

# **The Apparent Solubility of Aluminum(III) in Hanford High-Level Waste**

Prepared for the U.S. Department of Energy  
Assistant Secretary for Environmental Management

Contractor for the U.S. Department of Energy  
Office of River Protection under Contract DE-AC27-08RV14800

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*By Janis D. Aardal at 12:53 pm, Jun 20, 2012*

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11

12       **ABSTRACT**  
13

14       The solubility of aluminum in Hanford nuclear waste impacts on the processability of the waste  
15       by a number of proposed treatment options. For many years, Hanford staff has anecdotally noted  
16       that aluminum appears to be considerably more soluble in Hanford waste than the simpler  
17       electrolyte solutions used as analogues. There has been minimal scientific study to confirm these  
18       anecdotal observations, however. The present study determines the apparent solubility product  
19       for gibbsite in 50 tank samples. The ratio of hydroxide to aluminum in the liquid phase for the  
20       samples is calculated and plotted as a function of total sodium molarity. Total sodium molarity is  
21       used as a surrogate for ionic strength, because the relative ratios of mono, di and trivalent anions  
22       are not available for all of the samples. These results were compared to the simple

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28 NaOH-NaAl(OH)<sub>4</sub>-H<sub>2</sub>O system, and the NaOH-NaAl(OH)<sub>4</sub>-NaCl-H<sub>2</sub>O system data retrieved  
29 from the literature. The results show that gibbsite is apparently more soluble in the samples than  
30 in the simple systems whenever the sodium molarity is greater than two. This apparent enhanced  
31 solubility cannot be explained solely by differences in ionic strength. The change in solubility  
32 with ionic strength in simple systems is small compared to the difference between aluminum  
33 solubility in Hanford waste and the simple systems. The reason for the apparent enhanced  
34 solubility is unknown, but could include kinetic or thermodynamic factors that are not present in  
35 the simple electrolyte systems. Any kinetic explanation would have to explain why the samples  
36 are always supersaturated whenever the sodium molarity is above two. Real waste  
37 characterization data should not be used to validate thermodynamic solubility models until it can  
38 be confirmed that the apparent enhanced gibbsite solubility is a thermodynamic effect and not a  
39 kinetic effect.

40

41 **Keywords:** Hanford, nuclear waste, aluminum solubility, gibbsite.

42

## 43 INTRODUCTION

44

45 Aluminum(III) is one of the most prevalent constituents in high-level nuclear waste stored in  
46 177 tanks at the Hanford site in Richland, Washington. Given the large quantities of Al(III) in  
47 the waste, the solubility of Al(III) affects the performance of most waste treatment options.  
48 Indeed, several recent independent review groups noted that improved Al(III) solubility models  
49 were required to optimize tank waste cleanup activities.<sup>[1-4]</sup> The most direct impact of Al(III)  
50 solubility on waste separation processes is through caustic leaching of gibbsite [Al(OH)<sub>3</sub>] from

51 waste sludge.<sup>[5]</sup> Additionally, Al(III) precipitation can foul ion-exchange columns, evaporators,  
52 and filters. Any model of Al(III) solubility will eventually have to be validated against real waste  
53 behavior. A similar approach has been undertaken to validate Al(III) solubility models for the  
54 Bayer process,<sup>[6]</sup> which is a high-pH, high-ionic strength system like Hanford waste. An obvious  
55 way to start developing an improved model is to look at the apparent solubility of Al(III)  
56 exhibited by the waste itself. That is the purpose of the present study.

57

58 The dominant aluminum species in the aqueous phase of caustic Hanford high-level waste is  
59 believed to be the aluminate anion  $[\text{Al}(\text{OH})_4^-]$ , although the aluminate dimer is also present.<sup>[7, 8]</sup>  
60 Gibbsite and boehmite are assumed to be the most prevalent aluminum-bearing solids in Hanford  
61 tank waste.<sup>[9]</sup> Dawsonite  $[\text{NaAl}(\text{OH})_2\text{CO}_3]$  and aluminosilicates have recently been identified in  
62 a few tanks as well.<sup>[10, 11]</sup> Boehmite has extremely slow dissolution and precipitation kinetics at  
63 current tank temperatures,<sup>[12, 13]</sup> so boehmite is assumed to have a negligible impact on the liquid  
64 phase aluminum concentration. Therefore, gibbsite precipitation has historically been assumed to  
65 limit the liquid phase aluminum concentration of Hanford nuclear waste.

