Macromolecules 1991, *24*, 1249–1252

Different Protonation Behavior of Two Poly(methacrylic acid) Derivatives Containing *N*-Acylglycine and *N*-Acyl-β-alanine Residues: Thermodynamic and FT-IR Studies

Rolando Barbucci,* Mario Casolaro, Agnese Magnani, and Clelia Roncolini

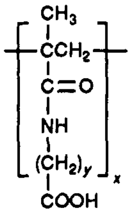
Dipartimento di Chimica, Università di Siena, Pian dei Mantellini 44, 53100 Siena, Italy

Received March 19, 1990; Revised Manuscript Received August 29, 1990

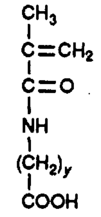
ABSTRACT: Two poly(methacrylic acid) derivatives containing amide and carboxyl groups in the side chain of the vinyl structure, namely poly(*N*-methacryloylglycine) (PMP_1) and poly(*N*-methacryloyl-β-alanine) (PMP_2), were studied in aqueous solution at 25 °C by thermodynamic (potentiometry, calorimetry, and viscometry) and spectroscopic (FT-IR) techniques together with the corresponding monomers. The basicity constants (log K) of both polymers were compared with those of other polymers with a vinyl backbone, and the anomalous behavior of PMP_2 was revealed by the n value of the modified Henderson–Hasselbalch equation. The slight exothermic enthalpy changes were pH-dependent and the $-\Delta H°$ increased with the degree of protonation α for PMP_1 but decreased for PMP_2. The extension of hydration of the amide group revealed by FT-IR spectra was the main factor controlling the different protonation behavior of the two polyelectrolytes.

Introduction

In previous papers[1,2] we reported the synthesis and thermodynamics of protonation of vinyl polymers with poly(acrylic acid) structure in which an amido group was inserted as the side substituent. It was shown that a lengthening of a polymethylene spacer between the amide and carboxyl groups increases the hydrophobicity superimposed on the hydrophilic quality imparted by the peptide linkage. Thus, we undertook the study of these polymers with special attention to the role played by their hydrophobicity in conformational changes.[3] Now we extend these studies to two methacrylic acid derivatives and their polymers.



PMP_1 ($y = 1$)
PMP_2 ($y = 2$)



MP_1 ($y = 1$)
MP_2 ($y = 2$)

The insertion of the α-methyl hydrophobic group on the skeleton of poly(methacrylic acid) drastically changes its behavior in aqueous solution.[4] The present paper deals with the protonation behavior of these polymers in aqueous solution, using viscometric, potentiometric, calorimetric, and FT-IR spectroscopic techniques, and their behavior has been compared with the other vinyl polymers previously studied.[5,6]

Results and Discussion

Viscometry and Potentiometry. The viscometric behavior and protonation equilibria of the two polymers and corresponding monomers were studied in aqueous solution at 25 °C and ionic strength 0.1 M NaCl (Figure 1 and Table I). In the ionized form, the higher reduced viscosity is related to the extended state of the macromolecular chain; the lowest viscosity value is achieved in the completely neutral form, i.e., at the degree of protonation $\alpha \approx 1$ (Figure 1).

The behavior of both polymers is well described by the Henderson–Hasselbalch equation modified by Katchalsky and Spitnik for polyelectrolytes.[7] The proton association constant, $K°$, is higher for the half-ionized polymer than the $K°$ for the corresponding monomer as a consequence of the charge of the polyelectrolyte chain. In fact when electrostatic repulsion is decreased, i.e., at a high degree of protonation α, when almost complete protonation is reached, the log K values of both the polymers approach those of the corresponding monomers (Figure 1). The increase in methylene groups from PMP_1 to PMP_2 and from MP_1 to MP_2 is reflected in a higher log $K°$ value for the compound with the longer chain, and the difference in log $K°$ values is the same ($\Delta \log K° = 0.6$) for both classes of compounds (Table I). This is an expected behavior that can be related to that of many amino acid derivatives.[8]

