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SOLUBILITY OF CALCIUM LACTATE IN AQUEOUS SOLUTION

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ABSTRACT

The presence of calcium lactate (CaL_2) crystals on the surface of Cheddar cheese is an appearance defect that consumer often attributes to mold. This problem is widespread in the cheese industry. The solubility of CaL_2 in water determines the likelihood of formation of CaL_2 crystals on the surface of Cheddar cheese. Very little fundamental work has been done to measure the solubility of CaL_2 at conditions appropriate to storage of cheese. CaL_2 is a moderately-soluble salt, with saturation concentrations that increase from about 2 g anhydrous CaL_2 /g water at 0°C to about 50 g anhydrous CaL_2 /g water at 60°C. The substantial variability in data available in the literature attests to the difficulties in accurately measuring solubility concentrations. The limited amount of data available in the literature suggest that a decrease in pH through addition of lactic acid leads to a decrease in solubility, and that addition of salt (NaCl) has no effect on solubility.

Key Words: Calcium lactate; Solubility; Crystallization; Salts

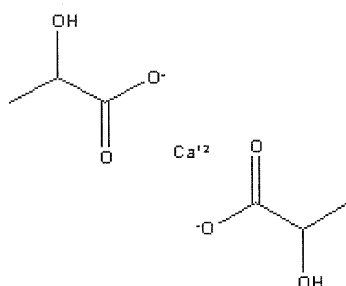
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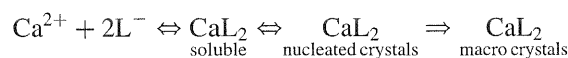
INTRODUCTION

In the research done in previous years, it was established that calcium lactate (CaL_2) is the main reason for the appearance of white crystals on the surface of Cheddar cheese.^[1,2] The formation of these crystals is promoted by the high lactose concentration in milk,^[3] an increase in lactic acid level in cheese, the racemization of lactic acid by nonstarter bacteria, and low temperature during storage after holding cheeses at a higher temperature.^[2]

It is known that CaL_2 is not a native component of milk; and therefore, free calcium and lactate ions should be present in cheese in order to form crystals. The chemical structure of CaL_2 is shown below.



The following equation can illustrate CaL crystal formation in cheese:^[1]

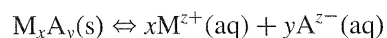


Typically, CaL_2 forms a pentahydrate crystal, where five molecules of water per CaL_2 are incorporated into the crystal lattice.

Each of the factors mentioned above have been extensively studied and reviewed in Refs. [1–3]. However, very few studies have evaluated the solubility of CaL_2 in terms of temperature, pH, and the addition of salts present in Cheddar cheese. The ability to determine the solubility of CaL_2 in terms of these factors will provide the necessary information to study the crystallization kinetics of CaL_2 and hence, this information will complete the knowledge needed to prevent the formation of CaL_2 crystals.

SOLUBILITY PRODUCT

The equilibrium of a solid M_xA_y with its saturated solution can be shown by the following equation:



For this equation, the thermodynamic equilibrium constant expression is:

$$K = ((a_+)^x(a_-)^y)/(a_{\text{solid}}),$$



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where (a_+) and (a_-) represent “effective” concentrations, called the ion activities, which account for interionic attractions.

Since the activity of a pure solid equals one and in dilute solutions ($10^{-5} N$) the ion activities are approximately the same as the molar concentrations,^[4] the above equation can be written as:

$$K_c = [M_+]^x[A_-]^y$$

Usually, instead of the equilibrium constant K_c , a special term called the solubility product constant, K_{sp} is used. The K_{sp} describes the equilibrium in a saturated solution of a slightly soluble ionic compound at a certain temperature and is the product of ionic concentration terms, with each term raised to an appropriate power.^[5]

