

Impacts of water advection and CO₂ exchanges on the carbon dioxide removal potential of ocean alkalinity enhancement

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Abstract. Ocean alkalinity enhancement (OAE) is a carbon dioxide removal strategy with high carbon dioxide (CO₂) uptake potential and moderate cost. Long-term implementation has been assessed in modelling studies, but most laboratory experiments have focused on shorter time frames with strong advection, which may not be representative of natural systems. Hence, the long-term fate of added alkalinity is has yet to be fully addressed in laboratory experiments. Also Furthermore, the role of CO₂ ingassing on overall efficiency is still largely overlooked. In a new setup, six6-month experiments were conducted using seawater amended with solid Ca(OH)₂ and Mg(OH)₂, and liquid NaOH, with a constant supply-atmospheric of CO₂ supply. Bottles were either kept stagnant or exposed to slow circular advection, similar-yielding gas-transfer-velocity equivalent wind speed to open-ocean-conditions with wind-speed of ~3 m·s⁻¹. All experiments revealed lower-smaller measured increases in total alkalinity thaneompared-to expected, but in the bottles subjected to advection, alkalinity generation was-led to nearly twice as much alkalinity-generationhigh in some cases and significantly faster CO₂ ingassing rates were significantly faster. This led to a quicker-faster equilibration, decreasing the high pH/low CO₂ period, lately-recently identified as a potential threat to living organisms. The concept-of-critical alkalinity period (CAP) is introduced here to represent such-this critical period of high-alkalinity-wherealtered carbonate chemistry during which unwanted-secondary calcium carbonate (CaCO₃) formation may occur, significantly altering the CDR potential of the approach. Yet, CaCO₃ precipitated in most bottles, either as needles with-when adding Ca(OH)₂ or as “broccoli-like structures” with-when adding Mg(OH)₂. Such broccolisThese broccoli-like structures most likely resulted most-likely from precipitation on slowly dissolving Mg(OH)₂ particles. Once-After the CAP was-overended, conditions remained stable and freshly precipitated CaCO₃ did-with-not redissolution-redissolveof the CaCO₃-precipitates. In addition, when using NaOH was used as the alkaline feedstock, Mg(OH)₂ precipitation-flakes was were observed but eventually redissolved in the bottles exposed to advection. The study emphasises the role of at-water advection on OAE’s efficiency-is-important-to-consider and the data presented here may be used to derive serve-as-a-lower thresholds-for OAE implementation in global models.

Introduction

The increase in atmospheric carbon dioxide (CO₂) concentrations since the Industrial Revolution of the 18th century has been identified as the main cause for-of global climate change (IPCC, 2022). While strong efforts to reduce CO₂ emissions are required, current projections report-indicate that these are insufficient to slow down global warming, with active CO₂ removal

from the atmosphere being the only solution (Hausfather, 2025; IPCC, 2021). Among the different carbon dioxide removal (CDR) strategies, ~~o~~Ocean ~~a~~Alkalinity ~~e~~Enhancement (OAE) has gained momentum over the past decade as it may ~~have offer~~ one of the ~~highest-largest~~ carbon capture potentials, ~~estimated at up to 100 Gt CO₂ (1 Gt CO₂ = 10¹⁵ g CO₂) per year (Fuss et al., 2026)~~. Through the ~~ds~~spreading-ispersal of alkaline materials in coastal areas or in the open ocean, for example, the increase

35 in seawater total alkalinity (TA) allows for more CO₂ to be captured by the ocean (Kheshgi, 1995). However, ~~understanding~~ whether added alkalinity remains in the surface ocean ~~is yet to be fully addressed~~, and especially whether such alkalinity remains stable long enough to allow for CO₂ equilibration ~~has yet to be fully addressed~~. ~~Some issues, Secondary effects~~-such as calcium carbonate (CaCO₃) precipitation, have ~~been~~ recently ~~been~~ identified and may ~~lead to decrease efficiency and CDR potential or possibly lead to even~~ CO₂ emissions from the ocean ~~itself~~ (Fuhr et al., 2022; Moras et al., 2022; Rigopoulos et al.,

40 ~~2018~~ Moras et al., 2022). Early work using olivine-rich minerals revealed that in the presence of particles, increasing alkalinity may trigger CaCO₃ formation on the undissolved particles despite slow dissolution kinetics (Fuhr et al., 2022; Rigopoulos et al., 2018). Studies using more reactive solid feedstocks such as calcium hydroxide (Ca(OH)₂) and magnesium hydroxide (Mg(OH)₂) have revealed CaCO₃ precipitation once alkalinity was increased by 500 μmol kg⁻¹ (Hartmann et al., 2023; Moras et al., 2022; Moras et al., 2024; Suitner et al., 2024; Varliero et al., 2024).

45 ~~When TA is increased by about 500 μmol kg⁻¹ using solid feedstocks such as calcium hydroxide (Ca(OH)₂) and magnesium hydroxide (Mg(OH)₂), the increase in alkalinity and the presence of solid surface area enhance CaCO₃ precipitation (Hartmann et al., 2023; Moras et al., 2022; Moras et al., 2024; Suitner et al., 2024; Varliero et al., 2024).~~

During the dissolution of ~~the~~ powdered alkaline minerals, the micro-environment formed around the alkaline particles likely reaches very high aragonite saturation ~~levels of CaCO₃ states~~ (Ω_A), ~~where it may precipitate in the form of~~ ~~under which~~ aragonite

50 ~~may precipitate~~. Aragonite precipitation may occur homogeneously from the water column or heterogeneously ~~on particles, using the yet to be using undissolved particles as nuclei (Marion et al., 2009). Once precipitation starts, it may enter a runaway pattern, with an increasing precipitation rate where the precipitation speed increases~~, sometimes reaching TA values lower than the starting conditions (Moras et al., 2022). Under some circumstances, the presence of inhibitors such as magnesium (Mg²⁺) or dissolved organic matter may delay or decrease the precipitation rate (Chave and Suess, 1970; Pan et al., 2021).

55 Hence, ~~one could suggest that~~ Mg(OH)₂ ~~may could~~ be a more favourable ~~mineral alkaline feedstock~~ than Ca(OH)₂, as the Mg²⁺ released from Mg(OH)₂ dissolution could inhibit CaCO₃, ~~and while~~ Ca(OH)₂ ~~may could~~ allow for early nucleation due to the lattice compatibility of CaCO₃ for such particles (Moras et al., 2022; Pan et al., 2021; Tang et al., 2020). Overall, the persistence of high alkalinity and Ω_A (micro-)environments ~~is expected to enhance will dictate the likeliness-likelihood of alkalinity loss to CaCO₃ formation, particularly through secondary precipitation in the diffusive boundary layer of dissolving particles or on~~

60 ~~suspended solids. Also, the concomitant, as well as high pH and low CO₂ conditions may, potentially impacting biological processes because CO₂ is essential for photosynthesis (Bach et al., 2019). This transient period following alkalinity addition, during which carbonate chemistry is substantially perturbed, is hereafter referred to as the critical alkalinity period (CAP). During the CAP, elevated pH and Ω_A together with reduced dissolved CO₂ concentrations increase the likelihood of secondary CaCO₃ formation and the duration of biological exposure to unfavourable conditions. The CAP ends once carbonate chemistry~~

65 progressively returns to conditions comparable to those prior to alkalinity addition, particularly via atmospheric CO₂ invasion and physical dilution. Hence, CAP duration is influenced ~~This period of high alkalinity~~ by the nature and amount of alkaline feedstock used, hydrodynamic conditions, i.e., mixing and dilution, particle settling and CO₂ equilibration, making it an important consideration ~~is therefore critical to the overall~~ for safe and effective implementation of OAE ~~to limit negative impacts of OAE and maintain high CDR potential, which is later referred to as the critical alkalinity period (CAP).~~

70 Previous OAE-focused laboratory experiments with Ca(OH)₂ and Mg(OH)₂ have been run in a variety of settings. First, these experiments were usually conducted over shorter periods of time, from minutes to weeks (Hartmann et al., 2023; Moras et al., 2022; Moras et al., 2024; Shaw et al., 2025; Yang et al., 2023), though longer experiments exist with slower-dissolving materials such as olivine-rich minerals (Flipkens et al., 2023; Fuhr et al., 2022; Fuhr et al., 2024; Fuhr et al., 2025; Montserrat et al., 2017). Hence, the longer-term alkalinity stability of hydroxide-based OAE is yet to be fully addressed. ~~One can~~

75 ~~hypothesise that g~~Given the time dependency of CaCO₃ precipitation kinetics, previously reported patterns for weeks may be different once monitored for months. ~~Similarly~~~~Second, these~~ experiments using hydroxides were mostly conducted ~~with~~ ~~using the use of~~ stir bars in closed reaction vessels. While these settings allow for easy and standardised comparison between the various dissolution rates of alkaline minerals, and possible precipitation of CaCO₃, they ~~usually omit the role~~ generally do not account the role of atmospheric of CO₂ ingassing ~~on for~~ the stability of alkalinity. ~~They also provide results for a water~~

80 ~~perturbation much stronger than what would occur in a natural setting.~~ Hence, ~~once could we hypothesise~~ argue that the patterns observed in closed laboratory experiments with stir bar may not fully represent open-ocean conditions. Here, in a new approach, the dissolution of Ca(OH)₂ and Mg(OH)₂ in natural seawater was ~~addressed~~ investigated both in static and advected reaction vessels for ~~6 six~~ months. Circular advection regimes were chosen to try and ~~represent reproduce open ocean conditions~~ turbulent conditions similar to those observed in the open ocean, allowing with for vertical diffusion but limiting disruption of

85 the water surface ~~tension disturbance~~, while gas exchange ~~s-were~~ as allowed to occur during the experiments. Carbonate chemistry was monitored throughout, particle analyses were conducted at the end of the experiments, and recorded changes in dissolved inorganic carbon (DIC) were used to calculate the gas transfer velocities ~~gas velocities for CO₂.~~ We later derived the ~~nd~~ equivalent open ocean wind speeds ~~were derived, allowing direct and comparison compared them between laboratory data and with values recorded in the open ocean global modelling studies.~~

90 **Material and methods**

Experimental setup, ~~and~~ selection of the revolutions per minute (RPM) and evaporation compensation

Two sets of 42 bottles (1 L wide-mouth polypropylene) each were used to conduct the experiments at room temperature (24.6 ± 1.4 °C). About 90 L of natural Indian Ocean seawater was collected from the surface ~400 km southeast of Sri Lanka (2°59.995'N, 77°59.997'E) and stored in five 18 L borosilicate 3.3 glass bottles. The bottles were stored in the dark in a

95 refrigerated container until delivery ~10 weeks after collection. The entire amount of seawater was sterile filtered using a Whatman Polycap 36 AS encapsulated filter with a 0.2 µm pore size. The seawater was stored in a cleaned and MilliQ-rinsed

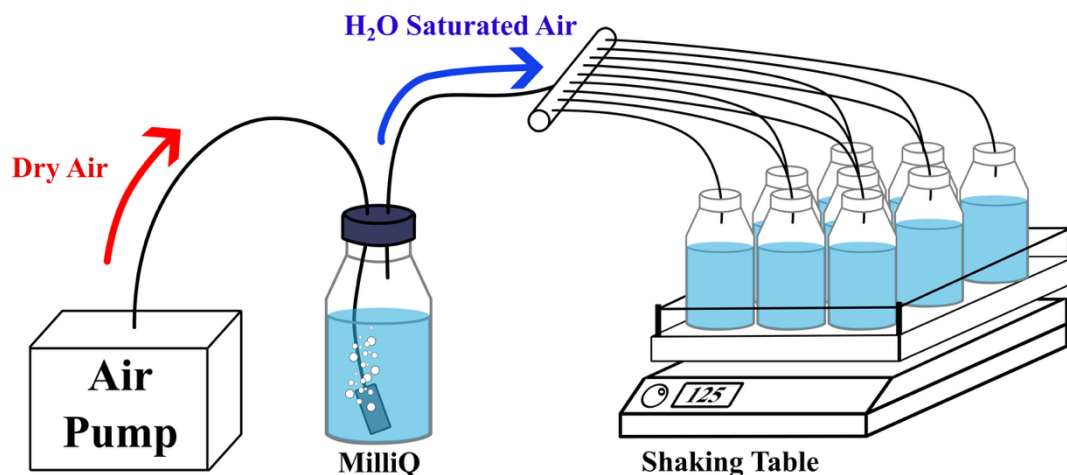
210 L PP rainwater tank (Garantia rainwater green round barrel). The seawater was bubbled for a week using MilliQ-saturated ambient air following Moras et al. (2023) to allow for the entire content of the tank to be equilibrated with atmospheric $p\text{CO}_2$. The experiments were conducted in 1 L PP bottles that were cleaned in a 3-step process. They were first soaked overnight in a 1% solution of NaOH with dishwashing tablets. The next day, the bottles were rinsed with tap water three times and soaked overnight in a 0.5% solution of HCl. Finally, all bottles were rinsed with tap water and MilliQ (18.2 M Ω) three times each, before being dried overnight in an oven at 60 °C. Each of the 84 cleaned bottles was filled with ~900 g of sterile-filtered seawater and the weight was accurately recorded (balance accuracy ± 0.01 g).

One set was placed on a table and kept static throughout the experiments, while the second set was placed on an Edmund Böhler 30 SM C shaking table, set to circular motion at 125 RPM. The choice of motion and RPM was made following a dye test where powdered pPhenolphthalein dye was added to seawater in a 1 L bottle. The dye changes from a colourless powder to a bright pinkish colour when dissolved in solutions of pH higher than 8. These experiments revealed that below 80 RPM, the vertical mixing was deemed insufficient, with most of the dye sinking and remaining at the bottom of the bottle. Past 130 RPM, the seawater contained in the bottles was reaching the edge-rim of the bottle and was deemed too high-great of a risk of spilling. Therefore, the RPM was selected slightly lower, with a slightly lower rotation rate of 125 RPM was selected, providing suitable particle resuspension yet avoiding water overflow. Similarly, the tests revealed that a circular motion was more appropriate than a linear motion to resuspend the dye without breaking water tension. Using linear motion produced, the dye resuspension was insufficient; particle resuspension and substantial disruption of the water tension-surface, including was significantly disturbed splashing, compared to the static bottles. Circular motion produced turbulent conditions with effective vertical mixing without visible splashing. Turbulence can be expressed as a dimensionless number which is also used in physical oceanography, known as the Reynolds number (Reynolds, 1883). To compare turbulence in our experimental setup to naturally occurring ocean turbulence, we derived the Reynolds number from the bottle diameter, tangential velocity, seawater density and dynamic viscosity, as per Griskey (2002). Chosen Our conditions yielded a Reynolds number approximately $\sim 1.7 \times 10^4$, above the threshold considered for turbulent systems (Millero and Poisson, 1981; Thorpe, 2007; Sharqawy, 2010). Such This setup was then used for a total of six months to assess the long-term response of alkalinity addition to seawater.

About 90 L of Indian Ocean natural seawater was collected ~400 km southeast of Sri Lanka (2°59.995'N, 77°59.997'E) and stored in five 18L borosilicate 3.3 bottles in the dark in the fridge at 4 °C. The total amount of seawater was sterile filtered using a Whatman Polycap 36 AS encapsulated filter with a 0.2 μm pore size. The seawater was stored in a cleaned and MilliQ rinsed 210 L PP rainwater tank (Garantia rainwater green round barrel). The seawater was bubbled for a week using MilliQ saturated ambient air as per Moras et al. (2023) to allow for the whole content of the tank to be equilibrated to air $p\text{CO}_2$. The bottles were cleaned in a 3-step process. They were first soaked overnight in a 1% solution of NaOH with dishwashing tablets. The next day, the bottles were rinsed with tap water three times and soaked overnight in a 0.5% solution of HCl. Finally, all bottles were rinsed with tap water and MilliQ (18.2 M Ω) three times each, before being dried overnight in an oven at 60 °C. About 90 L of Indian Ocean natural seawater was collected ~400 km southeast of Sri Lanka (2°59.995'N, 77°59.997'E) and

stored in five 18L borosilicate 3.3 bottles in the dark in the fridge at 4 °C. The total amount of seawater was sterile filtered using a Whatman Polycap 36 AS encapsulated filter with a 0.2 µm pore size. The seawater was stored in a cleaned and MilliQ rinsed 210 L PP rainwater tank (Garantia rainwater green round barrel). The seawater was bubbled for a week using MilliQ saturated ambient air as per Moras et al. (2023) to allow for the whole content of the tank to be equilibrated to air pCO₂.

135 ~~One major issue when it comes to open bottles is evaporation. To compensate for such loss minimize evaporation, the headspace of the open bottles was continuously replaced with water-saturated air. For this purpose, MilliQ saturated air was provided in all bottles using a similar process to the one described above for the seawater. Air was pumped bubbled through into MilliQ~~
contained in a sealed 1 L Schott bottle. Once the air was bubbled through the MilliQ, it was carried over into plastic air-flow dividers similar to those used for aquariums (Fig. 1). One ~~set of tubing~~ tube was ~~placed~~ positioned in the headspace of each ~~of~~
140 ~~the bottles and secured was maintained using binder elips connected to the bottle neck with binder clips (Fig. A1). The air was not allowed to~~ did not directly bubble the seawater, ~~directly~~ but rather continuously renewed ~~ensured that each of the bottles' headspaces was consistently provided with the same freshly MilliQ-saturated ambient air, allowing for easier comparison of results from bottles as they were all exposed to the same conditions. To compensate for evaporation, e~~
~~Each bottle was regularly also weighed, and accurate MilliQ additions were performed when regularly (every 2-3 days) to ensure that evaporation was not too important. If a difference decrease in weight of more than 5% was observed to be higher than 5% after these 2-3 days, a precise amount of fresh MilliQ was added by weighing to compensate for the slight evaporation. To reduce the potential growth of autotrophic organisms, Another issue that was raised before the start of the experiments was the risk of organisms growing in the bottles. Hence, experiments were mostly conducted in the dark except during sampling where laboratory lights were tuned on. To prevent or at least limit such growth, two germicidal UV lamps with ozone light of emitting at 254 nm were~~
145 ~~mounted on top of the setup and placed so that all bottles were exposed to the UV lights. These lamps were automatically switched on for an hour twice a week in the middle of the night, allowing for a long enough exposure to~~ prevent-limit
150 ~~microorganisms' microbial~~ microbial growth.



155 **Figure 41:** Schematic of the experimental setup. The air pump takes up dry air and pushes it into a closed Schott bottle where MilliQ is bubbled. The newly $\text{H}_2\text{O}/\text{H}_2\text{O}_2$ -saturated air is then carried to an air diffuser from which individual tubes are connected to each experimental bottle, constantly continuously and equally supplying each open bottle with fresh air equilibrated to atmospheric CO_2 .

TA manipulations and sampling procedure

160 Each 42-bottle experimental set consisted of three control bottles filled only with sterile-filtered seawater, 18 TA-enhanced bottles with $\text{Ca}(\text{OH})_2$ powder, 18 TA-enhanced bottles with $\text{Mg}(\text{OH})_2$ powder, and three TA-enhanced bottles with a stock solution of NaOH (Table 1). Every treatment was conducted in triplicate. Targeted alkalinity increases ranged from $+100 \mu\text{mol kg}^{-1}$ to $+600 \mu\text{mol kg}^{-1}$ in increments of $100 \mu\text{mol kg}^{-1}$ for both alkaline powders, and the NaOH-amended set consisted of a TA increase of $+600 \mu\text{mol kg}^{-1}$, leading to 3 control bottles, 18 TA-enhanced bottles using $\text{Ca}(\text{OH})_2$, 18 TA-enhanced bottles using $\text{Mg}(\text{OH})_2$, and 3 TA-enhanced bottles using NaOH. One bottle was lost from the 125 RPM NaOH treatment after day 6, and one was lost from the 0 RPM NaOH treatment after day 108. Subsequent values therefore represent the means of the two remaining replicates. The alkaline powders $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$ and the NaOH pellets used to make the 1 M NaOH solution were all reagent-grade, obtained from Carl Roth (Germany). The $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$ powders were analysed for grain size distribution with a Bettersizer S3 Plus, exhibiting similar D50 values, i.e., 31.98 and $29.79 \mu\text{m}$, respectively, as well as comparable specific surface areas, i.e., 58.90 and $62.82 \text{ m}^2 \text{ kg}^{-1}$. Both powders were dried overnight at 60°C , and specific amounts for TA amendments were weighed on weighing paper using a Kern ABT 120-5DNM ($\pm 0.01 \text{ mg}$) balance. $\text{Ca}(\text{OH})_2$ additions ranged from 3.67 to 22.26 mg L^{-1} , while $\text{Mg}(\text{OH})_2$ additions ranged from 2.88 to 17.54 mg L^{-1} to reach the desired ΔTA targets (Table A 1). The NaOH solution was added using an Eppendorf Research Plus 1000 pipette. Overall, the weighing uncertainty led to an offset between the targeted and estimated TA of less than 1%. Both the static and advected treatments consisted of the same 42 bottles set to allow for comparison. Each of the cleaned 84 bottles was filled with $\sim 900 \text{ g}$ of sterile filtered seawater and the weight was accurately recorded (balance accuracy $\pm 0.01 \text{ g}$). Two sets of 42 bottle experiments consisted of control bottles filled only with sterile filtered seawater, and TA-enhanced bottles, using either $\text{Ca}(\text{OH})_2$ powder,

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180 ~~Mg(OH)₂ powder, or 1 M NaOH solution. The enhanced alkalinity ranged from +100 $\mu\text{mol kg}^{-1}$ to +600 $\mu\text{mol kg}^{-1}$ with 100 $\mu\text{mol kg}^{-1}$ TA increments for both alkaline powders, and only one set of +600 $\mu\text{mol kg}^{-1}$ using NaOH. Every treatment was conducted in triplicate, leading to 3 control bottles, 18 TA enhanced bottles using Ca(OH)₂, 18 TA enhanced bottles using Mg(OH)₂, and 3 TA enhanced bottles using NaOH (one bottle was lost for the 125 RPM treatment from day 6 onwards and one was lost for the 0 RPM treatment from day 108 onwards, data reported are the average of the two remaining replicates). Both the static and advected treatments consisted of the same 42 bottles set to allow for comparison.~~

185 **Table 1: Experimental design for one set of 42 bottles, showing the treatment groups (control, Ca(OH)₂ powder, Mg(OH)₂ powder and NaOH stock solution), amendment levels (expressed in $\mu\text{mol kg}^{-1}$), and the number of replicate bottles per treatment.**

<u>Treatment</u>	<u>+0</u>	<u>+100</u>	<u>+200</u>	<u>+300</u>	<u>+400</u>	<u>+500</u>	<u>+600</u>
<u>Control</u>	<u>3</u>	=	=	=	=	=	=
<u>Ca(OH)₂</u>	=	<u>3</u>	<u>3</u>	<u>3</u>	<u>3</u>	<u>3</u>	<u>3</u>
<u>Mg(OH)₂</u>	=	<u>3</u>	<u>3</u>	<u>3</u>	<u>3</u>	<u>3</u>	<u>3</u>
<u>NaOH</u>	=	=	=	=	=	=	<u>3</u>

190 ~~The alkaline powders Ca(OH)₂ and Mg(OH)₂, and the NaOH pellets used to make the 1 M NaOH solution were all reagent-grade, obtained from Carl Roth (Germany). The Ca(OH)₂ and Mg(OH)₂ powders were analysed for grain size distribution with a Bettersizer S3 Plus, exhibiting similar D50 values, i.e., 31.98 and 29.79 μm , respectively, as well as comparable specific surface areas, i.e., 58.90 and 62.82 $\text{m}^2 \text{kg}^{-1}$. Both powders were dried overnight at 60 °C. The NaOH solution was added using an Eppendorf Research Plus 1000 pipette, while the Ca(OH)₂ and Mg(OH)₂ powder were weighed on a weighing paper using a Kern ABT 120 5DNM ($\pm 0.01 \text{ mg}$) balance. Overall, the weighing uncertainty led to an offset between the targeted and estimated TA of less than 1%.~~

195 ~~The pH, temperature and conductivity of each sample were measured before manipulation, after and one-, three- and six-~~ days ~~following~~ after alkalinity addition. Thereafter, measurements took place twice a week for the first ~~two~~ two months and once a week afterwards. pH and conductivity were measured using a multiparameter pocket measuring device MultiLine Multi 3630 IDS from WTW, with a SenTix 940 pH probe and a WTW TetraCon 925 conductivity probe. The pH probe was calibrated weekly using NIST-traceable buffers and measured values were converted to the total pH scale (pH_T) as using measured salinity and temperature and inputting the values into the equation derived by ~~per~~ Badocco et al. (2021). The conductivity probe was calibrated before the first measurement to determine the starting salinity of the seawater. Salinity was estimated ~~calculated as~~ 35.346 $\pm$ 0.077 on the practical salinity scale of 1978 (Lewis and Perkin, 1981).

