A DFT-based comparative equilibrium study of thermal dehydration and hydrolysis of CaCl$_2$ hydrates and MgCl$_2$ hydrates for seasonal heat storage

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Salt hydrates store solar energy in chemical form via a reversible dehydration–hydration reaction. However, as a side reaction to dehydration, hydrolysis (HCl formation) may occur in chloride based salt hydrates (specially in MgCl$_2$ hydrates), affecting the durability of the storage system. The mixture of CaCl$_2$ and MgCl$_2$ hydrates has been shown experimentally to have exceptional cycle stability and improved kinetics. However, the optimal operating conditions for the mixture are unknown. To understand the appropriate balance between dehydration and hydrolysis kinetics in the mixtures, it is essential to gain in-depth insight into the mixture components. We present a GGA-DFT level study to investigate the various gaseous structures of CaCl$_2$ hydrates and to understand the relative stability of their conformers. The hydration strength and relative stability of conformers are dominated by electrostatic interactions. A wide network of intramolecular homonuclear and heteronuclear hydrogen bonds is observed in CaCl$_2$ hydrates. Equilibrium product concentrations are obtained during dehydration and hydrolysis reactions under various temperature and pressure conditions. The trend of the dehydration curve with temperature in CaCl$_2$ hydrates is similar to the experiments. Comparing these results to those of MgCl$_2$ hydrates, we find that CaCl$_2$ hydrates are more resistant towards hydrolysis in the temperature range of 273–800 K. Specifically, the present study reveals that the onset temperatures of HCl formation, a crucial design parameter for MgCl$_2$ hydrates, are lower than for CaCl$_2$ hydrates except for the mono-hydrate.

1 Introduction

Solar thermal energy can be stored in three forms: sensible heat, latent heat, and in thermochemical form. In a thermochemical storage system, energy is stored by bond breaking and reforma-
tion in a reversible reaction.$^1$ Compared to sensible or latent heat, the energy storage density is higher in thermochemical form.$^2$ This explains the enormous interest in thermochemical materials (TCMs) for seasonal heat storage applications.$^3$ TCMs offer compact solar energy storage suitable for household applications. TCMs are widely available and inexpensive, and therefore appear to be an attractive material for seasonal heat storage. The cost of TCM is 30% of the total investment in the thermochemical heat storage system.$^3$ Thus, the selection of TCMs is a key aspect in designing the heat storage system.

Salt hydrates are one class of TCMs. Upon charging, these materials absorb solar energy and disintegrate into a lower hydrated or anhydrous form. Upon discharging, these dried salt hydrates recombine with H$_2$O, forming hydrates while dissipating energy.

\[ \text{Salt} \cdot n\text{H}_2\text{O} \rightleftharpoons \text{Salt} + n\text{H}_2\text{O} \]  

The forward reaction is endothermic while the backward reaction is exothermic. The major challenge in selection of these salt hydrates is to find materials with high storage capacity, cycle stability and fast kinetics. MgSO$_4$ hydrates have a high storage capacity (2–3 GJ m$^{-3}$). However, MgSO$_4$ hydrates have sluggish hydration kinetics due to the presence of strong hydrogen bonds (H-bonds).$^4$ Chloride based salt hydrates (MgCl$_2$,$^n$H$_2$O, CaCl$_2$,$^n$H$_2$O, $n = 0, 1, 2, 4, and 6) are also promising TCMs for seasonal heat storage. They have high storage capacity (2–3 GJ m$^{-3}$). However, hydrolysis is reported for chloride based salt hydrates. This is an irreversible endothermic reaction which competes with dehydration reaction, according to

\[ \text{MgCl}_2\cdot n\text{H}_2\text{O} \rightarrow \text{MOHCl} + (n - 1)\text{H}_2\text{O} + \text{HCl (M = Mg/Ca)} \]  

MgCl$_2$ hydrates have faster kinetics but also have the problem of hydrolysis as a side reaction.$^5$ Hydrolysis affects the cycle
stability of the storage system and generates HCl gas, which is harmful and corrosive. This poses a challenge for the application of MgCl$_2$ hydrate as a TCM.

Hydrolysis can be avoided in TCMs by mixing with other halides. Rammelberg et al. examined the performance of various salt mixtures. They observed that the mixture of CaCl$_2$ hydrates and MgCl$_2$ hydrates shows good cycle stability and improved kinetics compared to MgCl$_2$ and CaCl$_2$ hydrates alone. They have not described the chemical explanation for this extra stability. To assess the performance of the mixture of CaCl$_2$ hydrates and MgCl$_2$ hydrates, it is essential to investigate the dehydration and hydrolysis behavior of CaCl$_2$ hydrates and compare these to MgCl$_2$ hydrates.

CaCl$_2$ hydrate has been used as a TCM. It has a very high crystalline energy storage density (2.82 GJ m$^{-3}$), is readily available, inflammable and has high temperature lift during hydration. CaCl$_2$ hydrates are also used in mobile storage of industrial waste. In domestic heat storage systems, TCMs are used in the temperature range of 300–500 K while in mobile storage of industrial heat, a much higher temperature range is used. Thus, it is important to investigate the equilibrium products of dehydration and hydrolysis reactions of CaCl$_2$ hydrates over a wide temperature range. Next to the kinetics of dehydration/hydrolysis, the onset temperature of HCl formation in chloride-based salt hydrates is a crucial parameter in designing heat storage systems.

Rammelberg et al. and Zondag observed the kinetics of dehydration of CaCl$_2$ hydrates using Thermogravimetric analysis (TGA) and did not observe hydrolysis. Fraissler et al. observed that the formation of HCl gas from the hydrolysis of CaCl$_2$ occurs in a certain temperature range (683–1013 K). In these temperature regimes, the preference of hydrolysis over dehydration reaction is ambiguous from experiments.

Computational models can be successfully used to investigate the in-depth behavior of salt hydrates in different temperature regimes. The hydration of Ca$^{2+}$ ions has been successfully investigated using Density Functional Theory (DFT), molecular dynamics simulations and validated by experiments. Iype et al. characterized the H-bonds present in MgSO$_4$ hydrates using DFT calculations. They observed that strong H-bonds affect the hydration kinetics of MgSO$_4$ hydrates. Smeets et al. computationally investigated the dehydration and hydrolysis reactions for MgCl$_2$ hydrates. They have reported the equilibrium curves for dehydration/hydrolysis of MgCl$_2$ hydrates. We have explored the relative stability of various conformers of CaCl$_2$ hydrates, ideal operating conditions of CaCl$_2$ hydrates and compared them with MgCl$_2$ hydrates. Nevertheless, to the best of our knowledge, currently no computational studies exist on the hydrolysis reaction of CaCl$_2$ hydrates, the onset temperature of HCl formation and their comparison with MgCl$_2$ hydrates. The system level complexity of salt hydrate mixtures like grain boundaries, grain sizes and their effect on the reaction kinetics could be addressed by multiscale integration of DFT results to the system level.

