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# Effect of Pressure on the Rate of Alkaline Fading of Triphenylmethane Dyes in Cationic Micelles

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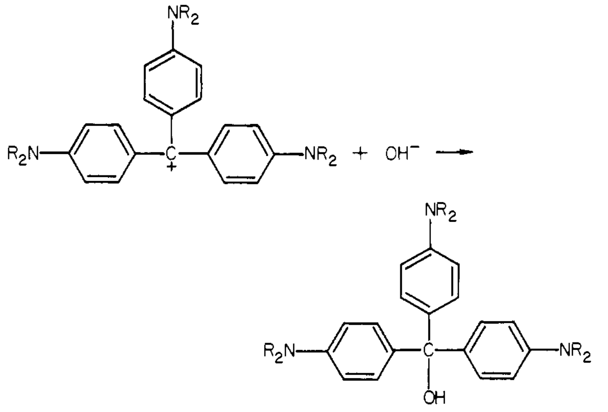
**Abstract:** The rates of alkaline fading of triphenylmethane dyes in cationic micelles have been measured at pressures up to about 2 kbar and at 25 °C in 0.1 M Tris buffer solution. These rates follow the kinetics of the Michaelis-Menten type. The volume changes for the incorporation of dyes in the micellar phase are 14-15 $cm^3$/mol for ethyl violet and brilliant green and 3-4 $cm^3$/mol for crystal violet and malachite green. It was confirmed that these differences are due to the hydrophobicity of dyes. The activation volumes for the fading reaction in micellar phases are positive except in the case of the crystal violet-cetyltrimethylammonium bromide micellar system. However, the activation volumes in nonmicellar aqueous solutions are negative. The differences of the activation volumes between micellar and nonmicellar systems are discussed from the point of view of the general base catalysis, which contains both the water and hydroxide ion pathways. It was clear that micelles change the water control into hydroxide ion control.

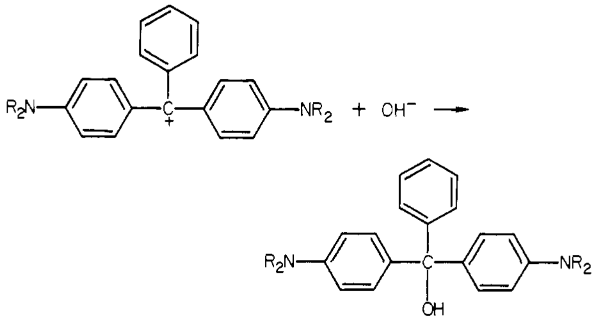
From high-pressure studies of chemical reactions, numerous data in terms of activation volumes have been accumulated, and some general rules for a number of reactions have been established.[1] One of the typical examples is the rule of solvation and desolvation around ions of reactants in aqueous solutions.[2] Ise et al.[3,4] have applied the rule to the interionic reactions catalyzed by polyelectrolytes. We focused on the hydrophobic effects on the hydrolysis of alkyl esters catalyzed by polyelectrolytes and micelles[5-7] and first found that the extent of hydrophobicity of esters corresponds to the large positive activation volumes. In general, the rates of the hydrolysis of phenyl esters catalyzed by hydrophobic polymers and micelles[8,9] are accelerated with increasing concentration of esters or catalysts and so follow the Michaelis-Menten kinetics.

In order to understand this reaction mechanism containing an acceleration effect, it is first necessary to get the data of the volume changes for the substrate binding in addition to the activation volumes for the process of the product formation. We have succeeded in separating the activation volumes of overall reactions into two contributions, volume changes for the substrate binding and activation volumes for the process of the product formation. The former is the process of the hydrophobic contributions between the substrates and catalysts. The latter corresponds to the solvation or desolvation processes of reactants between the initial and transition states with or without catalysts.