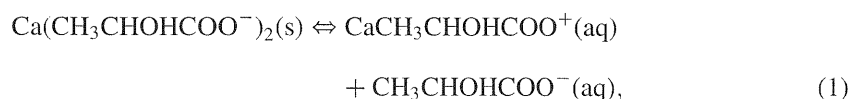
To determine whether a precipitate will form or not when the two solutions are mixed, the ion product (Q) is used. The ion product has the same form as K_{sp} for a given solid, except instead of the equilibrium concentrations, the initial concentrations are used. When Q is equal to K_{sp} , the system is saturated. If Q is larger than K_{sp} , the ion concentration is higher than the solution can absorb and a precipitate forms. If Q is smaller than K_{sp} , no precipitation occurs.^[5]

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CaL_2 exists in $\text{L}(+)$ and $\text{D}(-)$ isomeric forms. In cheese, starter bacteria ferment lactose to produce $\text{L}(+)$ -lactate, part of which may be converted to $\text{D}(-)$ -lactate by some strains of nonstarter bacteria during ripening.^[6] According to several studies,^[1,2,6,7] this racemic mixture of DL-CaL_2 appears to be one of the reasons for CaL_2 crystal formation. Thomas and Crow^[6] suggested that the moisture phase of Cheddar cheese is close to saturation with $\text{L}(+)$ - CaL_2 and therefore, the formation of less soluble DL-lactate may lead to crystal development. This appears to be the only source that relates the solubility of the two CaL_2 isomers and crystal formation.

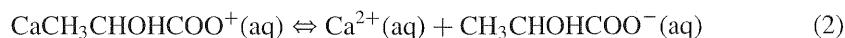
CaL_2 is a moderately-soluble compound whose solubility in water varies depending on the pH, temperature, and the presence of other compounds. Saturated solutions of moderately or highly soluble ionic compounds are simply much too concentrated to permit the assumption that activities and molar concentrations are equal. Without this assumption, much of the value of the solubility product concept is lost;^[5] therefore, the solubility product constant cannot be applied to CaL_2 .

Typically, solubility measurements involve dissolving pure crystals in water. The equilibrium existing in the saturated solution of CaL has two steps.^[8] The first step goes to completion:





whereas the second step, represented by the following equation:

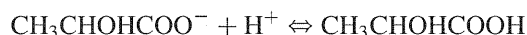


does not go completely to the end point. The dissociation constant for this ion was calculated by Davis^[9] using the following equation:

$$\log K = \log[\text{Ca}^{2+}][\text{L}^-]/[\text{CaL}^+] - 2.02\sqrt{I} + 2.8I,$$

where I is the ionic strength. The value for the dissociation constant is 0.034 at 25°C.^[9] The reason for the incomplete dissociation of CaL^+ is the presence of the negatively charged alcohol groups in the anion molecule, which cause the association of Ca^{2+} and L^- ions into complex CaL^+ .^[8]

Some of the lactate ions from the dissociation reactions associate with hydrogen ions of water and lactic acid will form, as shown below:



Lactic acid is a weak acid and, much like any other weak acid, it does not completely dissociate in solution and its ionization depends on the dissociation constant. The dissociation constant (K_a) for lactic acid is 1.4×10^{-4} . Lactic acid is highly soluble in water and readily dissociates to produce a negative ion (lactate) and a hydrogen ion (i.e., the reverse reaction occurs).

As mentioned above, CaL_2 can be found in L and D isomeric forms or a mixture of both isomers. The isomeric form of CaL_2 present in a solution determines which isomeric form of lactic acid forms.

Even though CaL_2 cannot be considered a slightly-soluble salt, we can still relate factors affecting the solubilities of slightly-soluble salts to CaL . These factors include the common ion effect, uncommon ion effect (salt effect), and ion-pair formation.

The addition of a small amount of salt with a common ion to a saturated solution of calcium lactate disrupts the established equilibrium and increases the concentration of one of the ions. According to Le Châtelier's principle, to compensate for this increase, the equilibrium shifts towards formation of solid CaL_2 and a precipitate forms.

The presence of the "uncommon" ions leads to an increase in the total ionic concentration of a solution and the stoichiometric concentration becomes larger compared to the ion activity. In order to reestablish the equilibrium, more solute (CaL_2) will be dissolved (i.e., the CaL_2 solubility increases). However, uncommon ions can cause the formation of products that are not soluble and precipitate. For example, addition of citrate or phosphate causes the formation of slightly-soluble calcium citrate or insoluble calcium phosphate that precipitates. To prevent any precipitation, highly-soluble salts should be used to study the uncommon ion effect on solubility of CaL_2 .