66

67 In addition to Al(III), Hanford waste has large quantities of sodium, nitrite, nitrate, carbonate,  
68 hydroxide, sulfate, phosphate, iron, manganese, uranium, zirconium, bismuth, many organics,  
69 and many other less prevalent constituents.<sup>[14]</sup> These components may affect the solubility of  
70 Al(III) in the tank waste. For instance, the soluble species would be expected to affect the ionic  
71 strength, which, in turn, has a major impact on the solubility of practically every other  
72 species.<sup>[15]</sup> Sodium, being the dominant cation in the liquid phase, typically has a concentration  
73 between one and 12 molar.

75 Given the large concentration of dissolved hydroxide in waste, the solubility of gibbsite in  
76 Hanford waste has traditionally been compared to gibbsite solubility in simple aqueous solution  
77 of sodium hydroxide. There is a large amount of gibbsite solubility data available in these simple  
78 sodium hydroxide solutions.<sup>[16]</sup> In 1976, Barney<sup>[17]</sup> discovered that the apparent solubility of  
79 gibbsite was as much as ten times higher in simulated Hanford waste supernatant solutions than  
80 in simple  $\text{NaOH}_{(\text{aq})}$  of the same hydroxide concentration. The reason for this difference has been  
81 hotly debated ever since.<sup>[18]</sup> Barney only studied simulants with sodium concentrations greater  
82 than 10 M Na, and did not study any real waste samples. Reynolds<sup>[19]</sup> determined that the  
83 solubility of gibbsite in Barney's study could not be modeled accurately without invoking new  
84 liquid phase aluminum species. Like Barney,<sup>[17]</sup> Reynolds<sup>[19]</sup> addressed waste simulants rather  
85 than real waste samples. Therefore, the apparent solubility of real waste needs to be evaluated, to  
86 determine if real waste also exhibits this apparent enhanced solubility. Since the time of Barney,  
87 Hanford laboratory staff has periodically noted dissolved aluminum concentrations that  
88 "seemed" elevated for real waste, but there has been minimal in-depth study of the enhanced  
89 solubility.

91 A more recent modeling study by Agnew<sup>[20]</sup> proposed that Al(III) concentrations in tank  
92 samples were more consistent with the solubility of dawsonite than gibbsite. Dawsonite is more  
93 soluble than gibbsite. Thus, Agnew<sup>[20]</sup> noted that the apparent solubility of real tank waste is  
94 considerably higher than simple  $\text{NaOH}_{(\text{aq})}$  solutions in equilibrium with gibbsite. These results  
95 are consistent with the simulant work performed by Barney<sup>[17, 18]</sup> more than 30 years ago. There  
96 were some limitations with Agnew's<sup>[20]</sup> work, however. For instance, he often determined the

97 dissolved hydroxide concentration in waste samples by charge balance. A recent study has  
98 determined that charge balance calculations on Hanford waste are frequently biased, and require  
99 a special methodology for reconciliation.<sup>[21]</sup> Charge balance calculations are also affected by  
100 species with unknown charge, such as “total organic carbon” measurements. Agnew<sup>[20]</sup>  
101 converted analytical data from mass per volume units to molality units using water contents  
102 calculated by the method of Reynolds and Carter.<sup>[22]</sup> That method depends heavily on the  
103 accuracy of the concentration of the constituents, including the hydroxide concentrations. Thus,  
104 errors in the hydroxide concentrations measured by charge balance are magnified molality units  
105 through the density conversion equation. The limitations with Agnew’s<sup>[20]</sup> study has left many in  
106 the scientific community skeptical of his results.<sup>[2-3]</sup> The purpose of the present study is to  
107 verify Agnew’s<sup>[20]</sup> conclusion that Al(III) is much more soluble in Hanford tank waste than in  
108 simple NaOH<sub>(aq)</sub> solutions of similar hydroxide concentration in equilibrium with gibbsite. This  
109 study will determine the apparent solubility product for gibbsite in tank waste as indicated by  
110 waste sample data. The term “apparent” is used because confirmation of equilibrium with  
111 gibbsite is not determinable from the present data (see discussion section).