200 ~~-TA and dissolved inorganic carbon (DIC) samples were taken before manipulation to establish background concentrations. Then, sampling occurred on, at days 1, 3 and 6 after alkalinity addition, and monthly afterwards. Seawater was sampled using~~

205 a peristaltic pump connected to a 0.2 μm filter from Sartorius. B24 mL borosilicate-glass 3.3 TOC vials (24 mL, from

Shimadzu) were ~~filled smoothly slowly filled~~ from the bottom ~~with a 14-gauge needle~~ and allowed to overflow by at least half of their volume (Dickson et al., 2007). The vials were then ~~closed sealed without headspace with using~~ screw caps ~~fitted~~ with ~~red~~ silicone ~~red/white~~ PTFE ~~white~~ double-layered ~~red~~ septa ~~without headspace~~. TA samples were collected ~~right immediately~~ after DIC samples, with ~~about ~320~~ mL of water sampled in 100 mL PP bottles. The TA bottles were closed tightly to prevent any evaporation, and both TA and DIC samples were stored ~~in the fridge~~ ~~refrigerated~~ in the dark until analysis (within one week of sampling). Up to 100 mL was removed at each sampling point, leaving less than 150 mL at the end of the experiment, as the sampled water was not replaced.

The remaining water was used for mineral microscopy and elemental composition analysis. For microscopy, about 20 mL of the remaining water was filtered through Whatman polycarbonate Nuclepore filters (0.2 μm) from bottles amended to a target ΔTA of +600 $\mu\text{mol kg}^{-1}$. The filters were briefly rinsed for several seconds with a few millilitres of fresh MilliQ to remove residual salts before being dried at 60 °C overnight. Another set of filters was used for elemental analysis. ~50 mL of water was accurately weighed and filtered on a Whatman polycarbonate Nuclepore filter (0.2 μm). The recorded weight was used to determine the exact amount of seawater filtered for the analysis. The filter was then placed into a 50 mL Falcon tube to dry overnight. The incubation bottles were rinsed three times with ~15 mL of 3% Suprapure nitric acid to remove precipitates from the bottle walls. The acid was then transferred into the Falcon tube with the filter to dissolve all collected particles before being made up to a final volume of 50 mL. Secondary mineral formation was expected and suspected to occur on the surface of added material, i.e., $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$. Scanning electron microscopy (SEM) analyses were conducted on a tabletop Hitachi TM4000 II, connected to a Bruker QUANTAX energy dispersive x-ray spectroscopy system (EDS). At the end of the experiments, about 20 mL of remaining water was collected on Whatman polycarbonate Nuclepore filters (0.2 μm) from the higher treatments to ensure a significant amount of sample. The samples were then rinsed with MilliQ before being dried at 60 °C overnight.

Carbonate chemistry measurements, mineral analyses and data processing and calculations

pH and conductivity were measured using a Multiparameter pocket measuring device, MultiLine Multi 3630 IDS from WTW, using a SenTix 940 pH probe and a WTW TetraCon 925 conductivity probe. The pH probe was calibrated weekly using NIST standards and converted to the pH on the total scale (pH_T) as per Badoeco et al. (2021). The conductivity was calibrated before the first measurement to determine the starting salinity of the seawater. The salinity was estimated at 35.346 ± 0.077 on the practical salinity scale of 1978 (Lewis and Perkin, 1981). The TA was determined by potentiometric titration ~~with using~~ an 888 Titrando connected to an 814 USB Sample Processor (~~from Metrohm~~), ~~with with aa~~ 0.02 M HCl ~~solution solution~~ as the titrant. The titrant ionic strength was adjusted to 0.72 mol kg^{-1} by NaCl addition, corresponding to a salinity of 35 (Dickson et al., 2007). The raw TA values were corrected against certified reference materials (CRM #187) from Dickson (2010). The overall instrument uncertainty was estimated at 0.1% via error propagation of all CRM measurements. DIC analyses were conducted on a Shimadzu TOC-L/CPH analyser calibrated using a CRM (#207). Samples were injected into the reactor filled

with fresh 25% H₃PO₄, and ~~DIC~~ the evolved CO₂ was carried ~~with by~~ CO₂-free air. The overall instrument uncertainty was also determined via error propagation of all CRM measurements and estimated at 3.5%.

During the ~~six~~6-month experiment, atmospheric CO₂ concentration, temperature and humidity were recorded every 15 minutes with a HOBO MX1102A CO₂ logger to control for any spike in atmospheric CO₂ and/or temperature. The pCO₂ logger was automatically calibrated to 400 ppm (as per manufacturer recommendations) by placing the logger in open air until a stable reading was recorded. ~~Such~~ This calibration ~~would not allow for~~ did not permit accurate pCO₂ measurement but allowed to rather control whether monitor fluctuations in spikes in pCO₂ ~~occurred~~ during the experiments. Overall, atmospheric pCO₂ (437.7 ± 34.6 ppm) ~~only~~ varied by less than 10% throughout the experiment, and the ~~average mean~~ temperature was recorded at 24.6 ± 1.4 °C (later rounded to 25 °C for further processing). To compensate for such atmospheric pCO₂ measurement uncertainty, the estimated pCO₂ from the control batches ~~will be used~~ was therefore used as the background concentration. Sample pCO₂ and Ω_A were calculated in CO2SYS To do so, using measured TA, DIC, salinity and temperature ~~measurements, and using with~~ the K₁ and K₂ dissociation constants from Schockman and Byrne (2021), KSO₄ dissociation constant from Dickson (1990), K_F dissociation constant from Perez and Fraga (1987) and the borate-to-salinity ratio from Lee et al. (2010), ~~the sample pCO₂ could be determined, as well as other carbonate chemistry parameters such as Ω_A.~~

For microscopic analysis, the dried Whatman filters were mounted on an aluminium holder using double-sided carbon sticky tabs. Scanning electron microscopy (SEM) analyses were conducted on a Hitachi TM4000 II tabletop SEM equipped with a Bruker QUANTAX energy-dispersive X-ray spectroscopy system (EDS). The second set of filters stored in Suprapure nitric acid was used to determine the concentrations of solid Ca and Mg, to assess whether CaCO₃ precipitated on undissolved Mg(OH)₂ particles. Analyses were performed using flame atomic absorption spectroscopy (F-AAS) on a Thermo Fisher Solaar S Series.

Changes in DIC attributed to CO₂ ingassing were used to derive the gas-transfer velocities and equivalent open-ocean wind speeds in each of the experiments. The DIC fluxes were estimated from the net changes in DIC between the control and the amended bottles. The change in DIC per day was estimated and inserted into the CO₂ flux equation from Wanninkhof (2014):

$$F = kK_0(pCO_{2w} - pCO_{2a}) \quad (1)$$

where F is the flux of CO₂ between the atmosphere and the seawater, K₀ is the solubility of CO₂ in seawater (obtained from Weiss, 1974, at 24.6 °C and salinity of 35.346), and pCO_{2w} and pCO_{2a} are the partial pressures of the seawater and air in atm, respectively. This equation, combined with the water depth on the sampling day, was solved to calculate the gas velocity terms k_{0 RPM} and k_{125 RPM} in cm h⁻¹, for the 0 RPM and 125 RPM experiments, respectively. Finally, these gas-transfer velocities can be correlated to open-ocean wind speed using the gas exchange wind speed relationship from Wanninkhof (2014):

$$k = 0.251 U^2 \left(\frac{Sc}{660} \right)^{-0.5} \quad (2)$$

where U^2 is the squared wind speed in m s^{-1} , and Sc is the Schmidt number at the measured temperature and salinity (Wanninkhof, 2014).

Statistical tests and TA budget

Throughout the experiment, laboratory conditions occasionally affected the performance of the DIC analyser. Therefore, measured DIC concentrations were routinely compared with DIC concentrations calculated independently from measured TA and pH_T measurements using CO2SYS. Because no changes in TA or pH_T accompanied randomly observed shifts in measured DIC, these deviations were attributed to analytical drift rather than changes in carbonate chemistry. For sampling campaigns showing systematic offset between measured and calculated DIC, measured DIC values were corrected using the mean ratio of all calculated to measured DIC values.

Changes in pCO_2 over time were analysed separately for the 0 and 125 RPM treatments using linear mixed-effects models implemented in MATLAB (MathWorks). The fixed effects included treatment, i.e., ΔTA targets, sampling time, and their interaction, while bottle identity was set as a random intercept to account for repeated measurements from the same experimental unit. Sampling time was treated as a categorical factor to avoid assuming a linear temporal response and to enable comparison among individual sampling dates. The control treatment and day 0 were specified as the reference levels for treatment and sampling time, respectively. Overall effects of treatment, sampling time, and their interaction were assessed using F-tests with Satterthwaite's approximation for the denominator degrees of freedom. Upon significant interaction between treatment and sampling time, planned one-sided post hoc contrasts were performed to compare each treatment with the control at each sampling time, testing the hypothesis that treatment pCO_2 was lower than that of the corresponding control. To account for multiple comparisons between treatments and the control at each sampling time, the one-sided p-values were adjusted using the Holm procedure. Treatment pCO_2 was considered significantly lower than the control when the estimated difference between treatment and control was negative and the corresponding one-sided adjusted p-value was < 0.05 . Model assumptions were evaluated by visual inspection of residual-versus-fitted plots and normal probability plots of the residuals.

A TA generation efficiency metric was calculated by dividing the measured ΔTA after 24 hours by the targeted ΔTA . This $\% \text{TA}_{\text{realised}}$ can be used to compare the dissolution efficiency of the minerals based on the treatment. After six months, Secondary mineral formation was expected and suspected to occur on the surface of added material, i.e., $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$. Scanning electron microscopy (SEM) analyses were conducted on a tabletop Hitachi TM4000 II, connected to a Bruker QUANTAX energy dispersive x-ray spectroscopy system (EDS). At the end of the experiments, about 20 mL of remaining water was collected on Whatman polycarbonate Nuclepore filters ($0.2 \mu\text{m}$) from the higher treatments to ensure a significant amount of sample. The samples were then rinsed with MilliQ before being dried at 60°C overnight. The Secondary

mineral formation was expected and suspected to occur on the surface of added material, i.e., $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$. Scanning electron microscopy (SEM) analyses were conducted on a tabletop Hitachi TM4000 II, connected to a Bruker QUANTAX energy dispersive x-ray spectroscopy system (EDS). At the end of the experiments, about 20 mL of remaining water was collected on Whatman polycarbonate Nuclepore filters (0.2 μm) from the higher treatments to ensure a significant amount of sample. The samples were then rinsed with MilliQ before being dried at 60 °C overnight. dried filters were mounted on an aluminium filter holder using double-sided carbon sticky tabs. The small remaining volume of the higher treatments using $\text{Mg}(\text{OH})_2$ was also collected for Ca^{2+} and Mg^{2+} analyses, to decipher whether undissolved $\text{Mg}(\text{OH})_2$ could have been used for CaCO_3 precipitation. In the case where CaCO_3 would have precipitated on yet to be dissolved $\text{Mg}(\text{OH})_2$ particles, a clear signal of Mg^{2+} would be observed. Otherwise, if CaCO_3 precipitated from the water column only, the Mg^{2+} signal would be absent. After 6 months, the remaining volume in the 1 L bottles was accurately weighed and filtered on a Whatman polycarbonate Nuclepore filter (0.2 μm). The weighing of the remaining liquid allowed to determine the exact amount of seawater filtered for the analysis, which is required to report on the concentration after analysis. The filter was then placed into a 50mL Falcon tube to dry overnight. The incubation bottle was then rinsed three times with ~15 mL of 3% Suprapure nitric acid to try and remove any precipitate on the wall. The acid was then transferred into the Falcon tube with the filter to dissolve all collected particles before being topped up to 50mL. Concentrations of Ca and Mg were analysed via Flame Atomic Absorption Spectroscopy (F-AAS) on a Thermo Fisher Solaar S Series. a TA budget was derived from the carbonate chemistry measurements and mineral analyses. Measured TA loss ($\text{Meas. TA}_{\text{loss}}$) was calculated as the difference between the TA value at day 0 and the lowest measured TA value. Similarly, a theoretical maximum ΔTA ($\text{Max. } \Delta\text{TA}_{\text{theo}}$) term was calculated using the mineral molar mass and the amount of material added in mg L^{-1} (assuming complete dissolution). From both estimates, a maximum TA_{loss} term ($\text{Max. TA}_{\text{loss}}$) was calculated from the difference between the $\text{Max. } \Delta\text{TA}_{\text{theo}}$ and the lowest recorded ΔTA . An estimate of particulate Ca concentration was derived from $\text{Max. TA}_{\text{loss}}$ with each mole of TA loss corresponding to 0.5 mol CaCO_3 . Using the total measured TA loss ($\text{Meas. TA}_{\text{loss}}$) and the estimated maximum ΔTA from the known amount of feedstock added in each bottle, an estimated maximum TA_{loss} term ($\text{Max. TA}_{\text{loss}}$) can be calculated. Similarly, using the $\text{Max. TA}_{\text{loss}}$ to CaCO_3 , an estimated particulate Ca^{2+} concentration can be determined (estimated Ca^{2+}). For each mole of TA lost, half a mole of CaCO_3 is formed, hence estimated Ca^{2+} is half of $\text{Max. TA}_{\text{loss}}$.

Throughout the experiments, quality control of the carbonate chemistry measurements was conducted using the TA, DIC and pH_T data from each set to prevent loss of data to unexpected instrument misbehaviour. The DIC measurements from months 1 and 2 were likely underestimated, while the DIC measurements from month 6 were likely overestimated (Fig. 2). While little to no changes were observed with TA and pH_T measurements, these changes in DIC suggested a measurement issue, which coincided with issues with the temperature control system in the laboratory where analyses were conducted. Therefore, a comparison of the offset between measured DIC and estimated DIC from measured TA and pH_T were conducted. The boxplot analysis (Fig. 2) revealed that the months 1 and 2 DIC measurements were significantly lower than the average offset, while the month 6 DIC measurements were significantly higher. Therefore, the DIC concentration for these three points was corrected

using the average offset of all measurements. Initial DIC values of the 1-month dataset were overestimated by $-101 \pm 5 \mu\text{mol kg}^{-1}$, those of the 2-month dataset were overestimated by $-47 \pm 3 \mu\text{mol kg}^{-1}$, and those of the 6-month dataset were underestimated by $-67 \pm 4 \mu\text{mol kg}^{-1}$. The TA dataset from day 6 of the 125 RPM experiments was lost during analysis. The TA datapoint was created assuming linear changes between day 3 and 35.

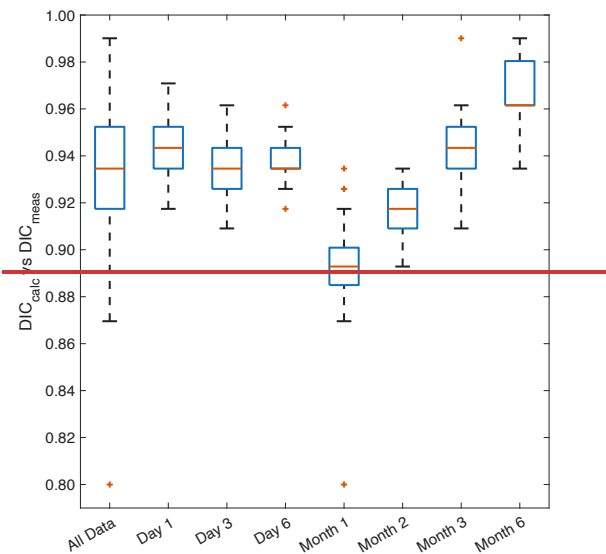


Figure 2: Boxplots of the ratio between the calculated DIC from TA and pH_T data and the measured DIC for all compiled data, and as a function of the dataset analysed. Ratios from the month 1 and 2 datasets clearly fall below the compiled data ratio, while the ratio from the month 6 dataset stands above the compile data ratio.

Results

Differences between expected and measured alkalinity using solid feedstocks

Alkalinity additions for each experiment were performed identically with the two solid feedstocks, with Δ TA targets varying from +100 to +600 $\mu\text{mol kg}^{-1}$. In all experiments, the expected maximum Δ TA was not reached, but positive Δ TA values were observed at day 1. Within these first 24 hours, s Stark differences emerged betweenin realised TA based- on the material used and the advection regime within the first 24 hours. The %TA_{realised}, i.e., net measured Δ TA in $\mu\text{mol kg}^{-1}$ divided by the targeted Δ TA in $\mu\text{mol kg}^{-1}$; The %TA_{realised}- values were significantly lower in bottles without- no advection (Fig. 32). In the aAdverted bottles, TA generation was higher, led to a with -%TA_{realised} values that were that was 2.5 $\pm$ 1.0 and 4.3 $\pm$ 2.1 times higher than those in the static ones treatments for Ca(OH)₂ and Mg(OH)₂, respectively. For example, in the advected bottles with a Δ TA target of +100 $\mu\text{mol kg}^{-1}$, TA increased by 55-58 $\mu\text{mol kg}^{-1}$ with Ca(OH)₂ and Mg(OH)₂, respectively, corresponding to %TA_{realised} values of 55-58%. However, TA generation in static bottles with a Δ TA target of +600 $\mu\text{mol kg}^{-1}$

only yielded TA increases of 20-45 $\mu\text{mol kg}^{-1}$ with $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$, respectively, corresponding to %TA_{realised} values ranging from 3 to 8% (Figs. 2 and 3).

In all cases, the %TA_{realised} decreased with increasing ΔTA targets. Using $\text{Mg}(\text{OH})_2$, the %TA_{realised} ~~appears to~~ consistently decreased~~s~~, regardless of the advection speed, while using $\text{Ca}(\text{OH})_2$, ~~the loss in~~ %TA_{realised} quickly decreased~~s~~ with advection and leads to converging values for the 0 and 125 RPM regimes at ΔTA targets of 500 and 600 $\mu\text{mol kg}^{-1}$. ~~It appears that the measured TA generation from $\text{Mg}(\text{OH})_2$ was generally without advection remains mostly lower than that from using $\text{Ca}(\text{OH})_2$, with values varying between 3-18% with $\text{Mg}(\text{OH})_2$, and from 3 to -30% with $\text{Ca}(\text{OH})_2$ compared relative to the expected TA increase. However, when exposed to advection, the TA generation of $\text{Mg}(\text{OH})_2$ was mostly higher than when using $\text{Ca}(\text{OH})_2$, and the loss-decrease in efficiency is not as was less pronounced compared to those of than for $\text{Ca}(\text{OH})_2$.~~

Advection led to a %TA_{realised} varying between from 10 to -55% using $\text{Ca}(\text{OH})_2$, and between from 29 to -58% using $\text{Mg}(\text{OH})_2$.