(1)

1a, R = $C_2H_5$
1b, R = $CH_3$



(2)

2a, R = $C_2H_5$
2b, R = $CH_3$

The data of the activation volumes from the effect of pressure on the rate of alkaline fading of malachite green did not conform to the general rules of the activation volumes, but recently le Noble et al.[10] succeeded in resolving this puzzle. When we apply the fading reaction to the micellar systems, we can easily change the reaction field in aqueous solution and we expect to get further information about this interesting reaction. From the effect of cationic micelles on the alkaline fading reaction of triphenylmethane dyes, ethyl violet (**1a**), crystal violet (**1b**), brilliant green (**2a**), malachite green (**2b**) the rate is markedly accelerated above the critical micelle concentration (cmc)[11-14] and with an increasing alkyl chain length of surfactants.[13] This seems to be one of the most typical examples of the rule of hydrophobic interactions in the kinetics of micellar

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Table I. Relative Ratio of cmc, Aggregation Numbers, and $\beta^P$ of 4 Micelle at Various Pressures and at 25 °C

| | P, kbar | | | | |
|---|---|---|---|---|---|
| | 0.001 | 0.5 | 0.75 | 1.0 | 1.5 |
| $(\mathrm{cmc})^P/(\mathrm{cmc})^1$ [a] | 1.00 | 1.04 | 1.14 | 1.15 | 1.13 |
| $N^P$ [c] | 75[b] | 51 | 47 | 45 | 47 |
| $\beta^P$ [d] | 0.75 | 0.63 | 0.60 | 0.58 | 0.60 |

[a] pH 8.0 (0.1 M Tris buffer). [b] Taken from ref 19. [c] Calculated from data in ref 20. [d] Ratio of the number of hydroxide ions to surfactants in micelles given by the equation of $\beta^P = 1 - q/N^P$, where $q$ is the effective charge on the micelle and the value estimated is 18.75 for the data of $q/N = 0.25$ (ref 21).

catalyzed reaction, and it can be applied to the Michaelis–Menten kinetics.

In the present paper, we applied the alkaline fading reaction of triphenylmethane dyes catalyzed by cationic micelles of cetyltrimethylammonium bromide **3** and myristyltrimethylammonium bromide **4** to the Michaelis–Menten kinetics under high pressure up to 2 kbar and at 25 °C. From the obtained volume change and the activation volumes we discussed the reaction mechanism from the points of view both of the hydrophobic interaction between dyes and micelles and of the solvation or desolvation of the reactions between initial and transition states with or without cationic micelles.

## Results and Discussion

Following the previous method,[9] we can apply the reaction of **1** and **2** in cationic micelles to the Michaelis–Menten type kinetics

$$\mathrm{M} + \mathrm{D} \overset{K}{\rightleftharpoons} \mathrm{MD} \xrightarrow{k_m} \mathrm{M} + \mathrm{DOH} \qquad (3)$$

$$\xrightarrow{k_s} \mathrm{DOH}$$

where M, D, MD, and DOH denote the micelles, dyes, the micelle–dye complexes, and carbinol, respectively. The relationship between the apparent rate constant $k_{app}$ and $k_s$ and $K$ values is given by eq 4 and 5

$$k_{app} = \frac{k_s + k_m[\mathrm{M}]\cdot K}{1 + [\mathrm{M}]\cdot K} \qquad (4)$$

$$[\mathrm{M}] = (C_D - \mathrm{cmc})/N \qquad (5)$$

where $C_D$ is the total concentration of cationic surfactants, and $N$ is the aggregation number of micelles. Applying eq 4 to the high-pressure experiment, we used the previous data[9] of cmc and $N$ for **3** and data of Table I for **4**. $K$ and $k_m$ values were determined by Lineweaver–Burk plots of eq 4 as shown in Figure 1 (**1b**) for **3** and in Figure 2 (**2b**) for **4**. They are summarized in Table II (**3**) and III (**4**). At each fading reaction at 1 bar, the $K$ values of **1a** and **2a** (rich hydrophobic dyes) are larger than those of **1b** and **2b**. The plots of the pressure vs. logarithms of $K$ are shown in Figures 3 for **3** which shows minima at about 1 kbar. The plots of the pressure vs. logarithms of $k_m$ are shown in Figures 4 for **3**. The inversion phenomena correspond to the appearance of maxima of the cmc of cationic micelles at 1 kbar