112

## 113 **METHODOLOGY**

114

115 Taking a statistically designed approach of sampling Hanford tanks is infeasible because of the  
116 high cost of taking samples of highly radioactive tank waste. Thus, this study uses existing  
117 sample data. The Hanford waste tanks have been sampled and analyzed for many different  
118 reasons and these data are used here opportunistically. The Hanford site maintains the Tank  
119 Waste Information Network System (TWINS) database that keeps all of the analytical data on

120 tank waste generated since year 1992. For this study, TWINS database was queried to find all  
121 samples where the sodium, aluminum, and hydroxide concentrations were measured on the same  
122 liquid samples. Exactly fifty data points in TWINS met these criteria, and these results are shown  
123 in Table 1. Aluminum was measured by Inductively Coupled Plasma Atomic Emission  
124 Spectroscopy in all cases, and was assumed to be in the form of the aluminate ion.

125

126 Most analytical data on nuclear waste liquids are reported by the laboratory in units of mass per  
127 volume (here converted to moles per liter). Ideally, a comparison of different Al(III) solubility  
128 datasets would be performed in units of molality (moles per Kg of water), because molality is  
129 independent of the concentration of other electrolytes in the mixture.<sup>[15, 23]</sup> Densities and water  
130 contents are required for the conversion from molarity to molality, however, and are not  
131 available for most of the data evaluated here. Consequently, volume units (moles/L) are used  
132 here to maximize the amount of data in the comparison. In order to overcome the limitation of  
133 using molarity units, hydroxide to Al(III) ratios were compared rather than concentrations.

134 Ratios are independent of concentration scale,<sup>[23]</sup> so comparisons of ratios are not degraded by  
135 using molarity units.

136

137 The apparent solubility product for aluminum will be expressed relative to the mineral gibbsite  
138  $[\text{Al}(\text{OH})_3]$ , because gibbsite is the most prevalent solid identified in Hanford waste.<sup>[9]</sup> The term  
139 “apparent” is used because confirmation of equilibrium with gibbsite is not determinable from  
140 the present data (see discussion section). The reaction between gibbsite and dissolved hydroxide  
141 to create aluminate is written as Equation 1.

142



144

145 From this reaction stoichiometry, the dissolved hydroxide to aluminate ratio defines the inverse  
146 of the apparent equilibrium constant (here denoted as 1/Q to distinguish it from a true  
147 equilibrium constant that is usually denoted K). This ratio also defines the number of hydroxide  
148 moles required to keep one mole of Al(III) soluble, a quantity of interest to engineers trying to  
149 prevent the precipitation of Al(III) during filtration and ion-exchange. The lower the 1/Q for a  
150 sample, the higher the apparent solubility of gibbsite exhibited. Table 1 provides this ratio for  
151 sample data taken from TWINS. The concentrations in Table 1 represent the means of duplicate  
152 analyses on each sample, or the ratio of the means. These results are compared to the solubility  
153 of gibbsite in simple  $NaOH_{(aq)}$  and  $NaCl_{(aq)}$ -  $NaOH_{(aq)}$  liquids, as described below.