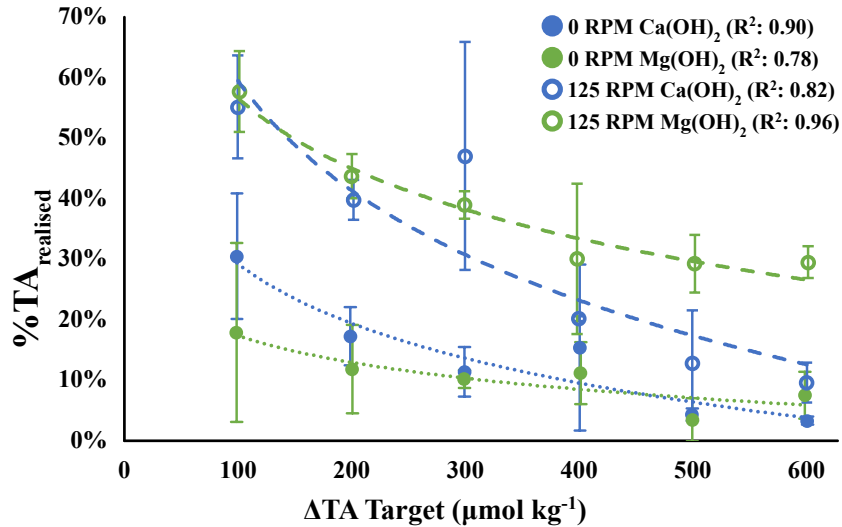


Figure 32: %TA_{realised} ~~from after 1 day, i.e., measured maximum ΔTA measured after one day~~, as a function of the ΔTA target (in $\mu\text{mol kg}^{-1}$). ~~Each point represents the average of each triplicate, with blue markers representing the $\text{Ca}(\text{OH})_2$ treatments and the green markers representing the $\text{Mg}(\text{OH})_2$ treatments. Dotted lines represent logarithmic fits of the 0 RPM treatments while the dashed lines represent the logarithmic fits of the 125 RPM treatments. Corresponding R^2 values are provided in the legend. Each point represents the mean of the triplicate measurements, with error bars representing the uncertainty based on the standard deviation of these three measurements.~~

Long-term stability of alkalinity

In most cases, ~~the~~ TA measured at 24 h was higher than ~~measurements obtained during the remainder of the the remaining six-month experiments data~~. Two clear patterns emerged between the 0 RPM and 125 RPM incubations (Fig. 43). The absence of advection mostly led to a negative ΔTA over time. With $\text{Ca}(\text{OH})_2$, a clear TA losses in TA followed the 500 and

600 $\mu\text{mol}\mu\text{mol kg}^{-1}$ additions, ~~for which, where~~ after ~~approximately two~~ 2 months, about 100 ± 12 and 190 ± 37 $\mu\text{mol}\mu\text{mol kg}^{-1}$ were lost, respectively. However, smaller additions led to slight changes in $\Delta\Delta\text{TA}$ but without a clear gain or loss over time. Using $\text{Mg}(\text{OH})_2$ led to a clear decrease in alkalinity over time, with ~~a higher loss in~~ alkalinity ~~loss increasing~~ with ~~higher the amount of~~ material ~~addition~~ added. After a month, the two ~~higher largest~~ additions led to a loss in TA of 150-180 $\mu\text{mol}\mu\text{mol kg}^{-1}$. The two medium additions led to a loss in TA of about 50 $\mu\text{mol}\mu\text{mol kg}^{-1}$, while lower additions led to losses around 10-20 $\mu\text{mol}\mu\text{mol kg}^{-1}$.

385 However, once exposed to advection, ~~the $\Delta\Delta\text{TA}$ is remained mostly~~ positive in ~~nearly~~ all cases, except for the 600 $\mu\text{mol}\mu\text{mol kg}^{-1}$ addition with $\text{Ca}(\text{OH})_2$. ~~With~~ When targeting ΔTA additions of 100, 200 and 300 $\mu\text{mol}\mu\text{mol kg}^{-1}$ with $\text{Ca}(\text{OH})_2$, $\Delta\Delta\text{TA}$ increased by ~ 50 , ~ 80 , and ~ 130 $\mu\text{mol}\mu\text{mol kg}^{-1}$, respectively, and remained relatively stable over time. The 400 and 500 $\mu\text{mol}\mu\text{mol kg}^{-1}$ additions did not generate more TA than the lower additions, ~~with a start~~. Instead, ~~their $\Delta\Delta\text{TA}$ values decreasing~~ decreased over time, from ~ 80 to 40 $\mu\text{mol}\mu\text{mol kg}^{-1}$ for the 400 $\mu\text{mol}\mu\text{mol kg}^{-1}$ target, and from ~ 60 to about -5 $\mu\text{mol}\mu\text{mol kg}^{-1}$ for the 500 $\mu\text{mol}\mu\text{mol kg}^{-1}$. Finally, the 600 $\mu\text{mol}\mu\text{mol kg}^{-1}$ treatment led to the lowest $\Delta\Delta\text{TA}$ increase of about 50 $\mu\text{mol}\mu\text{mol kg}^{-1}$ before decreasing ~~consistently steadily~~ over ~~6 six~~ months down to about -140 $\mu\text{mol}\mu\text{mol kg}^{-1}$. Conversely, the $\text{Mg}(\text{OH})_2$ treatments led to increasing $\Delta\Delta\text{TA}$ with increasing additions, and positive values over the ~~six~~ 6-month experiment. The +100, +200 and +300 $\mu\text{mol}\mu\text{mol kg}^{-1}$ treatments saw an increase over time in $\Delta\Delta\text{TA}$ despite a rather stable phase over the first week. The +400 and +500 $\mu\text{mol kg}^{-1}$ treatments did not lead to a secondary increase in alkalinity but ~~stayed~~ remained stable over the ~~six~~ 6-month experiments, following a peak in $\Delta\Delta\text{TA}$ on day 1 and slight loss until day 3. Finally, the highest ~~st~~ $\text{Mg}(\text{OH})_2$ treatment saw the ~~highest drop~~ greatest decrease in $\Delta\Delta\text{TA}$ over time, from nearly 180 $\mu\text{mol}\mu\text{mol kg}^{-1}$ on day 1, to 99 ± 7 $\mu\text{mol}\mu\text{mol kg}^{-1}$ after ~~two~~ 2 months, ~~after which, and it~~ remained stable until the end of the experiment.

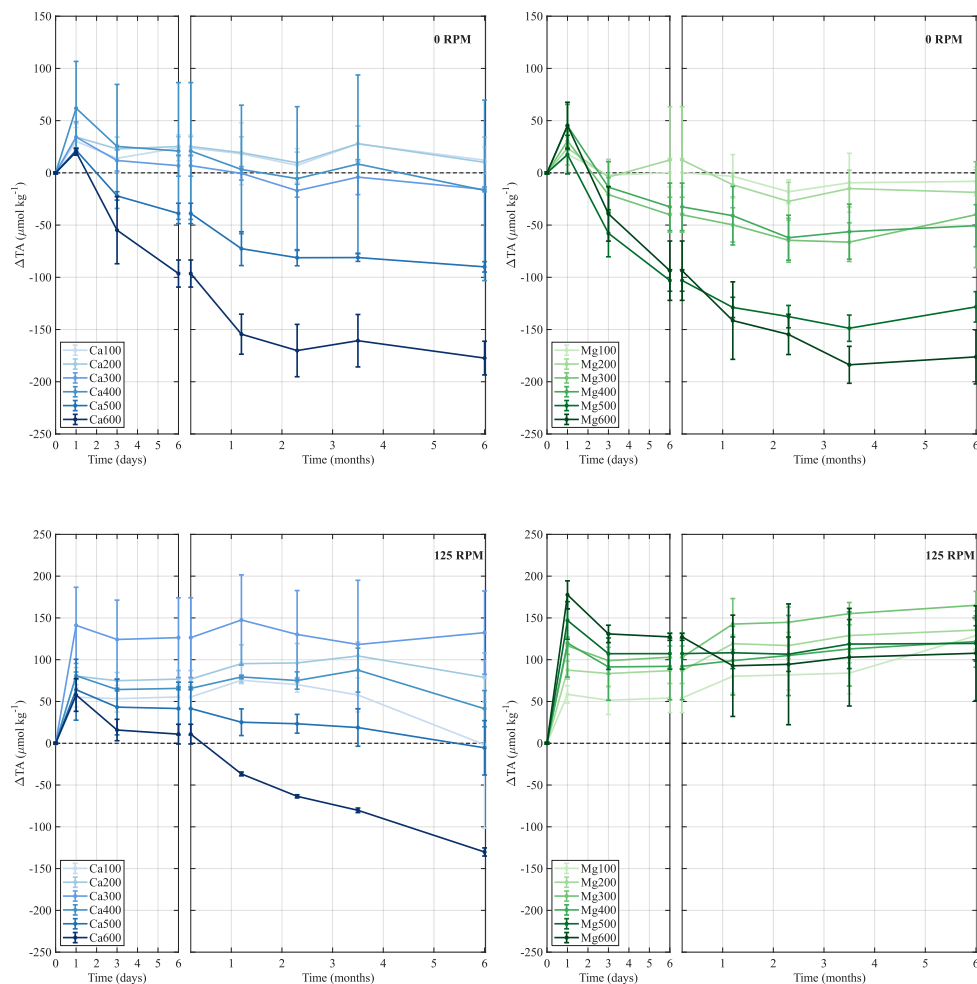


Figure 43: Changes in TA (Δ TA) as a function of time and expressed in $\mu\text{mol kg}^{-1}\text{kg}^{-1}$, calculated by subtracting the same-day control TA from measured TA (the controls are represented by the black dotted line). The 0 RPM experiments are represented in the top two tiles, while the 125 RPM experiments are represented in the bottom two tiles. $\text{Ca}(\text{OH})_2$ additions are represented in blue and $\text{Mg}(\text{OH})_2$ additions in green, with the shades representing the amount added, i.e., the darker the shade, the higher the alkaline material addition. Each point represents the average of triplicates, with error bars representing the uncertainty based on the standard deviation of these three measurements.

Effects of advection on pCO_2 changes

Following alkalinity additions, clear drops in the seawater pCO_2 were observed in all experiments (Fig. 54). In most cases, the decrease in pCO_2 was correlated to the targeted Δ TA, i.e., higher targeted Δ TA led to lower pCO_2 . This was the case in most treatments except for the 0 RPM $\text{Ca}(\text{OH})_2$ experiment, where the lowest pCO_2 was observed in the +400 $\mu\text{mol kg}^{-1}$

kg⁻¹ experiment. In the 0 RPM experiments, the lowest pCO₂ recorded was 250-280 μatm in the Ca(OH)₂ and Mg(OH)₂ experiments, respectively. With both Ca(OH)₂ and Mg(OH)₂, pCO₂ slowly increased over ~~one+~~ month ~~and-but~~ remained statistically significantly lower than the control pCO₂ values until day 3 on average in the Ca(OH)₂ treatments, and until day 6 on average in the Mg(OH)₂ and NaOH treatments. ~~After 2 months~~Between days 6 and 35, the difference was no ~~longer~~ statistically significant ~~anymore~~(95% confidence, ~~-(Table A21)~~). Over the following ~~four~~ 4-months, pCO₂ slightly varied, but followed fluctuations in the control ~~batchbottles-fluctuations~~. In the 125 RPM experiment, the sharp drop in pCO₂ ~~following~~ after alkaline material addition was ~~quickly-followed~~met with by a sharp increase ~~in pCO₂-over~~ the ~~following-next~~ days. In the case of Ca(OH)₂, the pCO₂ values were no ~~longer~~ ~~statistically~~ significantly lower than those of the control incubations after only 72 h. The Mg(OH)₂ incubations' pCO₂ also quickly increased, becoming ~~non-no longer statistically~~ significantly lower than the controls pCO₂ ~~after 3 dayson day 3~~ for additions of 100 and 200 ~~μmolμmol~~ kg⁻¹, and ~~after 6 dayson day 6~~ for additions of 300-600 ~~μmolμmol~~ kg⁻¹. In both cases, the ~~pCO₂pCO₂-~~remained ~~rather-consistent~~relatively stable over the following ~~six~~6 months, with variations following the same patterns as for the control incubations.

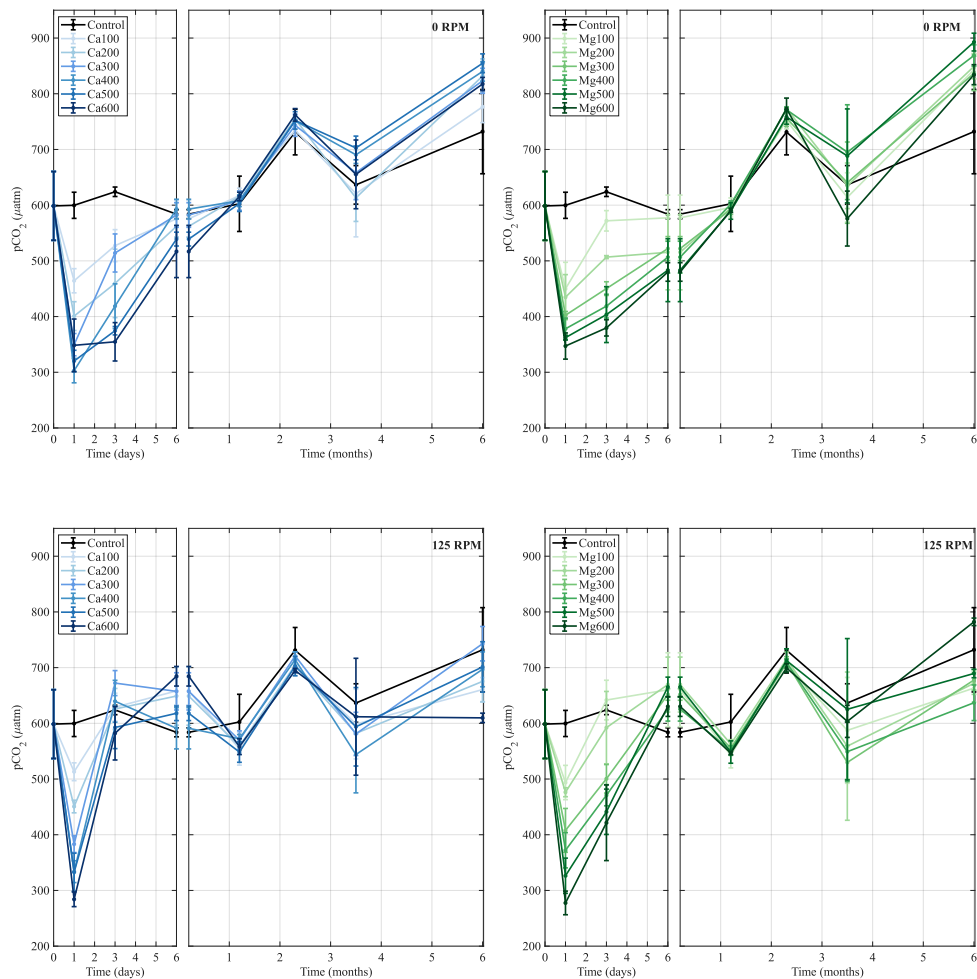


Figure 4:5: Estimated $p\text{CO}_2$ values from measured TA and DIC and reported in μatm , as a function of time. The 0 RPM experiments are represented in the top two tiles, while the 125 RPM experiments are represented in the bottom two tiles. $\text{Ca}(\text{OH})_2$ additions are represented in blue and $\text{Mg}(\text{OH})_2$ additions in green, with the shades representing the amount added, i.e., the darker the shade, the higher alkaline material addition. Each point represents the mean of the triplicate measurements, with error bars representing the uncertainty based on the standard deviation of these three measurements. Each point represents the average of triplicates, with error bars representing the uncertainty based on the standard deviation of these three measurements.

Calculated versus measured DIC

Throughout the experiment, DIC concentrations were measured directly and calculated from TA and pH_T measurements, as described in the “Material and methods” section. The ratio of calculated to measured DIC was determined for each sampling

time point and is presented as individual boxplots, together with the combined dataset (Fig. 5). The boxplot analysis showed that the month 1 and 2 datasets had lower DIC ratios than the combined dataset, with mean ratios of 0.896 and 0.922, respectively, both below the lower quartile of the combined dataset (0.928). In contrast, the month 6 dataset exhibited higher DIC ratios, with a mean of 0.969, exceeding the upper quartile of the combined dataset (0.957). In accordance with the criterion described in the “Material and methods” section, DIC concentrations for the month 1, 2, and 6 datasets were corrected using the average ratio derived from all measurements. Measured DIC values in the month 1 and 2 datasets were higher by $\sim 101 \pm 5 \mu\text{mol kg}^{-1}$ and $\sim 47 \pm 3 \mu\text{mol kg}^{-1}$ respectively, while the month 6 values were lower by $\sim 67 \pm 4 \mu\text{mol kg}^{-1}$.

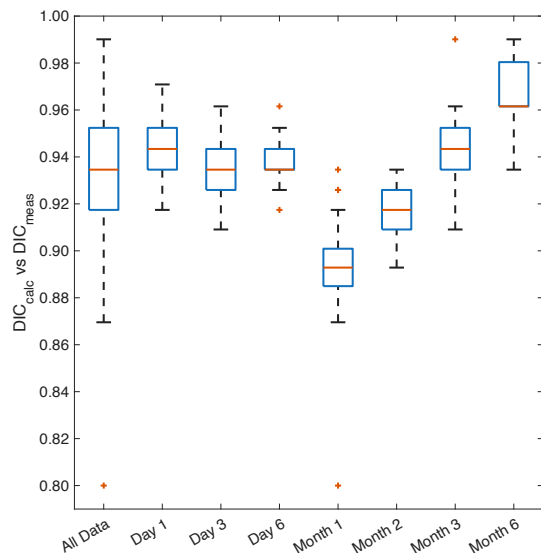


Figure 5: Boxplots of the ratio of DIC calculated from TA and pH_T measurements to measured DIC for the combined dataset and each sampling campaign. Ratios for the month 1 and 2 datasets fall below the ratio for the combined dataset, while the ratio for the month 6 dataset stands above.

CO₂ fluxes and derivation of equivalent wind speeds from DIC measurements

In almost all treatments, a decrease in DIC within the first 24 h was observed, followed by an increase in some cases. A decrease between days 1 and 3 was observed in the 0 RPM treatments and was particularly visible in the $\text{Ca}(\text{OH})_2$ treatment with the highest target addition, i.e., ΔTA target of $600 \mu\text{mol kg}^{-1}$, and in all $\text{Mg}(\text{OH})_2$ treatments with ΔTA targets higher than $200 \mu\text{mol kg}^{-1}$. Between days 1 and 3, rates of DIC change were -4.7 ± 21.2 and $38.2 \pm 25.3 \mu\text{mol kg}^{-1} \text{ d}^{-1}$ for the 0 and 125 RPM experiments, respectively. Using these rates of DIC change, Eq. (1) was solved, taking into account that the water depth in the bottle was $\sim 10.5 \text{ cm}$ at the time. For an average temperature of 24.6°C and salinity of 35.4, K_0 is approximated to $2.932 \times 10^{-2} \text{ mol L}^{-1} \text{ atm}^{-1}$ as per Weiss (1974). The differences between $\text{pCO}_{2\text{w}}$ and $\text{pCO}_{2\text{a}}$ were derived from Fig. 6 and averaged to

221.6 ± 71.2 and 256.5 ± 96.8 μatm for the 0 and 125 RPM treatments, respectively. Solving Eq. (1) for a bottle with an internal diameter of 9.6 cm yielded a $k_{0 \text{ RPM}}$ and $k_{125 \text{ RPM}}$ values of 0.4 ± 0.7 and $2.5 \pm 1.0 \text{ cm h}^{-1}$, respectively. Solving the second equation with a Schmidt number Sc of ~ 533.2 at 24.6°C and salinity 35.4 (Wanninkhof, 2014) yielded corresponding U wind speeds of 1.2 and 3.0 m s^{-1} .

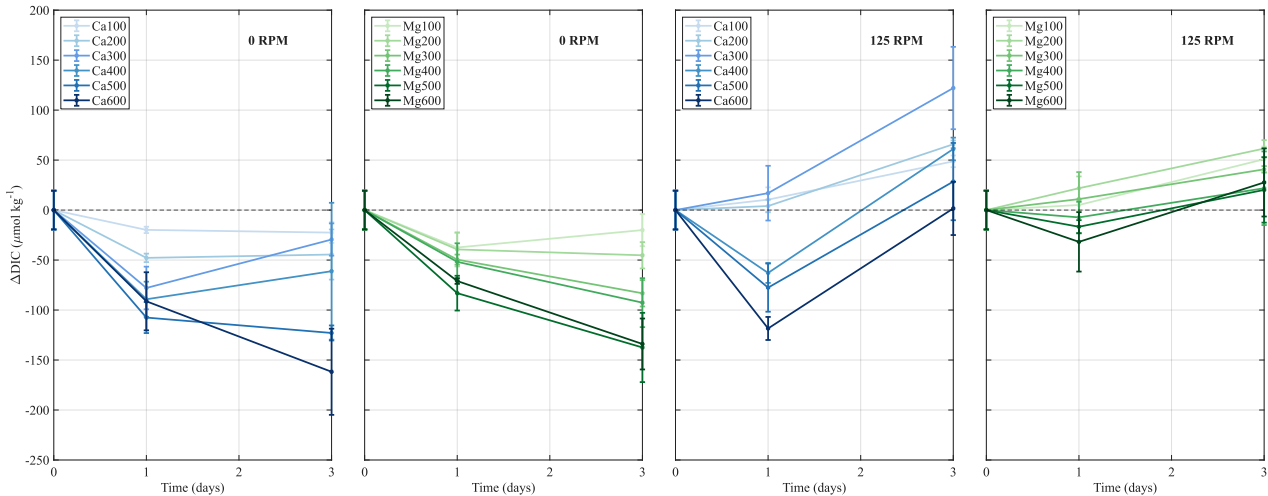


Figure 6: Changes in DIC within the first three days after $\text{Ca}(\text{OH})_2$ - $\text{Ca}(\text{OH})_2$ (blue) or $\text{Mg}(\text{OH})_2$ - $\text{Mg}(\text{OH})_2$ (green) additions with 0 RPM (two left-hand panels) and 125 RPM advection (two right-hand panels). The darker shades indicate higher alkaline material additions. Each point represents the mean of the triplicate measurements, with error bars representing the uncertainty based on the standard deviation of these three measurements.

NaOH experiments and differing patterns

The NaOH experiments revealed differing patterns to those from the $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$ incubations. While NaOH additions were performed identically in static and advected bottles. However, both setups are meant to be identical, a clear difference in ΔTA emerged between the two advection regimes. At 0 RPM, ΔTA only reached about $278 \mu\text{mol} \mu\text{mol kg}^{-1}$ within 24 h and peaked at $337 \mu\text{mol} \mu\text{mol kg}^{-1}$ after three days, corresponding to about 56% of the expected TA (Fig. 76). The 125 RPM experiments quickly reached about $552 \mu\text{mol} \mu\text{mol kg}^{-1}$ after one day, before reaching up to $598 \mu\text{mol} \mu\text{mol kg}^{-1}$ after three months, corresponding to about 100% of the expected TA increase (within measurement uncertainty). DIC followed a similar pattern in both cases, but the final ΔDIC was about approximately half as large in the 0 RPM experiments treatment as compared in to the 125 RPM treatment. Also, while ΔDIC was positive throughout the experiment at 125 RPM,

the 0 RPM experiment first showed a decline ~~at-on~~ day 1 of about 17 ~~$\mu\text{mol}\mu\text{mol}$~~ kg^{-1} before consistently increasing until the end of the experiment. pCO_2 decreased after NaOH addition down to similar levels for both 0 and 125 RPM experiments, i.e., 209 and 167 μatm , respectively. The increase that followed was quicker in the 125 RPM experiments, where, after only ~~six~~ 6 days, pCO_2 was not ~~statistically-longer~~ significantly lower than the controls, while it took at least one month for pCO_2 to be similar to the 0 RPM control (Table A2)-with 0 RPM.