In the present study, we would like to understand by means of DFT calculations, the temperature and pressure conditions under which hydrolysis precedes over dehydration. First we examine the molecular structure and vibrational frequencies of CaCl$_2$ hydrates and their conformers. We investigate the strength of H-bonds present in CaCl$_2$ hydrates, since strong H-bonds can result in sluggish hydration kinetics. We characterize the H-bonds using the Bond Valence (BV) sum rule. The BV sum rule is frequently used to correlate the structural properties of H-bonds and crystal structure determination.

The enthalpy of the dehydration and hydrolysis reactions of CaCl$_2$ hydrates is obtained from DFT calculations and compared to MgCl$_2$ hydrates. The change in Gibbs free energy for dehydration and hydrolysis of all hydrates is obtained, followed by the equilibrium product concentration. The equilibrium properties obtained from the present DFT calculations can be used to calculate safety limits for thermal degradation of CaCl$_2$ hydrates. These safety limits are compared with MgCl$_2$ hydrates.

## 2 Computational model and validation

### 2.1 Density functional theory calculations

The molecular structures of various CaCl$_2$ hydrates, their conformers, HCl, H$_2$O, and CaOHCl are fully optimized in DFT using the Perdew–Wang exchange and correlation functional (PW91) under generalized gradient approximation (GGA) implemented in the Amsterdam Density Functional (ADF) program with the double-polarized triple-$\zeta$ basis set. A spin restricted Kohn–Sham method is used thus keeping the maximum integration accuracy. DFT-GGA has been used to study the salt hydrates like CaCl$_2$, MgCl$_2$, MgSO$_4$, Iype et al. showed the applicability of GGA-DFT to study the H-bonds present in hydrates of MgSO$_4$.

### 2.2 Model validation

To validate the applicability of GGA-DFT using the PW91 functional, we have examined the binding enthalpy, Ca–O coordination length and vibrational frequency of [Ca(H$_2$O)$_6$]$^{2+}$ ions. The successive binding enthalpy of hexahydrated Ca$^{2+}$ ions (\(\Delta H_{6.5}^{\text{f}}\)) is defined as

\[
\Delta H_{6.5}^{\text{f}} = H_{\text{Ca(H}_2\text{O)}_6}^{\text{f}} - H_{\text{Ca(H}_2\text{O)}_5}^{\text{f}} - H_{\text{H}_2\text{O}}^{\text{f}}
\]

The optimized structure of [Ca(H$_2$O)$_6$]$^{2+}$ is shown in Fig. 1. We estimated the successive binding enthalpy of [Ca(H$_2$O)$_6$]$^{2+}$ ions from GGA-DFT to be 26.26 kcal mol$^{-1}$, which is in close agreement with previous DFT calculations (26.37 kcal mol$^{-1}$) and experimental results (25.3 kcal mol$^{-1}$).

In the present DFT study, the average coordination length of Ca–O in the first hydration shell of Ca$^{2+}$ is 2.39 Å, which is in excellent agreement with past DFT results (2.37 Å), high order Car–Parrinello molecular dynamics results (2.41 Å) and experimental results (2.39 Å, XRD). Furthermore, we have calculated the vibrational frequencies of [Ca(H$_2$O)$_6$]$^{2+}$. The IR peak position obtained from present calculations are in close agreement with Lei and Pan as shown in Fig. 2.
The equilibrium product concentration of dehydration and hydrolysis reactions of CaCl$_2$ hydrates:

The change in the Gibbs free energy of above reactions is calculated to quantify the vibrational contribution of energy and entropy. The information of vibrational spectra, ground state geometry and Gibbs free energy of all reactants and products are essential to calculate $\Delta G$ of dehydration and hydrolysis reaction. In the present study, we are interested in the following dehydration and hydrolysis reactions of CaCl$_2$ hydrates:

\[ \text{CaCl}_2 \cdot n\text{H}_2\text{O} \rightleftharpoons \text{CaCl}_2 \cdot (n-2)\text{H}_2\text{O} + 2\text{H}_2\text{O} \quad (n = 6, 4) \quad (4) \]
\[ \text{CaCl}_2 \cdot n\text{H}_2\text{O} \rightleftharpoons \text{CaCl}_2 \cdot (n-1)\text{H}_2\text{O} + \text{H}_2\text{O} \quad (n = 2, 1) \quad (5) \]
\[ \text{CaCl}_2 \cdot n\text{H}_2\text{O} \rightarrow \text{CaOHCl} + (n-1)\text{H}_2\text{O} + \text{HCl} \quad (n = 6, 4, 2, 1) \quad (6) \]

The change in the Gibbs free energy of above reactions is expressed as

\[ \Delta G = \sum G_{\text{product}} - \sum G_{\text{reactant}} \quad (7) \]

The equilibrium product concentration of dehydration and hydrolysis reactions can be obtained by equating $\Delta G$ to zero. For the $\Delta G$ calculation, the physical state of each reactant and product is important. Experimentally, CaCl$_2$ hydrates remain in the solid phase, and H$_2$O and HCl exist in the gaseous phase. However, the precise frequency calculation, subsequently the Gibbs free energy of solid CaCl$_2$ hydrates over a wide temperature range, is computationally challenging. The periodic unit cell of these CaCl$_2$ hydrates contains 36, 90, and 21 atoms for di, tetra, and hexahydrates respectively.\textsuperscript{31-35} Due to the large size of Ca and Cl atoms and a large number of atoms in unit cell, the Gibbs free energy calculation of the crystalline phase is expensive. Weck et al.\textsuperscript{30} obtained the Gibbs free energy of solid MgCl$_2$ hydrates under quasi-harmonic approximation (QHA) from GGA-DFT calculations. The Gibbs free energy of the crystalline phase obtained from QHA may breakdown at higher temperature. Despite the fact that these calculations were done for another solid salt hydrates (MgCl$_2 \cdot n$H$_2$O), it should also be valid for CaCl$_2 \cdot n$H$_2$O. Nevertheless, the solid reactant and product do not always remain in the crystalline phase and hydrolysis is expected to occur in the liquid phase.\textsuperscript{15} Thus for a solid–gas reaction, $\Delta G$ can be estimated by Gibbs free energy of the gaseous phase ($G_{\text{gas}}$). The equilibrium product concentration obtained under gas assumption will be considered as the safety limit of these reactions in seasonal heat storage systems.