154

155 The OH:Al ratio at equilibrium with gibbsite is influenced by solution ionic strength.<sup>[16, 24]</sup> The  
156 ionic strength range of Hanford waste varies widely. In order to account for ionic strength, 1/Q is  
157 plotted against the sodium molarity in Figure 1 of the tank waste samples, with sodium molarity  
158 being used as a surrogate for ionic strength. The data from Li et al.<sup>[25]</sup> at 25 °C are also plotted  
159 on Figure 1, as a dataset that has only  $NaOH_{(aq)}$  and no other electrolytes. The  $NaOH$ -  
160  $NaAl(OH)_4$ - $NaCl$ - $H_2O$  system data presented by Wesolowski<sup>[16]</sup> represents a slightly more  
161 complicated system, with an additional electrolyte influencing ionic strength besides  $NaOH$  and  
162  $NaAl(OH)_4$ . Wesolowski<sup>[16]</sup> presented his data in molality units, so his sodium concentrations  
163 are not directly comparable to the molarity units used on the X-axis in Figure 1. Thus, his data  
164 was converted to molarity units using densities calculated by the method of Reynolds and  
165 Carter.<sup>[22, 26]</sup>

166

167 **RESULTS AND DISCUSSION**

168

169 One/Q (one over the apparent solubility product for gibbsite) is plotted in Figure 1 as a function  
170 of sodium molarity for multi-electrolyte Hanford tank waste, along with the simple NaOH-  
171 NaAl(OH)<sub>4</sub>-H<sub>2</sub>O<sup>[25]</sup> and NaOH-NaAl(OH)<sub>4</sub>-NaCl-H<sub>2</sub>O systems.<sup>[16]</sup> The value for 1/Q is the  
172 number of moles of hydroxide that are in solution per mole of Al(III) dissolved. When these  
173 solutions are at equilibrium at standard state, 1/Q = 1/K, where K = the equilibrium constant for  
174 the reaction in Equation 1. Figure 1 indicates that the apparent gibbsite solubility is much higher  
175 in the overwhelming majority of the tank waste sample in this study than in the simpler NaOH-  
176 NaAl(OH)<sub>4</sub>-H<sub>2</sub>O and NaOH-NaAl(OH)<sub>4</sub>-NaCl-H<sub>2</sub>O systems in equilibrium with gibbsite. This is  
177 evidenced by the fact that most of the hydroxide to Al(III) ratios for Hanford waste are lower  
178 than those given by the simple systems at a given sodium molarity. The only exceptions are at  
179 sodium molarities below two. These results are consistent with Agnew's<sup>[20]</sup> conclusion that  
180 gibbsite appears more soluble in tank waste than in simple systems. This conclusion is also  
181 consistent with the original simulant data reported by Barney<sup>[17, 18]</sup> more than 30 years ago.  
182 Barney<sup>[17]</sup> only studied highly concentrated solutions (10.5 to 19 M). From the current results,  
183 however, we can discern that Barney's<sup>[17]</sup> conclusions extend down to the lower concentrations.

184

185 Several aspects of Figure 1 must be discussed in detail. The simple system data plotted in  
186 Figure 1 were collected at 25 °C.<sup>[16, 25]</sup> Temperature monitoring data in the Hanford's TWINS  
187 database indicate that the tank temperatures ranged from 18 to 45 °C, with all but three data  
188 points below 35 °C. Wesolowski<sup>[16]</sup> collected data at both 25 and 50 °C, which bound the upper

189 end of the tank temperature range, though most of the tank samples are closer to 25 °C than  
190 50 °C. The hot cell that the samples are stored in prior to analyses (for sometimes up to several  
191 months) is roughly 28 °C. Some of the Hanford data points do lay above the 50 °C  
192 Wesolowski<sup>[16]</sup> data on Figure 1, indicating that high tank temperature could potentially account  
193 for high solubility in some samples. Nonetheless, the majority of the data lays below the 50 °C  
194 Wesolowski<sup>[16]</sup> data in Figure 1, indicating that temperature is unlikely to explain the majority of  
195 the difference between the tank waste data and the simple system data.