With comparison of the n values of the modified Henderson–Hasselbalch equation of the present series of polymers containing a vinyl backbone,[2,5,6] a close relationship is observed between this parameter and the hydrophobic character of the polymers (Table II). High hydrophobicity corresponds to low polyelectrolyte behavior, bringing the value of n close to 1. Vinyl polymers with a lateral amide group always show $n > 1$. When the hydrophobic character of the monomer unit is increased, for example, by lengthening the aliphatic chain from poly[*N*-(2-aminoethyl)acrylamide] (P_1) to poly[*N*-(3-aminopropyl)acrylamide] (P_2), or from poly[*N*-acryloylglycine] (PAP_1) to poly[*N*-acryloyl-6-aminocaproic acid] (PAP_5), or by adding a methyl group from poly[1-(1-piperazinylcarbonyl)ethylene] (P_3), to poly[1-[(4-methyl-1-piperazinyl)carbonyl]ethylene] (P_4), or from PAP_1 to PMP_1, a lower value of n is obtained (Table II). On the contrary, PMP_2 seems to behave in a different manner: from PMP_1 to PMP_2 the n value increases even though a methylene group has been inserted in the chain. This suggests that in PMP_2 the protonation behavior is anomalous (see below).

FT-IR Spectrometry. The FT-IR spectroscopic results relative to PMP_1 and PMP_2 at two different pHs are shown in Figure 2. For comparison, the deconvoluted difference spectra of PAP_1 at the same pH[2] are also given. The chemical structure of PMP_1 differs from that of PAP_1 only by virtue of a methyl group on the skeleton of the polymer. Hence, the new band that appears at 1487 cm^{-1}

0024-9297/91/2224-1249$02.50/0

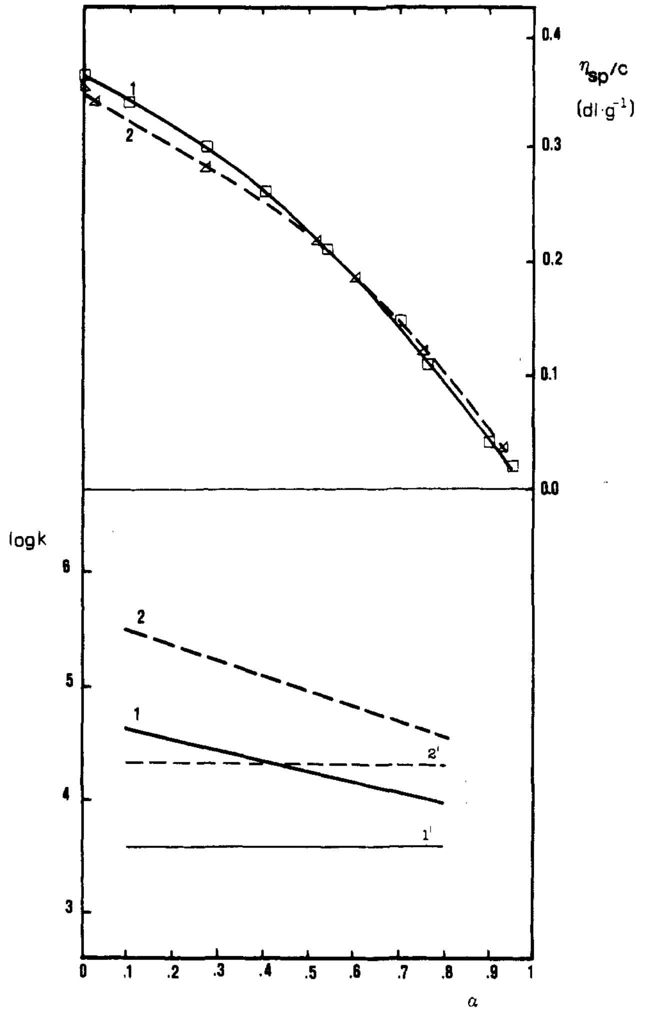


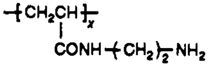
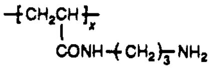
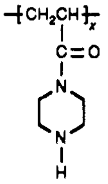
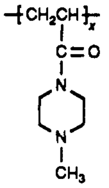
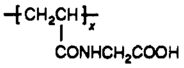
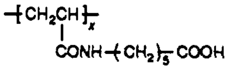
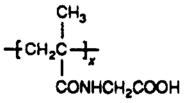
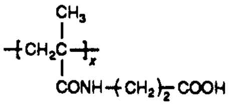
Figure 1. Reduced viscosity (η_{sp}/C) and basicity constants (log K) vs degree of protonation (α) for polymers PMP$_1$ (1, —) and PMP$_2$ (2, - - -) (with corresponding monomers 1′ and 2′) at 25 °C in 0.1 M NaCl.

