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# Are bacteria capable of precipitating magnesite?

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## Abstract

Laboratory experiments were conducted to investigate the possible (bio-) chemical pathways involved in the low-temperature nucleation of magnesite, $MgCO_3$. Two compounds known for their power to prevent hydrolysis (amorphous silica and magnesium trisilicate) were tested in contact with a slowly desiccating solution of magnesium bicarbonate. Amorphous silica gave rise to precipitates of nesquehonite instead of the more usual magnesium hydroxide carbonate, thereby illustrating its influence on hydrolysis. Magnesium tri-silicate led to X-ray amorphous precipitates. The presence of reducing compounds, in particular the strongly reducing hydrogen gas, does not lead to the low-temperature nucleation of magnesite.

*Key words:* magnesite; laboratory experiments; reducing conditions; microbial assemble.

## Introduction

In the present paper experiments will be discussed that pertain to the low-temperature formation of magnesite. From three different sides the suggestion has been made that microorganisms would be involved in the low-temperature formation of magnesite. In the first place there is the mineralogical analysis of meteorite ALH84001. This meteorite is thought to have been part of rocks occurring on planet Mars. In the second place there is the fact that the Recent magnesite found in caves on the island of Kauai (Hawaii) by Léveillé et al. (2000a, b), occurs in close association with algae. In the third place there is the observation by Nash et al. (2011) on the low-temperature formation of magnesite in modern-day corals of the Great Barrier Reef, Australia. Apparently all these occurrences of modern magnesite contradict the fact that until recently magnesite could be prepared in the laboratory only at high temperatures and/or high $CO_2$ pressures (Hänchen et al., 2008; Munz et al., 2009; Zhang et al., 2012). The low-temperature (30 to 50 ºC) syntheses of magnesite described by Felmy et al. (2012) involve carbon dioxide pressures of 90 atmosphere and more.

From a magnesium bicarbonate solution usually a hydrated form of magnesium carbonate or one of the more complex magnesium hydroxide carbonates precipitates. The apparent reluctance of anhydrous magnesium carbonate to precipitate under condition of low temperature and atmospheric pressure has been attributed to a dehydration barrier of magnesium cations in aqueous solution (e.g., Lippmann, 1973). However, the suggestion of a dehydration barrier preventing the formation of anhydrous magnesium carbonate needs to be reconsidered after the successful low-temperature syntheses of the anhydrous double salt magnesium sodium carbonate at 25 °C (Deelman, 1984).

Having noted that there really is no longer any need to study the possible effects of hydration phenomena, attention was focused on the tendency of magnesium cations to lead to the low-temperature formation of complex carbonates containing hydroxyl groups. Could it be that the slightly acidic conditions caused by microbiological activity such as for example sulphate reduction, are responsible for the removal of hydroxyl groups, and so favour the nucleation of anhydrous magnesium carbonate?

In the laboratory experiments described here, attention was given first to attempts to prevent, or at least to drive back, the effects of hydrolysis. Next was a series of tests conducted with compounds known as reducing agents (hydrogen iodide, glycol, hydroquinone, and ascorbic acid). Even the strongly reducing hydrogen sulphide gas and the ultimate reducing agent, hydrogen gas, were tested in combination with magnesium in solution. No magnesite formed in all of these experiments. Experiments conducted after the ones described here have led to the successful low-temperature (40 °C) synthesis of magnesite (Deelman, 1999). Alves dos Anjos et al. (2011) have demonstrated the low-temperature syntheses of magnesite to be reproducible. The present paper is based on the fact that only after 1999 the outcome of the previously conducted experiments could be fully appreciated.

## Magnesite and microbes

Many papers have appeared on meteorite ALH84001, in particular because this meteorite might well contain the first evidence of life on Mars. Carbonate globules consisting of calcite, magnesite, ankerite and siderite were found in the meteorite (Harvey and McSween Jr., 1996). In ALH84001 there are ultra-small structures (of some 20 to 100 nanometer) described as nanobacteria and considered to be evidence of life on Mars (McSween Jr., 1997). According to Kirschvink et al. (1997) paleomagnetic study showed the carbonates in the ALH84001 meteorite to have formed at temperatures below 110 °C. Two different sulphides (pyrrhotite and greigite) were recognized in the samples from ALH84001 by McKay et al. (1996) suggesting a possible origin by way of bacterial sulphate reduction. Controversy still exists over the temperature and the process of formation of the magnesite mineral in the Martian meteorite. It must be assumed that low temperatures (substantially lower than 100 °C) can ever have sustained forms of life on Mars; hence the magnesite must have formed at temperatures below 100 °C. McSween Jr. and Harvey (1998) as well as Warren (1998) claimed low-temperature formation of the magnesite on Mars and suggested a mechanism involving flooding by water followed by evaporation and precipitation from solution.