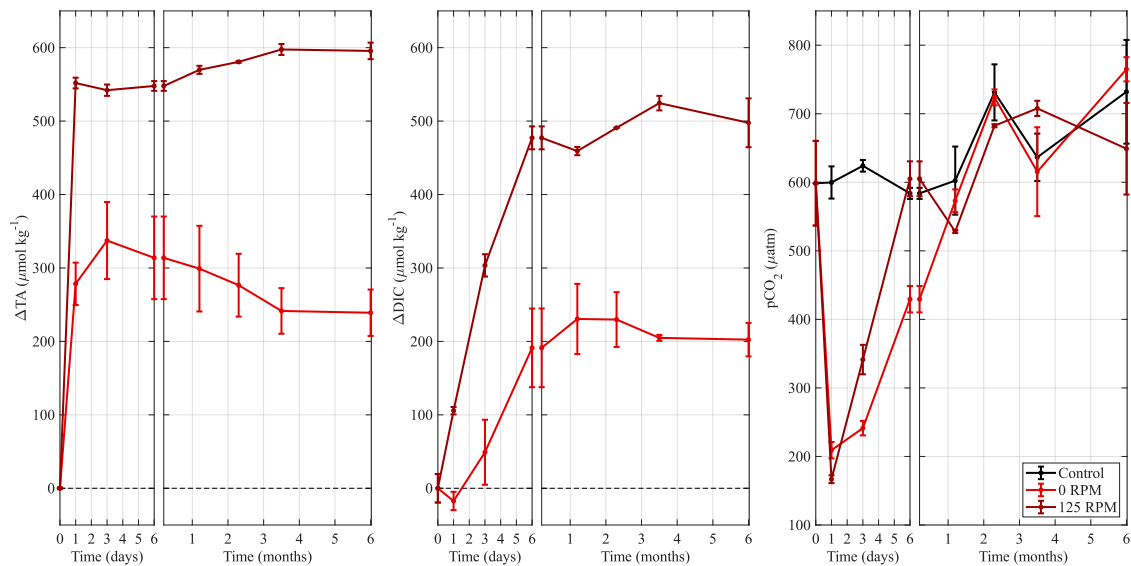


Figure 7: ΔTA , ΔDIC and pCO_2 changes as a function of time for the NaOH experiments. The black colour represents the control experiment, the lighter red colour represents the 0 RPM experiments, and the darker red colour represents the 125 RPM experiments. Each point represents the mean of the triplicate measurements, with error bars representing the uncertainty based on the standard deviation of these three measurements.

TA loss budget, pParticulate analysis and scanning electron microscopy

Particulate analysis was conducted using the precipitates from the highest ~~additions of~~ $\text{Mg}(\text{OH})_2$ addition, i.e., +600 ~~$\mu\text{mol}\mu\text{mol}$~~ kg^{-1} , at both 0 and 125 RPM, to ~~allow for a strong enough signal and decipher~~ provide sufficient material to determine whether CaCO_3 precipitated onto $\text{Mg}(\text{OH})_2$ particles ~~were used as surface area for CaCO_3 precipitation.~~ The sample for the third replicate of the 0 RPM $\text{Mg}(\text{OH})_2$ treatment was lost during analysis, and not all interpretations could be computed in Table 1 for this sample in particular. The data support the TA patterns observed in Fig. 4, where the The Max. TA_{loss} values for the 0 RPM treatment were higher than those for the 125 RPM by ~~experiments run at 0 RPM clearly show a stronger Max. TA_{loss} of~~ about 194 ~~$\mu\text{mol}\mu\text{mol}$~~ kg^{-1} ~~more than those run at 125 RPM.~~ (Table 2). The results obtained from F-AAS revealed higher Ca^{2+} and Mg^{2+} concentrations in the 0 RPM treatments, with values ~~correlate with such observations, where the measured~~

concentrations of particulate Ca^{2+} and Mg^{2+} are about 7.2 and 0.4 mg L^{-1} lower higher in the 125 RPM, respectively. On average, 62.7% $\pm 4.4\%$ and 58.2% $\pm 5.2\%$ of the estimated particulate Ca^{2+} is explained was accounted for by the measured Ca^{2+} , for in the 0 and 125 RPM treatments, respectively. Finally, another major difference emergeds from the measurements of Mg^{2+} , where the 0 RPM treatment yielded has nearly twice as much particulate Mg^{2+} as in the 125 RPM treatment. Overall, the particles collected from the 0 RPM experiments were composed of contained 3.8% $\pm 0.4\%$ of magnesium Mg^{2+} , while the particles collected from the 125 RPM experiments, were about 3.0% $\pm 0.3\%$ magnesium Mg^{2+} .

Table 24: Comparison of expected theoretical and measured changes in TA changes and particulate Ca and Mg concentrations in the 0 and 125 RPM $\text{Mg}(\text{OH})_2$ treatments with a target ΔTA of from the amount of material added, measured TA loss and estimated total loss assuming full dissolution. Ca^{2+} loss from the water column estimated from changes in TA and measured Ca^{2+} and Mg^{2+} from precipitates in the 0 and 125 RPM experiments with the highest addition of $\text{Mg}(\text{OH})_2$, i.e., $\Delta\text{TA} + 600 \mu\text{mol mol}^{-1} \text{ kg}^{-1}$. The table includes the theoretical maximum ΔTA assuming complete dissolution, measured and maximum TA losses, Ca^{2+} concentrations estimated from TA loss, measured Ca^{2+} and Mg^{2+} concentrations, and the A ratio of between measured to estimated Ca^{2+} and estimated Ca^{2+} is also provided. The third replicate of the 0 RPM treatment was lost during analysis. Hence, not all values could be calculated for this sample.

Sample replicate		Max. ΔTA estimated theo ($\mu\text{mol kg}^{-1}$)	Meas. TA_{loss} ($\mu\text{mol kg}^{-1}$)	Max. TA_{loss} ($\mu\text{mol kg}^{-1}$)	Estimated Ca^{2+} (mg L^{-1})	Measured Ca^{2+} Mg^{2+} (mg L^{-1})	% meas vs est. Measur ed/ estimated (%)
		($\mu\text{mol kg}^{-1}$)	($\mu\text{mol kg}^{-1}$)	($\mu\text{mol kg}^{-1}$)	(mg L^{-1})	(mg L^{-1})	
0 RPM	1	596.6	~~237.2~~ 172.6	769.1	38.5	25.3 1.1	65.8%
	2	601.5	~~231.9~~ 199.6	801.1	40.1	23.9 0.9	59.6%
	3	598.1	~~238.2~~ 16.4	814.5	40.7	LOST	LOST
125 RPM	1	605.2	~~177.6~~ 6.0	599.2	30.0	16.8 0.5	56.1%
	2	598.8	~~258.8~~ 66.7	665.5	33.3	18.1 0.5	54.4%
	3	600.6	~~96.9~~ 62.9	537.7	26.9	17.2 0.6	64.1%

From the same incubations, as well as from the highest r treatments of the $\text{Ca}(\text{OH})_2$ addition experiments at 0 and 125 RPM, samples for SEM analysis were taken and analysed. Precipitate analyses revealed clear and strong signals for calcium, oxygen and carbon in all samples, with traces of magnesium, strontium, sodium and chloride, i.e., $<1\%$ on average compared to the above-mentioned elements. While no clear difference in the composition of precipitates was observed, the morphology was clearly different. Precipitates from the $\text{Ca}(\text{OH})_2$ experiments mostly appeared were all needle-like, as shown in Fig. 78. A major difference emerge ds when comparing between these precipitates and those with the precipitates from the $\text{Mg}(\text{OH})_2$

experiments. In ~~both cases~~ the $\text{Mg}(\text{OH})_2$ treatments, the morphology of the crystals was clearly different, with crystals ~~appearing appearing~~ denser and displaying a broccoli-like structure. Although individual crystals were more difficult to distinguish. While it is harder to identify a single crystal, these precipitates were ~~mostly generally larger~~ bigger than those observed in the $\text{Ca}(\text{OH})_2$ experiments (Fig. 78).

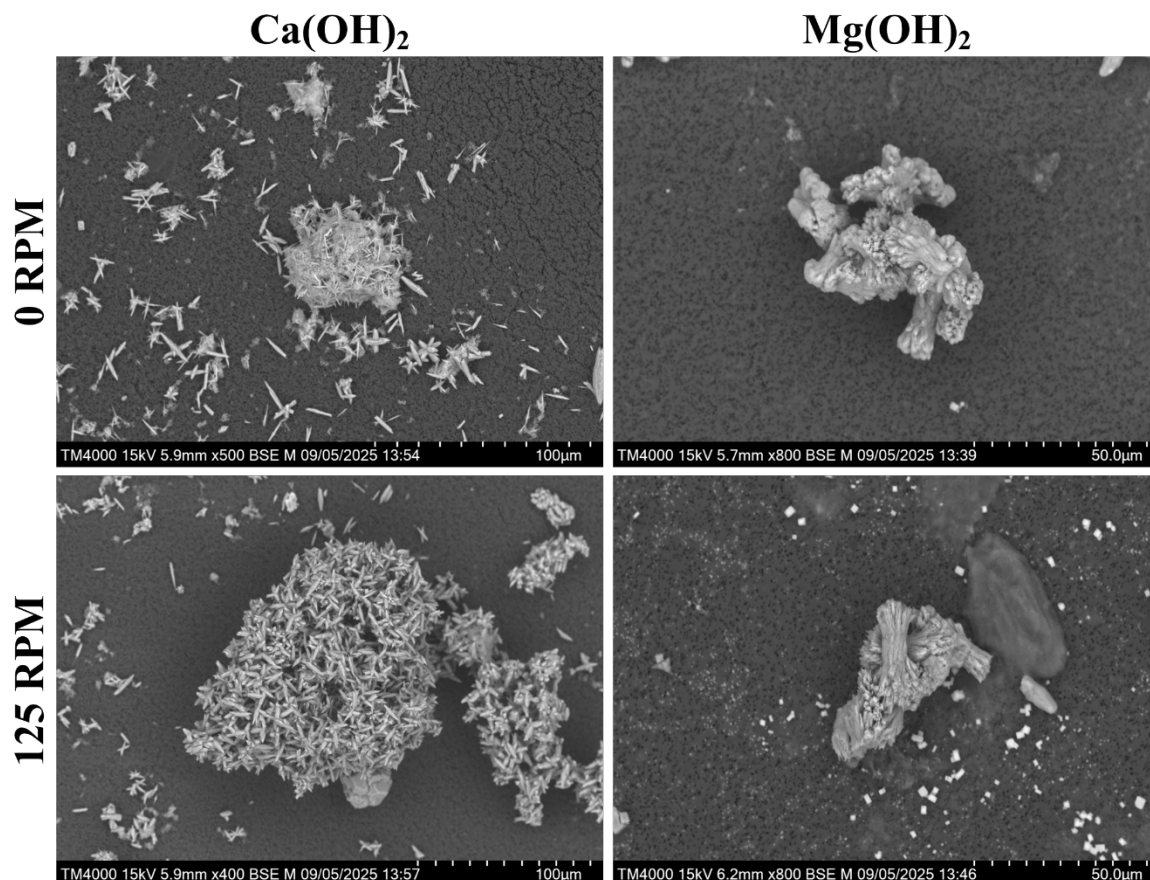


Figure 8: ~~7~~ SEM images of precipitates collected from the $\Delta\Delta\text{TA} + 600 \mu\text{mol kg}^{-1}$ experiments ~~additions amended with of~~ $\text{Ca}(\text{OH})_2$ (left) and $\text{Mg}(\text{OH})_2$ (right), both at 0 RPM (top) and 125 RPM (bottom).

Discussion

Advection affects alkaline material dissolution ~~kinetics~~, CaCO_3 precipitation ~~kinetics~~, and CO_2 ingassing rates

The results indicate that hydrodynamics influence several physicochemical processes in parallel, affecting mineral dissolution, carbonate chemistry, secondary precipitation and realised CDR. Considering real-world applications for OAE is of high

interest but is yet to be fully addressed. In an effort to incorporate natural processes in laboratory OAE experiments, the tests conducted here have deviated from the more classical experimental setup available in the literature and were conducted for an extended period of 6 months with intentional atmospheric CO₂ invasion. A very first clear general observation outcome is that regarding the alkalinity generated through the dissolution of Ca(OH)₂ and Mg(OH)₂ dissolution using a shaking table here was substantially lower than reported in the literature (Hartmann et al., 2023; Moras et al., 2022; Moras et al., 2024; Varliero et al., 2024; Yang et al., 2023). Such an outcome was to be expected as the weaker turbulence generated by the shaking table compared with that produced by a stir bar, for example, could have decreased fluid exchange around the dissolving particles and enhanced their settling. The milder approach of the shaking table, relying on water diffusion and mild vertical mixing compared to the vortex created by stir bars, decreases the dissolution effect and disturbance. Consequently, the dissolution was likely limited by mass transfer through the diffusive boundary layer surrounding particles (Lasaga, 2014; Levich, 1962; Saez Moreno et al., 2026; Sulpis et al., 2019). The reduced advection could have increased the thickness of the diffusive boundary layer surrounding particles, leading to micro-environments enriched in Ca²⁺, Mg²⁺ and OH⁻. These could have reduced the concentration gradient driving diffusion and ultimately decreased dissolution rates. In our case, the low (125 RPM, Reynolds number of $\sim 1.76 \times 10^4$) or absent (0 RPM) advection-turbulence coincided with a decrease in TA generation efficiency and suggest mass-transfer limitations (Drever, 1997; Saez Moreno et al., 2026). At 0 RPM, TA generation was up to 50% lower for the lower additions. At the same time, the persistence of such concentrated micro-environments increased the risk of heterogeneous and pseudo-homogeneous CaCO₃ precipitation (Marion et al., 2009). The competing processes of reduced dissolution and increased CaCO₃ precipitation likely explain the substantially lower measured alkalinity increase, at most half of reported values in literature, especially in the 0 RPM incubations. Although we did not directly measure such chemical gradients and micro-environments, the dissolution and precipitation patterns are consistent with the behaviour expected from thicker diffusive boundary layers surrounding dissolving particles (Lasaga, 2014; Levich, 1962). decreased the speed of diffusion around the particles. Hence, TA generation was up to 50% lower at 0 RPM for the lower additions, due to the slower dissolution and higher saturation in the diffusive boundary layer. Taken together, these findings and previous studies using hydroxide and olivine-rich feedstocks this supports the idea suggest that hydrodynamic conditions solid-based OAE implementation are a first-order control on realised alkalinity generation (Flipkens et al., 2023; Fuhr et al., 2025; Meysman and Montserrat, 2017). Expected alkalinity generation should therefore not be simply inferred from the alkaline material properties but rather a combined result of material properties and surrounding hydrodynamic regime. E should only be considered in environments with strong advection and/or physical disturbances, such as strong porewater exchange or water flow generating high bed shear stress, may therefore be more suitable, as previously suggested for olivine-rich materials (Flipkens et al., 2023; Meysman and Montserrat, 2017).

A This is also valid for similar influence of advection was also observed in the NaOH treatments, in which the two advection regimes produced noticeably different TA increases. Under advected conditions, based OAE. While these incubations were only used as controls for the higher TA additions, a significant difference in TA generations was observed. Once exposed to

575 ~~advection~~, TA reached about 92% of the expected TA within one day before ~~increasing~~ consistently increasing to reach 100%
on day 108 (Fig. 67). The ~~initially lack-reduced of~~ TA generation ~~at the start~~ was not accompanied by a decrease in DIC,
~~suggesting that CaCO₃ formation was unlikely to be the source of the missing TA. is likely due to secondary mineral formation,~~
~~but unlikely in the form of CaCO₃. If it had been the case, a drop in DIC would have been observed early on, and no increase~~
~~in TA later would have been expected as CaCO₃ does not dissolve naturally in open ocean seawater (Morse et al., 2007).~~
580 Similar patterns were observed in the literature and attributed to Mg(OH)₂ formation Hence, the lack of TA generation had to
~~come from the formation of a mineral that only affects TA and not DIC, which is the case of Mg(OH)₂. Mg(OH)₂ precipitation~~
following NaOH addition, ~~both experimentally at pH above ~10 (Haas, 1916; Kapp, 1928) and has been observed in the context~~
of OAE (Hashim et al., 2025; Shaw et al., 2025; Suitner et al., 2024; Varliero et al., 2024). The progressive increase in TA to
100% after three months is consistent with gradual dissolution of temporarily precipitated Mg(OH)₂ (Haas, 1916; Hashim et
585 al., 2025; Kapp, 1928). ~~In our experiments, the absence of advection likely led to Mg(OH)₂ precipitation, which ultimately~~
~~redissolved over time.~~ Secondary mineral formation has recently been considered a loss in TA efficiency for OAE (Lunstrum
et al., 2026). However, this transient Mg(OH)₂ formation may represent a temporary reduction in alkalinity generation rather
than a permanent loss of CDR potential. A contrasting pattern was observed in the static NaOH-amended bottles. The
maximum TA generated only reached about 55% on day 3 before decreasing further. The concomitant decrease in TA and
590 DIC suggests that CaCO₃ precipitation after NaOH addition became the dominant process, likely obscuring the signal of
~~subsequent Mg(OH)₂ dissolution (Fig. 7).~~ While these observations were not the focus of this study and therefore lack further
analyses, they may have implications for OAE efficiency. While initially the loss of TA may lead to a temporary lower
efficiency, it is later compensated by the redissolution of Mg(OH)₂, leading, in our case, to 100% of TA generation. However,
the effect of advection is clearly noticeable in the 0 RPM experiments. Maximum TA generated only reached about 55% on
595 day 3 before decreasing. The concomitant decrease in TA and DIC suggests CaCO₃ precipitation after NaOH addition and in
a runaway fashion during the first months (Fig. 6). While Mg(OH)₂ likely precipitated as well, the further loss of TA and DIC
in a runaway fashion, indicative of CaCO₃ formation, likely hid or undermined visible Mg(OH)₂ dissolution patterns.

Advection also directly affected CO₂ ingassing rates. In the advected bottles, pCO₂ was no longer significantly lower than the
600 control within three days in most treatments, and at most within six days for the higher treatments using Mg(OH)₂ and NaOH.
Without advection, the differences in pCO₂ between amended bottles and controls remained statistically different for at least
six days, and up to 35 days in some instances. The faster ingassing rapidly increased DIC, which in turn rapidly reduced
carbonate supersaturation (Zeebe and Wolf-Gladrow, 2001). This buffering likely decreased the risk of CaCO₃ precipitation.
Together, these observations indicate that hydrodynamic conditions strongly influence dissolution kinetics, atmospheric CO₂
605 ingassing rates and secondary mineral formation, and therefore the alkalinity generation efficiency of OAE.
CO₂ invasion and alkaline feedstock influence Like in the NaOH experiments, CaCO₃ precipitation was suspected in the
experiments with solid feedstocks. The creation of highly alkaline microenvironments led to inevitable secondary CaCO₃
precipitation. Furthermore, the low to inexistent advection led to longer CAP, allowing for more CaCO₃ to be formed. The

presence of such saturated microenvironments, particularly in the diffusive boundary layer of dissolving particles, enhanced heterogeneous CaCO_3 precipitation on feedstock particles (Marion et al., 2009). In all experiments, CaCO_3 formation was suspected, with a clear decrease in DIC after addition in the 0 RPM experiments, and a loss in DIC for all levels above $300 \mu\text{mol kg}^{-1}$ at 125 RPM. This was confirmed by SEM imaging (Fig. 7) and particulate analysis (Table 1). While the measured particulate CaCO_3 only explained about 60% of the TA loss, likely from solid sticking to the bottle's wall, the data support the hypothesis that CaCO_3 formed on dissolving feedstock. The signal of Mg was twice as high in particles from the 0 RPM experiments compared to those of the 125 RPM experiments. This likely resulted from coating of $\text{Mg}(\text{OH})_2$ particles by CaCO_3 , though it could not be successfully confirmed via SEM imaging. Hence, the lower TA generation in all experiments could be attributed to opposing mechanisms: the dissolution of alkaline minerals which increases TA, the loss of TA potential from the CaCO_3 coating of such minerals, and the precipitation of CaCO_3 , which decreases TA. While CaCO_3 precipitation has been well described in literature, a new external factor presented here makes the measurement of the rates harder. The continuous CO_2 ingassing occurring in all experiments maintains a positive increase in DIC over time. Measuring a decrease of TA and DIC in a 2:1 ratio would have given a good indication of CaCO_3 formation. Unfortunately, the CO_2 ingassing created from the added TA negates such a ratio as the loss of DIC from CaCO_3 is compensated by an increase from ingassing (Zeebe and Wolf-Gladrow, 2001). The continuous CO_2 ingassing occurring in all experiments maintains a positive increase in DIC over time. Measuring a decrease of TA and DIC in a 2:1 ratio would have given a good indication of CaCO_3 formation. Unfortunately, the CO_2 ingassing created from the added TA negates such a ratio as the loss of DIC from CaCO_3 is compensated by an increase from ingassing

Different observed CaCO_3 precipitation patterns and morphologies

Although CaCO_3 precipitation occurred in nearly all solid-amended bottles, its pattern and temporal evolution differed between advected and static bottles. CaCO_3 formation in the context of OAE has been described as a runaway process when using such feedstocks, in which the amount of alkalinity removed from the system accelerates over time and may become higher than the amount of alkalinity added (Hashim et al., 2025; Moras et al., 2022; Moras et al., 2024; Suitner et al., 2026). Here, such a process can be behaviour was observed in most of the 0 RPM experiments. However, the pattern loss of alkalinity loss attributed to CaCO_3 precipitation in the 125 RPM treatments differed. A seemingly linear and constant loss can be observed in the $\Delta\text{TA} + 600 \mu\text{mol kg}^{-1}$ of the $\text{Ca}(\text{OH})_2$ experiments with a target ΔTA of $+600 \mu\text{mol kg}^{-1}$, while whereas in for all the other 125 RPM experiments treatments, the early formation loss attributed of to CaCO_3 formation quickly stopped after a few days of incubation. To our knowledge, these patterns have not been reported in the literature for hydroxide-based OAE application before this study or the recent preprint by Saez Moreno et al. (2026).