Table I
Protonation Equilibria of Polymers and Related Monomers at 25 °C in 0.1 M NaCl Solution (Reaction: $L^- + H^+ \rightleftharpoons LH$; for Polymers, L^- Refers to the Monomer Unit in the Anionic Charged Form)

compd	$[\eta]^{DMF}_{30°}$	log K^a
PMP$_1$	0.27	4.26 (4) + 0.40 (2) log [(1 − α)/α]^b
MP$_1$		3.597 (4)^c
PMP$_2$	0.20	4.92 (3) + 0.61 (1) log [(1 − α)/α]^b
MP$_2$		4.33 (1)^c

[a] log K = log $K°$ + (n − 1) log [(1 − α)/α]. [b] Values in parentheses are the deviations obtained as: $d_i = x_i - \bar{x}$. [c] Values in parentheses are the standard deviations estimated as reported in ref 16.

Table II
Basicity Constants of Vinyl Polymers Following the Modified Henderson–Hasselbalch Equation
log K = log $K°$ + (n − 1) log [(1 − α)/α]

compd		log $K°$	n	ref
P$_1$	 $-\!\!\left[CH_2\underset{CONH(CH_2)_2NH_2}{CH}\right]_x\!\!-$	8.23	1.48	6
P$_2$	 $-\!\!\left[CH_2\underset{CONH(CH_2)_3NH_2}{CH}\right]_x\!\!-$	9.31	1.37	6
P$_3$		7.126	1.40	5
P$_4$		6.088	1.27	5
PAP$_1$	 $-\!\!\left[CH_2\underset{CONHCH_2COOH}{CH}\right]_x\!\!-$	4.32	1.50	2
PAP$_2$	 $-\!\!\left[CH_2\underset{CONH(CH_2)_5COOH}{CH}\right]_x\!\!-$	5.42	1.31	2
PMP$_1$	 $-\!\!\left[CH_2\overset{CH_3}{\underset{CONHCH_2COOH}{C}}\right]_x\!\!-$	4.26	1.40	this paper
PMP$_2$	 $-\!\!\left[CH_2\overset{CH_3}{\underset{CONH(CH_2)_2COOH}{C}}\right]_x\!\!-$	4.92	1.61	this paper

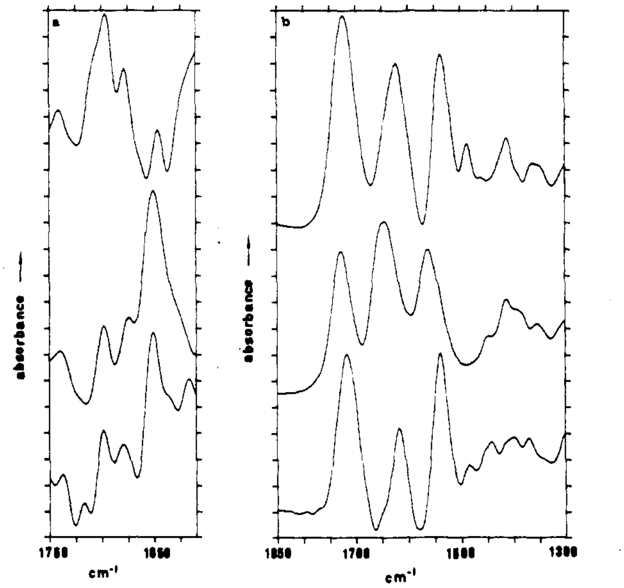


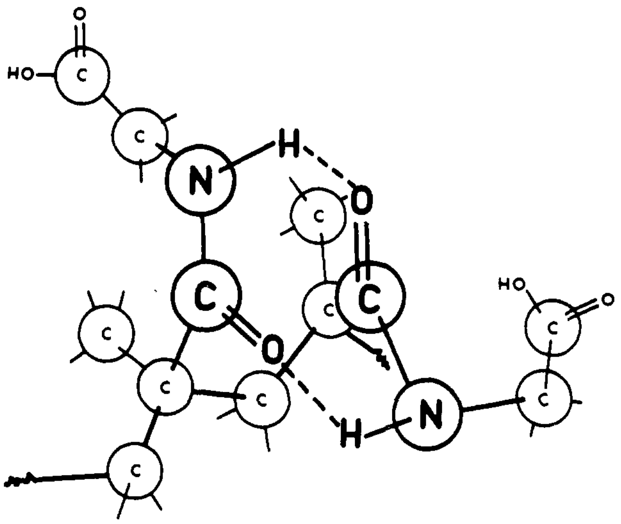
Figure 2. (a) Deconvoluted spectral region relative to the amide I vibration mode of PMP$_1$ at pH = 13.0 (top), of PAP$_1$ at pH = 10.5 (middle), and of PMP$_2$ at pH = 10.2 (bottom). (b) Difference spectra in the 1850–1300-cm^{-1} range of PMP$_1$ at pH = 2.9 (top), of PAP$_1$ at pH = 2.5 (middle), and of PMP$_2$ at pH = 2.9 (bottom).

