 0.5157
0.01500 0.2342	0.03039 0.338	0.02865 0.42	0.01119 0.5169
0.02001 0.2666			0.01469 0.5173
0.02501 0.2857			0.01889 0.5266
0.03126 0.3108			0.02378 0.5164
			0.02937 0.5213

2NA							
[OH]=0.00 M		[OH]=0.01 M		[OH]=0.03 M		[OH]=0.06 M	
[TTAOH] vs $k_{obs}$ , s <sup>-1</sup>		[TTAOH] vs $k_{obs}$ , s <sup>-1</sup>		[TTAOH] vs $k_{obs}$ , s <sup>-1</sup>		[TTAOH] vs $k_{obs}$ , s <sup>-1</sup>	
0.00356	0.00858	0.00356	0.0496	0.00238	0.0836	0.00238	0.2279
0.00475	0.01765	0.00475	0.0806	0.00356	0.1539	0.00475	0.33775
0.00594	0.0332	0.00594	0.1185	0.00475	0.21555	0.01188	0.41235
0.00713	0.0581	0.00713	0.1509	0.00831	0.2894	0.02672	0.4358
0.01069	0.117	0.01069	0.1917	0.01188	0.30985	0.04839	0.462
0.01639	0.1608	0.01639	0.2243	0.01639	0.3274	0.068412	0.47535
0.02672	0.207	0.02672	0.2646	0.02672	0.3389	0.097231	0.4925
0.04038	0.2443	0.04038	0.2867	0.04038	0.3558		
0.05611	0.2709	0.05777	0.305	0.05455	0.3753		
0.07981	0.3202	0.076014	0.3473	0.076014	0.41885		
		0.103987	0.37305	0.102063	0.42775		

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