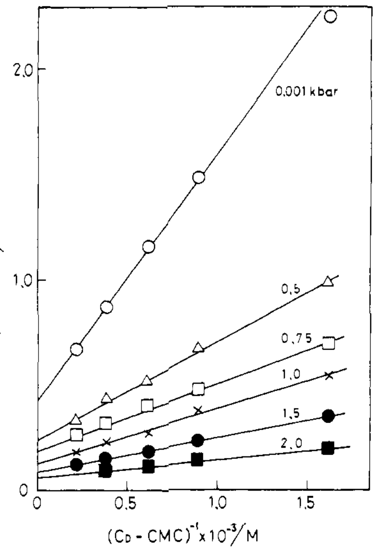


**Figure 1.** Lineweaver–Burk plots for **1b** in **3** micelle under six pressures: pH 9.5 (0.1 M Tris buffer), 25 °C, dye concentration $5 \times 10^{-6}$ M.

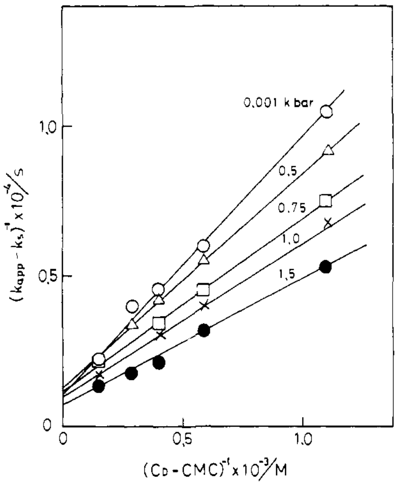


**Figure 2.** Lineweaver–Burk plots for **2b** in **4** micelle under five pressures: pH 8.0 (0.1 M Tris buffer), 25 °C, dye concentration $5 \times 10^{-6}$ M.

Table II. Kinetic Parameters of the Alkaline Fading Reactions Catalyzed by 3 Micelle at Various Pressures and at 25 °C

| | $10^{-4}K$, $\mathrm{M}^{-1}$ | | | | $10^4 k_m$, $\mathrm{s}^{-1}$ | | | |
|---|---|---|---|---|---|---|---|---|
| P, kbar | 1a | 1b | 2a | 2b | 1a | 1b | 2a | 2b |
| 0.001 | 9.12 | 3.21 | 7.40 | 1.22 | 4.36 | 2.68 | 15.18 | 11.91 |
| 0.5 | 6.53 | 2.96 | 6.04 | 1.15 | 4.40 | 3.11 | 12.94 | 10.80 |
| 0.75 | 6.33 | 2.87 | 5.63 | 1.10 | 4.50 | 3.40 | 12.15 | 10.30 |
| 1.0 | 9.73 | 2.71 | 7.73 | 0.95 | 4.28 | 3.80 | 10.27 | 7.68 |
| 1.5 | 14.22 | 3.49 | | | 4.14 | 3.99 | | |
| 2.0 | | 6.01 | | | | 4.04 | | |

Table III. Kinetic Parameters of the Alkaline Fading Reactions Catalyzed by 4 Micelle at Various Pressures and 25 °C

| | $10^{-4}K$, $\mathrm{M}^{-1}$ | | | | $10^4 k_m$, $\mathrm{s}^{-1}$ | | | |
|---|---|---|---|---|---|---|---|---|
| P, kbar | 1a | 1b | 2a | 2b | 1a | 1b | 2a | 2b |
| 0.001 | 9.74 | 4.16 | 6.81 | 0.890 | 1.64 | 3.88 | 5.98 | 11.16 |
| 0.5 | 7.00 | 4.10 | 4.67 | 0.840 | 1.50 | 3.51 | 4.87 | 8.58 |
| 0.75 | 6.41 | 4.07 | 4.38 | 0.837 | 1.46 | 3.40 | 4.39 | 8.40 |
| 1.0 | 5.95 | 4.05 | 4.00 | 0.825 | 1.44 | 3.32 | 4.15 | 7.22 |
| 1.5 | | | | 1.09 | | | | 5.64 |

Table IV. Volume Changes Accompanying the Formation of MD Complexes, and Activation Volumes in the Process of the Product Formation in Micelles at 1 bar and at 25 °C

| | 3 micelle | | | | 4 micelle | | | |
|---|---|---|---|---|---|---|---|---|
| vol changes, cm³/mol | 1a | 1b | **2a** | **2b** | 1a | 1b | **2a** | **2b** |
| $\Delta V \pm 2$ | 14 | 4 | 15 | 4 | 15 | 3 | 15 | 4 |
| $\Delta V^*_m \pm 2$ | 0 | −6 | 7 | 4 | 3 | 4 | 8 | 10 |
| $(\Delta V^*_{OH})_{micelle}$[a] ± 2 | −3 | −9 | 4 | 1 | −5 | −4 | 0 | 2 |

[a] The $\Delta V^*_m$ values were corrected by the pressure dependence of $\beta^p$ ($\Delta V_\beta^*$) to be 3.0 cm³/mol for 3 and 8.4 cm³/mol for 4.

