JOURNAL OF MATERIALS SCIENCE: MATERIALS IN MEDICINE 13 (2002) 155–158

The effect of sodium alginate on the crystal growth of calcium carbonate

F. MANOLI, E. DALAS*
University of Patras, Department of Chemistry, GR-26500 Patras, Greece

The effect of sodium alginate in supersaturated solutions of calcium carbonate was investigated under plethostatic conditions. The rates of crystal growth measured in the presence of sodium alginate at concentrations as low as 0.83×10^{-7} mol dm^{-3} were drastically reduced. Kinetic analysis according to a Langmuir-type adsorption isotherm led to the calculation of an affinity constant $K_{aff} = 999.8 \times 10^{-4}$ mol dm^{-3}. The apparent order found from kinetic data was 3.0 ± 0.2 suggesting a surface nucleation mechanism.

1. Introduction

Calcium carbonate is found as different polymorphs, consisting, in the order of increasing solubility, of calcite, aragonite, vaterite, calcium carbonate monohydrate and calcium carbonate hexahydrate. Under physiological conditions, the results of calcium carbonate deposition in biological systems can be seen in the formation of mollusc shells, egg shells, the exoskeleton of arthropods, pearls and corals [1, 2]. There are, however, also pathological aspects of biomineralization, including the formation of the human atherosclerotic aorta (9% carbonate mineral) [3, 4], kidney stones, pancreatic calcification and gallstones [5–7].

Alginic acid is found in all brown seaweeds at depths not exceeding 40 m, i.e., the depth to which sunlight can penetrate. The principal source is the giant kelp *Marcocystis pyrifera* along the coast of North and South America, New Zealand, Australia, and Africa. It is composed of varying proportions of D-mannuronic and L-guluronic acids, linked β-1.4. The molecular weight is about 120.000. The polymer replaces pectin in the brown algae, from which it is extracted by sodium hydroxide treatment. Alginates (usually as the sodium salt) are widely used in the food industry, in surgery as resorbable material, and in the pharmaceutical and cosmetic industries [8–10].

In the present work, an attempt was made to evaluate the antimineralization activity of sodium alginate in the calcium carbonate system by the constant composition method [11, 12]. The conditions of the experiments selected in the present work were such that the supersaturated solutions employed were stable for periods up to two days, their stability verified by the constant pH and calcium concentration. The low degree of supersaturation is a better representation of the physiological environment, where the free calcium concentration is rather low [13]. The supersaturated solution were seeded with synthetically prepared calcite crystals in the presence of sodium alginate and the calcium carbonate crystallization was followed.

2. Experimental procedure

All the experiments were done at 25 ± 0.1 °C, in a thermostated double-walled Pyrex vessel. Calcium carbonate supersaturated solutions of 0.2 dm^3 total volume were prepared from calcium nitrate, sodium bicarbonate and potassium nitrate stock solutions, (Merck pro Analysi) as described in detail elsewhere [14] along with sodium alginate (Sigma, from *Macrocystis pyrifera*). The arrangement was such that the air volume over the aqueous phase was kept at a minimum, so that the partial pressure of the carbon dioxide may be considered to be constant [12]. The pH in all experiments here was adjusted at 8.50 by the addition of standard potassium hydroxide solution (Merk, titrisol). Following verification of the stability of the supersaturated solution 100 mg of calcite synthetically prepared [12] was added to the solution. The BET specific surface area of calcite determined by N_2 adsorption (Perkin Elmer sorptometer 212D) was found to be 3.2 m^2 g^{-1}.

Precipitation reaction in all cases started immediately after the introduction of the seed crystals in the crystallization medium. The pH change (0.003 pH units), concomitant with the formation of calcium carbonate triggered the addition of titrants, with the stoichiometry of calcium carbonate from the coupled burettes of an appropriately modified pH stat (Metrohm, 614). The concentration of the titrant in the two burettes was calculated as follows:

$$\text{titrant 1} : (Cx + 2x)(Ca(NO_3))_2,$$
$$\text{titrant 2} : Cx(Na_2CO_3) + 2x(NaHCO_3) + 2y(KOH),$$

where x is the molar concentration of calcium nitrate or sodium bicarbonate in the working solution and y the

*Author to whom all correspondence should be addressed.