As mentioned above, the second step of the dissociation does not go to completion; therefore, in a CaL_2 solution with fully separated ions and cations, the



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partially ionized pairs of CaL^+ are also present. The presence of ion pairs reduces the free ion concentrations and, in order to maintain the equilibrium, the amount of solute increases and the solubility increases.

The ionic species participating in a solubility equilibrium are able to enter into acid–base reactions and, as a result, the solubility of a solute is affected by pH. Ionized lactic acid is extremely soluble in water; however, in cheese, the lactic acid interacts with calcium to form less soluble calcium salts. If there is residual lactose in the cheese, more and more lactic acid will be formed by bacteria during ripening and storage of cheese, the pH decreases and there is a greater chance for CaL_2 crystals to appear.^[10]

METHODS OF SOLUBILITY MEASUREMENT

There are two general methods used in the determination of solubility: dissolving crystals and precipitating crystals out of solution. The first method involves mixing a sufficiently large amount of crystalline solute with a solvent and bringing the prepared solution to a certain temperature. Dissolution of solute occurs continuously until the dynamic equilibrium is reached and the solution becomes saturated. The second method requires the preparation of a nearly saturated solution at the temperature where the solubility is higher, then adding a few crystals of solute and lowering the solution temperature where the solubility is lower (at this temperature the solubility will be determined). Since the quantity of solute in this case is greater than in a normal saturated solution at the given temperature (i.e., the solution is supersaturated), the excess solute crystallizes from the solution until the solution becomes saturated. The concentration of the saturated solution is referred to as the solubility of the solute in the given solvent at the given temperature.^[5]

To determine the solubility of CaL_2 in water, a method of measuring concentration must be chosen. For CaL_2 , generally the concentration of calcium ions is measured. A variety of methods for calcium determination has been employed, including (but not limited to) atomic absorption spectroscopy (AAS), complexometric (EDTA) titration, oxalic method, and the conductivity method. The AAS appears to be the most accurate method, but it involves complicated sample preparation and it is expensive. The AOAC standard method for the calcium determination (method 968.31) utilizes EDTA titration and this method is preferable to use.

FACTORS AFFECTING SOLUBILITY OF CALCIUM LACTATE

Effect of Temperature

In all the reported data, CaL_2 has been shown to increase in solubility with an increase in temperature. One of the earliest references for CaL_2 solubility was



reported in 1912 by Hill and Cocking, which was later published in “Solubilities, Inorganic and Metalorganic Compounds; Compilation of Solubility Data from the Periodical Literature” by Linke.^[11] Another common source for solubility data is “Solubilities of Inorganic and Organic Compounds” by Stephen and Stephen,^[12] which also provides CaL_2 solubility as a function of temperature. However, because the original source for this solubility data was cited from Seidell’s “Solubility of Inorganic, Metalorganic and Organic Compounds” and the numbers are identical (Table 1), we can assume that it is the same data. Neither of these two sources supplies any information as to which methods for measuring solubility were used to obtain the data. Table 1 shows the solubility of CaL_2 .

Pearce et al.^[3] published results on the solubility of CaL_2 measured by using two different methods to obtain solubility data: shaking solids of CaL_2 with water (method 1) and by precipitation from supersaturated solutions of CaL_2 by cooling (method 2). After 5 days of mixing, the two experiments gave different results (Table 2) indicating that the system had not reached equilibrium; however, both sets of data showed that the solubility of CaL_2 increased with increasing temperature.

Almost at the same time, a research group in Russia^[13] obtained a complete set of data for the solubility of CaL_2 in the temperature range between 12 and 80°C in 5°C intervals (Table 3). In this study, CaL_2 was mixed with water for 0.25–5 hr in a flask with three openings (for thermometer and mixing device) and the flask was immersed in a water bath with temperature control. To determine the amount of calcium in the solution, an oxalic method was used. The procedure for this method was not described in the article, but the usual procedure involves the precipitation of calcium ions in the presence of oxalate anions to form calcium oxalate, which is mixed with the acid and then titrated.^[14]

Additionally, Pateenko and Smirnov^[13] conducted some experiments by precipitating CaL_2 crystals to equilibrium from a supersaturated solution by cooling. The same values for solubility were obtained; however, a longer time was required to achieve equilibrium (data were not shown). Therefore, they concluded that the method of obtaining saturated solutions does not influence the solubility data.