In this paper, the $\Delta G$ in dehydration and hydrolysis reactions is computed under ideal poly-atomic gas assumption\textsuperscript{30} assuming each reactant and product to be in the gaseous state. Gibbs free energy of a gaseous molecule at given temperature $T$ and absolute pressure $p$ can be given as

\[ G(T, p) = U + pV - TS \quad (8) \]

The partition function of an ideal poly-atomic gas can be expressed as

\[ q = q_{\text{trans}}q_{\text{rot}}q_{\text{vib}}q_{\text{elec}} \quad (9) \]

where $q$ is the partition function and $q_{\text{trans}}, q_{\text{rot}}, q_{\text{vib}},$ and $q_{\text{elec}}$ are their translational, rotational, and electronic contributions.

Internal energy $U$ of poly-atomic gas can be expressed as

\[ U(T, p) = 3RT + R \sum_{j=1}^{3N-6} \left( \frac{\Theta_{\text{vib},j}}{2} + \frac{\Theta_{\text{rot},j}}{T} \right) + U_{\text{gr}} \quad (10) \]

where $R$ is the gas constant, $\Theta_{\text{vib},j}$ is vibrational temperature ($h\nu/k_b$) of the jth atom, $\nu$ is the frequency of the vibrational mode, $h$ is Planck’s constant, $k_b$ is the Boltzmann constant and $U_{\text{gr}}$ is the ground state energy of the molecule.

The entropy ($S$) of poly-atomic gas can be expressed as

\[ S(T, p) = R \left[ \ln \left( \frac{k_b T}{p} \right) \left( \frac{2\pi M k_b T}{\hbar^2} \right)^{3/2} \right] + \frac{3}{2} + R \sum_{j=1}^{3N-6} \left( \frac{\Theta_{\text{vib},j}}{T} \right) + S_{\text{rot}} \quad (11) \]

\[ + R \sum_{j=1}^{3N-6} \left[ \ln \left( 1 - e^{-\Theta_{\text{rot},j}/T} \right) \right] + S_{\text{rot}}. \]
where \( M \) is the total atomic mass and \( S_{\text{rot}} \) is the rotational contribution of entropy and can be expressed for the asymmetrical top molecule as
\[
S_{\text{rot}} = R \left[ \ln \left( \frac{\sqrt{\pi I_a I_b I_c \left( \frac{8\pi k_b T}{h^2} \right)^{3/2}}} {\sigma} \right) + \frac{3}{2} \right]
\] (12)

where \( I_a, I_b, \) and \( I_c \) are three principle moments of inertia and \( \sigma \) is the symmetry number of the molecule, which is the number of rotational modes of the molecule. From the optimized geometry of the molecule, we have obtained the \( U_{\text{gr}} \), principle moment of inertia and frequency of its vibrational mode. Eqn (7) and (8) are used to calculate \( \Delta G \) of a particular dehydration or hydrolysis reaction at given \( T \) and \( p \).

The H-bonds of CaCl\(_2\) hydrates are analyzed in the gaseous phase. This is a valid analysis as the locus of H-bond parameters can be generalized irrespective of the system being in a condensed phase or the gaseous phase. The presence of H-bonds is identified by the non-bonded distance \( X \cdot \cdot \cdot H \) (\( X = \text{Cl and O} \)) distance. When the non-bonded distance between O and H varies from 1.5 Å to 2.0 Å, we classify these H-bonds as OHO homonuclear type H-bonds, which is consistent with Iype et al. Similarly, when the non-bonded distance between Cl and H varies from 1.9 Å to 2.5 Å, we classify these H-bonds as OHCl heteronuclear type H-bond. Bader charges on various atoms in the molecule are calculated to understand the charge distribution. We compared the charge distribution in all hydrates to observe charge transfer, the effect of Coulombic interactions on their stability and enthalpy of dehydration.

### 3 Results and discussions

The gaseous structures of CaCl\(_2\) hydrates are investigated to understand the H-bond network, relative stability of conformers and enthalpy change in the hydrolysis and dehydration reactions. The optimized structures are used to obtain vibrational frequencies, and their entropic and energy contribution from DFT calculations. In this section, we present the optimized structures of CaCl\(_2\) hydrates and their conformers.

#### 3.1 Molecular structures of CaCl\(_2\) hydrates and their conformers

**3.1.1 Anhydrous CaCl\(_2\), CaOHCl, and CaCl\(_2\)-H\(_2\)O.** The gaseous anhydrous CaCl\(_2\) molecule is optimized with many initial configurations and all converged into the final CaCl\(_2\) structure as shown in Fig. 3a. The structure of CaCl\(_2\) is linear with the Ca–Cl bond length being 2.45 Å. The structural properties of CaCl\(_2\) are in agreement with the experimental crystalline structures of CaCl\(_2\). The optimized structure of CaOHCl is shown in Fig. 3b. CaOHCl is a non-linear molecule with Ca–Cl and Ca–O bond lengths of 2.49 Å and 1.99 Å. The optimized structure of CaCl\(_2\)-H\(_2\)O is shown in Fig. 3c. The Ca–O coordination length is 2.33 Å. The hydrated H\(_2\)O forms a H-bond with one of the Cl atoms of CaCl\(_2\) and stretched the Ca–Cl distance by 0.08 Å. The Bader charge of a Cl atom participating in H-bonding differs by 0.02 to that of a Cl atom not participating.

![Fig. 3](image-url)
3.1.2 \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \). Two stable conformers of \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) are observed. The optimized non-planar structure of \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) is shown in Fig. 3d and planar \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) is shown in Fig. 3e. The planar symmetric structure does not form a H-bond between H and Cl atoms while its conformer (non-planar) does. The intra-molecular H-bonds provide extra stability of 4.78 kcal mol\(^{-1}\) to the non-planar structure (Fig. 3d) over the planar structure (Fig. 3e) and are responsible for the loss in symmetry of the non-planar molecule. The Bader charge on Cl differs by 0.04 in planar and non-planar structures due to the presence of a H-bond. The average Ca–O coordination length is the same (2.36 Å) in both of these structures, which is in good agreement with the Ca–O coordination length (2.32 Å) in the experimental crystalline structures.\(^{31}\) There are two conformers reported for \( \text{MgCl}_2 \cdot 2\text{H}_2\text{O} \), which differ in energy by 4.53 kcal mol\(^{-1}\).\(^{32,33}\) Thus, H-bonds in \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) provide 0.25 kcal mol\(^{-1}\) extra stability over \( \text{MgCl}_2 \cdot 2\text{H}_2\text{O} \).