in PMP$_1$ (Figure 2b) can reasonably be assigned to the methyl deformation mode.[9]

The presence of the methyl group in the backbone of the vinyl chain does not particularly affect the absorption frequency of the carboxyl group in protonated or unprotonated forms. In fact the absorption frequencies relative to the stretching of the –COOH and –COO$^-$ groups in both the PMP$_1$ (1726 cm^{-1} COOH, 1582 cm^{-1} COO$^-$) and PAP$_1$ (1728 cm^{-1} COOH, 1580 cm^{-1} COO$^-$) are approximately the same.

Comparing the amide I spectral region of the two polymers at high pH (Figure 2a), we can see that, even if the amide I bands occur at the same frequencies for PMP$_1$ and PAP$_1$, the corresponding intensities behave in exactly the opposite way, in the sense that the band at 1650 cm^{-1} due to the formation of hydrogen bonding between the amide group and water molecules[2] has a lower intensity in the PMP$_1$ spectrum than in that of PAP$_1$, while the opposite trend is observed for the bands at 1696 and 1678 cm^{-1} due to the "free" amide group.[2] This means the effect

Chart I



of the methyl group is to decrease the hydration of the amidic group in the side chain.

On the other hand, the spectrum of PMP_1 at low pH shows an amide I band at 1620 cm^{-1} (Figure 2b). This frequency is lower than that in the PMP_1 spectrum at high pH (1650 cm^{-1}) and was previously assigned[2] to an amide group involved in strong hydrogen bonding with another amide group of a close pendant chain. Thus the methyl group on the skeleton removes the water molecules from the amide group favoring interaction between the pendant chains at low pH (Chart I). This occurs when the macromolecule assumes coil form, as seen by viscometric titration. As a matter of fact, in the spectrum of the PAP_1 polymer at low pH, the amide I occurs at 1650 cm^{-1} as at high pH,[2] showing that the amide groups remain strongly hydrated even in the protonated polymer.

The PMP_2 polymer differs from the PMP_1 polymer only in an additional methylene group between the amide and carboxyl groups. The lengthening effect of the alkyl chain causes a drop of the carboxylic (1728 cm^{-1} PMP_1, 1716 cm^{-1} PMP_2) and carboxylate (1582 cm^{-1} PMP_1, 1565 cm^{-1} PMP_2) C=O stretching modes as also observed in the previously studied polyacrylate analogues[2] and poly(amido amino acids).[10]

Comparison of the amide I spectral region of PMP_2 and PMP_1 at low pH (Figure 2b) shows that they behave in a similar manner in the protonated form, and thus the pendant chains also interact with each other in the neutralized PMP_2 polymer. This similarity is not shown by the IR spectrum at high pH: in the amide I region (Figure 2a) the band at 1650 cm^{-1} has greater intensity than those at 1698 and 1679 cm^{-1}. Since the absorption coefficients of the "free" and "hydrogen-bonded" C=O are not very different,[11] the greater intensity of the 1650-cm^{-1} band confirms that the amide group of PMP_2 is strongly hydrated, as suggested by the *n* parameter of the modified Henderson–Hasselbalch equation of PMP_2, which is also higher than that of PMP_2 (Table II).

Calorimetry. The calorimetric results ($-\Delta H°$) for the carboxyl protonation are reported in Table III together with the other thermodynamic functions $-\Delta G°$ and $\Delta S°$. It appears that the protonation of both polymers is exothermic like those of the corresponding monomers. The heat effects of PMP_1 and PMP_2 are quite similar to $\alpha \approx 0.5$ and lower than those of the MP_1 and MP_2 monomers, which are themselves similar.