TABLE I Crystallization of calcite on calcite seed crystals in the presence of sodium alginate, SA_t, pH 8.50, 25 °C, total calcium (Ca_t) = total carbonate (C_t)

Exp no.	$Ca_t(10^{-3}$mol dm$^{-3})$	$C_{KNO3}(10^{-2}$mol dm$^{-3})$	$SA_t(10^{-7}$mol dm$^{-3})$	$\Delta G_{calc.}$(KJ mol^{-1})	$R(10^{-6}$mol m^{-1}m$^{-2})$
blank	3	6.0	—	− 3.08	25.3
T-A7	3	6.0	1.67	− 3.08	8.4
T-A14	2.5	5.0	1.67	− 2.72	4.7
T-A16	2	4.0	1.67	− 2.27	1.4
T-A17	1.75	3.5	1.67	− 2.00	1.2
T-A10	3	6.0	0.83	− 3.08	14.3
T-A12	3	6.0	8.33	− 3.08	5.6
T-A21	3	6.0	16.70	− 3.08	1.5

amount of potassium hydroxide required for the pH adjustment in the working solution. For maintenance of the ionic strength constant an amount 2C of inert electrolyte (potassium nitrate) was added in the working solution where C is a constant (expressing how many times the titrants are more concentrated than the working solution). In our experiments, C was chosen as 10. The choice of the best value for C requires preliminary experiments.

Random sampling during the course of reaction verified that the solution supersaturation was kept constant [12]. Employing a constant solution composition has the advantage of determining the reaction rates very accurately, since the initial conditions are kept constant for a large part of the crystallization reaction. The samples withdrawn during the reaction were filtered through membrane filters (Gelman, 0.1 μm); the filtrates were analyzed for calcium by atomic absorption spectroscopy (Varian 1200) and the solid residues by powder X-ray diffraction (Phillips PW 1830/1840 using Cukα radiation Ni filter), scanning electron microscopy (Jeol GSM 5200), FT-IR spectroscopy (Perkin Elmer 16-PC FT-IR using KBr pellets) and thermogravimetric analysis (TGA, Du Pont 910). The rate of calcium carbonate crystallization were taken from the plots of titrant addition as a function of time, normalized for the total surface area of the seed crystals.

3. Results and discussion

The experimental conditions are summarized in Table I. The solution speciation was computed from the appropriate equilibria between calcium and carbonate species, mass-balance equations for calcium and carbonate and the electroneutrality condition by successive approximations for the ionic strength [15]. For the estimate of the activity coefficients the Davies equation [16] was used. The solid phases were found to be calcite from the examination of the powder X-ray diffraction spectra (h kl : 102, 104, 110, 113, 202, 108, 116, 212, 214, 300) [17], of the FT-IR spectra (bands : 1800, 1420, 876, 714 cm^{-1}) [18, 19]. The thermogravimetric analysis exclude the formation of hydrated calcium carbonate salts. Well grown calcite crystals in the presence of sodium alginate may be seen in the scanning electron micrographs in Fig. 1 [20].