3.1.3 \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \). Calcium chloride tetrahydrate (\( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \)) subsists into three crystalline structures.\(^{32–34}\) We have chosen the initial structure of \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) from these three crystalline phases and observed three stable structures shown in Fig. 4a–c. The optimized structure (ground state energy \((U_{gf}) = -1628.66\) kcal mol\(^{-1}\)) obtained from an initial gaseous tetrahydrate phase crystalline structure (2.8 Å).\(^{32}\) The initial molecular structure is chosen close to the optimized geometry. There are intermolecular H-bonds (OHCl type) present in the \( \beta \) phase of the experimental crystalline structures,\(^{33}\) while intramolecular H-bonds are observed in the gaseous molecule. This is the reason for the slight discrepancy in the structure of a gaseous molecule and the \( \beta \) phase crystalline structure.

\( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) also occurs in the \( \gamma \) phase.\(^{34}\) The initial molecular structure is chosen close to the \( \gamma \) phase structure. The optimized structure is symmetrical and has almost a planar structure as shown in Fig. 4c. This structure is 19.24 kcal mol\(^{-1}\) less stable than the \( \beta \) phase optimized geometry. The average Ca–O coordination length is 2.26 Å in the \( \gamma \) phase optimized structure, which is in agreement with the \( \gamma \) phase crystalline structures (2.33 Å).\(^{34}\) The two H atoms of surrounded \( \text{H}_2\text{O} \) molecules form a relatively strong H-bonds with the Cl compared to the \( \alpha \) and \( \beta \) phase optimized structures. The \( \text{H}_2\text{O} \) takes away the Cl atom from the Ca, hence the Ca–Cl distance becomes 4.06 Å. The Bader charge on Cl is 0.07 less electro-negative compared to the Cl of the \( \beta \) phase optimized structure. The Ca–Cl stretching and the lower atomic charge result in weaker electrostatic attraction between Ca–Cl pairs compared to \( \alpha \) and \( \beta \) phase optimized structures. Thus, the \( \gamma \) optimized structure has lower stability over other conformers (\( \alpha \) and \( \beta \) phase) despite having relatively strong H-bonds. The \( \angle \text{Cl–H–O} = 166.9^\circ \) which enables the greater availability of the Cl lone pairs for the anti-bonding Cl–H orbital overlap.

3.1.4 \( \text{CaCl}_2 \cdot 6\text{H}_2\text{O} \). We observe three stable gaseous structures of \( \text{CaCl}_2 \cdot 6\text{H}_2\text{O} \) shown in Fig. 5a–c. The first conformer of the optimized \( \text{CaCl}_2 \cdot 6\text{H}_2\text{O} \) \((U_gf = -2319.15\) kcal mol\(^{-1}\)) is shown in Fig. 5a. The average Ca–O coordination length is 2.86 Å and is 2.405 Å, which is in agreement with the experimental \( \beta \) phase crystalline structure (2.46 Å).\(^{13}\) Similarly, the average Ca–Cl coordination of the gaseous phase tetrahydrate is 2.68 Å, which is in close agreement with the \( \beta \) phase crystalline structure (2.74 Å).\(^{13}\) There are two relatively strong H-bonds (OHCl) in the \( \beta \) phase optimized geometry compared to the \( \alpha \) phase optimized geometry. There are intermolecular H-bonds (OHCl type) present in the \( \beta \) phase of the experimental crystalline structures,\(^{33}\) while intramolecular H-bonds are observed in the gaseous molecule. This is the reason for the slight discrepancy in the structure of a gaseous molecule and the \( \beta \) phase crystalline structure.

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Fig. 4 The optimized structure of \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) conformers. Bader charges (black, italic) and coordination lengths (in Å, blue) are shown. Color scheme: Ca = yellow, Cl = green, O = red, and H = white.
the average Ca–Cl coordination length is 2.72 Å. This optimized hexahydrate molecule (Fig. 5a) has four OHCl H-bonds (heteronuclear type) and three OHO H-bonds (homonuclear type). The average intramolecular H-bond length (2.19 Å) is in close agreement with the average intermolecular H-bond length (2.27 Å) present in the experimental γ phase crystalline structures.35

CaCl2·6H2O exists experimentally in the γ phase.35 There are OHCl type intermolecular H-bonds present in the γ phase crystalline structures.35 Those H-bonds are between the H of the H2O of the hexahydrate and the Cl of the neighboring molecule of the hexahydrate in the crystal. The initial molecular structure is chosen close to the γ phase crystalline structure. In the optimized structure, the H2O forms an octahedral with the Ca as shown in Fig. 5b. The optimized structure (Fig. 5b) forms six H-bonds (2.00 Å) of OHCl type. The average Ca–O coordination length of γ phase optimized CaCl2·6H2O is 2.37 Å. The average Ca–Cl coordination length is 3.84 Å. There is a discrepancy of 26.6% in the Ca–Cl coordination length of the gaseous hexahydrate and the γ phase crystalline hexahydrate. There are intramolecular H-bonds present in the gaseous γ phase of hexahydrate (Fig. 5b) while intermolecular H-bonds in the γ phase crystalline structures,35 which explains the discrepancy.

The second stable conformer of the optimized CaCl2·6H2O is shown in Fig. 5c. This conformer is 4.12 kcal mol⁻¹ less stable than the first conformer (Fig. 5a) and 4.41 kcal mol⁻¹ less stable than the γ phase optimized structure (Fig. 5b). The average Ca–O coordination length is 2.45 Å. There are five intramolecular OHCl H-bonds (heteronuclear type) and one OHO type H-bond present in this conformer (Fig. 5c). Three H2O molecules are forming H-bonds (OHCl type) with one Cl, one H2O forms a H-bond (OHO type) with one of these H2O molecules and two H2O molecules form H-bonds with the other Cl atom. The average H···Cl H-bond length is 2.05 Å in the second stable conformer (Fig. 5c) while in the more stable γ phase optimized structure (Fig. 5b) it is 2.00 Å and 2.19 Å in the first conformer (Fig. 5a). The average O···H H-bond length is 1.88 Å in the second conformer (Fig. 5c) and in the first conformer, it is 1.83 Å. The O···H H-bonds are stronger in the first conformer, which explains the extra stability of the first conformer over the second conformer. The detailed explanation of H-bonds will be discussed in the section “Hydrogen bond in CaCl2 hydrate system”.

No conformer for MgCl2·4H2O and MgCl2·6H2O has been reported from DFT calculations.15 To rationally compare the structure of CaCl2 hydrates with MgCl2 hydrates, the average Bader charge on each atom is compared in Table 1. The Bader charges on Ca and Cl atoms of CaCl2 are lower in magnitude when compared with Mg and Cl in their analogous MgCl2 hydrates. The Bader charge on Ca and Mg atoms increases monotonically with the hydration number. The magnitude of the Bader atomic charge on Cl in CaCl2 hydrates decreases with hydration number whereas it increases from MgCl2 to MgCl2·2H2O and decreases till MgCl2·6H2O. The magnitude of average Bader atomic charges on O and H in CaCl2 hydrates is lower than MgCl2 hydrates. The Ca–Cl, Ca–O distance in CaCl2 hydrates is higher than MgCl2 hydrates. Thus, electrostatic attractions
between Mg–O pairs in MgCl₂ hydrates are stronger than Ca–O pairs in CaCl₂ hydrates.