Table 1. Effect of Temperature on the Solubility of CaL_2 ^[12]

Temperature (°C)	Solubility of CaL_2 , g/100 g Solution		
	Pentahydrate Basis		
	Hill and Cocking	Linke	Anhydrous Basis
0	3.01	3.01	2.13
15	5.12	5.12	3.63
30	7.32	7.32	5.19



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Table 2. Variation in Solubility of CaL_2 with Temperature for Aqueous Solution^[3]

Temperature (°C)	Solubility of CaL (Method 1 ^a)		Solubility of CaL (Method 2 ^a)		
	Pentahydrate Basis		Anhydrous Basis		Anhydrous Basis g/100 g Solution
	mol/L	g/100 g Solution	mol/L	g/100 g Solution	
0	0.104	3.20	0.112	3.45	2.45
4.0	0.125	3.85	0.136	4.21	2.98
9.3	0.138	4.27	0.154	4.74	3.36
16.7	0.160	4.94	0.170	5.24	3.71

Data is estimated from the graph.

^a Method 1: mixing crystals of CaL_2 with water and dissolving to equilibrium. Method 2: precipitation of CaL_2 crystals from supersaturated solution by cooling.

*Table 3.* Effect of Temperature on the Solubility of CaL_2 ^[13]

Temperature (°C)	Solubility of CaL_2 (g Anhydrous CaL_2 /100 g Solution) (Experimental)	Solubility of CaL_2 (g Anhydrous CaL_2 /100 g Solution) (Calculated)
12	2.90	2.96
15	3.15	3.27
20	4.08	3.85
25	4.50	4.54
30	5.19	5.31
35	6.54	6.30
40	8.51	8.39
45	9.77	9.80
50	11.32	11.73
55	14.25	14.34
60	17.94	17.89
65	23.20	22.70
70	29.20	29.23
75	38.00	38.10
80	49.10	50.10

Pateenko and Smirnov^[13] also proposed a model that can be used to calculate CaL_2 solubility at different temperatures. Because of computer limitations at that time, they had to use two different formulas for the temperature ranges 12–35 and 40–80°C. The formula for the first interval was:

$$y = 2.001 \times 1.102^T$$

and for the second interval:

$$y = 0.3442e^{0.0611T} + 4.418$$

where y is solubility at the given temperature T . The agreement between experimental data and calculated solubilities is shown in Table 3.

Xie et al.^[15] also used the dissolution technique (method 1) and the oxalic method to determine the amount of CaL_2 present in a saturated solution. Diluted sulfuric acid was used to dissolve precipitated calcium oxalate and potassium permanganate used for the titration. The results are given in Table 4.

In these experiments it was found that the solubility curve had a transformation (break) point at 41°C. Xie et al.^[15] suggested that the break point in the solubility curve was due to the different number of waters of hydration in crystal lattice of CaL_2 , but gave no further details. This study also proposed formulas for the calculation of CaL solubility at different temperatures. Below 41°C, the solubility formula was given as:

$$C = 4.0824 - 1.4087 \times 10^{-3}T + 3.2514 \times 10^{-3}T^2$$



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Table 4. Effect of Temperature on the Solubility of CaL_2 ^[15]

Temperature (°C)	Solubility of CaL_2 (g Anhydrous $\text{CaL}/100$ g Solution) (Experimental)	Solubility of CaL_2 (g Anhydrous $\text{CaL}/100$ g Solution) (Calculated)
22.3	5.67	5.67
25.0	6.07	6.08
30.0	6.99	6.97
35.0	8.00	8.02
39.0	8.96	8.97
40.3	9.32	9.31
41.0	9.76	9.49
41.5	9.74	9.56
42.8	9.81	9.85
43.8	9.88	10.09
44.6	10.05	10.29
45.0	10.17	10.39
50.4	12.10	11.98
55.2	13.81	13.69
60.0	15.66	15.69
65.0	18.21	18.08
70.0	20.66	20.78

and above 41°C, the solubility was described as:

$$C = 11.3690 - 0.3025T + 6.2411 \times 10^{-3}T^2$$

where C is solubility at the given temperature T .