196

197 In Figure 1, sodium molarity was used on the X-axis as an indicator of ionic strength. In the  
198 simple system data,<sup>[16, 25]</sup> all of the ions were monovalent, so the ionic strength would be  
199 identical at a given sodium molarity (assuming no ion-pairing). Real tank waste samples have  
200 some divalent and trivalent anions (sulfate, carbonate, phosphate, and some other trace ions), so  
201 the ionic strength will be slightly larger in the real tank waste samples at a given sodium  
202 molarity. The (relatively) low solubility of these divalent and trivalent salts limits their  
203 concentrations at high sodium concentrations.<sup>[27-29]</sup> The difference between 1/Q for the simple  
204 systems and real waste was much larger than the difference between the simple systems as a  
205 function of ionic strength over nearly the entire sodium molarity range. Between one and seven  
206 molar sodium concentration, the OH:Al ratio only changed about 10% in the Li et al.<sup>[25]</sup> data,  
207 but differed by as much as an order of magnitude from the tank waste samples. Thus, the  
208 difference between real waste and the simple systems is not solely due to an ionic strength  
209 difference at a given sodium molarity (at least between two and seven molar). The Li et al.<sup>[25]</sup>  
210 data are actually trending closer to the real waste data at higher sodium molarities (Fig.1). From

211 the present data, it cannot be determined if the simple system data would trend onto the tank  
212 waste data at higher sodium concentrations.

213

214 One interesting note is that the OH:Al ratios as a function of sodium molarity were nearly  
215 identical in the Li et al.<sup>[25]</sup> and Wesolowski<sup>[16]</sup> studies, even though Wesolowski<sup>[16]</sup> altered the  
216 sodium molarity by adding NaCl. The maximum OH<sup>-</sup> molarity in the Wesolowski<sup>[16]</sup> data is  
217 0.1 M, with the rest of the sodium being made up by NaCl. In contrast, OH<sup>-</sup> is the most prevalent  
218 Na-charge balancing anion in the Li et al.<sup>[25]</sup> data, with hydroxide concentrations greater than  
219 six molar in some cases. Thus, the effect of monovalent sodium-bearing electrolytes on the  
220 OH/Al ratio is roughly the same regardless of the anion identity.

221

222 Now that it has been established that the apparent aluminum solubility exhibited by tank waste is  
223 much higher than in the simple system reported in the literature,<sup>[16, 25]</sup> some explanation is  
224 required. Reynolds and Reynolds<sup>[18]</sup> evaluated Barney's<sup>[17]</sup> simulant data and concluded that  
225 there are at least two contributing factors to the high solubility observed by Barney. One of those  
226 is that the aluminate dimer and the sodium aluminate ion-pair are more prevalent species in the  
227 liquid phase Barney<sup>[17]</sup> simulants than the simple system data. They concluded this because  
228 Barney's simulants had much lower water content than a simple sodium hydroxide solution of  
229 the same free hydroxide concentration, which would favor the formation of these species. It is  
230 not clear if this is a major driver in the present study, because the apparent enhanced aluminum  
231 solubility was observed even at relatively low sodium molarities (two to four molar Na). In  
232 contrast, Barney studied 10.5 to 19 M Na solutions. The aluminate dimer and the sodium-

233 aluminate ion-pair may be more important for the tank waste samples with the highest sodium  
234 molarities.

235

236 Reynolds and Reynolds <sup>[18]</sup> also concluded that the Barney <sup>[17]</sup> data was affected by slow kinetics.  
237 Barney equilibrated his samples only from the supersaturation direction, and only for seven days.  
238 Reynolds and Reynolds <sup>[18]</sup> concluded that Barney's <sup>[17]</sup> studies likely had not reached a true  
239 equilibrium point, because seven days was a short time relative to the equilibration times  
240 required by other studies in the literature. <sup>[30, 31]</sup> Slow kinetics may affect the current study as  
241 well. The time available for equilibration varied widely for the 50 data points used in the present  
242 study, ranging from 21 days to 23 years (Table 1). Given that we cannot confirm that all, or any,  
243 of the characterization data used in the present study had come to equilibrium, we have chosen to  
244 denote 1/Q as the "apparent" equilibrium ratio rather than the true equilibrium value.