The decrease and halt in the precipitation process coincided with CO_2 invasion. Under advected conditions, CO_2 invasion occurred more rapidly. When CO_2 dissolves in water, it reacts with a carbonate ion, CO_3^{2-} , to form two bicarbonate ions, HCO_3^- (Zeebe and Wolf-Gladrow, 2001). This consumption of CO_3^{2-} decreases Ω_A , which directly influences CaCO_3 formation rates (Suitner et al., 2026). These observations suggest that rapid CO_2 ingassing may have reduced the risk of runaway CaCO_3

precipitation. Consequently, more rapid CO₂ ingassing likely shortened the CAP, reducing or, in some cases, preventing runaway precipitation. Conversely, slower CO₂ ingassing in the 0 RPM experiments was associated with a longer CAP and runaway CaCO₃ precipitation (Fig. 4). Hence, the risk of loss of TA through CaCO₃ precipitation appears to depend not only on the initial amount of alkalinity and form of alkalinity added, but also on the rate of CO₂ ingassing. These findings complement earlier studies showing that secondary CaCO₃ formation is influenced by mineral surface area, realised alkalinity, saturation state and precipitation inhibitors (Fuhr et al., 2022; Moras et al., 2022; Pan et al., 2021; Rigopoulos et al., 2018; Suitner et al., 2026). It also demonstrates that CO₂ invasion and hydrodynamic conditions regulate CAP duration and directly influence alkalinity generation.

The different rates of CO₂ ingassing resulted in a broader range of Ω_A at which precipitation stopped. Previous OAE experiments with similar materials reported that precipitation ceased at $\Omega_A \sim 2$ (Hashim et al., 2025; Moras et al., 2022; Moras et al., 2024). In our experiments, the observed final Ω_A values were higher, particularly in the 125 RPM experiments, with values ranging between 2.2 and 3.6 (Fig. A2). This discrepancy most likely resulted from differences in experimental setups, particularly the fact that each bottle was open to the air. The continuous atmospheric CO₂ ingassing allowed in the system by the open-bottle design progressively reduced the CaCO₃ supersaturation. These observations suggest that OAE deployments may benefit from settings in which physical and/or air-sea gas exchange rapidly reduce local supersaturation. In natural systems, physical dilution and hydrodynamics in general are expected to reduce local supersaturation on shorter timescales than atmospheric CO₂ equilibration alone. Atmospheric CO₂ invasion should therefore be considered together with physical transport processes. Such settings could include areas with relatively rapid atmospheric CO₂ invasion or horizontal dilution, provided that alkalised water remains near the surface long enough for atmospheric equilibration (Jones et al., 2014; Lunstrum et al., 2026; Zeebe and Wolf-Gladrow, 2001). In the absence of advection, however, particle settling was likely more important, which may have reduced dissolution and CO₂ invasion, eventually promoting CaCO₃ precipitation. Coupled to weaker CO₂ invasion, the absence of fluid exchange around particles could have maintained higher local concentrations of Ca²⁺ and possibly Mg²⁺ around particles, leading to further CaCO₃ precipitation and lower final Ω_A values, i.e., from 1.8 to 2.6.

SEM-EDS analyses were consistent with the formation of CaCO₃ and revealed differing morphologies depending on the alkaline material used (Fig. 8). All crystal morphotypes were morphologically consistent with previously reported aragonite precipitates formed in comparable seawater conditions, i.e., natural seawater at 25 °C, while EDS analyses did not reveal significant incorporation of other measured elements (Morse et al., 1997). Precipitates in using Ca(OH)₂-amended seawater consisted of needle-like CaCO₃ crystals morphologically consistent with aragonite ~~crystals~~ crystals similar to those described in Moras et al. (2022). However, using Mg(OH)₂-amended seawater, precipitates formed spherulitic, broccoli-shaped ~~aragonite~~ CaCO₃ -structures, morphologically consistent with aragonite ~~similar to those~~ synthesised by Suitner et al. (2024). Synthesis of aragonite has been undertaken in a variety of experimental setups, with morphology influenced by various factors such as including varying Mg:Ca ratios and dissolved organic matter (Chave and Suess, 1970; Morse et al., 1997; Pan et al., 2021; Pytkowicz, 1965; Rushdi et al., 1992). ~~Given that~~ Because all of our experiments presented here were conducted in parallel using the same batch of seawater, differences in the starting Mg:Ca ratio or dissolved organic matter concentrations

were negligible. Instead, the differences were more likely related to the alkaline feedstock used. The dissolution of $\text{Mg}(\text{OH})_2$ may have resulted in higher concentrations of Mg^{2+} in the diffusive boundary layer of particles compared to $\text{Ca}(\text{OH})_2$. The role of Mg in delaying and restricting CaCO_3 formation is well-documented in the literature (Morse et al., 1997; Pan et al., 2021; Pytkowicz, 1965). Prolonged exposure to Mg-rich micro-environments has previously been associated with slower aragonite formation, favouring broccoli-shaped over needle-like crystals. Although these micro-environments were not directly quantified here, these observations are consistent with previously reported heterogeneous precipitation occurring within micro-environments with differing chemical compositions (Beck and Andreassen, 2010; Gránásy et al., 2005; Tracy et al., 1998; Zubovic et al., 2025). However, the slower dissolution of $\text{Mg}(\text{OH})_2$ compared to $\text{Ca}(\text{OH})_2$ and its natural CaCO_3 inhibitory effect may be at play (Morse et al., 1997; Pan et al., 2021; Pytkowicz, 1965). In the context of slow dissolving $\text{Mg}(\text{OH})_2$ particles and the low (125 RPM) or inexistent (0 RPM) water movement on the particle surface, higher Ω_A and $[\text{Mg}^{2+}]$ in the diffusive boundary layers were maintained. Precipitation could therefore occur for longer while the inhibitory effect of Mg forced CaCO_3 to crystallise at a slower pace. Such conditions have been reported as favouring broccoli shaped crystals over needle like structures, which could explain the observed structures in the $\text{Mg}(\text{OH})_2$ experiments (Beck and Andreassen, 2010; Gránásy et al., 2005; Tracy et al., 1998; Zubovic et al., 2025).

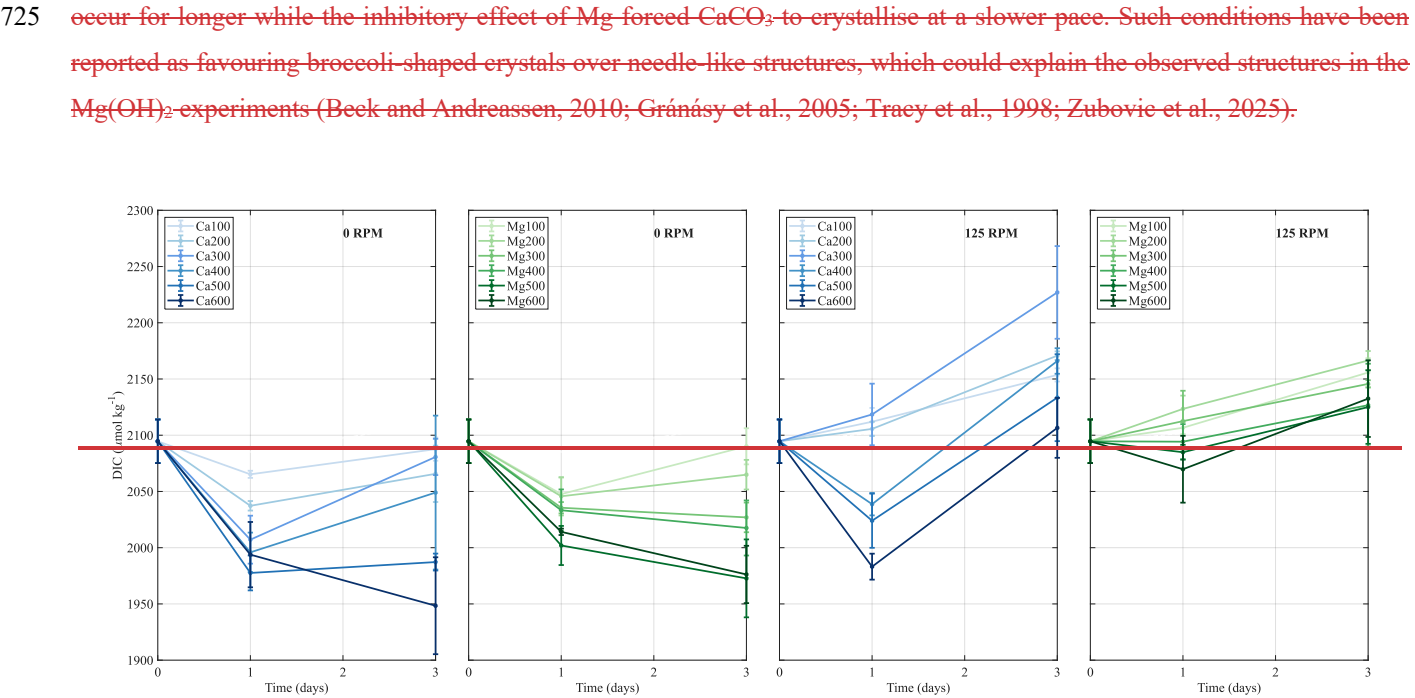
A new CaCO_3 precipitation pattern emerged from the experiments with $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$. CaCO_3 formation in the context of OAE has been described as a runaway process when using such feedstocks, where the amount of alkalinity removed from the system accelerates over time and may become higher than the amount of alkalinity added. Here, such a process can be observed in most of the 0 RPM experiments. However, the loss of alkalinity to CaCO_3 precipitation in the 125 RPM differs. A linear and constant loss can be observed in the $\Delta\text{TA} +600 \mu\text{mol kg}^{-1}$ of the $\text{Ca}(\text{OH})_2$ experiments, while for all the other 125 RPM experiments, the early formation of CaCO_3 quickly stops after a few days of experiments. This unusual halt of precipitation is correlated with the quick increase in DIC following CO_2 ingassing. The duration of the CAP mentioned earlier is likely responsible for such a halt in the precipitation process.

Recent studies have identified thresholds for runaway CaCO_3 precipitation and have correlated its trigger to the time at which the CAP is maintained, Ω_A values, and surface area available (Suitner et al., 2025). Furthermore, precipitation observed here stopped at varying values of Ω_A , while previous experiments observed the stop to occur at $\Omega_A \sim 2$ (Hashim et al., 2025; Moras et al., 2022; Moras et al., 2024). Under such consideration, any process decreasing the CAP would decrease precipitation. One external factor not reported yet in the literature is the CO_2 ingassing. When CO_2 dissolves in water, Ω_A decreases due to the conversion of dissolved CO_2 and carbonate ions, CO_3^{2-} , into two bicarbonate ions, HCO_3^- (Zeebe and Wolf Gladrow, 2001). Hence, allowing for fast CO_2 ingassing limits and decreases the risk for runaway CaCO_3 precipitation by shortening the CAP. This is clearly visible here, where when a low CO_2 ingassing occurs, such as in the 0 RPM experiments (Fig. 5), the CAP is maintained for longer and CaCO_3 precipitates in a runaway fashion. When CO_2 quickly ingasses as in the 125 RPM experiments, the duration of the CAP is significantly shorter, preventing CaCO_3 runaway precipitation. Under such considerations, OAE application should be considered in areas with higher CO_2 invasion rates to minimise the risks of CaCO_3 runaway precipitation (Figs. 3 and 4).

710 Nevertheless, CaCO_3 formation likely occurred in all experiments. DIC fluxes appear smaller between days 0 and 1 than between days 1 and 3 in the 100–300 $\mu\text{mol kg}^{-1}$ additions at 125 RPM, particularly in the $\text{Ca}(\text{OH})_2$ experiments (Fig. 8). The discrepancy likely results from the concomitant DIC loss from CaCO_3 precipitation, and DIC increase from CO_2 ingassing. Such an assumption was confirmed by SEM analyses, where CaCO_3 crystals appeared in two distinctive morphologies and varied based on the alkaline material used (Fig. 7). All crystal morphotypes identified were aragonite, which was expected given that they were formed in seawater, and EDS analyses did not reveal any significant incorporation of other ions. Using $\text{Ca}(\text{OH})_2$ resulted in needle-like aragonite crystals similar to those described in Moras et al. (2022). However, using $\text{Mg}(\text{OH})_2$ resulted in spherulitic (broccoli-shaped) aragonite crystals, similar to those synthesised by Suitner et al. (2024). Synthesis of aragonite has been undertaken in a variety of experimental setups, with various factors such as varying Mg:Ca ratios and dissolved organic matter (Chave and Suess, 1970; Morse et al., 1997; Pan et al., 2021; Pytkowicz, 1965; Rushdi et al., 1992).

715 Given that all experiments presented here were conducted in parallel and using the same batch of seawater, differences in the starting Mg:Ca ratio or dissolved organic matter concentration were negligible. However, the slower dissolution of $\text{Mg}(\text{OH})_2$ compared to $\text{Ca}(\text{OH})_2$ and its natural CaCO_3 inhibitory effect may be at play (Morse et al., 1997; Pan et al., 2021; Pytkowicz, 1965). In the context of slow dissolving $\text{Mg}(\text{OH})_2$ particles and the low (125 RPM) or inexistent (0 RPM) water movement on the particle surface, higher Ω_A and $[\text{Mg}^{2+}]$ in the diffusive boundary layers were maintained. Precipitation could therefore

720 occur for longer while the inhibitory effect of Mg forced CaCO_3 to crystallise at a slower pace. Such conditions have been reported as favouring broccoli-shaped crystals over needle-like structures, which could explain the observed structures in the $\text{Mg}(\text{OH})_2$ experiments (Beck and Andreassen, 2010; Gránásy et al., 2005; Tracy et al., 1998; Zubovic et al., 2025).



730 **Figure :** Changes in DIC within the first three day after $\text{Ca}(\text{OH})_2$ (blue) or $\text{Mg}(\text{OH})_2$ (green) additions with 0 RPM (two left side panels) and 125 RPM advection (two right side panels) advection. The darker shades indicate higher alkaline material additions. Each point represents the average of triplicates, with ranges representing the uncertainty based on the standard deviation of these three measurements.

Environmental implications of the decreased CAP

A primary outcome of this research is the impact of the CAP duration on various geochemical processes. This duration will also dictate potential environmental effects on organisms. It is important to note that any impact of OAE on biology does not result from alkalinity itself but rather from the concomitant changes in the carbonate system (Bach et al., 2019). In fact, the depletion of dissolved CO_2 concentration and the increase in pH following alkalinity addition are most likely responsible for the observed effects. Hence, the duration of the CAP will likely dictate how biology responds to alkalinity addition in an OAE context. Several studies have explored the impacts of OAE on biology, with a focus ranging from phytoplankton to fish larvae (Faucher et al., 2025; Ferderer et al., 2022; Ferderer et al., 2024; Goldenberg et al., 2024; Kousoulas et al., 2025; Paul et al., 2025; Ramirez et al., 2025). Generally, experiments conducted using an equilibrated approach, i.e., alkalinity increase occurring in parallel to DIC increase, resulted in low to negligible pH and pCO_2 changes (Kousoulas et al., 2025; Paul et al., 2025; Ramirez et al., 2025). Without depletion of the CO_2 concentration, organisms' response appeared mild to negligible in most cases, suggesting that equilibrated OAE is safer for implementation. However, using non-equilibrated approaches, i.e., with alkaline materials that only increase TA and not DIC, such as with $\text{Ca}(\text{OH})_2$, $\text{Mg}(\text{OH})_2$ and NaOH , may result in observable effects. In batch cultures, *Emiliana huxleyi* growth was not optimal when exposed to pCO_2 values below $\sim 100 \mu\text{atm}$ following NaOH addition (Faucher et al., 2025). This is consistent with further research on the response of other key phytoplankton groups to OAE, i.e., diatoms (Ferderer et al., 2024; Ferderer et al., 2025). There, the optimal growth of diatoms was estimated at pCO_2 values above $\sim 120 \mu\text{atm}$, where, once exposed to lower values, some negative effects could be observed. However, these observations mostly occurred after days of experiments, which omitted CO_2 invasion in nearly all cases. The absence of CO_2 exchange with the atmosphere allowed for prolonged exposure to low pCO_2 conditions. Such a process would be less likely to occur in natural settings where CO_2 exchanges will take place as soon as alkalinity is released. One can therefore argue that the quick invasion of CO_2 as observed in our experiments may decrease or even negate any effect of OAE on organisms' response. A mesocosms study reports that pCO_2 of $\sim 250 \mu\text{atm}$ led to optimal coccolithophore calcification (Schneider et al., 2026). In our experiments with no advection, low pCO_2 conditions were observed over 1 week to 1 month, though pCO_2 only decreased down to $\sim 200 \mu\text{atm}$ with NaOH (Figs. 5 and 6). This suggests that under the carbonate chemistry conditions observed in our experiments, little to no effect would have been observed on the coccolithophores' response. Furthermore, once exposed to advection, the time of exposure to low pCO_2 levels decreased to a few days, with pCO_2 levels in all experiments promptly returning to background levels after up to 6 days. The decreased CAP duration led to limited maintenance of high pH and low CO_2 levels. The experiments presented here highlight that despite a CO_2 transfer velocity lower than the average in the open ocean (see below), the quick decrease of pH and rapid replenishment of dissolved CO_2 after alkalinity addition could efficiently decrease risks for living organisms. This is still largely overlooked in laboratory scale experiments. Based on observations and assumptions made here, one could argue that OAE effects on biology in laboratory experiments without CO_2 exchanges serve as an upper limit. In open ocean application, the quick dilution and subsequent CO_2

ingassing will most likely significantly decrease the CAP duration, decreasing OAE effects. However, such hypothesis is yet to be tested as open ocean OAE trials are very scarce.

Altered CDR potential of $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$ and implications for MRV

770 The observations reported above have direct implications for estimates of the realised CDR potential of OAE, as both incomplete mineral dissolution and secondary precipitation determine realised alkalinity and therefore the total CO_2 uptake potential. Atmospheric CO_2 ingassing invasion also is a major determinant of OAE efficiency, and particularly its realised CDR potential, of such an approach, but although it is also dependent strongly on the material used. $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$ have appeared in several laboratory scale studies and are among the most widely investigated feedstocks minerals that have been strongly considered for OAE (Eisaman et al., 2023; Moras et al., 2022; Moras et al., 2024; Shaw et al., 2025; Varliero et al., 2024; Yang et al., 2023). From a thermodynamic point of view Thermodynamically, both minerals generate 2 two moles of TA alkalinity per mole of mineral dissolved, corresponding to a theoretical CDR potential of. Assuming the efficiency of 0.75–0.9 moles of CO_2 captured per mole of TA (Schulz et al., 2023), $\text{Ca}(\text{OH})_2$ would allow for 0.9–1.1 tonne of CO_2 per tonne of $\text{Ca}(\text{OH})_2$, and $\text{Mg}(\text{OH})_2$ would allow for 1.1–1.4 t CO_2 t $^{-1}$ tonne of CO_2 per tonne of $\text{Mg}(\text{OH})_2$. Assuming a CO_2 uptake the efficiency of 0.75–0.9 moles of CO_2 captured per mole of TA (Schulz et al., 2023), $\text{Ca}(\text{OH})_2$ would allow for. Such This potential is already decreasesd drastically if, for example, the CO_2 emitted during the calcination of limestone to produce $\text{Ca}(\text{OH})_2$ is not captured (Caserini et al., 2021; Foteinis et al., 2022; Kheshgi, 1995). Our results also point to an additional limitation, Another CDR potential pitfall using either of these hydroxides is the incomplete TA generation presented here. With Under a conservative a theoretical safe OAE implementation deployment scenario, where only a low amount of alkalinity was to be added, e.g., $\Delta\text{TA} + 100 \mu\text{mol} \mu\text{mol} \text{ kg}^{-1}$ (Bach et al., 2025), and moderate exposed to some advection, e.g., 125 RPM, the actual realised TA only reacheds 55–58% (Fig. 42). The realised CDR potential of solid-based OAE therefore depends not only on the chemistry and quantity of material added, but also on hydrodynamic conditions, the fraction of TA realised, and the amount of atmospheric CO_2 ultimately taken up. Although these processes are While more easily readily measurable quantified in a controlled lab setting controlled laboratory setting, they become increasingly difficult to distinguish under open ocean conditions. In principle, incomplete dissolution can be differentiated from competing processes of CO_2 ingassing and CaCO_3 precipitation by measuring both TA and DIC in the system. make the actual measurement of TA generation complicated. DIC is usually used to decipher between incomplete mineral dissolution and CaCO_3 precipitation Limited TA generation without a DIC decrease may indicate incomplete dissolution, while a concomitant decrease in TA and DIC in an approximate 2:1 $\Delta\text{TA}:\Delta\text{DIC}$ ratio may indicate CaCO_3 precipitation. Once but is not suitable when CO_2 ingassing is allowed in the system occurs, the resulting DIC increase hides this stoichiometric relationship, making the quantification of realised TA more difficult. This represents an important challenge for is therefore true if OAE was to be implemented in the ocean, which constitutes major uncertainties for monitoring, reporting and verification (MRV) efforts during field-scale OAE deployments protocols. Current MRV protocols recognise loss terms, such as biotic calcification, additionality and secondary mineral formation, but the distinction between incomplete dissolution

800 and CaCO_3 precipitation becomes more difficult to constrain once atmospheric CO_2 ingassing alters DIC (Bach, 2024; Isometric, 2026; Lunstrum et al., 2026). To date, the first independently verified carbon credits generated from coastal OAE deployment were issued under the Isometric protocol to the Halifax-based company Planetary (Isometric, 2026; Kauk, 2025). The protocol requires monitoring of effluent characteristics, particularly pH, to ensure compliance with regulatory thresholds that must not be exceeded. In the United States, historical Environmental Protection Agency aquatic life criteria established an upper pH guideline of 8.5 and 9.0 for saltwater and freshwater ecosystems, respectively. However, the European Union does not prescribe a single pH limit for surface waters (European Parliament and Council of the European Union, 2000; U.S. Environmental Protection Agency, 1986, 2026). Our observations support the need for such a process-based approach and identify processes that should be considered in a future iteration of MRV protocols.