The most recently published data for the CaL_2 solubility is that of Cao et al.^[16] In this study, the solubility of L(+)- CaL_2 and DL(\pm)- CaL_2 (the mixture of L(+)- and D(-)- CaL isomers) in water was determined in the temperature range between 5 and 80°C in 5°C intervals. Salts were added to the solutions until they formed precipitates and then the solutions were stirred for 2–4 hr until equilibrium was reached. After equilibrium had been achieved, a small amount of supernatant was taken out, diluted, and the conductivity of each sample was measured. Their results are given in Table 5.

Cao et al.^[16] proposed models for the determination of CaL solubility at different temperatures.

L(+)- CaL :

$$S = 4.53 - 0.37T + 0.02T^2 - 4.24 \times 10^{-4}T^3 + 3.31 \times 10^{-6}T^4$$

DL(\pm)- CaL :

$$S = 1.93 + 0.45T - 0.019T^2 + 3.51 \times 10^{-4}T^3 - 1.28 \times 10^{-6}T^4$$

**Table 5.** Effect of Temperature on the Solubility of CaL_2 ^[16]

Temperature (°C)	Solubility of L(+)- CaL_2 (g Anhydrous CaL_2 /100 g Solution)		Solubility of DL(±)- CaL_2 (g Anhydrous CaL_2 /100 g Solution)	
	Experimental (from Graph)	Calculated	Experimental (from Graph)	Calculated
5	4.20	3.75	2.90	3.13
10	4.50	4.87	3.00	2.44
15	5.00	5.52	3.00	2.22
20	6.00	5.93	3.40	3.07
25	6.25	6.29	3.60	3.32
30	8.00	6.77	4.00	3.72
40	10.00	8.72	6.20	3.07
45	10.70	10.44	8.40	3.32
50	13.60	12.81	9.80	3.72
55	16.20	15.89	11.00	4.43
60	21.80	19.76	14.10	5.64
65	26.90	24.45	17.30	7.62
70	32.00	29.99	—	10.67
75	39.00	36.38	26.30	15.14
80	46.00	43.61	36.40	21.42

where S is solubility at the given temperature T . Significant deviations between experimental and predicted solubilities were observed, as seen in Table 5.

The data of Cao et al.^[16] shows that the solubility of L(+)- CaL_2 is higher than that of DL(±)- CaL_2 throughout the broad temperature range and therefore, these results support the hypothesis that the L form is more soluble than the DL form of CaL_2 .

Chan and Hartel (unpublished data) also measured CaL_2 solubility in water. The L(+)- CaL_2 was dissolved in water, and when equilibrium was achieved (3–21 days), the EDTA titration method was used to determine the amount of CaL_2 in the solution. These results are given in Table 6.

Table 6. Effect of Temperature on the Solubility of CaL_2 (Chan and Hartel, Unpublished)

Temperature (°C)	Solubility of L(+)- CaL_2		
	Pentahydrate Basis		Anhydrous Basis g/100 g Solution
	g/100 g Water	g/100 g Solution	
4	3.60	3.47	2.46
10	4.32	4.14	2.93
24	6.57	6.16	4.36



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Effect of pH

The effect of pH on CaL_2 solubility was experimentally studied in cheese by Pearce et al.^[3] An increase in lactose concentration in cheese milk resulted in a lower pH (bacteria fermented lactose to the lactic acid) and a decrease in CaL_2 solubility. Furthermore, they indicated that lactose was more important than calcium in determining cheese pH. Cheesemaking experiments demonstrated that dilution of the cheese milk with water would reduce the lactose content and therefore, would increase the pH of the finished cheese and reduce CaL_2 precipitation (i.e., increase CaL_2 solubility).