### 3.2 Reaction enthalpies of CaCl₂ hydrates

For reaction enthalpy calculations, we have considered the structure having minimum energy and no imaginary frequency. Table 2 describes the electronic ground state energy and various energy terms for all the studied hydrates. The binding energy ($\Delta E_{\text{Binding}}$) of a hydrate is defined as

$$\Delta E_{\text{Binding}} = E_{\text{CaCl}_2+n\text{H}_2\text{O}} - (E_{\text{CaCl}_2} + nE_{\text{H}_2\text{O}})$$  (13)

The enthalpy change during the dehydration reaction ($\Delta E_{\text{Dehydration}}$) is defined as

$$\Delta E_{\text{Dehydration}} = 0.5 \times [E_{\text{CaCl}_2(n-2)\text{H}_2\text{O}} + 2E_{\text{H}_2\text{O}} - E_{\text{CaCl}_2+n\text{H}_2\text{O}}] \quad (n = 4, 6)$$  (14)

$$\Delta E_{\text{Dehydration}} = [E_{\text{CaCl}_2+n\text{H}_2\text{O}} - E_{\text{CaCl}_2(n-1)\text{H}_2\text{O}} - E_{\text{H}_2\text{O}}] \quad (n = 1, 2)$$  (15)

Similarly, the enthalpy change during the hydrolysis reaction ($\Delta E_{\text{Hydrolysis}}$) is defined as

$$\Delta E_{\text{Hydrolysis}} = [E_{\text{CaCl}_2+n\text{H}_2\text{O}} - E_{\text{CaCl}_2}(n-1)\text{H}_2\text{O} - E_{\text{H}_2\text{O}} - E_{\text{HCl}} - nE_{\text{H}_2\text{O}}] \quad (n = 1, 2, 4, 6)$$  (16)

The H₂O and HCl molecules are separately optimized using GGA-PW91 and TZ2P basis sets. The binding energy of CaCl₂ hydrates and their conformers depends on the Ca–Cl, Ca–O distance and on the strength of the intramolecular H-bond formed. The binding energy is the energy released on hydration per mole of salt, whereas the dehydration enthalpy is per mole of H₂O. For the monohydrate, the binding energy is the same as the enthalpy change in dehydration, but opposite in sign, and it is equal to $-21.9 \text{ kcal mol}^{-1}$. The H₂O molecule attached with CaCl₂ forms a monohydrate. The attached H₂O forms a H-bond with the Cl of CaCl₂ (see Fig. 3c). Similarly, another H₂O molecule attached with the monohydrate and 20.48 kcal mol⁻¹ energy is released. The strength of the H-bond in monohydrate (2.24 Å, Fig. 3c) and in the dihydrate (2.205 Å, Fig. 3d) are of similar magnitude so the dehydration enthalpy is very close. The dehydration enthalpy released per H₂O molecule decreases from 21.9 kcal mol⁻¹ to 14.56 kcal mol⁻¹ for the increase in the hydration number from 1 to 6. The average negative atomic charge induced on the O of the H₂O molecule is decreased in magnitude from CaCl₂·H₂O (−1.202) to CaCl₂·6H₂O (−1.184). This suggests that dehydration enthalpy is dominated by electrostatic attraction. The dehydration of two H₂O molecules from CaCl₂·6H₂O is easier than the removal of the two H₂O molecules from CaCl₂·4H₂O and so on. The enthalpy change in the hydrolysis reaction is very high for CaCl₂·6H₂O compared to the enthalpy change in dehydration as given in Table 2. The enthalpy change in hydrolysis decreases from higher hydrates to lower hydrates.

The enthalpy change during the dehydration and hydrolysis reaction of CaCl₂ hydrates is compared with MgCl₂ hydrates, as shown in Table 2. The dehydration enthalpy of CaCl₂·6H₂O, CaCl₂·4H₂O, and CaCl₂·2H₂O is slightly lower than analogous MgCl₂ hydrates. The dehydration enthalpy for CaCl₂·2H₂O is 26.8% higher than MgCl₂·2H₂O, thus CaCl₂·2H₂O dehydrates at higher temperature compared with MgCl₂·2H₂O. The CaCl₂ hydrates (except mono hydrate) have high enthalpy of hydrolysis compared to MgCl₂ hydrates (as given in Table 2) thus, CaCl₂ hydrates are more resistant to hydrolysis reaction.

### 3.3 Hydrogen bond in the CaCl₂ hydrate system

CaCl₂ hydrates have the OHCl type heteronuclear H-bond and the HOH type homonuclear H-bond. The strong H-bond has a predominant covalent character while the moderate H-bond has mostly electrostatic character. The H-bond energy is a strong function of $r_{X-H}$, $r_{X-H}$, and $r_{X-Y}$ in XHX and XHY H-bonded systems. The Cl atom is larger in size compared to O and less electronegative thus the availability of the lone pair for H-bond formation is lower compared to the O. Thus the OHO type H-bond is stronger than the OHCl type H-bond.

In the present study, we observe the OHO type homonuclear H-bond in CaCl₂·6H₂O isomers (Fig. 5a and c). The strength of H-bonds depends on the distance between the donor H and the acceptor O distance (H-bond distance) and the $\angle \text{O-H-H}$ (H-bond angle). The $\angle \text{O-H-H}$ of 165° provides the complete lone pair availability in the O for the anti-bonding O-H orbital overlap therefore stabilizing the H-bond. The non-bonded O–O distance varies from 1.64 Å to 1.95 Å and $\angle \text{O-H-H}$ of 135.9° to 165.4°. The lengthening of the O–H bond varies from 0.017 Å to 0.041 Å. These OHO type H-bonds are moderate in strength and predominantly electrostatic in nature.