As it can be seen from Table I, the presence of sodium alginate in the supersaturated solution retarded the crystallization of calcite even at concentration as low as 0.83×10^{-7} mol dm^{-3}. It is interesting to note that the retardation effect becomes more significant in comparison to pure calcite system, as sodium alginate concentration increased from 0.83×10^{-7} to 16.7×10^{-7} mol dm^{-3}. It is possible that, as in the case with most of the ionic inhibitors [21, 22], the active molecules occupy active growth sites of the substrate thus making them unavailable. The assumption that the inhibitory effect of sodium alginate is mainly due to adsorption and subsequent blocking of the active growth site was tested by assuming Langmuir-type adsorption. Inherent to the Langmuir formalism is the assumption

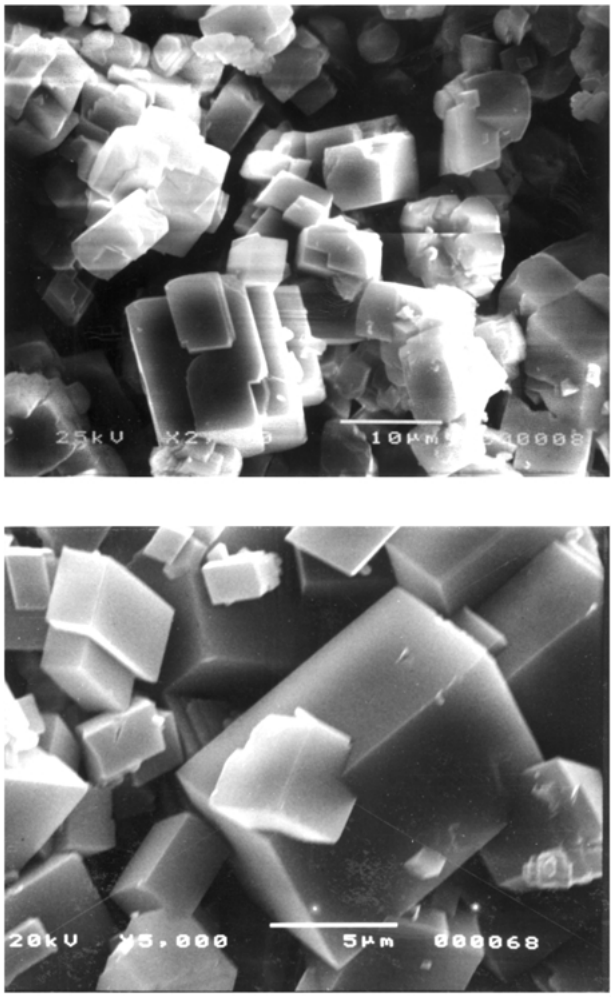

Figure 1 Scanning electron microscopy of calcite crystallization on calcite seed crystal in the presence of 16.7×10^{-7} M sodium alginate.

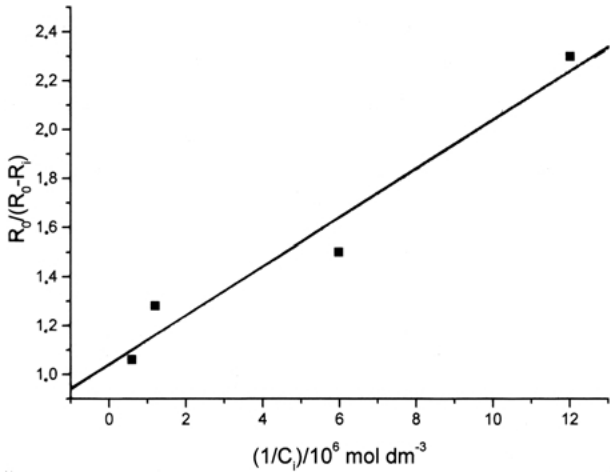

Figure 2 Kinetics of calcite crystal growth in the presence of various concentrations of sodium alginate, according to a Langmuir-type kinetic model; 25 °C, pH 8.50, $Ca_t = 10^{-3}$ M.