Much more relevant than the preceding cosmochemical interpretations are actual descriptions of occurrences of modern magnesite. Undoubtedly one of the most curious occurrences of such modern magnesite is that on the Clipperton atoll (Bourrouilh-Le Jan et al., 1985). The coralline rim of the atoll encloses a lagoon filled with a distinctly layered body of water. The top layer consists of brackish water, turbid from high amounts of organic matter in suspension, with a pH of 8.7 to 9.3. The deeper parts of the lagoon water contain high concentrations of *Thiobacillus*, anaerobic bacteria and marine yeasts. The water of the deeper parts of the lagoon is quite different from the upper layer: it is crystal clear, of normal salinity, and pH = 6.2.

The slightly acidic water exerts a strong influence on the contacting carbonates, in that dissolution (notably of aragonite) takes place. Calcareous fragments dredged from the deeper parts of the Clipperton lagoon were strongly leached, discoloured and very friable. Modern magnesite was found disseminated within these leached carbonate fragments.

According to Perthuisot et al. (1990), who had found modern magnesite in the Holocene sediments of sabkha El Melah (near Zarzis, Tunisia), notably sulphate reducing bacteria must be responsible for the formation of magnesite. This conclusion was based on the absence of any dissolved sulphate in the interstitial brines of the organic-detrital deposits in which the magnesite had been found. The black-coloured sediments continuously evolved hydrogen sulphide and methane.

The papers by Léveillé et al. (2000a, b) are of considerable interest, because they record deposits of modern magnesite from caves in *basaltic* rocks on the island of Kauai, Hawaii. A number of the deposits of magnesite (along with aragonite, calcite, monohydrocalcite, dolomite and hydromagnesite) are associated with living algal mats. "A common feature of some of the thicker, completely mineralized speleothems is a laminated texture, resembling microstromatolites" (Léveillé et al., 2000a, p. 615). The fact that the magnesite found on the island of Kauai, Hawaii is of authigenic origin can not be doubted, since the mineral magnesite (much like magnesium hydroxide carbonate and monohydrocalcite) do not occur in the surrounding rocks. [In this paper the chemical term "magnesium hydroxide carbonate" is used to replace the confusing term "hydromagnesite"]. Even so the exact relation to the microorganisms of the algal mat was difficult to prove ("Desiccated and mineralized microbialites contain few preserved cells, and instead are dominated by mineralized material, which is typically laminated on a micrometer scale": Léveillé et al., 2000a, p. 351).

A comparable observation has been made by Nash et al. (2011), but there coralline algae apparently were involved in the low-temperature formation of magnesite. Crustose algae (*Hydrolithon onkodes*) were collected at a depth of 3 to 5 m on the Great Barrier Reef near Heron Island, Australia. The algae had actually been living up to the moment of sample preparation. Nash et al. (2001), using X-ray diffraction as well as electron microscopy, found pure magnesite in the cell spaces of these crustose algae.

## Experiments

Various means are known to prevent hydrolysis or to reduce its effects. For example the use of compounds capable of base exchange could be considered. In my experiments M-99 and M-101 pure magnesium bicarbonate would dry out while in contact with 0.25 gram (g) of amorphous silica, $SiO_2.nH_2O$ (BAKER Analyzed art. no. 0254). The magnesium bicarbonate solution contained 0.2 g/dm$^3$ magnesium hydroxide carbonate: MERCK art. no. 5828. Gasometric analysis of a random sample of this particular lot of magnesium hydroxide carbonate showed that it contains 71 wt.% $MgCO_3$, whereas titration showed it to contain 72 wt.% $MgCO_3$. The choice of amorphous silica was based on the paper by Bogue (1920), in which it was shown that colloidal silica is capable of reducing the amount of hydroxyl ions during the hydrolysis of sodium silicate.