The lower enthalpy changes in both polymers compared to the corresponding monomers can be attributed to the larger desolvation effect, which generates an endothermic contribution produced by the macromolecular chain during the protonation process. In fact the hydration shell of a polyelectrolyte embraces several repeating units, leading to a more pronounced release of water molecules during the neutralization process than that obtained with the monomer. This behavior is also observed in other polymers.[2,10] Figure 3 shows the trend of the three thermodynamic functions ($-\Delta G°$, $-\Delta H°$, and $\Delta S°$) against the degree of protonation α of the polymers PMP_1 and PMP_2. The variation of $-\Delta H°$ versus α follows an opposite trend in the two polymers: the slope of $-\Delta H°/\alpha$ is positive

Table III
Thermodynamic Changes for the Protonation of Polymers and Monomers in 0.1 M NaCl at 25 °C (Reaction: $L^- + H^+ \rightleftharpoons LH$)[d]

compd	$-\Delta G°$[a]	$-\Delta H°$[a]	$\Delta S°$[c]
PMP_1	5.81 (5)[b]	0.26 (15)[b]	18.60 (5)[b]
MP_1	4.905 (5)	0.95 (2)	13.27 (7)
PMP_2	6.71 (4)[b]	0.37 (8)[b]	21.30 (4)[b]
MP_2	5.90 (1)	0.96 (2)	16.59 (10)

[a] In kcal·mol^{-1}. [b] Values calculated at $\alpha = 0.5$. [c] In cal·mol^{-1}·K^{-1}. [d] Values in parentheses are the estimated standard deviations of the last significant figure.

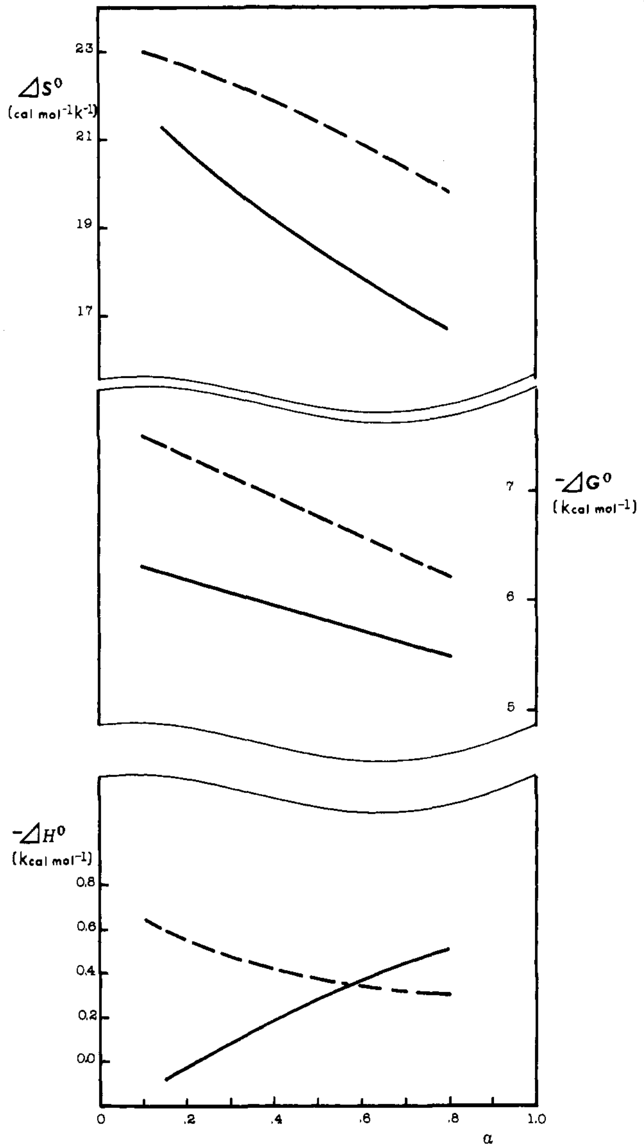
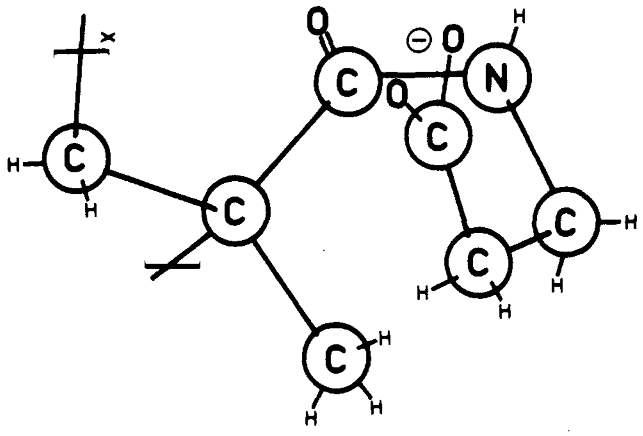


Figure 3. Thermodynamic functions ($-\Delta G°$, $-\Delta H°$, and $\Delta S°$) for protonation of PMP_1 (—) and PMP_2 (- - -) at 25 °C in 0.1 M NaCl.