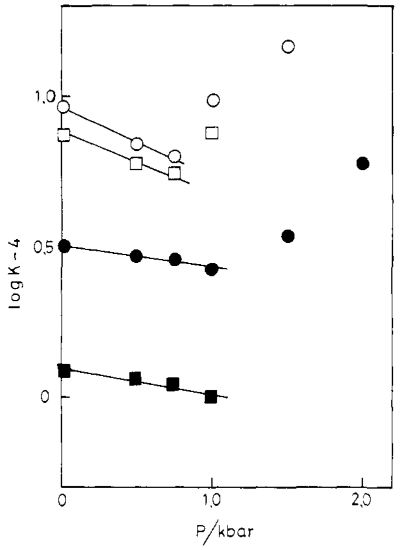


**Figure 3.** Pressure dependence of log *K* in **3** micelle at 25 °C: **1a** (○); **2a** (□); **1b** (●); **2b** (■).

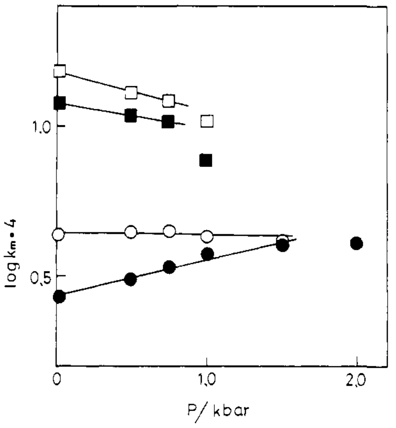


**Figure 4.** Pressure dependence of log $k_m$ in **3** micelle at 25 °C: **1a** (○); **2a** (□); **1b** (●); **2b** (■).

(Table I).[9] The volume changes accompanying the formation of dye–micelle complexes and the activation volumes for the process of the product formation determined from the initial slopes of the plots of log *K* vs. pressure at 1 bar, which are summarized in Table IV.

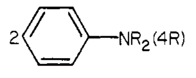
**Dye Binding.** It is generally accepted that acceleration of the alkaline fading reactions of cationic dyes catalyzed by cationic micelles[11-14] is due to the incorporation of dyes into the micellar phase by the hydrophobic attraction in spite of the electrostatic repulsion between dyes and micelles. This is also supported by the facts that the binding constants of **1a** and **2a** are larger than those of **1b** and **2b** in each micelle and that the inversion phenomena of log *K* under high pressure for **3** are in line with the general observation of the maximum of the cmc of surfactants. Therefore, we can conclude that the volume changes for the incorporation of dyes into micellar phases is the only hydrophobic contribution.

The magnitude of $\Delta V$ values per methyl or methylene group for the formation of the hydrophobic interaction from studying model systems[15,16] was obtained at 1–2 cm³/mol per methyl or methylene group, 3–4 cm³/mol per ethyl or ethylene group, and nearly 0 cm³/mol for the benzene ring, respectively. When the

Table V. Activation Volumes for the Alkaline Fading Reaction in Aqueous Solution at 1 bar and at 25 °C

| vol changes, cm³/mol | 1a | 1b | **2a** | **2b** |
|---|---|---|---|---|
| $\Delta V^*_s \pm 1$ | −1.7[a] | 0.7[a]<br>−2[c]<br>0[e] | −10[b]<br>−11.0[d] | −14[b]<br>−11.0[d]<br>−12[e] |
| $\Delta V^*_{H_2O} \pm 1$ | | −15[g] | −12[g] | −11,[g] −12[h] |
| $\Delta V^*_{OH} \pm 1$ | | −0[g] | −1[g] | −1,[g] 0[h] |