805 As mentioned earlier, current MRV protocols account for loss of TA to secondary mineral formation through a loss-term (Isometric, 2026). Hence, reliably quantifying secondary ~~Despite high-quality measurement instruments, competing processes likely hide the impacts of each other and increase overall measurement uncertainties, as highlighted by the somewhat large standard error from these experiments (Figs. 4, 5 and 6).~~

810 CaCO_3 precipitation remains is undoubtedly one of the major uncertainties for OAE efficiency and MRV—a major pitfall for MRV protocols for OAE. Although precipitated CaCO_3 may contribute to long-term CDR, as it decreases its formation in surface waters consumes alkalinity that would otherwise remain available for atmospheric CO_2 uptake, and therefore reduces the ~~its overall near-term~~ CDR potential of OAE and may lead to unwanted seawater changes (Lunstrum et al., 2026; Moras et al., 2022). While it may occur both with particulate and liquid feedstocks, as well as in both equilibrated and unequilibrated approaches, the fate of the precipitates is still hard to comprehend and safely report on. A general assumption is that, Although secondary given the oversaturation of seawater with respect to aragonite, any precipitated CaCO_3 is unlikely to redissolve rapidly aragonite crystal will not dissolve in natural open-ocean surface waters, it (Morse et al., 2007). Hence, the crystals will ultimately may sink and be transported settle onto coastal sediments or be exported to into the deeper ocean, where it may later dissolve (Suitner et al., 2025). Such sinking may come at hand as the removal of surface area from the surface layer may prevent further precipitation. Once transported in the deeper waters, CaCO_3 particles may redissolve in areas of undersaturated seawater (like below the carbonate compensation depth) or be come buried in sediments where further dissolution may occur until transported below the reactive layer (Dale et al., 2024; Fuhr et al., 2023; Fuhr et al., 2024; Fuhr et al., 2025; Suitner et al., 2026; Fuhr et al., 2025). This temporary loss of CDR potential could become a second source of alkalinity overtime, though it is yet to be fully addresses. However, Nevertheless, because secondary CaCO_3 in our experiments precipitation mostly formed heterogeneously occurred on added particles, export of coated undissolved feedstock could reduce further alkalinity generation within the surface layer, constituting an additional near-term, as explained earlier, another loss in CDR potential

825 results from the coating of alkaline feedstock particles and their inevitable removal from the surface layer via sinking. It is acknowledged, though, that in bottle experiments, the sinking of CaCO_3 does not happen, which leads to a higher surface area remaining available for further precipitation to occur.

830

The morphology of secondary CaCO_3 precipitates may also influence such transport. Yet, the persistence of undissolved alkaline material in the ocean's surface layer is a prerequisite for higher CDR potential. Sinking of particles needs to be thoroughly understood, and coupling CaCO_3 formation rates with sinking speed is yet to be addressed. Residence times of particles in the surface layer of the ocean are correlated to the size and density of said particles (Baltruschat et al., 2026; Suitner et al., 2025). Using Based on modelling considerations from Suitner et al. (2025/2026), the porous one could hypothesise that the amount of $\text{Ca}(\text{OH})_2$ coated by needle-like structures observed in the $\text{Ca}(\text{OH})_2$ -amended bottles are unlikely to inhibit dissolution because seawater can seep between the needle structures and reach undissolved particles. In contrast, the slower dissolution of $\text{Mg}(\text{OH})_2$ seems to favour slow-growing, denser broccoli-shaped structures. The higher Ca^{2+} and Mg^{2+} concentrations measured in the precipitates from the 0 RPM experiments support the interpretation that CaCO_3 coated $\text{Mg}(\text{OH})_2$ particles (Table 2). Such coatings CaCO_3 , if any, is negligible. Their uneven formation allows for water to seep between the needle structures and reach undissolved $\text{Ca}(\text{OH})_2$. However, such needles may remain in the surface layer for weeks to months which offers surface area for further CaCO_3 precipitation to occur. Using $\text{Mg}(\text{OH})_2$ particles would lead to the opposite. Given that the slower dissolving $\text{Mg}(\text{OH})_2$ allows for a higher CAP in the diffusive boundary layer, coating of such particles may occur more efficiently. This was identified by the F-AAS analyses conducted here, which resulted in a loss of TA generation. The broccoli-shaped structures led to denser and larger particles enhancing their sinking velocity. These may could simultaneously reduce dissolution efficiency while increasing particle density and therefore thereby transport promoting export of the coated undissolved $\text{Mg}(\text{OH})_2$ particles out from of the surface mixed layer. Conversely, rapid particle export, though the quick sinking occurring within days may be sufficient could to significantly decrease the available surface area available for further precipitation, limiting the loss of TA to CaCO_3 . (Baltruschat et al., 2026; Suitner et al., 2026). Although highly relevant to OAE applications, such physical processes are poorly represented in bottle experiments. In natural systems, vertical and horizontal transport, tidal mixing and dilution would further reduce local supersaturation and available surface area for precipitation, as may occur in point-source alkalinity release (Lee et al., 2026). While this was not tested here, these processes could shorten the CAP duration more rapidly than atmospheric CO_2 ingassing alone because CO_2 equilibration occurs over longer timescales in the open ocean and varies substantially with latitude and season (Zhou et al., 2025). Better constraining these coupled physical and geochemical processes should therefore be a priority for future OAE research. Dedicated experiments incorporating realistic hydrodynamics, atmospheric CO_2 equilibration and particle transport would improve estimates of realised CDR and provide valuable constraints for future MRV protocols. These statements are only hypotheses and require further work.

Implications of the CAP duration for CO_2 uptake and environmental impacts **Linking laboratory data with modelling studies and future OAE research**

A primary outcome of this research is the close coupling between the CAP duration and multiple geochemical processes. Hydrodynamics influenced mineral dissolution and CO_2 invasion, while CaCO_3 precipitation and CO_2 invasion affected carbonate chemistry and determined the duration of the CAP. During the first few days after alkalinity addition, DIC

measurements yielded estimated net CO₂ fluxes inferred from DIC changes of -4.7 ± 21.2 and 38.2 ± 25.3 $\mu\text{mol kg}^{-1} \text{d}^{-1}$ for the 0 and 125 RPM experiments, respectively. However, these rates coincided with a decrease in TA during that time, most likely resulting from CaCO₃ formation, as indicated by the simultaneous decrease in TA and the negative apparent flux of CO₂ in the 0 RPM treatments. This simultaneous removal of DIC likely biased the inferred CO₂ fluxes. Estimated DIC loss to CaCO₃ can be inferred from TA loss, where for each mole of TA lost to CaCO₃, 0.5 mole of DIC is removed. This resulted in about 10.9 ± 6.3 and 4.8 ± 4.0 $\mu\text{mol kg}^{-1} \text{d}^{-1}$ in the 0 and 125 RPM treatments respectively, corresponding to about 1.1 and 0.5 mg CaCO₃ $\text{kg}^{-1} \text{d}^{-1}$. Correcting for this loss substantially increased the estimated CO₂ fluxes, especially in the 0 RPM treatments, yielding 5.3 ± 15.5 and 43.0 ± 25.1 $\mu\text{mol kg}^{-1} \text{d}^{-1}$ for the 0 and 125 RPM experiments, respectively. This supports the need to combine carbonate chemistry measurements with direct CO₂ flux estimates, because in field applications, such uncertainty could lead to underestimation of the realised CDR.

The corrected fluxes correspond to gas-transfer velocities $k_{0 \text{ RPM}}$ and $k_{125 \text{ RPM}}$ of 0.8 ± 0.6 and 2.8 ± 0.9 cm h^{-1} , respectively. These values are lower than typical mean open-ocean gas-transfer velocities, commonly averaging between 14 and 20 cm h^{-1} , though records range from nearly 0 to roughly 50 cm h^{-1} . Although derived from simplified laboratory experiments, these gas-transfer velocities provide a useful first-order approximation for comparison with empirical air-sea gas exchange parametrisations. The derived wind speeds should therefore be interpreted as gas-exchange equivalents rather than direct hydrodynamic analogues of open-ocean conditions (Frew et al., 2007; Yu et al., 2013; Zhou et al., 2023). Using the parametrisation of Wanninkhof (2014), corrected equivalent wind speeds were estimated to be 1.1 and 3.5 m s^{-1} for the 0 and 125 RPM experiments, respectively. Current understandings of the global CDR potential of OAE mostly rely on models and rarely implement laboratory experiment results. Correlating results obtained in bottle experiments with global modelling remains largely overlooked and is still a major shortcoming in OAE research. Using the DIC fluxes and pCO₂ data collected in these experiments, a first attempt to put ours into modelling studies context can be undertaken. Using the first steps after alkalinity addition, i.e., between days 1 and 3, DIC increases were 21.3 ± 14.3 and 39.9 ± 25.3 $\mu\text{mol kg}^{-1} \text{d}^{-1}$ for the 0 and 125 RPM experiments, respectively (Fig. 8).

It is acknowledged that uncertainties arise from CaCO₃ precipitation that occurred at varying rates based on the mineral and the amount added. Yet, using those data and the difference in pCO₂ between background and seawater at day 1, an estimate of the gas transfer velocity k (in cm h^{-1}) can be derived and compared to natural systems using the air-sea gas fluxes equation from Wanninkhof (2014):

$$F = kK_0(p\text{CO}_{2w} - p\text{CO}_{2a})$$

where F is the flux of CO₂ between the atmosphere and the seawater bottle between day 1 and 3, expressed in $\text{mol m}^{-2} \text{d}^{-1}$, K_0 is the solubility of CO₂ in seawater (obtained from Weiss (1974), estimated at $2.932 \cdot 10^{-2} \text{ mol L}^{-1} \text{atm}^{-1}$ at 24.6°C and for a salinity of 35.4 in our experiments), and $p\text{CO}_{2w}$ and $p\text{CO}_{2a}$, the partial pressures of the seawater and air in atm, respectively, at day 1 following alkalinity addition. Given that F , K_0 and the $p\text{CO}_2$ difference are known, as well as the depth of the water at day 3, i.e., -10.5 cm, and the density of seawater of salinity 35.346 at 24.6°C (Millero and Poisson, 1981), the gas velocities

900 for the 0 and 125 RPM experiments can be estimated, i.e., $k_{0 \text{ RPM}}$ and $k_{125 \text{ RPM}}$, respectively (the authors emphasize that even though these may not fully reflect natural conditions, they provide a first useful initial approximation):

$k_{0 \text{ RPM}}$ was estimated at $1.4 \pm 0.6 \text{ cm h}^{-1}$, while exposing the bottles to an advection speed of 125 RPM led to a $k_{125 \text{ RPM}}$ of $2.7 \pm 1.1 \text{ cm h}^{-1}$. These gas velocities are significantly lower than those observed naturally in the open ocean (averaging between 14 and 20 cm h^{-1} , though ranging from nearly 0 to roughly 50 cm h^{-1}) but may be correlated to wind speed (Frew et al., 2007; Yu et al., 2013; Zhou et al., 2023). Estimates of the corresponding wind speed in our experiments can be derived from the gas exchange wind speed relationship from Wanninkhof (2014):

$$k = 0.251 U^2 \left(\frac{Sc}{660} \right)^{-0.5}$$

910 where U^2 is the wind speed at 10 m height in m s^{-1} , and Sc is the Schmidt number at 24.6°C and salinity 35.4. Resolving this equation at 24.6°C corresponds to wind speeds of about $2.2\text{--}1.7 \text{ m s}^{-1}$ and $3.4\text{--}1 \text{ m s}^{-1}$, for the 0 and 125 RPM experiments, respectively. Wind speeds in the open ocean vary greatly, from less than 1 m s^{-1} to over 20 m s^{-1} during heavy wind and storm periods, with a mean of approximately averaging 6 m s^{-1} (Ho et al., 2006; Miller et al., 2010; Risien and Chelton, 2008). Hence, the data presented here provide a first-order approximation to gas transfer velocities associated can be compared

915 to with low to moderate wind scenarios and may provide useful constraints be used for modelling low-wind or weak gas-exchange in specific geographical areas conditions, mostly around the Equator (Tokinaga and Xie, 2011). Stronger winds may enhance wave-driven turbulence, in turn enhancing mineral dissolution and air-sea CO_2 exchange. However, their overall influence on the CAP duration will also mostly depend on mixed-layer deepening, dilution, and vertical transport, which will determine the residence time of alkalised water near the surface (Cushman-Roisin and Beekers, 2011; England and Bach, 2025; Flipkens et al., 2023; Tokinaga and Xie, 2011).

920 Geographical areas where lower wind speeds are observed coincide with lower wave heights (Tokinaga and Xie, 2011). While the effect of wave height on the dissolution of alkaline materials is not fully constrained, results presented here and published research (England & Bach, 2025; Flipkens et al., 2023) support that the stronger the water movement and disturbance, the quicker the mineral dissolution. Our data would then represent the lower end of open ocean OAE application. Also, given that

925 surface water subduction speed is not correlated to wave height, alkaline surface water may remain in contact with the atmosphere for longer (Cushman-Roisin and Beekers, 2011). Hence, regions with stronger winds and higher waves may allow for faster mineral dissolution and CO_2 ingassing rates than presented here, yielding higher efficiency. That also suggests that CaCO_3 precipitation from microenvironments may occur less likely, or at least not as fast. Implementing the data presented here in a global model may yield lower overall CDR estimates, but most likely conservative, which can therefore serve as a

930 base for further work and considerations in MRV protocols.

Although the present study did not focus on biological responses to OAE perturbations, the observed carbonate chemistry changes may provide a useful framework to interpret previously published OAE-focused laboratory and mesocosm research.

Biological responses to OAE are generally expected to arise primarily from associated changes in carbonate chemistry, including increased pH and reduced dissolved CO₂ concentration, rather than from TA increase in isolation (Bach et al., 2019).

Several studies have explored the impacts of OAE on a wide range of organisms, from single-species laboratory incubations to large-scale mesocosm campaigns encompassing natural phytoplankton communities and fish larvae (Faucher et al., 2025; Ferderer et al., 2022; Ferderer et al., 2024; Goldenberg et al., 2024; Kousoulas et al., 2025; Paul et al., 2025; Ramírez et al., 2025; Schneider et al., 2026). Experiments conducted using an equilibrated approach, i.e., a TA increase accompanied by a DIC increase, resulted in low to negligible pH and pCO₂ changes (Kousoulas et al., 2025; Paul et al., 2025; Ramírez et al., 2025). Without depletion of the CO₂ concentration, biological responses were small to negligible in most cases. However, non-equilibrated approaches with materials such as Ca(OH)₂, Mg(OH)₂ and NaOH increased TA without an accompanying initial increase in DIC, resulting in longer periods of elevated pH and reduced pCO₂. Under such conditions, reduced growth of *Gephyrocapsa (Emiliana) huxleyi* was reported at pCO₂ values below ~100 µatm following NaOH addition (Faucher et al., 2025). Similar thresholds were observed for other key phytoplankton groups, particularly diatoms (Ferderer et al., 2024; Ferderer et al., 2025). Optimal growth occurred at pCO₂ values above ~120 µatm, with some observed negative effects at lower values. Together, these studies indicate that ecological responses to depend less on TA perturbation than on the magnitude and duration of high pH/low dissolved CO₂ perturbations.

Importantly though, these observations resulted from batch and mesocosm experiments with limited CO₂ ingassing. Limited CO₂ invasion likely prolonged exposure to low pCO₂ conditions for nearly 50 days in some cases (Ferderer et al., 2024; Goldenberg et al., 2024; Schneider et al., 2026). Such prolonged exposure may be representative of point-source alkalinity release, although dilution and CO₂ invasion will most likely decrease such duration. Our experiments support this assumption. In the absence of advection, reduced pCO₂ conditions persisted between one week and one month, although minimum pCO₂ values remained around 200 µatm following NaOH addition (Figs. 4 and 7). These concentrations remained close to reported optimal values for coccolithophore calcification in a large-scale mesocosm study (Schneider et al., 2026), suggesting, based on previously reported thresholds, that severe biological effects would have been unlikely under our experimental carbonate chemistry conditions. When advection was introduced, the CAP duration decreased substantially, which resulted in reduced pCO₂ conditions for only a few days. With pCO₂ returning to background values within three to six days, we hypothesise that, under higher gas-transfer velocities and corresponding equivalent wind speeds, atmospheric CO₂ invasion may decrease pH and replenish dissolved CO₂ concentrations more rapidly, substantially reducing environmental stress from OAE.

When considered together with previous (semi)-closed bottle and mesocosm experiments, our results indicate that studies conducted with limited CO₂ exchange may serve as an upper limit for quantifying biological responses to non-equilibrated OAE. In the open ocean, CO₂ ingassing and physical processes, particularly dilution and mixing, are expected to substantially decrease the CAP duration, thereby reducing the magnitude and duration of biological exposure to altered carbonate chemistry. Although this hypothesis is consistent with the observations reported here, it remains to be tested under realistic field conditions. Such testing will be important for assessing the environmental feasibility of large-scale OAE deployment. Several limitations should be considered when interpreting these findings. First, the experiments only reproduced simplified

hydrodynamic conditions and did not fully capture the complexity of ocean turbulence, wave breaking, or horizontal and vertical transport. Second, repeated sampling progressively reduced the volume in each bottle, changing the headspace and potentially influencing gas exchange over time, although all treatments underwent the same sampling schedule. In addition, airflow rate and relative humidity were not measured at the individual outlets, and differences in headspace renewal among bottles therefore cannot be excluded. Nevertheless, disregarding biological processes, particle transport and large-scale dilution allowed physicochemical processes to be isolated. The results should therefore be interpreted as process-based laboratory observations that inform, rather than reproduce, field-scale OAE deployment.

Conclusion

This study experiments presented presents here report on the first long term (six months) laboratory incubations for OAE with using varying alkalinity increases using $\text{Ca}(\text{OH})_2$, $\text{Mg}(\text{OH})_2$ and NaOH , across a range of alkalinity additions and contrasting hydrodynamic conditions in stagnant and advected seawater. Our results indicate that hydrodynamic conditions are a primary control on the realised efficiency of hydroxide-based OAE. Reduced advection limited mineral dissolution, favoured the persistence of supersaturated micro-environments around the dissolving particles, and reduced atmospheric CO_2 invasion through lower gas-transfer velocities. Ultimately, this resulted in lower realised alkalinity generation, a lower OAE efficiency and reduced CDR potential. In contrast, moderate advection enhanced mineral dissolution, accelerated CO_2 ingassing, shortened the CAP and limited efficiency losses due to secondary mineral formation. The CAP duration emerged as a useful indicator linking hydrodynamic conditions with the persistence of altered carbonate chemistry, the likelihood of secondary precipitation, and the duration of biological exposure to altered carbonate chemistry. Longer CAPs coincided with slower mineral dissolution, reduced gas exchange and prolonged supersaturation, thereby increasing the risk of CaCO_3 formation. Although the laboratory-derived gas-transfer velocities may not adequately represent natural conditions, our observations suggest that stronger mixing and more efficient gas exchange may reduce the CAP duration and improve realised alkalinity generation, although this hypothesis requires validation under realistic field conditions. These results also have implications for MRV protocols, because realised alkalinity generation cannot simply be inferred from feedstock dissolution alone. Processes such as mineral dissolution, CO_2 ingassing and secondary mineral formation all alter carbonate chemistry in different and sometimes opposite directions. Incorporating these coupled processes into future MRV frameworks is of primary importance for accurately quantifying realised CDR in field-scale deployments. Finally, the contrasting morphologies of CaCO_3 precipitates following $\text{Ca}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$ additions suggest feedstock-specific precipitation pathways and particle transport behaviour. These may further influence dissolution efficiency, secondary precipitation and alkalinity stability in the mixed layer, all relevant for CDR accounting. These processes remain poorly constrained in laboratory settings and should be studied under realistic hydrodynamic conditions, including water mixing, dilution, particle sinking and CO_2 equilibration. Better quantifying these coupled geochemical and physical processes will improve confidence in process-based models, reduce uncertainty in CDR estimates, and support the safe, effective and verifiable. A first important outcome was the lower measured

TA compared to the expected one. Maximums of ~20-30% at 0 RPM and ~60% at 125 RPM of expected TA were measured following low alkalinity additions, before decreasing down to less than 10% at 0 RPM and ~10-35% at 125 RPM. However, all experiments with solid feedstocks revealed CaCO_3 precipitation, with more formation occurring without advection. The CaCO_3 formation was directly correlated to the amount of $\text{Mg}(\text{OH})_2$ added at 125 RPM, while above targets of $300 \mu\text{mol kg}^{-1}$ with $\text{Ca}(\text{OH})_2$, the TA generation decreased while CaCO_3 precipitation rate increased at 125 RPM. These precipitated CaCO_3 particles did not eventually redissolve during the 6-month experiments. Using NaOH revealed co-precipitation of CaCO_3 and $\text{Mg}(\text{OH})_2$ at 0 RPM, which did not redissolve, yielding a maximum TA generation of ~55%. At 125 RPM, CaCO_3 precipitation, if any, was negligible, and the early $\text{Mg}(\text{OH})_2$ eventually redissolved after 3 months, leading to 100% of the TA release. CaCO_3 precipitation occurred in two distinct morphologies based on the solid feedstock used. Needle-like aragonite crystals were formed when using $\text{Ca}(\text{OH})_2$, while broccoli-shaped aragonite was formed with $\text{Mg}(\text{OH})_2$, most likely coating $\text{Mg}(\text{OH})_2$ particles. In open ocean applications, the permanence of the needle-like structure in the surface layer may trigger further CaCO_3 precipitation, while the sinking of broccoli-like crystals would remove surface area for further precipitation in the surface layer. Finally, the 125 RPM experiments corresponded to open ocean wind speeds of about 3.1 m s^{-1} , which is about half of the average open ocean wind speeds. The patterns presented here may be significantly decreased as stronger CO_2 ingassing rates in the open ocean will limit the persistence of the CAP around particles. Thus, this may not significantly impact the overall CDR potential of OAE. The quick CO_2 ingassing would also allow for a fast decrease in pH and replenishing of the CO_2 concentration after alkalinity addition, decreasing side effects on living organisms. implementation of OAE at the field scale. Eventually, once the role of sinking particles and mixing with water below the surface has been addressed in further experiments, the interaction between CO_2 ingassing, saturation change and creation of surface area will allow to constrain the efficiency of the different OAE approaches. This will become possible using process-based models, once all relevant factors addressed here can be parameterized.

Appendix A



Figure A 1: Photograph of the setup with tubing maintained to each bottle rim with a binder clip. The air splitter apparatus can be distinguished in the background.