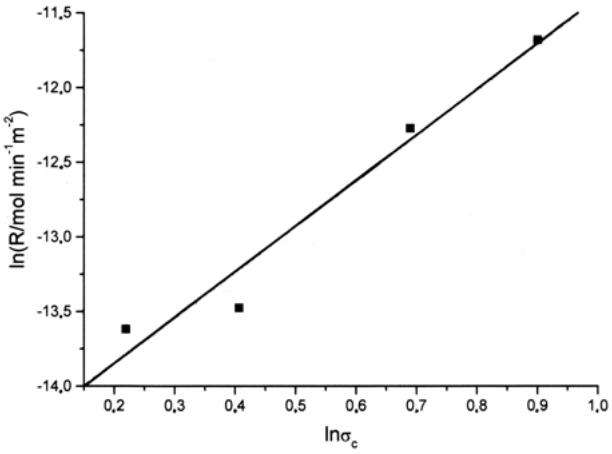

Figure 3 Kinetics of calcite crystallization on calcite seed crystals in the presence of 1.67×10^{-7}M sodium alginate at pH 8.50, 25 °C.

that the adsorption free energy is constant over the entire adsorbent surface. Thus the relationship between the crystallization rates in the absence, R_o, and in the presence of the sodium alginate inhibitor, R_i, as a function of the ratio of the inhibitor concentration c_i in the supersaturated solutions is given by [23]

$$R_o/(R_o - R_i) = 1 + \left(\frac{k_d}{k_a}\right)\frac{1}{c_i} \tag{1}$$

where k_a and k_d are the specific rate constants for adsorption and desorption respectively. The ratio of the above mentioned rate constants is defined as the affinity constant K_{aff}. Plots according to Equation 1 shown in Fig. 2 resulted in a straight line, from the slope of which a value of 999.8×104 mol dm^{-3} was obtained for the affinity constant. A list of other values of affinity constants for other inhibitors for the crystal of calcite are given in Table II.

The degree of supersaturation with respect to calcite is defined as:

$$\Omega_c = (\mathrm{IP})/\mathrm{K}_{s,c} \tag{2}$$

where IP is the activity product of the phase considered at the experimental conditions and $K_{s,c} = 3.311 \times 10^{-9}$ [26] at equilibrium. The driving force for the formation of calcite is the change in Gibbs free energy for going from the supersaturated to equilibrium:

$$\Delta \mathrm{G}_{\mathrm{calc.}} = -\frac{R_g T}{\upsilon} \ln \Omega_c \tag{3}$$

where ν is the number of ions in the lattice; for calcite ν = 2. In all cases, the measured crystal-growth rate, R, was proportional to the relative supersaturation, σ, with respect to calcite:

$$\sigma_c = \Omega_c^{1/2} - 1 \tag{4}$$

$$R = k\sigma_c^n \tag{5}$$

where k is the precipitation rate constant and n the apparent order of reaction. Kinetics plots according to Equation 5 gave a satisfactory fit, as can be seen in Fig. 3. From the linear plots, a value of $n = 3 \pm 0.2$ was obtained for the crystallization of calcite in the presence of sodium alginate which is indicative of a surface nucleation controlled mechanism [27]. This value is in contrast to earlier results where $n = 2 \pm 0.2$ [14, 18–21]. As may be seen, the presence of sodium alginate in the supersaturated solution not only inhibits the crystal growth process but also alters the apparent order of crystallization, suggesting a different mechanism from the classical spiral growth mechanism ($n \leq 2$). In conclusion, sodium alginate may be a useful drug not only as an antitumor agent [28] but in many other cases of pathological calcification. Similar results were obtained for metallocene dichlorides [29, 30].

TABLE II Comparative data on the inhibition of calcium carbonate precipitation

Inhibitor	K_{aff} ($\times 10^{-4}$ mol dm^{-3})	Ref.
Phosphate	5842	14
Melitic acid	200	23
Ethylenediamini-tetra-bis-methylene phosphonic acid	1000	24
2-Dihydroxyphosphonyl-2-hydroxy-propionic acid	1350	25
1.3-bis[(1-phenyl-1-dihydroxy-phosphonyl)methyl]-2-imidiazolidenon	1580	25
sodium alginate	999.8	This work