[a] pH 9.5, 0.1 M Tris buffer. [b] pH 8.0, 0.1 M Tris buffer. [c] pH 14.0, NaOH solution. [d] pH 12.3, NaOH solution. [e] Reference 2, pH 11.3 for 1b, pH 10.6 for **2b**. [f] Reference 3, pH 12.0. [g] Extrapolated values from the pH dependence in 0.1 M Tris buffer solution. [h] Reference 10.



groups of each dye transfer from the aqueous media to the hydrophobic micellar phase, there is no contribution of benzene ring to $\Delta V$ values. Therefore, the experimental values for the dye binding are reasonable for 3–4 cm³/mol (= 1 cm³/mol × 4) for **1b** and **2b** (R = $CH_3$), and for 14–15 cm³/mol (= 3 cm³/mol × 4) for **1a** and **2a** (R = $C_2H_5$) in **3** and **4** micelles.

**Fading Reaction in Aqueous Solution.** The activation volumes for the alkaline fading reaction of triphenylmethane dyes in 0.1 M Tris buffer solutions are summarized in Table V. The effect of pressure on Tris buffer is negligible.[17] There is a dramatic difference of $\Delta V_s^*$ values between 0.7 and −2.0 cm³/mol for **1a** and **1b** and −10 and −14 cm³/mol for **2a** and **2b**. This was first observed for **1b** and **2b** by Chen and Laidler.[2] Recently, le Noble et al.[10] have studied the pressure effect on the reaction of **2b** over a range of pH values and confirmed that the reaction was due to general base catalysis given by

$$k_s = k_{H_2O} + k_{OH}[OH] + k_{Tris}[Tris] \qquad (6)$$

It was confirmed that the $k_{Tris}$[Tris] term was negligible from the studies of the concentration dependence of Tris buffer on the rate constant of the alkaline fading at the constant pH. The values of $\Delta V^*_{H_2O}$ and $\Delta V^*_{OH}$ (uncorrected values for the dissociation of $H_2O$) in Table V were calculated from the pressure dependence on $k_{OH}$ and $k_{H_2O}$, which were determined from the intercept and the slope at the plots of $k_s$ vs. [OH] at various pressures. The values of $\Delta V^*_{H_2O} = -11$ cm³/mol and $\Delta V^*_{OH} = -1$ cm³/mol for **2b** (corrected for the compression of water) agreed with those of le Noble et al.[10] As the hydroxide ion concentration for the ionization of water in the Tris buffer was increased by the compression, a volume change of 13 cm³/mol corresponding to half of the ionization volume given by $\partial \ln K_w/\partial P = 26$ cm³/mol[18] ($K_w$: ionization constant of water) must be added to the $\Delta V^*_{OH} = 0$

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to −1 cm$^3$/mol. Therefore, the $\Delta V^*_{H_2O}$ values of −11 to −15 cm$^3$/mol are due to the bond formation, and $\Delta V^*_{OH}$ value of 12–13 cm$^3$/mol (corrected value) is the charge neutralization.[18] When we compare the $\Delta V^*_s$ values with $\Delta V^*_{H_2O}$ and $\Delta V^*_{OH}$ values, it is interesting to note that $\Delta V^*_s$ values of reaction 1 contain both water and hydroxide ion reaction pathways, and for reaction 2 mainly a water reaction pathway. We estimated the composite of the effects of two reaction pathways to be more than 90% of water reaction at pH 8.0 for reaction 2 and be less than 30% of water reaction at pH 9.5 for reaction 1.

**Fading Reaction in Micelle.** As dye molecules and hydroxide ions are incorporated into the cationic micellar phase, the reaction field is rich in hydroxide ion and poor in water atmosphere. Therefore, the water-controlled reaction 2 in bulk water is expected to change dramatically into hydroxide ion control in the micellar system, which is identified by the volume transfer from $\Delta V^*_s$ = −10 to −14 cm$^3$/mol in water to $\Delta V^*_m$ = 4–10 cm$^3$/mol in micellar systems. But in reaction 1, the hydroxide ion controlled reaction, we cannot observe any big volume differences between a bulk water and a micellar system except **1b**. As the concentration of hydroxide ion in micelles is larger than in bulk water at the reaction field, i.e., $K_{OH}[OH] \gg k_{H_2O}$ in eq 6, the rate constant for hydroxide ion $(k_{OH})_{micelle}$ in the micellar system is given approximately by eq 7