245

246 One problem with the assumption that slow kinetics account for all of the apparent  
247 supersaturation in Figure 1 is that all of the samples above two molar sodium concentrations  
248 appear supersaturated. Given the large number of tanks evaluated in this study, it is surprising  
249 that they are all supersaturated and none under-saturated when the sodium molarities are above  
250 two molar. The effect of tank history on gibbsite supersaturation could not be determined, but  
251 there is nothing in the tank history that would suggest that equilibrium would have always come  
252 from the supersaturation direction. Thus, it would be surprising if there were no samples below  
253 or near saturation if kinetics were the sole reason for the elevated aluminum concentrations  
254 observed. Each tank sample has its own unique history. Some of the samples were created by  
255 evaporating supernatants, some by blending supernatants; some were created straight from the

256 process facilities. Most samples are a result of some combination of these three processes.

257 A more detailed description of tank waste history can be found in Agnew. [32]

258

259 Agnew [20] believed that the apparent supersaturations in tank waste with respect to gibbsite  
260 results from the wastes are equilibrated with solid-phase dawsonite instead of gibbsite. This  
261 hypothesis does not explain, however, why the liquids are still supersaturated with gibbsite.  
262 Dawsonite is more soluble than gibbsite under most conditions. If the samples are supersaturated  
263 with gibbsite, then gibbsite should still precipitate until equilibrium is reached with respect to  
264 gibbsite as well as dawsonite. Gibbsite has been identified in many Hanford tanks, [9] so it is  
265 reasonable to conclude that gibbsite can precipitate from Hanford waste liquids. A possible  
266 explanation of the apparent supersaturation of gibbsite is the simultaneous slow precipitation of  
267 gibbsite coupled with the slow dissolution of a less stable phase. If the tanks have a less stable  
268 phase, like dawsonite, than this less stable phase would slowly dissolve and re-precipitate the  
269 more stable phase. If the dissolution of the less stable phase was faster than the precipitation of  
270 the more stable phase, the solution would always be supersaturated with the more stable phase,  
271 until the less stable phase has been completely consumed. The liquid would be poised at the  
272 saturation concentration of the less stable phase until all of the less stable phase was consumed.  
273 This, if true in the case of Hanford tank waste, would explain Agnew's [20] conclusion that many  
274 tanks appear to be near the dawsonite saturation point. The present study cannot confirm this  
275 hypothesis. Dawsonite has been observed in some Hanford tank waste samples [11] but it is not  
276 clear if dawsonite is present in enough tank samples to account for the wide-spread elevated  
277 aluminum concentrations observed.

278

279 Until it is possible to determine if the large differences in solubility observed are due to kinetic  
280 or thermodynamic factors, it is unreasonable to use real tank waste data to validate  
281 thermodynamic-based solubility models. This is because a system that is not at equilibrium  
282 would be inconsistent with an equilibrium model. Similarly, it may not be appropriate to use a  
283 thermodynamic model to predict aluminum solubility if the waste never comes to equilibrium  
284 over the time frame of interest. Thus, determining the reason for the elevated aluminum  
285 concentrations is important for determining the appropriate way to model aluminum distribution  
286 during waste processing.

287

## 288 CONCLUSION

289

290 This study has concluded that Al(III) appears to be more soluble in Hanford waste than would  
291 have been expected based on the ionic strength and measured hydroxide concentrations. This  
292 conclusion has been determined by comparing real tank waste data to simple system data that has  
293 frequently been used as waste analogues. Both thermodynamic and kinetic factors have been  
294 discussed to explain this phenomenon, but no conclusions could be made conclusively in this  
295 regard with the present data. If real waste samples are to be used to validate Al(III) models, then  
296 determining if thermodynamic or kinetic factors are driving the elevated Al(III) concentrations in  
297 the tanks sample is of paramount importance. Testing an equilibrium model against data that is  
298 not actually at equilibrium would be inappropriate. Furthermore, an equilibrium model may not  
299 be an appropriate model to use if the kinetics is so slow that equilibrium is never realistically  
300 reached.