### Table 1

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Ca [Mg]</th>
<th>Cl</th>
<th>O</th>
<th>H</th>
<th>Ca–O [Mg–O]</th>
<th>Ca–Cl [Mg–Cl]</th>
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</thead>
<tbody>
<tr>
<td>CaCl₂·6H₂O</td>
<td>1.633</td>
<td>(1.762)</td>
<td>-0.696 (−0.760)</td>
<td>−1.193 (−1.210)</td>
<td>0.577 (0.575)</td>
<td>2.37 (2.1)</td>
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<tr>
<td>CaCl₂·4H₂O</td>
<td>1.573</td>
<td>(1.700)</td>
<td>-0.755 (−0.802)</td>
<td>−1.185 (−1.188)</td>
<td>0.582 (0.583)</td>
<td>2.40 (2.15)</td>
</tr>
<tr>
<td>CaCl₂·2H₂O</td>
<td>1.538</td>
<td>(1.696)</td>
<td>-0.762 (−0.837)</td>
<td>−1.192 (−1.200)</td>
<td>0.587 (0.590)</td>
<td>2.36 (2.08)</td>
</tr>
<tr>
<td>CaCl₂·H₂O</td>
<td>1.554</td>
<td>(1.650)</td>
<td>-0.770 (−0.826)</td>
<td>−1.202 (−1.204)</td>
<td>0.594 (0.605)</td>
<td>2.33 (2.05)</td>
</tr>
<tr>
<td>CaCl₂ [MgCl₂]</td>
<td>1.553</td>
<td>(1.621)</td>
<td>-0.776 (−0.811)</td>
<td>N.A.</td>
<td>N.A.</td>
<td>2.49 (2.21)</td>
</tr>
</tbody>
</table>

### Table 2

<table>
<thead>
<tr>
<th>Molecule</th>
<th>$U_p$/kcal mol⁻¹</th>
<th>$\Delta E_{\text{Binding}}$/ kcal mol⁻¹</th>
<th>$\Delta E_{\text{Dehydration}}$/ kcal mol⁻¹</th>
<th>$\Delta E_{\text{Hydrolysis}}$/ kcal mol⁻¹</th>
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<tr>
<td>CaCl₂·6H₂O</td>
<td>−2319.15</td>
<td>14.56 (14.8)</td>
<td>121.68 (119.3)</td>
<td></td>
</tr>
<tr>
<td>CaCl₂·4H₂O</td>
<td>−1630.28</td>
<td>16.57 (16.7)</td>
<td>92.53 (89.7)</td>
<td></td>
</tr>
<tr>
<td>CaCl₂·2H₂O</td>
<td>−937.40</td>
<td>21.19 (15.7)</td>
<td>59.41 (46.3)</td>
<td></td>
</tr>
<tr>
<td>CaCl₂·H₂O</td>
<td>−587.05</td>
<td>21.9 (22.4)</td>
<td>38.93 (40.3)</td>
<td></td>
</tr>
</tbody>
</table>
conformer of CaCl$_2$·6H$_2$O (Fig. 5a), three H-bonds of this type are present while in the second conformer of CaCl$_2$·6H$_2$O (Fig. 5c) only one H-bond is present. The first conformer of CaCl$_2$·6H$_2$O has one relatively strong but moderate strength H-bond (1.64 Å) thus provides 4.12 kcal mol$^{-1}$ stability over the second conformer.

The CaCl$_2$ hydrates have mostly OHCl type heteronuclear H-bonds. The H atoms of the hydrated CaCl$_2$·xH$_2$O molecule are attracted by adjacent Cl atoms and form OHCl type H bonds. In the present case, the Cl···H distance varies from 1.82 Å to 2.48 Å and the Cl···O distance varies from 2.86 Å to 3.15 Å. The lengthening of the OH bond varies from 0.01 Å to 0.08 Å. A similar bond length of H-bonds is observed in organic and organometallic crystals. Tommaso $et$ al. also observed the OHCl type H-bond in the hydration of Ca$^{2+}$ ions in salt solution. The bond valence is proportional to the electron density. The bond valence of the atom is distributed between the bonds which it forms. Each bond involves the same number of electrons. The BV sum rule is described as

$$S_{\text{H-O}} + S_{\text{H-Cl}} = 1 \quad (17)$$

In heteronuclear H-bonds, the BV sum rule (eqn (17)) has four parameters (two of O-H and two of H···Cl). From Fig. 6, it is evident that heteronuclear H-bonds follow the BV sum rule in close agreement ($r^2 = 0.98$). Therefore, the BV sum rule is applicable for this class of H-bonds. The bond parameters $a$ and $b$ for H···Cl are 0.96 Å, 0.80 Å and for the H···O pair 0.97 Å, 0.38 Å, which is in agreement with the literature. The strength of the H-bond is inversely related to the H···Cl distance. The non-bonded H···Cl distance depends on $\angle$ O-H···Cl, which varies from 166.9° to 123° in CaCl$_2$ hydrates. The larger $\angle$ O-H···Cl (> 165°) facilitates anti-bonding Cl···H orbital overlap and favors hydrogen bonding. Most of these (~ 77%) OHCl type H-bonds are moderate and weak in strength and are dominated by electrostatic interactions and dispersion. In CaCl$_2$·4H$_2$O, the H-bonds destabilize the structure by 19.24 kcal mol$^{-1}$ while stabilize the CaCl$_2$·6H$_2$O and CaCl$_2$·2H$_2$O by 4.41 kcal mol$^{-1}$ and 4.78 kcal mol$^{-1}$ respectively. This suggests that the formation of strong H-bonds in CaCl$_2$·4H$_2$O could form metastable states and may result in sluggish hydration kinetics.

### 3.4 Equilibrium product concentrations

The $\Delta G$ of a system is a function of pressure, temperature, and the system characteristics. At chemical equilibrium, the $\Delta G$ (from eqn (7)) should be zero and this will provide information about the equilibrium concentration of products.

#### 3.4.1 Dehydration reaction of CaCl$_2$ hydrates and their comparison with MgCl$_2$ hydrates

In the dehydration reaction, the higher hydrate absorbs energy and disintegrates into lower hydrates and H$_2$O molecules. In the experimental dehydration set-up only the water vapor pressure ($p_{\text{H}_2\text{O}}$) and the HCl pressure ($p_{\text{HCl}}$) are the controlled variables. The partial pressure (concentration) of hydrates is not important. In the present study, the partial pressure of water and system temperature are the control variables, while the partial pressure of hydrates is kept constant at 1 atm. The equilibrium water vapor pressure depends on the system temperature for all dehydration reactions of both hydrates (CaCl$_2$ and MgCl$_2$) as shown in Fig. 7. It is recognizable from Fig. 7 that dehydration from hexahydrate to tetrahydrate is easier (favored at lower temperature) than from tetra to dihydrate and from dihydrate to monohydrate for CaCl$_2$ and MgCl$_2$ hydrates. This trend of dehydration of various CaCl$_2$ hydrates is similar to the experimental results and outlined as well by the enthalpy of dehydration in Table 2, where the enthalpy of dehydration increases from hexa to monohydrate.