In experiment M-99 ten (10) ml of the described magnesium bicarbonate solution would stream once per hour (in an automatically controlled set-up) into a slowly rotating Petri dish containing the amorphous silica, and dry completely during the rest of each hour at a constant temperature of 27 °C. Heating was controlled by way of a heat lamp above the rotating Petri dish (Figure 1). After ten days X-ray diffraction showed that nesquehonite $MgCO_3.3\ H_2O$ had been formed. This

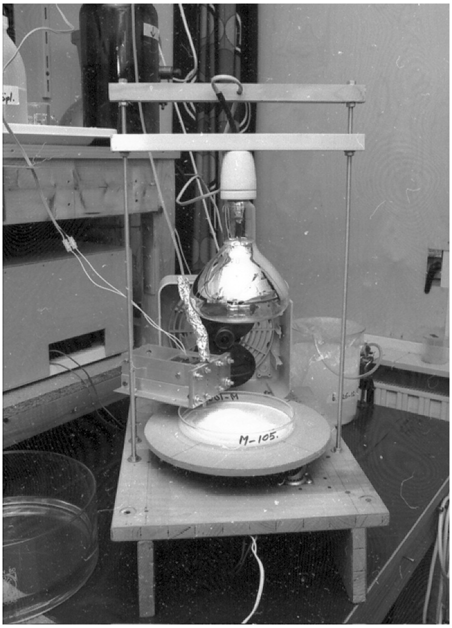

Figure 1. Automatic set-up delivering once per hour 10 ml of a magnesium bicarbonate solution onto a slowly rotating Petri dish.

observation proves that the amorphous silica does have an effect, because without it magnesium hydroxide carbonate would have formed. The set-up of the rotating Petri-dish is such that at the very site of the desiccating bicarbonate solution low values of $pCO_2$ exist. As shown for example by Langmuir (1965) (his Figure 5, p. 748) such low values of $pCO_2$ lead to the precipitation of magnesium hydroxide carbonate. In experiment M-101 two $dm^3$ of the previously described magnesium bicarbonate plus 0.25 g amorphous silica were desiccated at a constant temperature of 30 °C in a large glass bowl placed on an electric heater. In this case too, the compound formed upon desiccation was nesquehonite. The formation of nesquehonite in experiment M-101 was not unexpected, because the magnesium bicarbonate solution slowly desiccating in a (high model) glass bowl will give rise to nesquehonite (unless high temperatures are involved). Because carbon dioxide is heavier than air, geometry of the test set-up determines to a large degree the local $pCO_2$. A large glass bowl will tend to store the heavy carbon dioxide, but the low rim of a Petri dish ensures rapid escape of the carbon dioxide.

In experiment M-105 an attempt was made to study the possible effect of another compound known for its base exchange capacity: magnesium tri-silicate. One gram magnesium tri-silicate (purum, FLUKA 63144) was spread out in a slowly rotating Petri dish, and each hour 10 ml of the previously described magnesium bicarbonate solution would be added automatically and desiccate at 35 °C. No crystalline form of a magnesium carbonate resulted. The precipitate formed was X-ray amorphous. (According to Mignardi et al., 2011 the formation of nesquehonite instead of an amorphous carbonate from a magnesium bicarbonate solution would depend on the $CO_2$: Mg ratio of the solution.)

From these three experiments the conclusion can be drawn, that apparently the two compounds with base exchange capacity, are not actively involved in the low-temperature nucleation of magnesite.

Perhaps compounds known as reducing agents would be more effective in preventing hydrolysis. The following methods of preventing the formation of hydroxyl ions are known: adding hydrazine, adding carbon monoxide, or adding saturated aliphatic alcohols such as propanol, butanol, glycol, or propanediol to the aqueous solution. Well-known reducing agents $H_2$, HI, $H_2S$, $Na_2SO_3$ and aldehydes such as hydroquinone (1,4 dihydroxybenzole) may exert much the same power in preventing the precipitation of $Mg(OH)_2$ and therefore preventing the formation of magnesium hydroxide carbonates. In this manner the possible formation of $MgCO_3$ might be favoured.