References

1. K. SIMKISS, *Am Zool*. **24** (1984) 847.
2. G. KRAMPITZ and G. GRASER, *Angew. Chem. Int. Ed. Engl*. **27** (1988) 1145.
3. B. B. TOMAZIC, W. E. BROWN and F. J. SCHOEN, *J. Biomed. Mater. Res*. **28** (1994) 35.
4. K. SCHMID, W. O. MCSHARRY, C. H. PAMEIJER and J. P. BINETTE, *Atherosclerosis* **37** (1980) 199.
5. H. SARLES, *Gastroenterology* **66** (1974) 604.
6. H. J. VERINE, *Bull Comp. Pathol*. **5(3)** (1973).
7. D. K. AGARWAL, G. CHOUDHURT and J. KUMAR, *Scand. J. Gasteoenterol*. **28(7)** (1993) 613.
8. M. BREWER and T. SCOTT, ''Concise Encyclopedia of Biochemistry'' (de Gruyter, Berlin, 1983) p. 17.
9. H. F. MARK, N. G. GAYLORD and N. M. BIKALES, ''Encyclopedia of Polymer Science and Technology'' (Wiley, New York, 1969) vol. 11, pp. 409–411.
10. H. F. MARK, N. M. BIKALES, C. G. OVERBERGER, G. MENGES and J. I. KROSCHWIT, ''Encyclopedia of Polymer Science and Engineering'' (Wiley, New York, 1988) vol. 13, pp. 121–122.
11. M. B. TOMSON and G. H. NANCOLLAS, *Science* **200** (1977) 1059.
12. T. F. KAZMIERCZAK, M. B. TOMSON and G. H. NANCOLLAS, *J. Phys. Chem*. **86** (1982) 103.
13. D. GLEESON, K. A. HOOD and G. M. MURPHY, *Gastreonterology* **102(5)** (1992) 1707.
14. E. GIANNIMARAS and P. G. KOUTSOUKOS, *J. Colloid Int. Sci*. **116** (1987) 423.
15. G. H. NANCOLLAS, ''Interactions in Electrolyte Solutions'' (Elsevier, Amsterdam, 1966).
16. C. W. DAVIES, ''Ion Association'' (Butterworths, London, 1962).
17. ASTM Card file No. 5-0586.
18. F. MANOLI and E. DALAS, *J. Cryst. Growth* **204** (1999) 369.
19. F. MANOLI, S. KOUTSOPOULOS and E. DALAS, *ibid.* **182** (1997) 116.
20. E. DALAS and P. G. KOUTSOUKOS, *Langmuir* **4** (1998) 907.
21. S. J. ZAWACKI and G. H. NANCOLLAS, *Ind. Cryst*. **61** (1984) 132.
22. Z. AMZAD, P. G. KOUTSOUKOS and G. H. NANCOLLAS, *J. Colloid Int. Sci*. **101** (1984) 250.
23. Z. AMZAD, *Langmuir* **3** (1987) 224.
24. M. M. REDDY, in ''Chemistry of Wastewater Technology'' edited by A. J. Rubin (Ann Arbor Science Pub., Ann Arbor, Mich., 1978) p. 31.
25. A. G. XYLA, J. MIKROYANNIDIS and P. G. KOUTSOUKOS, *J. Colloid Int. Sci*. **153(2)** (1992) 537.
26. N. L. PLUMMER, T. M. L. WIGLEY and D. C. PARKHURST, *Am. J. Sci*. **278** (1978) 179.
27. H. E. L. MADSEN, *J. Cryst. Growth* **80** (1987) 450.
28. F. MICHIO and N. TERUKAZU, *Carbohydr. Res*. **224** (1992) 342.
29. E. DALAS, N. KLOURAS and C. MANIATIS, *Langmuir* **8** (1992) 343.
30. S. KOUTSOPOULOS, A. KONTOGEORGOU, J. PETROHEILOS and E. DALAS, *J. Mater. Sci.: Mater. Med.* **9** (1998) 421.

Received 4 May 2000
and accepted 27 February 2001