Rammelberg $et$ al. reported from thermal gravimetric analysis/differential scanning calorimetry (TGA/DSC) measurements 30% conversion of CaCl$_2$·6H$_2$O to CaCl$_2$·4H$_2$O at 336 K. In the present study, we observed 30% conversion of CaCl$_2$·6H$_2$O to CaCl$_2$·4H$_2$O only at around 340 K. There is an offset of 4 K is observed. The effect of temperature on the equilibrium dehydration curve of CaCl$_2$·6H$_2$O (slope = 0.035) is in agreement with the experiments (slope = 0.032) as shown in Fig. 7. An offset of 20 K is observed in the low vapor regime (< 0.03 atm) for CaCl$_2$·6H$_2$O. The offset may be a consequence of the ideal polyatomic gas phase assumption. Another reason could be that the equilibrium product concentrations obtained from $\Delta G$ (DFT calculations) reports the static equilibrium properties while the experimental (TGA/DSC)
kinetics reports the dynamic equilibrium properties. There is a similar offset observed in the equilibrium dehydration curve for \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) to \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) as well. The equilibrium partial vapor pressure (\( p_{\text{HCl}} \)) of 1 atm is observed at 350.5 K for \( \text{CaCl}_2 \cdot 6\text{H}_2\text{O} \) while at 490.5 K and 586.5 K for \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) and \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) as shown in Fig. 7.

The gradient of the dehydration curve for \( \text{CaCl}_2 \cdot 6\text{H}_2\text{O} \) is 0.002 higher than \( \text{MgCl}_2 \cdot 6\text{H}_2\text{O} \) while 0.007 lower for \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) and \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) when compared with their analogous \( \text{MgCl}_2 \) hydrates. Thus, \( \text{CaCl}_2 \cdot 6\text{H}_2\text{O} \) has the shortest range of dehydration temperature while \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) and \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) have a higher temperature range of operation. The dehydration enthalpy of \( \text{CaCl}_2 \cdot 6\text{H}_2\text{O} \) is 0.25 kcal mol\(^{-1}\) lower than \( \text{MgCl}_2 \cdot 6\text{H}_2\text{O} \) (as given in Table 2) thus it dehydrates at lower temperature compared to \( \text{MgCl}_2 \cdot 6\text{H}_2\text{O} \). The dehydration enthalpy of \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) is 0.13 kcal mol\(^{-1}\) lower than \( \text{MgCl}_2 \cdot 4\text{H}_2\text{O} \) while \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) dehydrates at higher temperature than \( \text{MgCl}_2 \cdot 4\text{H}_2\text{O} \). This behavior can be explained from the role of the meta-stable conformer formed due to presence of strong H-bonds present in the \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) isomer (Fig. 4c). The dehydration enthalpy of \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) is 5.3 kcal mol\(^{-1}\) higher than \( \text{MgCl}_2 \cdot 2\text{H}_2\text{O} \) (see inset graph of Fig. 7) so it dehydrates at higher temperature compared with \( \text{MgCl}_2 \cdot 2\text{H}_2\text{O} \).

3.4.2 Hydrolysis reaction of \( \text{CaCl}_2 \) hydrates and their comparison with \( \text{MgCl}_2 \) hydrates.

Hydrolysis starts at higher temperature (> 396 K) in comparison to dehydration for all the hydrates of \( \text{CaCl}_2 \), which is consistent with the enthalpy change in dehydration (38.93 kcal mol\(^{-1}\) to 121.68 kcal mol\(^{-1}\)) and dehydration (21.9 kcal mol\(^{-1}\) to 14.56 kcal mol\(^{-1}\)) as given in Table 2.

The equilibrium vapor pressure of 1 atm is observed from \( \text{CaCl}_2 \cdot 6\text{H}_2\text{O} \) at 350.5 K and hydrolysis above the safety limit (\( p_{\text{HCl}} = 1 \times 10^{-3} \) atm) starts at 396 K. There is no overlapping temperature region between hydrolysis and dehydration for \( \text{CaCl}_2 \cdot 6\text{H}_2\text{O} \). The \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) dehydrates into \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) at 490.5 K (\( p_{\text{HCl}} = 1 \) atm) while hydrolysis starts (above safety limit, \( p_{\text{HCl}} > 1 \times 10^{-3} \) atm) at 475.8 K. There is an overlap of around 14.7 K (from Fig. 7 and 8). Similarly the overlap for the \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) is around 32.5 K. Fraissler et al.\(^{14}\) reported experimentally from TGA measurements that the HCl formation from thermal decomposition of \( \text{CaCl}_2 \) salt hydrates occurs in the temperature range of 683–1013 K. We observe the starting point (\( p_{\text{HCl}} = 1 \times 10^{-3} \) atm, onset) of HCl formation for \( \text{CaCl}_2 \cdot 2\text{H}_2\text{O} \) and \( \text{CaCl}_2 \cdot \text{H}_2\text{O} \) at 660 K and 845 K (\( p_{\text{HCl}} = 1 \times 10^{-3} \) atm) respectively. \( \text{CaCl}_2 \cdot 4\text{H}_2\text{O} \) and \( \text{CaCl}_2 \cdot 6\text{H}_2\text{O} \) will not be present in this temperature range. The offset in the onset of HCl formation temperature may be the consequence of the dynamic and static equilibrium comparison and the effect of the ideal polyatomic gas assumption. The hydrolysis of \( \text{CaCl}_2 \cdot \text{H}_2\text{O} \) is independent of the \( \text{H}_2\text{O} \) vapor pressure as shown in Fig. 8.

Hydrolysis of \( \text{MgCl}_2 \) hydrates starts at lower temperature in comparison to \( \text{CaCl}_2 \) hydrates for hexa, tetra, and dihydrates, which is consistent with the enthalpy change in hydrolysis of \( \text{MgCl}_2 \) and \( \text{CaCl}_2 \) hydrates as given in Table 2. The slope of the \( \text{MgCl}_2 \) hydrate hydrolysis curve is slightly higher compared to \( \text{CaCl}_2 \) hydrates so the hydrolysis of \( \text{CaCl}_2 \) hydrates has a longer operating temperature range. Hydrolysis in \( \text{MgCl}_2 \cdot 6\text{H}_2\text{O} \), \( \text{MgCl}_2 \cdot 4\text{H}_2\text{O} \) and \( \text{MgCl}_2 \cdot 2\text{H}_2\text{O} \) starts at 388, 446, and 616 K (\( p_{\text{HCl}} = 1 \times 10^{-3} \) atm) respectively, which is lower than analogous \( \text{CaCl}_2 \) hydrates. Hydrolysis in \( \text{CaCl}_2 \cdot \text{H}_2\text{O} \) starts at lower temperature (845 K) than \( \text{MgCl}_2 \cdot \text{H}_2\text{O} \) (976 K). Thus, \( \text{CaCl}_2 \) hydrates have improved hydrolysis resistance if the operating temperature is less than 800 K.