Chart II



for PMP_1 but negative or almost constant for PMP_2 (Figure 3). In both polymers two contributions have to be taken into account to understand the protonation process: the desolvation, due to the neutralization process, which becomes progressively lower with the decrease in charge of the macromolecule, and the interaction through hydrogen bonds between the amide moieties of the pendant groups.

The FT-IR spectra reveal that the PMP_2 polymer is highly solvated at basic pH on both the hydrophilic, i.e., amide and carboxylate groups, more so than PMP_1, probably because the longer aliphatic chain of PMP_2 renders the polymer more flexible and the two methylene groups can interact hydrophobically with the methyl group on the main chain. This structure renders the amide moiety more open and free to hydration (Chart II). Thus the neutralization process of PMP_2 always provokes a longer desolvation effect than that of PMP_1. In the latter case the exothermic effect increases with increasing α because during neutralization the release of water molecules becomes progressively smaller and smaller and the exothermic effect greater. For PMP_2 the heat is almost constant because the two thermal effects equalize; this means that PMP_2 releases more water molecules than PMP_1, which causes α to increase. This is evident from the trend of $\Delta S°/\alpha$ for PMP_1 and PMP_2: both $\Delta S°$'s decrease with increasing α, but the curve of the more solvated PMP_2 remains higher than that of PMP_1, and the slope is smaller.

Experimental Section

Synthesis. The monomers and corresponding polymers were synthesized according to a Winston procedure.[12]

***N*-Methacryloylglycine[13] (MP_1).** Elem anal. Calcd for $C_6H_9NO_3$: C, 50.3; H, 6.3; N, 9.8. Found: C, 49.9; H, 6.3; N, 9.7.

Poly(*N*-methacryloylglycine) (PMP_1). Intrinsic viscosity (DMF, 30 °C) $[\eta] = 0.27$ dL·g^{-1}. Elem anal. Calcd for $(C_6H_9NO_3{\cdot}H_2O)_x$: C, 44.7; H, 6.9; N, 8.7. Found: C, 44.7; H, 6.7; N, 8.6.

***N*-Methacryloyl-*β*-alanine[14] (MP_2).** Elem anal. Calcd for $C_7H_{11}NO_3$: C, 53.5; H, 7.0; N, 8.9. Found: C, 53.1; H, 7.0; N, 8.7.

Poly(*N*-methacryloyl-*β*-alanine) (PMP_2). Intrinsic viscosity (DMF, 30 °C) $[\eta] = 0.20$ dL·g^{-1}. Elem anal. Calcd for $(C_7H_{11}NO_3{\cdot}{}^3/_4H_2O)_x$: C, 49.3; H, 7.4; N, 8.2. Found: C, 49.5; H, 6.6; N, 8.2.

Spectroscopic Measurements. A Varian XL 200 spectrometer was used to record proton NMR spectra on samples of ca. 20 mg of each compound dissolved in 0.5 mL of Me_2SO-d_6. Chemical shifts were measured, taking the residual Me_2SO peak as the internal reference signal at 2.48 ppm. A Perkin-Elmer M1800 spectrophotometer was used with a Data Station 7500 professional computer to record FT-IR spectra of the compounds by a previously reported procedure.[2] The concentration of the polymer solution was about 4×10^{-2} mol dm^{-3}. The pH values of the solution were adjusted by addition of small amounts of aqueous HCl or aqueous NaOH solution.

The difference spectra are obtained by subtracting the spectrum of water from the spectrum of the polymer solution according to the null criterion.

The spectral digital processing adopted to obtain the deconvoluted difference spectra has previously been described.[2]

Viscometric Measurements. A Cannon-Ubbelohde 50 E 998 viscometer was used to obtain viscometric data. For the determination of the intrinsic viscosities the polymers were dissolved in DMF (31 mg of PMP_1 in 15 mL of solvent; 34 mg of PMP_2 in 20 mL of solvent) and held at 30 °C. The flow time for DMF solvent was 170 s.

