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**Adsorption of Aqueous Silicate on Hematite**

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ABSTRACT

During radioisotope sorption studies, adsorption of silicate from synthetic groundwaters by synthetic hematite was observed. To further investigate this observation, the adsorption of silicate onto hematite ( $\alpha$ -Fe<sub>2</sub>O<sub>3</sub>) powder from a neutral, aqueous NaCl solution (0.1 mol/dm<sup>3</sup>), containing  $2.56 \times 10^{-4}$  mol/dm<sup>3</sup> of Si added as Na<sub>2</sub>SiO<sub>3</sub>·9H<sub>2</sub>O, was measured at ~21°C. Equilibrium adsorption of silicate amounted to ~1.93  $\mu$ mol/m<sup>2</sup> (one Si(O,OH)<sub>4</sub> moiety per 86 Å<sup>2</sup>). It is important to take this adsorption into account when evaluating the ability of iron oxides to adsorb other species, especially anions, from groundwaters. Silicate adsorption is known to diminish the ability of iron oxides to adsorb other anions.

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RÉSUMÉ

Au cours des études de sorption des radio-isotopes, on a observé l'adsorption du silicate des eaux souterraines synthétiques par l'hématite synthétique. Pour corroborer cette observation, on a mesuré l'adsorption du silicate sur la poudre d'hématite ( $\alpha\text{-Fe}_2\text{O}_3$ ) d'une solution de NaCl aqueuse neutre ( $0,1 \text{ mol/dm}^3$ ), contenant  $2,56 \times 10^{-4} \text{ mol/dm}^3$  de Si ajouté sous la forme de  $\text{Na}_2 \text{SiO}_3 \cdot 9\text{H}_2\text{O}$  à  $\sim 21 \text{ }^\circ\text{C}$ . L'adsorption à l'équilibre du silicate s'est chiffrée à  $\sim 1,93 \text{ } \mu\text{mol/m}^2$  (une fraction de  $\text{Si}(\text{O},\text{OH})_4$  par  $86 \text{ \AA}^2$ ). Il est important de prendre en compte cette adsorption quand on évalue la capacité des oxydes de fer à adsorber d'autres espèces, spécialement les anions, des eaux souterraines. On sait que l'adsorption de silicate diminue la capacité des oxydes de fer à adsorber d'autres anions.

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## 1. INTRODUCTION

Fine-grained iron oxide minerals have a high capacity to adsorb both cationic and anionic species from aqueous solution, and can thus immobilize a variety of soil nutrients and environmental contaminants (Parfitt 1978). For example, iron oxide minerals may help to retard the migration of radionuclides from a nuclear fuel waste disposal vault to the biosphere (Davison et al. 1994, Vandergraaf et al. 1993, Vandergraaf and Ticknor 1994). In addition, iron minerals such as magnetite and biotite are expected to be important in the restoration of reducing conditions in a waste vault (Johnson et al. 1994; Taylor and Owen 1993, 1995, 1997; Jobe et al. 1997).

Dissolved silicon (as  $\text{Si(OH)}_4$  and  $\text{OSi(OH)}_3^-$ ) occurs at concentrations around  $2 \times 10^{-4} \text{ mol/dm}^3$  in deep granitic groundwaters (Gascoyne 1988); therefore, it is important to understand how it may affect both the sorption properties and the redox chemistry of iron oxides in a waste disposal system. For example, adsorbed silicate might account for the relatively low affinity for iodate adsorption by a natural goethite ( $\alpha\text{-FeOOH}$ ), compared with a pure, synthetic hematite ( $\alpha\text{-Fe}_2\text{O}_3$ ), as reported by Ticknor and Cho (1990). The literature on adsorption of silicate on iron oxides, and its influence on their sorption properties and transformation reactions, has been reviewed elsewhere (Taylor 1995). Here, we report measurements of the adsorption of dissolved silicate from simulated groundwaters and dilute NaCl solutions onto synthetic hematite powder. In the experiments with simulated groundwaters, the sorption of silicate by synthetic hematite was observed during conditioning of the solid and solution in preparation for sorption experiments with radioisotopes of nickel and uranium (Ticknor 1994a,b). Some of the data illustrating silicate uptake by synthetic hematite have, therefore, been documented previously but were not discussed at length because these data were of secondary interest at the time. Hereafter, we use the term "silicate" broadly to include silicic acid ( $\text{Si(OH)}_4$ ) and its conjugate anions in both aqueous and adsorbed forms.