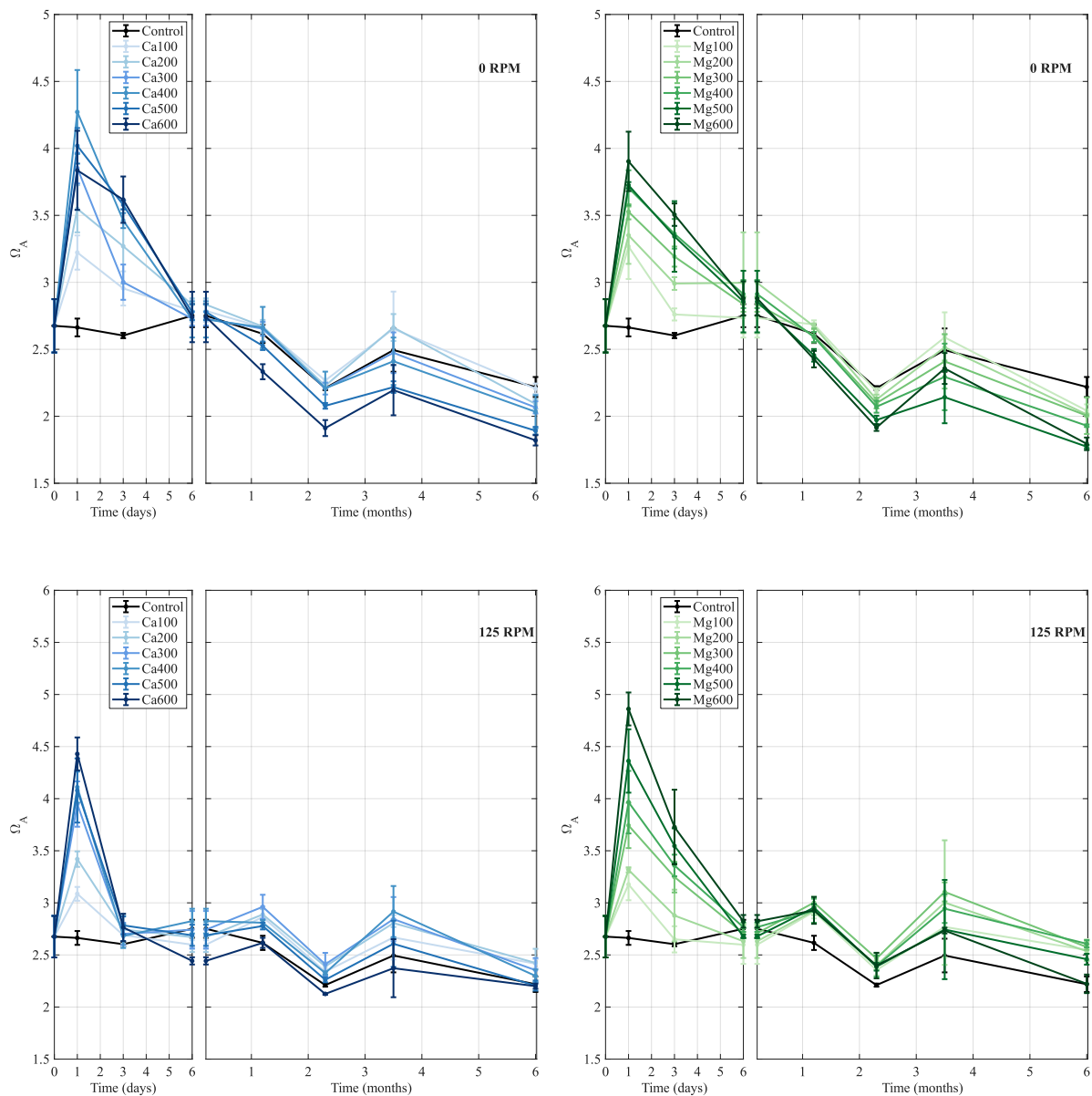


Figure A-24: Changes in Ω_A over time following Ca(OH)_2 (blue) or Mg(OH)_2 (green) additions with 0 RPM (top) and 125 RPM advection (bottom). The darker shades indicate higher alkaline material additions. Each point represents the **average mean** of triplicates, with **ranges-error bars** representing the **uncertainty-based-on-the-one** standard deviation of these three measurements.

1035 **Table A 1: Amount of material added per treatment and target TA (mean of triplicates ±standard deviation) in mg L⁻¹.**

Treatment	+100	+200	+300	+400	+500	+600
0 RPM Ca(OH) ₂	3.67 ±0.06	7.36 ±0.11	11.09 ±0.03	14.85 ±0.05	18.51 ±0.07	22.26 ±0.09
0 RPM Mg(OH) ₂	2.88 ±0.05	5.85 ±0.09	8.72 ±0.06	11.71 ±0.07	14.57 ±0.04	17.46 ±0.07
125 RPM Ca(OH) ₂	3.70 ±0.08	7.47 ±0.01	11.10 ±0.08	14.81 ±0.02	18.51 ±0.06	22.24 ±0.08
125 RPM Mg(OH) ₂	2.95 ±0.04	5.84 ±0.10	8.73 ±0.06	11.61 ±0.05	14.63 ±0.12	17.54 ±0.10

1040 **Table A- 24: ~~Compiled Holm-adjusted one-sided p-values from planned post hoc treatment-control contrasts of the linear mixed-effects model, testing two-samples t-test to determine whether a given treatment pCO₂ in each treatment was statistically significantly lower than that of the corresponding control pCO₂ for at each sampling time-specific day. The Holm adjustment was applied across all treatment-control comparisons within each sampling time. Values in bold indicate the treatments statistically significant differences (adjusted p < 0.05). Adjusted p-values below 0.01 are reported as < 0.01, statistically significantly lower with a confidence level of 95% (p-value < 0.05) are bolded, and those with a confidence level of 90% (p-value ≤ 0.1) are indicated by a *.~~**

0 RPM	Ca(OH)₂						Mg(OH)₂						NaOH
RPM	Ca(OH)₂						Mg(OH)₂						NaOH
ΔTA	100	200	300	400	500	600	100	200	300	400	500	600	600
ΔTA	00	00	00	00	00	00	00	00	00	00	00	00	0
Day 1	<0.01	<0.01	<0.01	<0.01	<0.01	<0.01	<0.01	<0.01	<0.01	<0.01	<0.01	<0.01	<0.01
1	0.02	0.01	0.00	0.00	0.00	0.00	0.01	0.01	0.01	0.01	0.01	0.00	.00
Day 3	<0.01	<0.01	<0.01	<0.01	<0.01	<0.01	0.02	<0.01	<0.01	<0.01	<0.01	<0.0	<0.01
3	0.02	0.01	0.01	0.00	0.00	0.00	0.02	0.02	0.01	0.00	0.00	10.0	.00
							0.07					00.0	
							*					0	
Day 6	1.00	0.79	1.00	1.00	0.20	0.03	1.00	0.03	0.05	0.01	<0.01	<0.01	<0.01
6	0.30	0.16	0.40	0.55	0.03	0.00	0.33	0.00	0.01	0.00	0.04	0.00	.00
	0.27	0.06	0.33	0.60	0.00	0.05	0.33	0.10	0.01	0.01			
		*											
Day 35	0.71	0.71	0.71	0.71	0.67	0.71	0.55	0.55	0.50	0.67	0.49	0.49	0.130
35	0.23	0.21	0.14	0.14	0.10	0.21	0.07	0.06	0.05	0.10	0.04	0.04	010.0
		0.04	0.04	0.02	0.02	0.04	0.02	0.01	0.01	0.01	0.01	0.01	0

	<u>0.07</u> *												
<u>Day</u> <u>69</u> Day 69	<u>1.00</u> <u>0.18</u> <u>0.06</u> *	<u>1.00</u> <u>0.42</u> <u>0.32</u>	<u>1.00</u> <u>0.27</u> <u>0.17</u>	<u>1.00</u> <u>0.38</u> <u>0.24</u>	<u>1.00</u> <u>0.39</u> <u>0.30</u>	<u>1.00</u> <u>0.53</u> <u>0.56</u>	<u>1.00</u> <u>0.38</u> <u>0.29</u>	<u>1.00</u> <u>0.69</u> <u>0.84</u>	<u>1.00</u> <u>0.44</u> <u>0.39</u>	<u>1.00</u> <u>0.66</u> <u>0.80</u>	<u>1.00</u> <u>0.47</u> <u>0.45</u>	<u>1.00</u> <u>0.68</u> <u>0.79</u>	<u>1.000</u> <u>0.0</u> <u>3</u>
<u>Day</u> <u>108</u> Day 108	<u>1.00</u> <u>0.62</u> <u>0.57</u>	<u>1.00</u> <u>0.52</u> <u>0.52</u>	<u>1.00</u> <u>0.95</u> <u>0.88</u>	<u>1.00</u> <u>1.00</u> <u>1.00</u>	<u>1.00</u> <u>1.00</u> <u>1.00</u>	<u>1.00</u> <u>0.94</u> <u>0.83</u>	<u>1.00</u> <u>0.53</u> <u>0.53</u>	<u>1.00</u> <u>0.80</u> <u>0.78</u>	<u>1.00</u> <u>0.85</u> <u>0.71</u>	<u>1.00</u> <u>1.00</u> <u>0.88</u>	<u>1.00</u> <u>1.00</u> <u>0.87</u>	<u>1.00</u> <u>0.10</u> <u>0.17</u>	<u>1.000</u> <u>540.5</u> <u>2</u>
<u>Day</u> <u>174</u> Day 174	<u>1.00</u> <u>0.37</u> <u>0.32</u>	<u>1.00</u> <u>0.95</u> <u>0.95</u>	<u>1.00</u> <u>0.91</u> <u>0.96</u>	<u>1.00</u> <u>0.98</u> <u>0.99</u>	<u>1.00</u> <u>0.99</u> <u>1.00</u>	<u>1.00</u> <u>0.88</u> <u>0.99</u>	<u>1.00</u> <u>0.98</u> <u>0.97</u>	<u>1.00</u> <u>0.99</u> <u>0.99</u>	<u>1.00</u> <u>0.97</u> <u>0.95</u>	<u>1.00</u> <u>1.00</u> <u>1.00</u>	<u>1.00</u> <u>1.00</u> <u>1.00</u>	<u>1.00</u> <u>0.96</u> <u>0.99</u>	<u>1.000</u> <u>250.1</u> <u>6</u>

<u>125</u> <u>RPM</u> <u>125</u> <u>RPM</u>	<u>Ca(OH)₂</u> <u>Ca(OH)₂</u> <u>Ca(OH)₂</u>						<u>Mg(OH)₂</u> <u>Mg(OH)₂</u> <u>Mg(OH)₂</u>						<u>NaO</u> <u>HNa</u> <u>OH</u>
<u>ΔTA</u> <u>ΔTA</u>	<u>1004</u> <u>00</u>	<u>2002</u> <u>00</u>	<u>3003</u> <u>00</u>	<u>4004</u> <u>00</u>	<u>5005</u> <u>00</u>	<u>6006</u> <u>00</u>	<u>1004</u> <u>00</u>	<u>2002</u> <u>00</u>	<u>3003</u> <u>00</u>	<u>4004</u> <u>00</u>	<u>5005</u> <u>00</u>	<u>6006</u> <u>00</u>	<u>60060</u> <u>0</u>
<u>Day 1</u> <u>Day</u> <u>1</u>	<u><0.01</u> <u>0.01</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.01</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.00</u>	<u><0.010</u> <u>.00</u>
<u>Day 3</u> <u>Day</u> <u>3</u>	<u>1.00</u> <u>0.62</u> <u>0.68</u>	<u>1.00</u> <u>0.59</u> <u>0.60</u>	<u>1.00</u> <u>0.94</u> <u>0.97</u>	<u>1.00</u> <u>0.73</u> <u>0.76</u>	<u>1.00</u> <u>0.23</u> <u>0.21</u>	<u>1.00</u> <u>0.15</u> <u>0.17</u>	<u>1.00</u> <u>0.75</u> <u>0.79</u>	<u>1.00</u> <u>0.23</u> <u>0.29</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.00</u>	<u><0.01</u> <u>0.01</u>	<u><0.010</u> <u>.00</u>
<u>Day 6</u> <u>Day</u> <u>6</u>	<u>1.00</u> <u>0.99</u> <u>0.75</u>	<u>1.00</u> <u>0.98</u> <u>0.73</u>	<u>1.00</u> <u>0.99</u> <u>0.75</u>	<u>1.00</u> <u>0.64</u> <u>0.55</u>	<u>1.00</u> <u>0.87</u> <u>0.64</u>	<u>1.00</u> <u>1.00</u> <u>0.81</u>	<u>1.00</u> <u>0.99</u> <u>0.75</u>	<u>1.00</u> <u>1.00</u> <u>0.78</u>	<u>1.00</u> <u>0.98</u> <u>0.74</u>	<u>1.00</u> <u>0.91</u> <u>0.66</u>	<u>1.00</u> <u>0.99</u> <u>0.77</u>	<u>1.00</u> <u>0.93</u> <u>0.68</u>	<u>1.000</u> <u>750.6</u> <u>0</u>
<u>Day</u> <u>35</u> Day 35	<u>0.71</u> <u>0.36</u> <u>0.31</u>	<u>0.71</u> <u>0.41</u> <u>0.34</u>	<u>0.71</u> <u>0.53</u> <u>0.55</u>	<u>0.71</u> <u>0.56</u> <u>0.61</u>	<u>0.67</u> <u>0.29</u> <u>0.19</u>	<u>0.71</u> <u>0.40</u> <u>0.32</u>	<u>0.55</u> <u>0.24</u> <u>0.15</u>	<u>0.55</u> <u>0.45</u> <u>0.40</u>	<u>0.50</u> <u>0.37</u> <u>0.27</u>	<u>0.67</u> <u>0.32</u> <u>0.21</u>	<u>0.49</u> <u>0.30</u> <u>0.20</u>	<u>0.49</u> <u>0.28</u> <u>0.16</u>	<u>0.130</u> <u>160.0</u> <u>7*</u>

<u>Day</u> ~~69~~ Day 69	<u>1.00</u> <u>0.53</u> 0.55	<u>1.00</u> <u>0.52</u> 0.53	<u>1.00</u> <u>0.70</u> 0.80	<u>1.00</u> <u>0.63</u> 0.71	<u>1.00</u> <u>0.50</u> 0.51	<u>1.00</u> <u>0.42</u> 0.36	<u>1.00</u> <u>0.56</u> 0.60	<u>1.00</u> <u>0.64</u> 0.72	<u>1.00</u> <u>0.58</u> 0.63	<u>1.00</u> <u>0.58</u> 0.63	<u>1.00</u> <u>0.62</u> 0.67	<u>1.00</u> <u>0.47</u> 0.44	<u>1.000</u> <u>310.1</u> 9
<u>Day</u> ~~108~~ Day 108	<u>0.19</u> <u>0.04</u> 0.13	<u>0.16</u> <u>0.02</u> 0.06	<u>0.10</u> <u>0.01</u> 0.05	<u>0.01</u> <u>0.00</u> 0.04	<u>0.17</u> <u>0.03</u> 0.13	<u>0.24</u> <u>0.08</u> 0.26	<u>0.15</u> <u>0.02</u> 0.16	<u>0.02</u> <u>0.00</u> 0.16	<u><0.01</u> <u>0.01</u> 	<u>0.01</u> <u>0.00</u> 0.03	<u>0.31</u> <u>0.15</u> 0.34	<u>0.20</u> <u>0.05</u> 0.10	<u>0.880</u> <u>880.8</u> 7
<u>Day</u> ~~174~~ Day 174	<u>1.00</u> <u>0.19</u> 0.28	<u>1.00</u> <u>0.46</u> 0.47	<u>1.00</u> <u>0.97</u> 0.87	<u>1.00</u> <u>0.70</u> 0.64	<u>1.00</u> <u>0.74</u> 0.66	<u>0.33</u> <u>0.03</u> 0.12	<u>1.00</u> <u>0.33</u> 0.38	<u>1.00</u> <u>0.43</u> 0.45	<u>1.00</u> <u>0.49</u> 0.49	<u>1.00</u> <u>0.12</u> 0.22	<u>1.00</u> <u>0.62</u> 0.59	<u>1.00</u> <u>1.00</u> 0.93	<u>1.000</u> <u>230.3</u> 4

1045 Data ~~a~~Availability

The data will be made available in a public repository upon acceptance of the manuscript for publication.

Author ~~c~~Contributions

CAM designed the experiments with inputs from JH. CAM and MSM ran the experiments. CAM and PB performed the analyses. CAM wrote the first draft of the manuscript, and inputs from JH₁, ~~as well as further inputs from~~ MSM₂ and PB₃ led to the final version of the manuscript.

Competing ~~i~~Interests

JH ~~is~~ serves as a consulting consultant for the Planeteers GmbH. All authors declare that they have no competing interests.

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1065 References

- Bach, L. T., Tyka, M. D., Wang, B., and Fennel, K.: Lethal by design? Guiding environmental assessments of ocean alkalinity enhancement toward realistic contextualization of the alkalinity perturbation, CDRXIV, 10.70212/cdrxiv.2025457.v1, 2025.
- Bach, L. T.: The additionality problem of ocean alkalinity enhancement, Biogeosciences, 21, 261-277, 10.5194/bg-21-261-2024, 2024.
- Bach, L. T., Gill, S. J., Rickaby, R. E. M., Gore, S., and Renforth, P.: CO₂ removal with enhanced weathering and Ocean Alkalinity Enhancement: Potential risks and co-benefits for marine pelagic ecosystems, *Frontiers in Climate*, 1, 10.3389/fclim.2019.00007, 2019.
- Badocco, D., Pedrini, F., Pastore, A., di Marco, V., Marin, M. G., Bogialli, S., Roverso, M., and Pastore, P.: Use of a simple empirical model for the accurate conversion of the seawater pH value measured with NIST calibration into seawater pH scales, *Talanta*, 225, 122051, 10.1016/j.talanta.2020.122051, 2021.
- Baltruschat, S., Hartmann, J., Saitner, N., Moras, C. A., Lim, C., Bastianini, L., and Renforth, P.: Assessment of solid ikaite release into seawater – implications for ocean alkalinity enhancement, *Applied Geochemistry*, 203, 106781, 10.1016/j.apgeochem.2026.106781, 2026.
- Beck, R. and Andreassen, J.-P.: Spherulitic Growth of Calcium Carbonate, *Crystal Growth & Design*, 10, 2934-2947, 10.1021/cg901460g, 2010.
- Caserini, S., Pagano, D., Campo, F., Abbà, A., De Marco, S., Righi, D., Renforth, P., and Grosso, M.: Potential of maritime transport for ocean liming and atmospheric CO₂ removal, *Frontiers in Climate*, 3, 10.3389/fclim.2021.575900, 2021.
- Chave, K. E. and Suess, E.: Calcium carbonate saturation in seawater: Effects of dissolved organic matter, *Limnology and Oceanography*, 15, 633-637, 10.4319/lo.1970.15.4.0633, 1970.
- Cushman-Roisin, B. and Beckers, J.-M.: Chapter 20 - Oceanic General Circulation, in: *International Geophysics*, edited by: Cushman-Roisin, B., and Beckers, J.-M., Academic Press, 657-699, 10.1016/B978-0-12-088759-0.00020-1, 2011.
- Dale, A. W., Geilert, S., Diercks, I., Fuhr, M., Perner, M., Scholz, F., and Wallmann, K.: Seafloor alkalinity enhancement as a carbon dioxide removal strategy in the Baltic Sea, *Communications Earth & Environment*, 5, 452, 10.1038/s43247-024-01569-3, 2024.
- Dickson, A. G.: Standard potential of the reaction: $\text{AgCl(s)} + 12\text{H}_2\text{(g)} = \text{Ag(s)} + \text{HCl(aq)}$, and the standard acidity constant of the ion HSO_4^- in synthetic sea water from 273.15 to 318.15 K, *The Journal of Chemical Thermodynamics*, 22, 113-127, 10.1016/0021-9614(90)90074-Z, 1990.
- Dickson, A. G.: Standards for ocean measurements, *Oceanography*, 23, 34-47, 10.5670/oceanog.2010.22, 2010.
- Dickson, A. G., Sabine, C. L., and Christian, J. R.: Guide to best practices for ocean CO₂ measurements, PICES Special Publication 3; IOCCP Report 8, Sidney, British Columbia, North Pacific Marine Science Organization, 191 pp., 10.25607/OBP-1342, 2007.
- Drever, J. I.: The Geochemistry of Natural Waters, 3rd, Surface and Groundwater Environments, Prentice Hall, 1997.
- Eisaman, M. D., Geilert, S., Renforth, P., Bastianini, L., Campbell, J., Dale, A. W., Foteinis, S., Grasse, P., Hawrot, O., Löscher, C. R., Rau, G. H., and Rønning, J.: Assessing the technical aspects of ocean-alkalinity-enhancement approaches, *Guide to Best Practices in Ocean Alkalinity Enhancement Research*, 2-oae2023, 3, 10.5194/sp-2-oae2023-3-2023, 2023.
- England, P. I. and Bach, L. T.: Influence of Wave Action on Applications of Olivine-Based Ocean Alkalinity Enhancement on Sandy Beaches, Geophysical Research Letters, 52, e2025GL114922, 10.1029/2025GL114922, 2025.
- European Parliament and Council of the European Union: Directive 2000/60/EC of the European Parliament and of the Council of 23 October 2000 establishing a framework for Community action in the field of water policy, Official Journal of the European Communities OJ L 327, 1–73, 22/12/2000 (consolidated version: 10/05/2026), 2000.
- Faucher, G., Haunost, M., Paul, A. J., Tietz, A. U. C., and Riebesell, U.: Growth response of *Emiliania huxleyi* to ocean alkalinity enhancement, *Biogeosciences*, 22, 405-415, 10.5194/bg-22-405-2025, 2025.
- Ferderer, A., Chase, Z., Kennedy, F., Schulz, K. G., and Bach, L. T.: Assessing the influence of ocean alkalinity enhancement on a coastal phytoplankton community, *Biogeosciences*, 19, 5375-5399, 10.5194/bg-19-5375-2022, 2022.
- Ferderer, A., Schulz, K. G., Riebesell, U., Baker, K. G., Chase, Z., and Bach, L. T.: Investigating the effect of silicate- and calcium-based ocean alkalinity enhancement on diatom silicification, *Biogeosciences*, 21, 2777-2794, 10.5194/bg-21-2777-2024, 2024.