The viscometric titration point data were obtained at 25 °C by dissolving a known amount of polymer (30 mg of PMP_1; 40 mg of PMP_2) in 25 mL of 0.1 M NaCl containing an excess of 0.1 M sodium hydroxide. Hydrochloric acid solution (0.1 M) was used as titrant in the stepwise neutralization of the carboxylate groups of the polymers. The FITH program running on an Olivetti M24 computer[15] was used to calculate α from the basicity constants obtained above.

Potentiometric Measurements. The basicity constants of the two polymers and the corresponding monomers were determined in aqueous solution by potentiometric techniques using a previously reported procedure.[15] Duplicate experiments were carried out for both polymers and monomers in 0.1 M NaCl to determine the basicity constants with the previously described calculation programs: APPARK on the Olivetti M20 computer[15] for polymers and SUPERQUAD on an Olivetti M28 computer[16] for monomers.

Calorimetric Measurements. Continuous calorimetric titrations in the isothermal mode were carried out with a previously described procedure[15] using a Tronac calorimeter (Model 1250). The $-\Delta H°$ values were calculated at each α value and then averaged, ensuring that they fell within the 0.05 α-range. The standard deviations were calculated on each group of $-\Delta H°$ values chosen in the range of α considered.

Acknowledgment. We thank the Italian Ministry of Education and the National Research Council (CNR) for financial support.

References and Notes

(1) Barbucci, R.; Casolaro, M.; Ferruti, P.; Magnani, A.; Roncolini, C. *Polymer* **1989**, *30*, 1751.
(2) Barbucci, R.; Casolaro, M.; Magnani, A. *Makromol. Chem.* **1989**, *190*, 2627.
(3) Borden, K. A.; Eum, K. M.; Langley, K. H.; Tirrell, D. A. *Macromolecules* **1987**, *20*, 454.
(4) Crescenzi, V.; Quadrifoglio, F.; Delben, F. *J. Polym. Sci., Polym. Phys. Ed.* **1972**, *10*, 357.
(5) Barbucci, R.; Casolaro, M.; Ferruti, P.; Tanzi, M. C.; Grassi, L.; Barozzi, C. *Makromol. Chem.* **1984**, *185*, 1525.
(6) Barbucci, R.; Casolaro, M.; Nocentini, M.; Reginato, G.; Ferruti, P. *Makromol. Chem.* **1986**, *187*, 1953.
(7) Katchalsky, A.; Spitnik, P. *J. Polym. Sci.* **1947**, *2*, 432.
(8) Zief, M.; Edsall, J. T. *J. Am. Chem. Soc.* **1937**, *59*, 2245. King, E. J.; King, G. W. *J. Am. Chem. Soc.* **1956**, *78*, 1089. Barbucci, R.; Casolaro, M.; Nocentini, M.; Corezzi, S.; Ferruti, P.; Barone, V. *Macromolecules* **1986**, *19*, 37.
(9) Bellamy, L. J. In *The Infrared Spectra of Complex Molecules*; Chapman and Hall, Ltd.: London, 1975.
(10) Barbucci, R.; Casolaro, M.; Ferruti, P.; Nocentini, M. *Macromolecules* **1986**, *19*, 1856.
(11) Coleman, M. M.; Lee, K. H.; Skrovanek, D. J.; Painter, P. C. *Macromolecules* **1986**, *19*, 2149.
(12) Winston, A.; Kirchner, D. *Macromolecules* **1978**, *11*, 597.
(13) Ioffe, D. V.; Kuznetsov, S. G. *Zh. Obshch. Khim.* **1959**, *29*, 3804.
(14) Winston, A.; McLaughlin, G. R. *J. Polym. Sci., Polym. Chem. Ed.* **1976**, *14*, 2155.
(15) Barbucci, R.; Casolaro, M.; Danzo, N.; Barone, V.; Ferruti, P.; Angeloni, A. *Macromolecules* **1983**, *16*, 456.
(16) Gans, P.; Sabatini, A.; Vacca, A. *Inorg. Chim. Acta* **1983**, *79*, 219.

Registry No. PMP_1, 30602-15-4; PMP_2, 131618-71-8; MP_1, 23578-45-2; MP_2, 131618-73-0.