301

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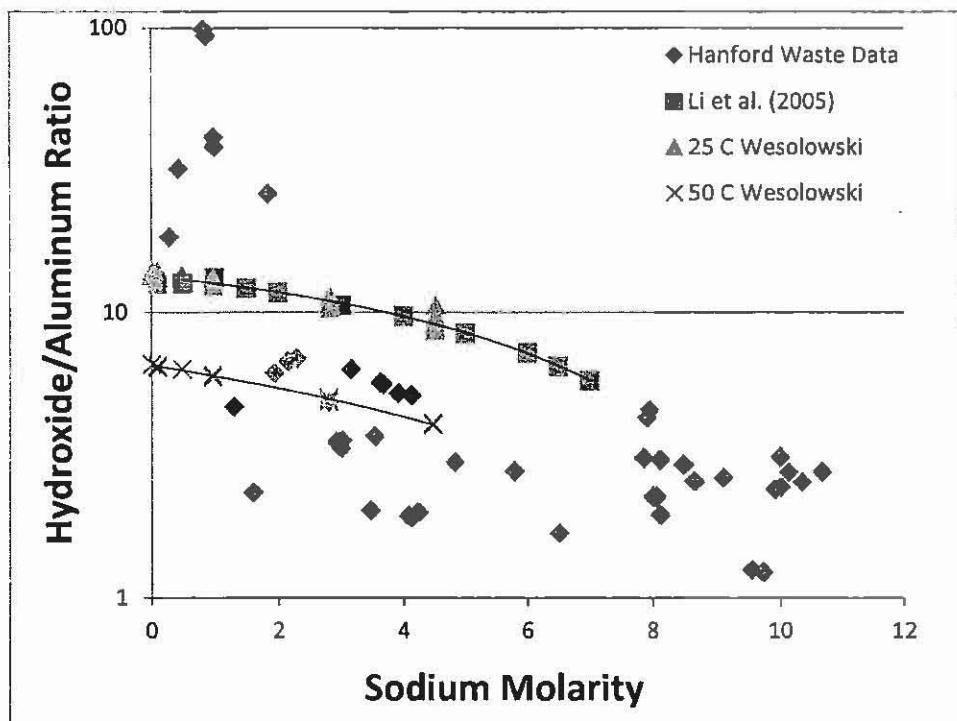
402

403 **FIGURE CAPTIONS**

404  
405 **Figure 1.** OH:Al Mole ratio as a function of sodium Molarity for tank waste and simple  
406 systems.

407

408



410 Fig.1

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 412  
 413 **Table 1.** List of samples investigated in Figure 1, ordered from lowest to highest sodium  
 414 molarity

Sodium Molarity	OH:Al Mole Ratio (1/Q)	Time Before Last Change to Tank and Sample
0.29	18.30	2 years
0.44	31.84	2 years
0.83	98.67	3 years
0.87	93.24	3 years
0.99	41.18	11 years
1.00	37.94	11 years
1.00	41.07	11 years
1.31	4.66	21 days
1.60	2.34	6 months
1.84	25.94	1 year

Sodium Molarity	OH:Al Mole Ratio (1/Q)	Time Before Last Change to Tank and Sample
1.94	6.12	21 days
2.16	6.72	21 days
2.30	6.89	21 days
2.81	4.84	21 days
2.92	3.50	3 months
2.96	3.43	3 months
3.00	3.34	3 months
3.02	3.55	3 months
3.16	6.31	6 months
3.46	2.02	3 months
3.54	3.66	2 years
3.62	5.68	6 months
3.67	5.59	6 months
3.92	5.21	6 months
4.07	1.94	5 years
4.13	1.91	5 years
4.13	5.12	6 months
4.23	1.99	5 years
4.83	2.98	3 months
5.79	2.77	4 years
6.51	1.68	17 years
7.85	3.08	23 years
7.91	4.27	3 months
7.96	4.54	3 months
8.01	2.26	3 years, 3 months
8.06	2.23	3 years, 3 months
8.07	2.27	3 years, 3 months
8.12	3.04	23 years
8.12	1.94	1 year

Sodium Molarity	OH:Al Mole Ratio (1/Q)	Time Before Last Change to Tank and Sample
8.49	2.92	10 months
8.65	2.55	10 months
9.11	2.63	10 months
9.55	1.25	17 years
9.73	1.23	17 years
9.91	2.40	10 months
10.00	3.11	16 years
10.01	2.44	16 years
10.14	2.75	16 years
10.36	2.54	16 years
10.68	2.76	15 years

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