- 115 Ferderer, A., Schulz, K. G., Willis, A., Baker, K. G., Chase, Z., and Bach, L. T.: Carbonate chemistry fitness landscapes inform diatom resilience to future perturbations, *Science Advances*, 11, eadu8024, [doi:10.1126/sciadv.adu8024](https://doi.org/10.1126/sciadv.adu8024), 2025.
- Flipkens, G., Fuhr, M., Fiers, G., Meysman, F. J. R., Town, R. M., and Blust, R.: Enhanced olivine dissolution in seawater through continuous grain collisions, *Geochimica et Cosmochimica Acta*, 359, 84-99, [10.1016/j.gca.2023.09.002](https://doi.org/10.1016/j.gca.2023.09.002), 2023.
- Foteinis, S., Andresen, J., Campo, F., Caserini, S., and Renforth, P.: Life cycle assessment of ocean liming for carbon dioxide removal from the atmosphere, *Journal of Cleaner Production*, 370, 133309, [10.1016/j.jclepro.2022.133309](https://doi.org/10.1016/j.jclepro.2022.133309), 2022.
- 120 Frew, N. M., Glover, D. M., Bock, E. J., and McCue, S. J.: A new approach to estimation of global air-sea gas transfer velocity fields using dual-frequency altimeter backscatter, *Journal of Geophysical Research: Oceans*, 112, 10.1029/2006JC003819, 2007.
- [Fuhr, M., Geilert, S., Schmidt, M., Liebetrau, V., Vogt, C., Ledwig, B., and Wallmann, K.: Kinetics of Olivine Weathering in Seawater: An Experimental Study, *Frontiers in Climate*, 4, 10.3389/fclim.2022.831587, 2022.](https://doi.org/10.3389/fclim.2022.831587)
- 125 [Fuhr, M., Wallmann, K., Dale, A. W., Diercks, I., Kalapurakkal, H. T., Schmidt, M., Sommer, S., Böhnke, S., Perner, M., and Geilert, S.: Disentangling artificial and natural benthic weathering in organic rich Baltic Sea sediments, *Frontiers in Climate*, 5, 2023, 10.3389/fclim.2023.1245580, 2023.](https://doi.org/10.3389/fclim.2023.1245580)
- [Fuhr, M., Wallmann, K., Dale, A. W., Kalapurakkal, H. T., Schmidt, M., Sommer, S., Deusner, C., Spiegel, T., Kowalski, J., and Geilert, S.: Alkaline mineral addition to anoxic to hypoxic Baltic Sea sediments as a potentially efficient CO₂-removal technique, *Frontiers in Climate*, 6, 2024, 10.3389/fclim.2024.1338556, 2024.](https://doi.org/10.3389/fclim.2024.1338556)
- 130 Fuhr, M., Dale, A. W., Wallmann, K., Bährle, R., Kalapurakkal, H. T., Sommer, S., Spiegel, T., Dobashi, R., Buchholz, B., Schmidt, M., Perner, M., and Geilert, S.: Calcite is an efficient and low-cost material to enhance benthic weathering in shelf sediments of the Baltic Sea, *Communications Earth & Environment*, 6, 106, [10.1038/s43247-025-02079-6](https://doi.org/10.1038/s43247-025-02079-6), 2025.
- [Fuss, S., Lunstrum, A., Prütz, R., and Gruner, F.: Chapter 10: Costs and Potentials, in: *The State of Carbon Dioxide Removal 3rd Edition 2026* \(eds. Edwards, M. R. et al.\), 10.17605/OSF.IO/FBZ4X, 2026.](https://doi.org/10.17605/OSF.IO/FBZ4X)
- 135 Goldenberg, S. U., Riebesell, U., Brüggemann, D., Börner, G., Sswat, M., Folkvord, A., Couret, M., Spjelkavik, S., Sánchez, N., Jaspers, C., and Moyano, M.: Early life stages of fish under ocean alkalinity enhancement in coastal plankton communities, *Biogeosciences*, 21, 4521-4532, [10.5194/bg-21-4521-2024](https://doi.org/10.5194/bg-21-4521-2024), 2024.
- Gránásy, L., Pusztai, T., Tegze, G., Warren, J. A., and Douglas, J. F.: Growth and form of spherulites, *Physical Review E*, 72, 011605, [10.1103/PhysRevE.72.011605](https://doi.org/10.1103/PhysRevE.72.011605), 2005.
- 140 [Griskey, R. G.: Fluid Flow Basic Equations, in: *Transport Phenomena and Unit Operations: A Combined Approach*, John Wiley & Sons, 23-54, 10.1002/0471722057, 2002.](https://doi.org/10.1002/0471722057)
- [Haas, A. R.: The effect of the addition of alkali to sea water upon the hydrogen ion concentration, *Journal of Biological Chemistry*, 26, 515-517, 10.1016/S0021-9258\(18\)87433-6, 1916.](https://doi.org/10.1016/S0021-9258(18)87433-6)
- 145 Hartmann, J., Suitner, N., Lim, C., Schneider, J., Marín-Samper, L., Aristegui, J., Renforth, P., Taucher, J., and Riebesell, U.: Stability of alkalinity in ocean alkalinity enhancement (OAE) approaches – consequences for durability of CO₂ storage, *Biogeosciences*, 20, 781-802, [10.5194/bg-20-781-2023](https://doi.org/10.5194/bg-20-781-2023), 2023.
- Hashim, M. S., Marx, L., Klein, F., Dean, C. L., Burdige, E., Hayden, M., McCorkle, D. C., and Subhas, A. V.: Mineral formation during shipboard ocean alkalinity enhancement experiments in the North Atlantic, *Biogeosciences*, 22, 7149-7165, [10.5194/bg-22-7149-2025](https://doi.org/10.5194/bg-22-7149-2025), 2025.
- 150 Hausfather, Z.: An assessment of current policy scenarios over the 21st century and the reduced plausibility of high-emissions pathways, *Dialogues on Climate Change*, 2, 26-32, [10.1177/29768659241304854](https://doi.org/10.1177/29768659241304854), 2025.
- Ho, D. T., Law, C. S., Smith, M. J., Schlosser, P., Harvey, M., and Hill, P.: Measurements of air-sea gas exchange at high wind speeds in the Southern Ocean: Implications for global parameterizations, *Geophysical Research Letters*, 33, [10.1029/2006GL026817](https://doi.org/10.1029/2006GL026817), 2006.
- 155 [IPCC: Summary for Policymakers. In: *Climate Change 2021: The Physical Science Basis. Contribution of Working Group I to the Sixth Assessment Report of the Intergovernmental Panel on Climate Change* \[Masson-Delmotte, V., P. Zhai, A. Pirani, S.L. Connors, C. Péan, S. Berger, N. Caud, Y. Chen, L. Goldfarb, M.I. Gomis, M. Huang, K. Leitzell, E. Lonnoy, J.B.R. Matthews, T.K. Maycock, T. Waterfield, O. Yelekçi, R. Yu, and B. Zhou \(eds.\)\], Cambridge University Press, Cambridge, United Kingdom and New York, NY, USA, pp. 3-32, 10.1017/9781009157896.001, 2021.](https://doi.org/10.1017/9781009157896.001)
- 160 [IPCC: Summary for Policymakers. In: *Climate Change 2022: Impacts, Adaptation and Vulnerability. Contribution of Working Group II to the Sixth Assessment Report of the Intergovernmental Panel on Climate Change* \[H.-O. Pörtner, D.C. Roberts, M. Tignor, E.S. Poloczanska, K. Mintenbeck, A. Alegría, M. Craig, S. Langsdorf, S. Löschke, V. Möller, A. Okem, B. Rama](https://doi.org/10.1017/9781009157896.001)

- (eds.)), Cambridge University Press, Cambridge, United Kingdom and New York, NY, USA, pp. 3–33, 10.1017/9781009325844.001, 2022. IPCC: Summary for Policymakers. In: Climate Change 2022: Impacts, Adaptation and Vulnerability. Contribution of Working Group II to the Sixth Assessment Report of the Intergovernmental Panel on Climate Change [H. O. Pörtner, D.C. Roberts, M. Tignor, E.S. Poloczanska, K. Mintenbeck, A. Alegria, M. Craig, S. Langsdorf, S. Löschke, V. Möller, A. Okem, B. Rama (eds.)], Cambridge University Press, Cambridge, United Kingdom and New York, NY, USA, 2022.
- Isometric: Ocean Alkalinity Enhancement from Coastal Outfalls, Isometric, Version 1.1, 2026.
- Jones, D. C., Ito, T., Takano, Y., and Hsu, W.-C.: Spatial and seasonal variability of the air-sea equilibration timescale of carbon dioxide, *Global Biogeochemical Cycles*, 28, 1163–1178, 10.1002/2014GB004813, 2014.
- Kapp, E. M.: The Precipitation of Calcium and Magnesium from Sea Water by Sodium Hydroxide, *Biological Bulletin*, 55, 453–458, 10.2307/1536800, 1928.
- Kauk, S.: World's first verified Ocean Alkalinity Enhancement credits: <https://isometric.com/writing-articles/worlds-first-verified-ocean-alkalinity-enhancement-credits> (last access: 14 July 2026), 2025.
- Kheshgi, H. S.: Sequestering atmospheric carbon dioxide by increasing ocean alkalinity, *Energy*, 20, 915–922, 10.1016/0360-5442(95)00035-F, 1995.
- Kousoulas, K., Ferderer, A., Eriksen, R., and Bach, L. T.: Winners and losers under hydroxide-based ocean alkalinity enhancement in a Tasmanian plankton community, *Journal of Phycology*, 61, 989–1006, 10.1111/jpy.70052, 2025.
- Lasaga, A. C.: *Kinetic Theory in the Earth Sciences*, Princeton University Press, Princeton, 10.1515/9781400864874, 2014.
- Lee, K., Kim, T.-W., Byrne, R. H., Millero, F. J., Feely, R. A., and Liu, Y.-M.: The universal ratio of boron to chlorinity for the North Pacific and North Atlantic oceans, *Geochimica et Cosmochimica Acta*, 74, 1801–1811, 10.1016/j.gca.2009.12.027, 2010.
- Lee, C.-H., Subhas, A. V., Kim, J.-H., and Lee, K.: Ocean Carbon Dioxide Removal and Storage, *Chemical Reviews*, 126, 1110–1144, 10.1021/acs.chemrev.5c00433, 2026.
- Levich, V. G.: *Physicochemical hydrodynamics*, Prentice-Hall, 1962.
- Lewis, E. L. and Perkin, R. G.: The practical salinity scale 1978: conversion of existing data, *Deep Sea Research Part A. Oceanographic Research Papers*, 28, 307–328, 10.1016/0198-0149(81)90002-9, 1981.
- Lunstrum, A., Keller, D., Hartmann, J., and Moras, C. A.: A framework for understanding efficiency losses of Ocean Alkalinity Enhancement, *CDRXIV [preprint]*, 10.70212/cdrxiv.2026517.v1, 2026.
- Marion, G. M., Millero, F. J., and Feistel, R.: Precipitation of solid phase calcium carbonates and their effect on application of seawater $S_A - T - P$ models, *Ocean Science*, 5, 285–291, 10.5194/os-5-285-2009, 2009.
- Meysman, F. J. R. and Montserrat, F.: Negative CO₂ emissions via enhanced silicate weathering in coastal environments, *Biology Letters*, 13, 20160905, 10.1098/rsbl.2016.0905, 2017.
- Miller, S. D., Marandino, C., and Saltzman, E. S.: Ship-based measurement of air-sea CO₂ exchange by eddy covariance, *Journal of Geophysical Research: Atmospheres*, 115, 10.1029/2009JD012193, 2010.
- Millero, F. J. and Poisson, A.: International one-atmosphere equation of state of seawater, *Deep Sea Research Part A. Oceanographic Research Papers*, 28, 625–629, 10.1016/0198-0149(81)90122-9, 1981.
- Montserrat, F., Renforth, P., Hartmann, J., Leermakers, M., Knops, P., and Meysman, F. J. R.: Olivine dissolution in seawater: Implications for CO₂ sequestration through enhanced weathering in coastal environments, *Environmental Science & Technology*, 51, 3960–3972, 10.1021/acs.est.6b05942, 2017.
- Moras, C. A., Bach, L. T., Cyronak, T., Joannes-Boyau, R., and Schulz, K. G.: Ocean alkalinity enhancement – avoiding runaway CaCO₃ precipitation during quick and hydrated lime dissolution, *Biogeosciences*, 19, 3537–3557, 10.5194/bg-19-3537-2022, 2022.
- Moras, C. A., Bach, L. T., Cyronak, T., Joannes-Boyau, R., and Schulz, K. G.: Preparation and quality control of in-house reference materials for marine dissolved inorganic carbon and total alkalinity measurements, *Limnology and Oceanography: Methods*, 10.1002/lom3.10570, 2023.
- Moras, C. A., Cyronak, T., Bach, L. T., Joannes-Boyau, R., and Schulz, K. G.: Effects of grain size and seawater salinity on magnesium hydroxide dissolution and secondary calcium carbonate precipitation kinetics: implications for ocean alkalinity enhancement, *Biogeosciences*, 21, 3463–3475, 10.5194/bg-21-3463-2024, 2024.

- Morse, J. W., Arvidson, R. S., and Lüttge, A.: Calcium carbonate formation and dissolution, *Chemical Reviews*, 107, 342-381, 10.1021/cr050358j, 2007.
- 215 Morse, J. W., Wang, Q., and Tsio, M. Y.: Influences of temperature and Mg:Ca ratio on CaCO₃ precipitates from seawater, *Geology*, 25, 85-87, 10.1130/0091-7613(1997)025<0085:lotamc>2.3.Co;2, 1997.
- Pan, Y., Li, Y., Ma, Q., He, H., Wang, S., Sun, Z., Cai, W.-J., Dong, B., Di, Y., Fu, W., and Chen, C.-T. A.: The role of Mg²⁺ in inhibiting CaCO₃ precipitation from seawater, *Marine Chemistry*, 237, 104036, 10.1016/j.marchem.2021.104036, 2021.
- 220 Oschlies, A., Bach, L. T., Rickaby, R. E. M., Satterfield, T., Webb, R., and Gattuso, J. P.: Climate targets, carbon dioxide removal, and the potential role of ocean alkalinity enhancement, *Guide to Best Practices in Ocean Alkalinity Enhancement Research*, 2-oae2023, 1, 10.5194/sp-2-oae2023-1-2023, 2023.
- Paul, A. J., Haunost, M., Goldenberg, S. U., Hartmann, J., Sánchez, N., Schneider, J., Suitner, N., and Riebesell, U.: Ocean alkalinity enhancement in an open-ocean ecosystem: biogeochemical responses and carbon storage durability, *Biogeosciences*, 22, 2749-2766, 10.5194/bg-22-2749-2025, 2025.
- 225 Perez, F. F. and Fraga, F.: Association constant of fluoride and hydrogen ions in seawater, *Marine Chemistry*, 21, 161-168, 10.1016/0304-4203(87)90036-3, 1987.
- Pytkowicz, R. M.: Rates of inorganic calcium carbonate nucleation, *The Journal of Geology*, 73, 196-199, 10.1086/627056, 1965.
- Ramírez, L., Pozzo-Pirotta, L. J., Trebec, A., Manzanares-Vázquez, V., Díez, J. L., Arístegui, J., Riebesell, U., Archer, S. D., and Segovia, M.: Ocean alkalinity enhancement (OAE) does not cause cellular stress in a phytoplankton community of the subtropical Atlantic Ocean, *Biogeosciences*, 22, 1865-1886, 10.5194/bg-22-1865-2025, 2025.
- 230 Reynolds, O.: XXIX. An experimental investigation of the circumstances which determine whether the motion of water shall be direct or sinuous, and of the law of resistance in parallel channels, *Philosophical Transactions of the Royal Society of London*, 935-982, 10.1098/rstl.1883.0029, 1883.
- 235 Rigopoulos, I., Harrison, A. L., Delimitis, A., Ioannou, I., Efstathiou, A. M., Kyratsi, T., and Oelkers, E. H.: Carbon sequestration via enhanced weathering of peridotites and basalts in seawater, *Applied Geochemistry*, 91, 197-207, 10.1016/j.apgeochem.2017.11.001, 2018.
- Risien, C. M. and Chelton, D. B.: A Global Climatology of Surface Wind and Wind Stress Fields from Eight Years of QuikSCAT Scatterometer Data, *Journal of Physical Oceanography*, 38, 2379-2413, 10.1175/2008JPO3881.1, 2008.
- 240 Rushdi, A. I., Pytkowicz, R. M., Suess, E., and Chen, C. T.: The effects of magnesium-to-calcium ratios in artificial seawater, at different ionic products, upon the induction time, and the mineralogy of calcium carbonate: A laboratory study, *Geologische Rundschau*, 81, 571-578, 10.1007/BF01828616, 1992.
- Saez Moreno, M., Hartmann, J., Börker, J., Bartsch, P., and Moras, C. A.: The role of temperature and salinity on ocean alkalinity enhancement performance, *EGUsphere*, 2026, 1-30, 10.5194/egusphere-2026-3277, 2026.
- 245 Schneider, J., Riebesell, U., Moras, C. A., Marín-Samper, L., Kittu, L. R., Ortíz-Cortes, J., and Schulz, K. G.: Carbon fixation of a temperate plankton community in response to calcium- and silicate-based Ocean Alkalinity Enhancement using air-sea gas exchange measurements, *Biogeosciences*, 23, 137-153, 10.5194/bg-23-137-2026, 2026.
- Schockman, K. M. and Byrne, R. H.: Spectrophotometric determination of the bicarbonate dissociation constant in seawater, *Geochimica et Cosmochimica Acta*, 300, 231-245, 10.1016/j.gca.2021.02.008, 2021.
- 250 Schulz, K. G., Bach, L. T., and Dickson, A. G.: Seawater carbonate system considerations for ocean alkalinity enhancement research, *State Planet Discuss.*, 2023, 1-24, 10.5194/sp-2023-12, 2023.
- Sharqawy, M. H., Lienhard, J. H., and Zubair, S. M.: Thermophysical properties of seawater: a review of existing correlations and data, *Desalination and Water Treatment*, 16, 354-380, 10.5004/dwt.2010.1079, 2010.
- 255 Shaw, C., Ringham, M. C., Carter, B. R., Tyka, M. D., and Eisaman, M. D.: Using magnesium hydroxide for ocean alkalinity enhancement: elucidating the role of formation conditions on material properties and dissolution kinetics, *Frontiers in Climate*, ~~Volume 7~~, – 2025, 10.3389/fclim.2025.1616362, 2025.
- Suitner, N., Hartmann, J., Varliero, S., Faucher, G., Suessle, P., and Moras, C. A.: Surface area and Ω -aragonite oversaturation as controls of the runaway precipitation process in ocean alkalinity enhancement, *Biogeosciences*, 23, 3965-3980, 10.5194/bg-23-3965-2026, 2026.
- 260 ~~Suitner, N., Hartmann, J., Varliero, S., Faucher, G., Suessle, P., and Moras, C. A.: Surface area and Ω -aragonite oversaturation as controls of the runaway precipitation process in ocean alkalinity enhancement, *EGUsphere* [preprint], 2025, 1-26, 10.5194/egusphere-2025-381, 2025.~~

Suitner, N., Faucher, G., Lim, C., Schneider, J., Moras, C. A., Riebesell, U., and Hartmann, J.: Ocean alkalinity enhancement approaches and the predictability of runaway precipitation processes: results of an experimental study to determine critical alkalinity ranges for safe and sustainable application scenarios, *Biogeosciences*, 21, 4587-4604, 10.5194/bg-21-4587-2024, 2024.

Sulpis, O., Mucci, A., Boudreau, B. P., Barry, M. A., and Johnson, B. D.: Controlling the diffusive boundary layer thickness above the sediment–water interface in a thermostated rotating-disk reactor, *Limnology and Oceanography: Methods*, 17, 241-253, 10.1002/lom3.10309, 2019.

Tang, H., Wu, X., Xian, H., Zhu, J., Wei, J., Liu, H., and He, H.: Heterogeneous nucleation and growth of CaCO₃ on calcite (104) and aragonite (110) surfaces: Implications for the formation of abiogenic carbonate cements in the ocean, *Minerals*, 10, 294, 10.3390/min10040294, 2020.

Thorpe, S. A.: *An Introduction to Ocean Turbulence*, Cambridge University Press, Cambridge, 10.1017/CBO9780511801198, 2007.

Tokinaga, H. and Xie, S.-P.: Wave- and Anemometer-Based Sea Surface Wind (WASWind) for Climate Change Analysis, *Journal of Climate*, 24, 267-285, 10.1175/2010JCLI3789.1, 2011.

Tracy, S. L., Williams, D. A., and Jennings, H. M.: The growth of calcite spherulites from solution: II. Kinetics of formation, *Journal of Crystal Growth*, 193, 382-388, 10.1016/S0022-0248(98)00521-1, 1998.

U.S. Environmental Protection Agency: National Recommended Water Quality Criteria - Aquatic Life Criteria Table: <https://www.epa.gov/wqc/national-recommended-water-quality-criteria-aquatic-life-criteria-table> (last access: 14 July 2026), 2026.

U.S. Environmental Protection Agency: Quality Criteria for Water 1986, U.S. Environmental Protection Agency, Washington, DC, EPA 440/5-86-001, 1986.

Varliero, S., Buono, A., Caserini, S., Raos, G., and Macchi, P.: Chemical Aspect of Ocean Liming for CO₂ Removal: Dissolution Kinetics of Calcium Hydroxide in Seawater, *ACS Engineering Au*, 4, 422-431, 10.1021/acsengineeringau.4c00008, 2024.

Wanninkhof, R.: Relationship between wind speed and gas exchange over the ocean revisited, *Limnology and Oceanography: Methods*, 12, 351-362, 10.4319/lom.2014.12.351, 2014.

Weiss, R. F.: Carbon dioxide in water and seawater: the solubility of a non-ideal gas, *Marine Chemistry*, 2, 203-215, 10.1016/0304-4203(74)90015-2, 1974.

Yang, B., Leonard, J., and Langdon, C.: Seawater alkalinity enhancement with magnesium hydroxide and its implication for carbon dioxide removal, *Marine Chemistry*, 253, 104251, 10.1016/j.marchem.2023.104251, 2023.

Yu, T., He, Y., and Yan, X.: Global air-sea surface carbon dioxide transfer velocity and flux estimated using 17 a altimeter data and a new algorithm, *Acta Oceanologica Sinica*, 32, 24-33, 10.1007/s13131-013-0362-9, 2013.

Zeebe, R. E. and Wolf-Gladrow, D.: *CO₂ in seawater: equilibrium, kinetics, isotopes*, -65, Gulf Professional Publishing, 360 pp., 2001.

Zhou, X., Reichl, B. G., Romero, L., and Deike, L.: A Sea State Dependent Gas Transfer Velocity for CO₂ Unifying Theory, Model, and Field Data, *Earth and Space Science*, 10, e2023EA003237, 10.1029/2023EA003237, 2023.

Zhou, M., Tyka, M. D., Ho, D. T., Yankovsky, E., Bachman, S., Nicholas, T., Karspeck, A. R., and Long, M. C.: Mapping the global variation in the efficiency of ocean alkalinity enhancement for carbon dioxide removal, *Nature Climate Change*, 15, 59-65, 10.1038/s41558-024-02179-9, 2025.

Zubovic, K. P., Horvath, A., Brien, D. M., Rateau, R., Terribili, L., Winters, S., Docaigne, E., Guyett, P. C., and Rodriguez-Blanco, J. D.: Crystallisation of CaCO₃ polymorphs induced by layered PET-based microplastic particles, *Environmental Sciences Europe*, 37, 52, 10.1186/s12302-025-01090-0, 2025.