Because of practical difficulties in obtaining pure hydrogen iodide, an intermediate reaction was used in experiment M-109. In that experiment a combination of KI and $I_2$ was used. At least part of the KI would change into $I_2$ plus HI in the presence of weak acids such as carbonic acid. Both KI and $I_2$ were used, because in addition to the power of reduction of HI, a reaction between iodine and hydroxyl groups according to

$$I_2 + 2OH^- = I^- + IO^- + H_2O$$

will take place, capable of driving back the amounts of hydroxyl groups. However no magnesite at all could be found. The precipitate from experiment M-109 was X-ray amorphous.

A number of additional experiments in which glycol, hydroquinone or ascorbic acid were added to a pure magnesium bicarbonate solution similarly met with no success (hydroquinone and ascorbic acid are “Reductone” in the sense of Von Euler et al., 1957). Perhaps the stronger reducing agents $H_2$ or $H_2S$, in the absence of any oxygen, would be capable of inducing the low-temperature nucleation of magnesite. Therefore a new kind of experiment was needed, in which carbon dioxide from the magnesium bicarbonate solution would be removed gradually from the solution in the absence of any oxygen. In the presence of any oxygen, hydrogen sulphide will quickly decompose. The prerequisite “in the absence of oxygen” might well be substantiated by comparing the formation of $MgCO_3$ and $FeCO_3$ in the presence and absence of oxygen. The formation of iron hydroxide carbonates can be prevented only, when excluding all oxygen from possible contact with the iron bicarbonate solution. Perhaps much the same mechanism would be taking place in the case of the formation of magnesium hydroxide carbonate. My next experiment had to involve the gradual removal of carbon dioxide from a magnesium bicarbonate solution in the absence of any oxygen. What was needed was a set-up, which is essentially air tight (which in itself is difficult to achieve). After deciding that nitrogen gas could be used to displace all of the air in the equipment to be used, the problem might well be solved. A first attempt was made using a large plastic aquarium, fitted with a thick plexiglass lid, resting on what was thought to be an air tight rubber seal. In the aquarium a glass beaker was placed with a pure magnesium bicarbonate solution. The actual experiment consisted of bubbling pure nitrogen gas through this solution. An overpressure valve ensured the gradual escape of the ensuing carbon dioxide - nitrogen mixture. Indeed a precipitate formed after 14 days, but it turned out to be X-ray amorphous.

Perhaps using nitrogen was not a really good idea. Therefore in my next experiment (M-117) the beaker with the magnesium bicarbonate solution was placed in the carefully closed plastic aquarium together with an open beaker with alkaline solution of pyrogallol (benzene -1, 2, 3-triol: a well-known reducing agent) as well as an open beaker with KOH pellets (to absorb the carbon dioxide from the magnesium bicarbonate solution). After 7 days a precipitate was discerned in the glass beaker with the magnesium bicarbonate solution, but this too turned out to be X-ray amorphous.

An adequate method to withdraw carbon dioxide from water saturated with carbon dioxide while introducing more and more of the extremely powerful reducing agent hydrogen gas was used in my experiment M-126. Into a 1.5 litre jam jar partially filled with 1 litre distilled water through which carbon dioxide had been bubbling for 72 hours, 0.48 g finely powdered pure magnesium metal (MERCK art. no. 5815) was poured. The reaction between magnesium metal and the bicarbonate solution initiates at once the formation of hydrogen gas. Quickly a broad-rimmed beer glass containing 20 g KOH tablets (MERCK art. no. 5033) was placed inside the jam jar in such a way that it would float on the solution. The inside diameter of the jam jar in combination with the outside diameter of the beer glass were such that the floating beer glass with the potassium hydroxide could not tumble over (Figure 2). Thus the KOH pellets would start absorbing carbon dioxide (and water) from the solution as soon as the gas-tight lid of the jam jar was closed, but the KOH pellets would not be able to contact the solution itself. After 4 weeks at room temperature a precipitate had formed, adhering the inside of the jam jar. It was removed and subjected to X-ray diffraction. What had been formed was the mineral dypingite $Mg_5(CO_3)_4(OH)_2.8H_2O$ and not magnesite. Apparently not only hydrolysis, but also hydration was involved in this experiment.