The only data showing the effect of pH on the solubility of CaL_2 is available from Chan and Hartel (unpublished data). These data (Table 7) show the trend of increasing solubility of CaL_2 with increasing pH. These experiments were performed by adding lactic acid to water and adjusting the pH to a certain level. Then, CaL_2 crystals were added to this solvent and when the equilibrium of the solution was reached, the solubility of the CaL_2 was measured. From Table 7, it can be seen that at 10°C the solubility of CaL is lower at a lower pH. This can be explained using Le Châtelier's principle. The addition of lactic acid led to the increase of lactate ions. In order to compensate for this increase, the equilibrium shifted towards the formation of solid CaL_2 so that less CaL_2 was able to dissolve at the equilibrium state.

Uncommon Ion Effect

Very little data has been published on the effects of other ions on CaL_2 solubility despite the importance of controlling CaL_2 crystallization on cheese. Pearce et al.^[3] briefly mentioned the influence of sodium chloride on the solubility of CaL_2 . Sodium chloride in concentrations varying from 0 to 2.0 M was added to the CaL_2 solution at 19°C. The effect was insignificant and the results are given in Table 8.

Table 7. Effect of pH on Solubility of CaL_2 in Water at 10°C (Chan and Hartel, Unpublished)

pH	Solubility of L(+)- CaL_2		
	Pentahydrate Basis		Anhydrous Basis g/100 g Solution
	g/100 g Water	g/100 g Solution	
5.00	2.72	2.65	1.88
5.12	3.41	3.30	2.34
6.00	4.08	3.92	2.78
7.00 (no lactic acid added)	4.32	4.14	2.93

**Table 8.** Effect of Salt on the Solubility of CaL at 19°C ^[3]

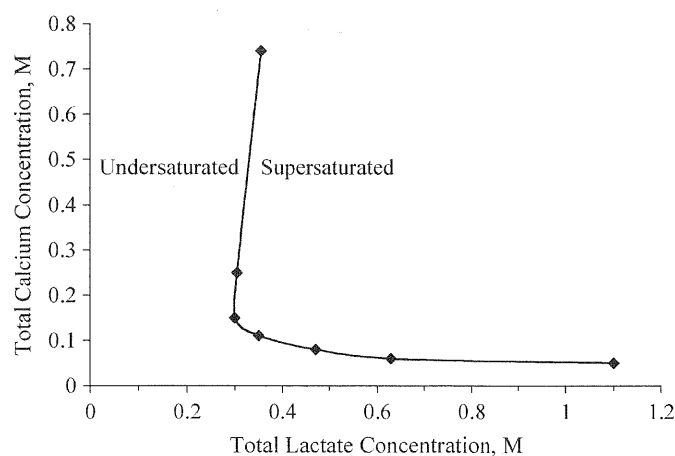
Concentration of NaCl (<i>M</i>)	Solubility of CaL_2		
	Pentahydrate Basis		Anhydrous Basis g/100 g Solution
	mol/L	g/100 g Solution	
0	0.165	5.10	3.61
0.5	0.18	5.56	3.94
2.0	0.17	5.25	3.72

Effects of Calcium and Lactate Ions

Pearce et al.^[3] studied the effect of varying calcium and lactate content on the solubility of CaL_2 at 18°C (Fig. 1). The concentration of lactate should be higher than 0.3 M to maintain a supersaturated solution. The positive curve at high calcium concentrations was explained by the formation of CaL^+ ions. The effects of calcium or lactate ions on the solubility of CaL_2 can be seen in Fig. 1; however, it is not clear how the calcium and lactate ions were added without causing other effects (e.g., uncommon ion effects).

DISCUSSION

All the available data for CaL_2 solubility provides evidence that an increase in temperature results in an increase in CaL_2 solubility. However, values for solubility vary in the different studies. A summary of all the experimental data on CaL_2 solubility is given in Table 9.