$$(k_{OH})_{micelle} = k_m/\langle OH\rangle_{micelle} \tag{7}$$

where $\langle OH\rangle_{micelle}$ is the average concentration of the hydroxide ions near the micellar phase, which was determined by the ratio of the number of hydroxide ions to that of surfactant ions $\beta^p$ given by the previous method.[9] The $\beta^p$ values used for **3** were taken from the previous data,[9] and those for **4** are listed in Table I.[19,20] The values of $(\Delta V^*_{OH})_{micelle}$ in Table IV were corrected by $-RT(\partial \ln \beta^p/\partial P) = \Delta V^\beta$, which are 8.4 cm$^3$/mol for **4** and 3.0 cm$^3$/mol for **3**, respectively. The values of $(\Delta V^*_{OH})_{micelle}$ for reaction 2 agree well with those of $\Delta V^*_{OH}$ (uncorrected by the water dissociation). The values of $(\Delta V^*_{OH})_{micelle}$ = 13–16 cm$^3$/mol for reaction 2 corrected by the concentration effect accompanying the water dissociation ($\Delta V_{H_2O}$ = 13 cm$^3$/mol) are a reasonable explanation of the charge neutralization between cationic dyes and hydroxide ions. The corrected $(\Delta V^*_{OH})_{micelle}$ for reaction 1 are in the range from 4 to 10 cm$^3$/mol. The smaller $(\Delta V^*_{OH})_{micelle}$ value for reaction 1 compared with that of reaction 2 is explained by the fact that both reactants of dyes and hydroxide ion are desolvated partly by strong hydrophobic interactions between the dye molecules and micelles. It is very interesting that micelles dramatically change the reaction mechanism of water control of reaction 2 into the hydroxide ion control, and also are a useful tool in understanding the reaction mechanism of general base catalysis.

## Experimental Section

**Materials.** The **3** (Wako Chemicals Ind.) and **4** (Aldrich Chemical Co.) were extracted with petroleum ether and then recrystallized 3–4 times with ethanol. The cmc values of **3** under high pressure used the previous data.[9] The cmc value of **4** in aqueous solution at 25 °C by the method using *p*-nitrophenoxide ion ($\lambda_{max}$ = 400 nm) as a probe is 3.5 × 10$^{-3}$ M. This corresponds to 3.55 × 10$^3$ M from the surface tension measurement[21] and 3.82 × 10$^{-3}$ M from the conductance measurement.[21] The cmc values of **4** in the experimental conditions (0.1 M Tris buffer), with each triphenylmethane dye at 25 °C, are 1.5 × 10$^{-3}$ M for **2a** and **2b** (pH 8.0) and 3.0 × 10$^{-3}$ M for **1a** and **1b** (pH 9.5), respectively. They were determined by the measurements of the alkaline fading reaction of each dye at various concentrations of **4**. The relative ratio of cmc values ($(cmc)^p/(cmc)^1$) at 0.1 M Tris buffer solution (pH 8.0) and at 25 °C up to 2 kbar in Table I were determined by the dye method using phenolphthalate ion ($\lambda_{max}$ = 560). Dyes (**1a**, **1b**, **2a**, and **2b**) of Wako (special grade) were used without further purification.

**Apparatus and Procedure.** The high-pressure apparatus[22] used was the clamp type made of SUS 630 stainless steel, which took 2–3 min at least to compress to a certain experimental pressure after mixing dyes and catalysts in the high-pressure cell. The reaction in the micelle system was performed at pH 8.0 (**2a** and **2b**) and pH 9.5 (**1a** and **1b**) in 0.1 M Tris buffer solutions up to 2 kbar at 25 ± 0.5 °C. The reaction at pH 14.0 (**1a** and **1b**) and pH 12.3 (**2a** and **2b**) was in a system of sodium hydroxide solution (no buffer system). The reaction under high pressure was monitored directly by observing the change in the optical density at 620 nm (**2a** and **2b**) and at 595 nm (**1a** and **1b**) by means of a Hitachi 340 spectrophotometer. The dyes of (5 × 10$^6$ M) used in the present