Although the rubber sealing ring in combination with the steel closure device must have been able to guarantee the exclusion of any oxygen from the jam jar, a second attempt was made in order to be sure about the outcome of the reaction. Experiment M-129 was an exact duplication of M-126, and the outcome was also the same: dypingite. Apparently, the presence of abundant hydrogen gas had not been able to prevent hydrolysis (see for a review of the results Table 1).

Figure 2. Hermetically sealed jam-jar filled with $CO_2$-saturated water to which magnesium (metal) powder had been added; the floating beer glass contains KOH pellets needed to absorb the dissolved carbon dioxide.

Ultimately the successful low-temperature (40 ºC) synthesis of magnesite first took place in my experiment M-211 (see the X-ray diffractogram of Figure 3), and has been described in the paper Deelman (1999).

## Discussion

It is possible to dissolve powdered magnesite in carbon dioxide-saturated water. When withdrawing the carbon dioxide from the ensuing magnesium bicarbonate solution, or when evaporating the solution as a whole, hydrated magnesium carbonates or the hydroxyl group-containing magnesium hydroxide carbonates will form. Whatever the approach chosen, it is impossible to precipitate anhydrous magnesium carbonate again from the bicarbonate solution at low temperatures (around 25 ºC) and atmospheric pressure. As a result Halla (1962) concluded that apparently an *irreversible* reaction is involved in the dissolution/precipitation of magnesite (and dolomite as well). Once the observation concerning irreversibility has been made, it should inevitably lead to serious consequences. Foremost stands the observation, that in the absence of equilibrium the ”solubility” concept can no longer be used. As stated by Lewin (1960, p.41): “… the solubility product principle can be derived for stable equilibrium conditions only”. As soon as the strict requirement of chemical equilibrium is no longer given, as is the case with the low-temperature formation of magnesite, the concept of solubility can no longer be applied. The occurrence of metastable phases (in this case the hydrous magnesium carbonates and the various forms of magnesium hydroxide carbonates: see Langmuir, 1965) effectively rules out the use of Nernst’s (1889) concept of

Table 1. Summary of experiments and precipitates formed.

| Experiment number | Mg bicarbonate solution + .... | Temperature | Precipitate |
|---|---|---|---|
| M-99 | amorphous silica | 27 ºC | nesquehonite |
| M-101 | amorphous silica | 33 ºC | nesquehonite |
| M-105 | Mg trisilicate | room temperature | amorphous |
| M-109 | KI + $I_2$ | room temperature | amorphous |
| M-116 | bubbling $N_2$ throughthe Mg bicarbonate solution | room temperature | amorphous |
| M-117 | Mg bicarbonate solution next to KOH and pyrogallol | room temperature | amorphous |
| M-126 | Mg powder in $CO_2$-rich water next to KOH | room temperature | dypingite |
| M-129 | Mg powder in $CO_2$-rich water next to KOH | room temperature | dypingite |