## 2. EXPERIMENTAL

Two types of experiments are documented here. Firstly, silicate uptake was observed on synthetic hematite during the conditioning of solids and solutions preparatory to performing sorption studies with radioisotopes of other elements (Ticknor 1994a,b). The data collected on silicate sorption by hematite in these experiments is discussed. Secondly, saturation sorption experiments were conducted to determine the sorptive capacity of synthetic hematite for silicate in neutral NaCl solutions. These studies are described in detail.

### 2.1 MATERIALS

Reagent-grade  $\text{Fe}_2\text{O}_3$  powder was obtained from Fisher Scientific. Powder X-ray diffraction analysis confirmed that it was hematite. The specific surface area, determined by the BET method ( $\text{N}_2$  adsorption), was  $11.8 \pm 0.7 \text{ m}^2/\text{g}$ . Chemical analysis by X-ray fluorescence confirmed that the solid was 98.8%  $\text{Fe}_2\text{O}_3$ , with 0.04 wt.% MnO as the only metallic impurity. The weight loss on ignition at  $1000^\circ\text{C}$  was 0.6%.

Gascoyne (1988) described a representative groundwater for the Whiteshell Research Area near Pinawa, Manitoba and provided a methodology for preparing a synthetic variant for use in laboratory experiments. This synthetic groundwater has a total dissolved solids content of about  $10\,000 \text{ mg/dm}^3$  and has been used in sorption experiments at Whiteshell Laboratories over a number of years. In the sorption experiments with U (Ticknor 1994a), four variants of this solution were used, two prepared to full strength and two dilute variants. The data obtained for

Si sorption onto synthetic hematite from the two full-strength variants are discussed here. In the sorption experiments with Ni (Ticknor 1994b), a number of variants of the reference solution composition were used. The data obtained for Si sorption on synthetic hematite from eight of these solutions is described here.

The solution used in the saturation sorption experiments was  $0.1 \text{ mol/dm}^3$  NaCl, with  $\sim 7.2 \text{ mg/dm}^3$  ( $2.56 \times 10^{-4} \text{ mol/dm}^3$ ) of Si added as  $\text{Na}_2\text{SiO}_3 \cdot 9\text{H}_2\text{O}$ . This silicon concentration is towards the high end of the range of values reported for saline groundwater samples from boreholes in plutonic rocks, most of which lie between 1 and  $7 \text{ mg/dm}^3$  (Gascoyne et al. 1987). It is well below the solubility limit of amorphous silica (about  $2 \times 10^{-3} \text{ mol/dm}^3$  in dilute saline solutions at  $25^\circ\text{C}$ , Marshall 1980), and close to the solubility of quartz. The salinity of this solution is comparable to moderately deep ( $\sim 500 \text{ m}$ ) groundwaters, but the salt composition of groundwaters is much more complex (e.g., Table 3 of Gascoyne et al. 1987).

The NaCl solutions were made nominally neutral by adjustment with dilute HCl following the addition of the  $\text{Na}_2\text{SiO}_3 \cdot 9\text{H}_2\text{O}$ . Twelve pH measurements on blank solutions, made over the course of the experiments, varied between 6.5 and 7.9, with an average of 7.0. Ten pH measurements on the solutions from the experiments with  $V:m = 120:1$  (see below) varied between 6.5 and 7.0, with an average of 6.8.