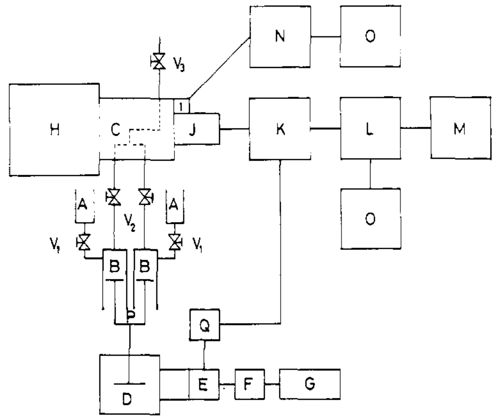

Figure 5. HPSF diagram: (A) sample reservoir; (B) high-pressure vessel; (C) high-pressure optical and mixing chamber; (D) intensifierer; (E) solenoid valve; (F) gas controller; (G) gas reservoir; (H) light source and monocrometer; (I) strain gauge; (J) photomultiplier; (K) amplifier and trigger circuit; (L) digital memory; (M) monitor scope; (N) dynamic strain amplifier; (O) recorder; (P) piston; (Q) valve controller; (V) valve.

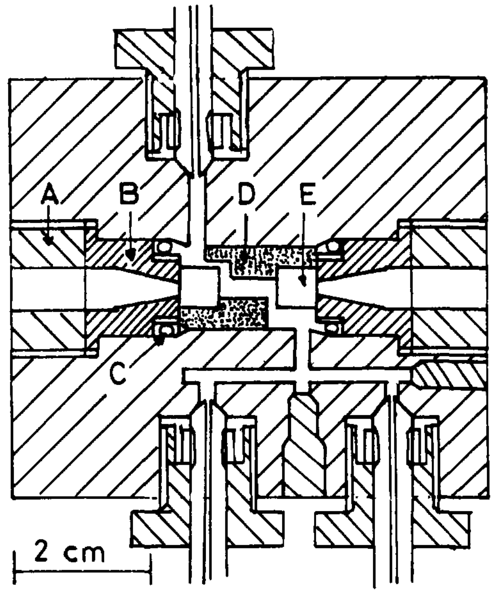

Figure 6. The high-pressure optical and mixing chamber: (A) screw plug; (B) window support plug; (C) O-ring; (D) Teflon spacer; (E) quartz window (6 mm thickness).

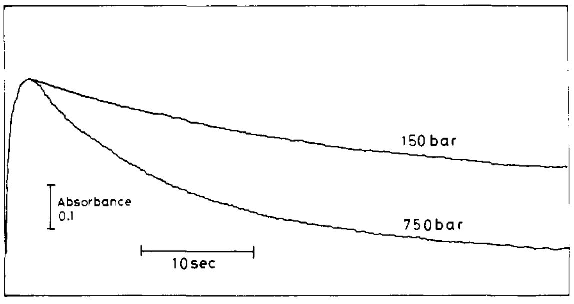

Figure 7. Typical stopped-flow traces of the reaction **2** at 25 °C; [**2b**] = 2 × 10$^{-6}$ M, pH = 12.3.

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investigation were found to obey Beer's law up to a concentration of $10^{-5}$ M. For the rapid fading reaction of less than 2–3 min, we used the high-pressure stopped flow apparatus.

**High-Pressure Stopped Flow Technique.** The schematic diagram of the high-pressure stopped flow (HPSF) apparatus is shown in Figure 5. The high-pressure optical and mixing chamber (SUS 630 stainless steel) was a T-jet type with two quartz windows (Figure 6). The spectrophotometric system was a Union Giken stopped flow apparatus. Pressure was measured by a strain gauge (Kyowa KPC-5C1-11) which was calibrated against a Heise precision bourdon gauge (full scale 50 000 psi, unit scale 50 psi). The temperature was kept constant by circulating the thermostated water surrounding the solution reservoirs and the mixing chamber. The temperature of the mixing chamber was monitored by the thermocouple ±0.1 °C.