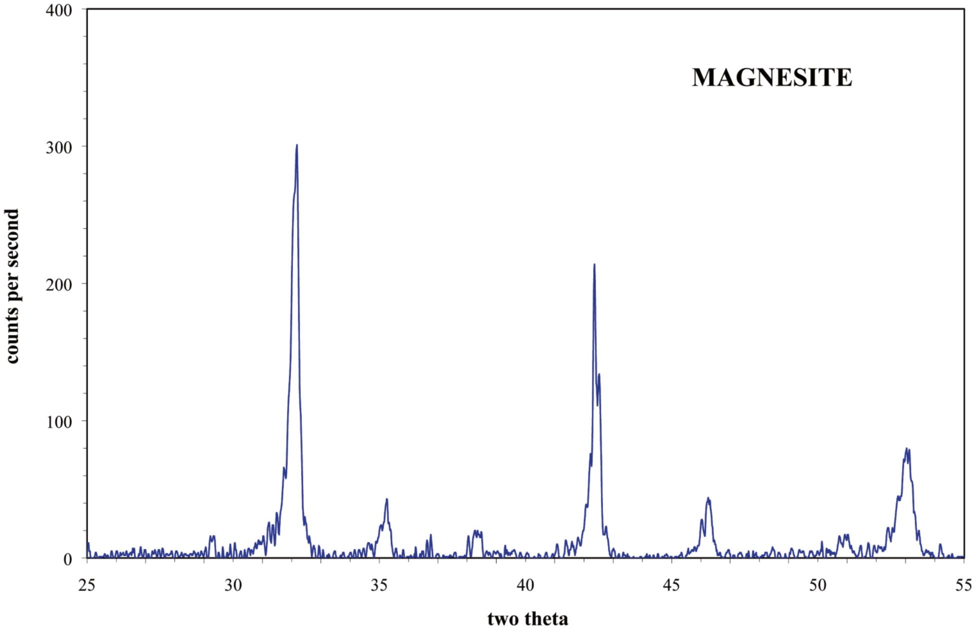

Figure 3. X-Ray diffractogram (Cu-Kα radiation) of magnesite synthesized in an inorganic laboratory experiment conducted at a temperature of 40 °C (copyright: J.C. Deelman, 2003).

solubility. According to Bénézeth et al. (2011) the available information on the solubility of magnesite is sparse and contradictory in itself. Like so many authors before them, Bénézeth et al. (2011) did not consider the possibility that the dissolution / precipitation of magnesite consists of an irreversible reaction.

According to Saldi et al. (2009) the low-temperature formation of magnesite would be limited by very slow step advancement rates during the nucleation process. However, those observations were based on experiments conducted at temperatures between 80 and 120 $^o$C. Most likely the barriers preventing the low-temperature nucleation of magnesite, which according to Saldi et al. (2009) involve slow layer formation *frequencies*, can be overcome in open systems only with pronounced fluctuations in p, T, or x (Deelman, 2001).

As shown in my laboratory experiments, the low-temperature synthesis of magnesite requires the occurrence of fluctuations (of p, T, x) in an open system (Deelman, 1999). It must be remembered that Liebermann's (1967) experiment no. 57 (on which those experiments were based) involve two fundamentally different steps to be repeated 14 times. The first step consists of bubbling carbon dioxide through artificial seawater for 12 hours. (The artificial seawater contained 0.918 mol NaCl, 0.0316 mol $MgCl_2.6\ H_2O$, 0.018 mol $MgSO_4.7\ H_2O$ and 0.020 mol KCl dissolved in 330 ml distilled water; added to it were 2 mMol calcium carbonate). The next step consists of titration with dilute ammonia until pH = 8.0. After that an interval of 60 hours of gradual escape of the dissolved carbon dioxide took place, while the solution was being heated to a constant temperature of for example 40 $^o$C. As described by Liebermann (1967) the whole experiment consisted of repeating 14 times the dissolution interval, the titration and the heating/gradual escape of carbon dioxide interval. In other words, the necessary fluctuations took the form of a finite number of intervals of precipitation alternating with intervals of dissolution. An open system is involved, because the two different intervals require repeated titrations with ammonia respectively bubbling carbon dioxide through the solution at different times.

Additional experiments demonstrated unequivocally the need for fluctuations in an open system. Performing three different *static control* experiments did not result in any magnesite at all, thereby illustrating the actual need for fluctuations in an open system (Deelman, 2003). For example when adding all of the ammonia used in the 14 different titrations in one action into the solution, the precipitate consisted of magnesium hydroxide carbonate. But the most convincing static control experiment is based on the (patent) claim of Waeser (1923/1926), who stated that the reaction

$$MgCl_2 + 2NH_3 + CO_2 + H_2O \rightarrow MgCO_3 + 2NH_4Cl$$

would start at temperatures above 30 $^o$C. However, in my static control experiments it was seen how after saturating a magnesium chloride solution with carbon dioxide and adding a large quantity of ammonia solution (both solutions equilibrated at 45 $^o$C), the precipitate was magnesium hydroxide and not magnesite (for details see Deelman, 2003). Therefore, the patent claim of Waeser (1923/1926) that the reaction between magnesium chloride solution and ammonia plus carbon dioxide gas would start at 30 $^o$C (under static conditions, that is) cannot be maintained.

At the same time my experiments have shown the definite need for the presence of ammonia. In contrast to the suggestion by Liebermann (1967) that either ammonia or sodium carbonate could be used to increase the pH of the bicarbonate solution in the titrations, I have found that successful synthesis of magnesite requires ammonia. The present observation underlines the importance of the statements by Gómez de Llarena (1953) and Lesko (1972) that ammonia would be involved in the low-temperature formation of magnesite. As a free gas ammonia is rare in the atmosphere, but it will be set free upon the decomposition of organic compounds. Notably bacteria capable of nitrate ammonification will be able to free ammonia.