**Figure 1.** Solubility curve for CaL_2 in aqueous solution at 18°C .^[3]



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Table 9. Summary of Experimental for Solubilities of CaL_2

Temperature (°C)	Solubility of CaL_2 , g Anhydrous $\text{CaL}/100$ g Solution							
	Hill and Cocking, Linke	Pearce		Pateenko Method 1	Xie Method 1	Chan Method 1 $\text{L}(+)\text{-CaL}_2$	Cao	
		Method 1	Method 2				Method 1 $\text{L}(+)\text{-CaL}_2$	Method 1 $\text{DL}(\pm)\text{-CaL}_2$
0	2.13	2.27	2.45					
4		2.72	2.98			2.46		
5							4.20	2.90
9.3		3.02	3.36					
10.0						2.93	4.50	3.00
12				2.90				
15.0	3.63			3.15			5.00	3.00
16.7		3.50	3.71					
20.0				4.08			6.00	3.40
22.3					5.67			
24.0						4.36		
25.0				4.50	6.07		6.25	3.60
30.0	5.19			5.19	6.99		8.00	4.00
35.0				6.54	8.00			
39.0					8.96			
40.0				8.51			10.00	6.20
40.3					9.32			
41.0					9.76			
41.5					9.74			
42.8					9.81			
43.8					9.88			
44.6					10.05			
45.0				9.77	10.17		10.70	8.40
50.0				11.32			13.60	9.80
50.4					12.10			
55.0				14.25			16.20	11.00
55.2					13.81			
60.0				17.94	15.66		21.80	14.10
65.0				23.20	18.21		26.90	17.30
70.0				29.20	20.66		32.00	
75.0				38.00			39.00	26.30
80.0				49.10			46.00	36.40

Method 1: mixing crystals of CaL with water and dissolving to equilibrium. Method 2: precipitation of CaL crystals from supersaturated solution by cooling.

There are several potential causes of variability in CaL_2 solubility results. The initial form of CaL_2 used for the experiments may affect solubility. The CaL_2 may occur in either isomeric form (L or D) or as a mixture of the two isomers. On the basis of the data of Cao et al.,^[16] the use of the racemic mixture for dissolution experiments leads to lower solubility of CaL_2 . In addition, the impurities that are present in the initial CaL_2 material can affect the solubility and possibly create interferences during sample analyses. Since there is little detailed information on the starting materials in the reviewed studies (especially the earlier work), we can only speculate about possible differences in solubility due to the starting materials. Another possible cause of variability in literature results is the approach to equilibrium. Methods 1 and 2 give different solubility results if the solution has not reached equilibrium and, in the reviewed studies, there is no evidence that the



system was at equilibrium at the time of sampling. In some studies, the mixing time for equilibration appears to be very short (15 min) and it is difficult to believe that the solution had sufficient time to achieve equilibrium, especially at the lower temperatures. Dissolution during filtration may also have had an effect on the measured solubility value of CaL_2 in some of the previous studies. Rapid dissolution of CaL_2 crystals may occur if filtration temperature is higher than equilibration temperature, as may have been the case when low equilibration temperatures were used but the slurry was filtered at room temperature. This may have led to erroneous results. Methods of sample analyses were also different among the studies and may have contributed to the variability observed in Table 9.

In all of the studies, except Cao et al.,^[16] CaL_2 pentahydrate was used as the starting material, but no further specifications (e.g., purity) were available. Cao et al.^[16] used both $\text{L}(+)\text{-CaL}_2$ dihydrate and $\text{DL}(\pm)\text{-CaL}_2$ pentahydrate as starting material, with $\text{DL}(\pm)\text{-CaL}_2$ giving lower solubility than $\text{L}(+)\text{-CaL}_2$. Solubility data for all other experiments appear to be close to the average of these two values at the given temperature. Solubility data provided by Xie et al.^[15] were higher than other data at lower temperatures and lower at the higher temperatures (above 55°C).

CONCLUSION

Several researchers have studied the factors affecting the solubility of CaL_2 . Substantial discrepancies are observed among different studies, although certain trends were consistently noted. For example, an increase in temperature leads to an increase in CaL_2 solubility. The pH effect is similar to the temperature effect where an increase in pH causes an increase in CaL_2 solubility. However, there was no well-defined relationship between the presence of other salts and CaL_2 solubility. Further studies under carefully controlled conditions are needed to accurately determine the solubility of CaL_2 . This information is critical to control of CaL_2 crystal formation on the surface of Cheddar cheese during storage.

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