The operation was carried out as follows. First, the valves $V_1$ and $V_3$ were closed and the valve $V_2$ was opened after filling the high-pressure sample reservoir B with each reactant solution. Then, as soon as the solenoid valve E was opened in order to push the piston P upward through the intensifier D, each reactant solution was mixed up in the mixing chamber and was immediately compressed. The desired pressure was set up by controlling the gaseous pressure in the gas reservoir G and it took 2 s at least to generate a constant pressure in the mixing chamber. The temperature of the solution accompanying the compression up to 1 kbar increased about 2 °C more than in the experimental condition and 2 s was needed before the constant temperature was recovered. With this operation, the reliable kinetic data were obtained 2 s after the operation of E. Figure 7 is a typical example of the alkaline fading reaction of **2b** at 150 and 750 bar.

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**Registry No. 1a**, 2390-59-2; **1b**, 548-62-9; **2a**, 633-03-4; **2b**, 569-64-2; $OH^-$, 14280-30-9; $H_2O$, 7732-18-5.

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# Clarification of the Hole-Size Cation-Diameter Relationship in Crown Ethers and a New Method for Determining Calcium Cation Homogeneous Equilibrium Binding Constants

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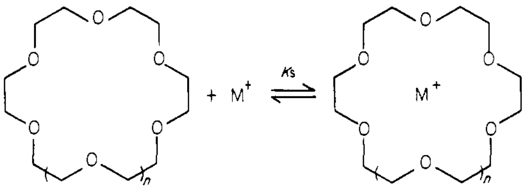
**Abstract:** The equilibrium stability constants for the homologous series of crown ethers ranging from 12-crown-4 to 24-crown-8 with sodium, potassium, ammonium, and calcium cations have been determined in anhydrous methanol solution. The key finding of this study is that the widely recounted "hole-size-selectivity" principle is not applicable to this series of simple macrocycles. Instead, it is observed that the $K^+$ cation is bound most strongly by all of the macrocycles in this series and that the strongest binding for all of the cations in this study is observed with 18-crown-6. In order to determine the stability constants for $Ca^{2+}$ in MeOH by ion selective electrode methods, it was necessary to develop a new, competitive technique utilizing the $Na^+$-selective glass electrode. We report such a method here which is rapid, reliable, and applicable in organic solvents that are incompatible with commercial $Ca^{2+}$-selective electrodes.

Not long after Pedersen's discovery of cation binding by crown ethers,[1] he and Frensdorff reported that picrate extraction constants showed an apparent relationship between cation diameter and crown ether hole size.[2] The same sort of relationship was observed for the relatively inflexible, three-dimensional cation binders devised by Lehn and called cryptands.[3] The hole-size concept was widely accepted and led to the design and synthesis of a remarkable variety of cation and molecule hosts including the spherands[4] and cavitands.[5] These inflexible cation binders appeared to affirm the hole-size concept as well. In contrast to this, Reisse and Michaux recently reported that the enthalpies of binding between $Na^+$ and $K^+$ and the 12–18-membered crown rings do not correlate with hole sizes.[6]

Some of our previous work with the macrocycles we have called "lariat ethers" has suggested that the more flexible polyether systems do not abide by this "hole-size rule".[7] We therefore felt it would be useful to survey stability constants, $K_s$ (see eq 1), for as broad a range as possible of homologous crowns. Since much of the binding data thus far accumulated have been taken in anhydrous MeOH solution, we wished to conduct our studies in this solvent. Furthermore, we were anxious to have a full set of $Ca^{2+}$ binding, as well as $Na^+$, $K^+$, and $NH_4^+$, data in this solvent. To date, virtually all of the $Ca^{2+}$ stability constants that are available have been determined by calorimetric methods[8] since commercial, $Ca^{2+}$-selective ion selective electrodes (ISE) are incompatible with organic solvents.[9] We have therefore developed a convenient method for determining $Ca^{2+}$ stability constants in solvents other than water. We have applied the method to the 12-crown-4 to 24-crown-8 series of crown ethers in anhydrous



(1)

$n = -1, 0, 1, 2, 3$
$M = Na^+, K^+, NH_4^+, Ca^{2+}$

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(9) This is due to their construction. The epoxy body of the electrode is unstable to the nonaqueous solvent, and the ionophore contained within it dissolves away.

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