One important conclusion from the above described laboratory experiments concerns the role of bacterial sulphate reduction. No active role seems to be played by living bacteria such as *Desulfovibrio* sp. in the low-temperature formation of magnesite. As a result the suggestion of McKay et al. (1996) that because of the observed paragenesis of magnesite, ankerite, iron sulphides and magnetite, anaerobic conditions must have reigned during the formation of the carbonates found in the ALH84001 meteorite, must be refuted. Similarly the suggestion made by Sanz-Montero and Rodríguez-Aranda (2012) that the mere association between (Miocene) magnesite and abundant microfossils embedded in sheets of organic material would point to an active role of bacteria in the formation of this magnesite, cannot be upheld. When considering only the aspect of reducing conditions, the laboratory experiments involving the various reducing agents do not support any such role. However, it is not at all unlikely that when re-examining for example all of the microorganisms present in the sediments of sabkha El Melah de Zarzis (where Perthuisot et al., 1990 have found Recent magnesite), numerous genera of bacteria different from the sulphate reducers will be found. Possibly such a mixed *ensemble* of microorganisms may well be responsible for the low-temperature nucleation of magnesite.

Algae may well play a crucial role in the low-temperature nucleation of magnesite, in that these microorganisms will withdraw dissolved carbon dioxide from the water during daylight, but will turn to respiration (uptake of oxygen and release of carbon dioxide) during the night. This diel change in amounts of carbon dioxide dissolved in water will inevitable lead to changes in pH (Dubinsky and Rotem, 1974) required for the low-temperature formation of magnesite. In conclusion then, the low-temperature syntheses of magnesite suggest how most probably a variety of different microorganims will be involved in the formation of most of the present-day magnesite deposits.

The mode of formation of magnesite described here is of crucial importance to all discussions on the possible occurrence of life on Mars. As shown in laboratory syntheses, magnesite may well form at temperatures around 40$^o$ C (Deelman, 1999). The mechanism involved is definitely based on *nonequilibrium* conditions. As explained before (Deelman, 2001) *fluctuations* between precipitation and dissolution taking place in an *open* system are capable of accumulating more and more of a stable mineral phase instead of one or more of its metastable equivalents, provided only that the reaction involved is an irreversible one. It will be difficult to underestimate the possible significance of the successful low-temperature syntheses of magnesite. The more so because Oelkers et al. (2008) have pointed out that reliable long-term storage of carbon dioxide requires the (low-temperature) production of dolomite and/or magnesite. The Liebermann-Deelman syntheses are founded on the use of a brine resembling seawater (it contained sodium chloride, magnesium chloride, magnesium sulphate, calcium sulphate and potassium chloride). Actual large-scale production of dolomite and/or magnesite could be based on the use of seawater and take the form of a modified ammonia-soda (Solvay method) process, but with the introduction of the necessary fluctuations.

Laboratory syntheses of magnesite at 40 $^o$C cast doubts on any interpretation of the mineral magnesite as being indicative of elevated temperatures or even hydrothermal reactions. At the same time these inorganic laboratory experiments do not invalidate the possible participation of organic compounds as such, because only a few organic compounds were tested.

## Conclusions

From the experiments described here, the following conclusions can be drawn. Amorphous silica is capable of preventing the incorporation of hydroxyl groups into magnesium carbonates being precipitated (under conditions of room temperature and atmospheric pressure). Minerals such as nesquehonite ($MgCO_3.3\ H_2O$) or lansfordite ($MgCO_3.5\ H_2O$) will form in the presence of amorphous silica instead of one or more of the complex hydroxide-containing magnesium carbonates.

The presence of the strongly reducing hydrogen gas does not lead to the precipitation of anhydrous magnesium carbonate. Therefore, the conclusion that reducing conditions as such do not lead to the low-temperature nucleation of magnesite is justified. On the other hand, the participation of ammonia has been shown to be essential for the low-temperature nucleation of magnesite.

The possible role of microorganisms such as algae may well be rather indirect, in that they will cause the required fluctuations (for example fluctuations of dissolved carbon dioxide). Another possible role played by microorganisms such as bacteria might be found in their ability to free ammonia by deamination.

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