Furthermore, the partial pressure of HCl (\( p_{\text{HCl}} \)) is plotted with their equilibrium temperature at fixed \( \text{H}_2\text{O} \) vapor pressure (\( p_{\text{H}_2\text{O}} = 4 \times 10^{-3} \) atm) for \( \text{MgCl}_2 \) and \( \text{CaCl}_2 \) hydrates in Fig. 9. For any temperature (< 761 K), \( \text{MgCl}_2 \) will always have high partial pressure of HCl (\( p_{\text{HCl}} \)) compared to \( \text{CaCl}_2 \) hydrates. The slope of the equilibrium hydrolysis curve (Fig. 9) decreases from the hexahydrate to the monohydrate for both \( \text{CaCl}_2 \) and \( \text{MgCl}_2 \) hydrates. The slope of \( \text{CaCl}_2 \) is slightly lower than analogous \( \text{MgCl}_2 \) hydrates. Hence, the HCl concentration could be better controlled with temperature in \( \text{CaCl}_2 \) hydrates over \( \text{MgCl}_2 \) hydrates.

For a given change in partial pressure of HCl (\( p_{\text{HCl}} \)), the temperature changes maximum for the monohydrate. The reason for the change in the slope of the hydrolysis curve can be explained by the stoichiometric ratio between HCl and \( \text{H}_2\text{O} \). The stoichiometric ratio of HCl to \( \text{H}_2\text{O} \) increases from 0.2 to 1 for the hexahydrate to the monohydrate. Thus, the molar ratio between HCl and \( \text{H}_2\text{O} \) decreases from mono to hexahydrate, therefore, temperature varies maximum for the monohydrate for a given change in \( p_{\text{HCl}} \). The similar trend of slope is observed for \( \text{MgCl}_2 \) hydrates.
The equilibrium water vapor pressure ($p$) increases with an increase in the HCl partial pressure. The equilibrium vapor pressure for hydrolysis and dehydration of CaCl$_2$ hydrates at different temperatures is obtained from DFT and compared with MgCl$_2$ hydrates. The equilibrium composition of dehydration and hydrolysis reactions is obtained by equating the $\Delta G$ to zero over a wide range of temperature and pressure conditions. The trend in the dehydration reaction is similar to the experiments for all the CaCl$_2$ hydrates. The effect of temperature on the dehydration of CaCl$_2$-6H$_2$O ($\beta$ phase) is in close agreement with experiments (slope = 0.032). We compared the equilibrium composition of dehydration and hydrolysis reactions for CaCl$_2$ and MgCl$_2$ hydrates. The CaCl$_2$-6H$_2$O dehydrates at lower temperature when compared to their analogous MgCl$_2$ hydrates. The CaCl$_2$ hydrates have better temperature control with the partial pressure of the products ($p_{CaCl_2}, p_{H_2O}$) when compared to the MgCl$_2$ hydrates. We investigated dehydration and hydrolysis at different temperature, pressure regimes for CaCl$_2$ and MgCl$_2$ hydrates. The CaCl$_2$ hydrates have better hydrolysis resistance over MgCl$_2$ hydrates in the temperature range 273 to 800 K, which explains the higher stability in the mixture of CaCl$_2$ and MgCl$_2$ hydrates. Hydrolysis is very rare for CaCl$_2$-6H$_2$O because the onset of hydrolysis reaction at the safety limit ($p_{CaCl_2} = 1 \times 10^{-3}$ atm) is higher than the dehydration temperature of CaCl$_2$-6H$_2$O. Similarly CaCl$_2$-4H$_2$O and CaCl$_2$-2H$_2$O can undergo hydrolysis ($p_{CaCl_2} = 1 \times 10^{-3}$ atm) at 475.8 and 660 K respectively. The similar temperatures for MgCl$_2$ hydrates are 446 and 616 K. CaCl$_2$-H$_2$O may undergo hydrolysis above 850 K. These temperature ranges should be

### 4 Conclusions

To understand the equilibrium product concentrations of the dehydration and hydrolysis reactions of CaCl$_2$-nH$_2$O in comparison to MgCl$_2$-nH$_2$O, DFT calculations have been carried out. The structural properties of CaCl$_2$ hydrates along with their atomic charges are analyzed. This study reveals the formation of many conformers due to the presence of various H-bonds. The relative stability of a conformer and dehydration enthalpy is dominated by electrostatic attraction and strength of H-bond formed. We observed a homonuclear H-bond (OHO) and a heteronuclear H-bond (OHCl) in the various hydrates of CaCl$_2$. Most of the H-bonds (~77%) in CaCl$_2$ hydrate are of moderate strength. CaCl$_2$-2H$_2$O has two conformers due to the presence of an OHCl type H-bond. Similarly, three conformers of CaCl$_2$-4H$_2$O, and CaCl$_2$-6H$_2$O are observed. The most stable conformers are non-planar CaCl$_2$-2H$_2$O (Fig. 3d), $\beta$ phase optimized CaCl$_2$-4H$_2$O (Fig. 4b), and $\gamma$ phase optimized CaCl$_2$-6H$_2$O (Fig. 5b). These conformers are 4.78, 19.24, and 4.41 kcal mol$^{-1}$ more stable than their lowest stable conformer. The relatively strong OHCl H-bonds provide 4.41 kcal mol$^{-1}$ stability compared to its lowest stable conformer in CaCl$_2$-6H$_2$O. The H-bonds in CaCl$_2$-4H$_2$O destabilize the $\gamma$ phase optimized structure by 19.24 kcal mol$^{-1}$ compared with the $\beta$ phase optimized structure. This implies that CaCl$_2$-4H$_2$O could form a metastable stage, and may have a sluggish hydration kinetics. Both CaCl$_2$-2H$_2$O and CaCl$_2$-6H$_2$O have less stable conformers, consequently they do not form metastable stages. Thus, the H-bonds should not markedly affect their hydration kinetics. The BV sum rule parameters for OHCl type H-bonds are 0.96 Å, 0.80 Å (for H-Cl pair) and 0.97 Å, 0.38 Å (for H-O pair). These parameters can be used to characterize the OHCl H-bonded system.

Fig. 9 Equilibrium vapor pressure for the hydrolysis reactions of CaCl$_2$ hydrates (green) at various temperatures and fixed H$_2$O pressure $p_{H_2O} = 4 \times 10^{-3}$ atm and $p_0 = 1$ atm. The dashed red lines represent the hydrolysis curves of MgCl$_2$ hydrates. The inset graph shows the comparison of enthalpy change during the hydrolysis reactions of CaCl$_2$ and MgCl$_2$ hydrates.

Fig. 10 Equilibrium water vapor pressure for hydrolysis and dehydration reactions at different temperatures. $p_{HCl}$ is varied from $10^{-7}$ to $10^{-2}$ atm and $p_0 = 1$ atm.
treated as the safety limit range for hydrolysis reactions. The CaCl₂ hydrates could be mixed with MgCl₂ hydrates to improve the hydrolysis resistance and therefore durability of the system.

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References