## 2.2 PROCEDURE

### 2.2.1 Conditioning Experiments

Preparatory to conducting sorption experiments with radioisotopes, the aqueous and solid phases are often placed in contact with each other to allow the main solution constituents to approach equilibrium prior to the introduction of the radioisotope(s) of interest. In two such studies (Ticknor 1994a,b), it was observed that the Si content of the aqueous phase was substantially lowered following contact with synthetic hematite. In these experiments, the hematite was given several successive washes with fresh aliquots of the initial synthetic groundwater in the ratio of 1 g of hematite to  $10 \text{ cm}^3$  of synthetic groundwater. At the end of each wash, the solution was carefully removed by pipette and the aliquot filtered through a  $0.40 \mu\text{m}$  polycarbonate filter prior to analysis for major cations and anions.

### 2.2.2 Saturation Sorption Experiments

Experiments were run at four different initial ratios of solution volume to solid mass ( $V \text{ (cm}^3\text{)} : m \text{ (g)}$ ). These were 20:1 ( $20 \text{ cm}^3$  of solution in contact with 1 g of solid), 60:1 ( $18 \text{ cm}^3$  and 0.3 g), 120:1 ( $120 \text{ cm}^3$  and 1 g), and 240:1 ( $240 \text{ cm}^3$  and 1 g). The reaction vessels were polypropylene bottles, varying in capacity from 22 to  $250 \text{ cm}^3$ . All experiments were run in triplicate. The reaction vessels were agitated manually each working day to ensure contact between the solids and solutions. Solutions were sampled for elemental analysis between 8 and 13 times over a period of 80 days; the sampling schedules for each  $V:m$  value are indicated in the tables and figures..

Two different sampling procedures were used. For  $V:m = 20:1$ , 60:1 and 120:1, the solution was decanted and a portion removed for analysis. Each aliquot was filtered through  $0.40 \mu\text{m}$  polycarbonate filters prior to analysis. Sampled solutions, including blanks, were analysed for Na and K by atomic absorption spectrophotometry (AA), and for Ca, Mg, Si, Al, and Fe by

inductively coupled plasma spectrophotometry (ICPS); some solutions were also analysed for chloride and sulfate by ion chromatography (IC).

The small mass of solution still in contact with the solid was weighed to determine the residual amount of contacting solution retained after each contact. An appropriate aliquot of the initial contacting solution was then added to the solid for the continuing experiment. For  $V:m = 240:1$ , solution aliquots of  $15 \text{ cm}^3$  were removed at each sampling interval, but the solution was not replenished. Thus, a total of  $120 \text{ cm}^3$  of solution was removed for analysis over the course of the experiment, and  $V:m$  varied accordingly. This procedure was used because a large solution volume would otherwise be needed for replenishing. Also, the results from this experiment helped to distinguish between slow, continuous adsorption of silicate from solution and renewed adsorption prompted by replenishment of the solution.

### 3. RESULTS AND DISCUSSION

#### 3.1 CONDITIONING EXPERIMENTS

Data illustrating the removal of Si from synthetic groundwaters (SGWs) by hematite are given in Table 1. More complete analyses of SGW-A and -B are provided in Ticknor (1994a). These two SGWs were based on WN-1M and were of near identical composition with the exception that SGW-B was prepared under an  $\text{N}_2$  atmosphere and without the addition of carbonate. More complete analyses of SGW-1 through -8 are provided in Ticknor (1994b). These SGWs were also based on WN-1M and differed in their Na, Ca, Mg and carbonate concentrations. The data illustrate that substantial amounts of Si were removed from each solution by the hematite. Where available, the data also illustrate that, in general, the amount of Si removed from each solution decreased slightly with successive washes. Saturation of the hematite surface with Si was not achieved in these experiments; therefore a further series of experiments (described in Section 3.2) was conducted to measure saturation sorption.

The solutions prepared in the absence of a source of carbonate (SGW-B and SGW-1 through -4) showed slightly greater losses of Si from solution compared to like solutions containing carbonate. This implies that there may be some weak competition by carbonate for surface sites occupied by silicate. Several authors have reported evidence for the adsorption of  $\text{CO}_2$ , bicarbonate, and carbonate anions on hydrous iron oxide surfaces, which tends to support this suggestion (Harrison and Berkheiser 1982, Rochester and Topham 1979, Zeltner and Anderson 1988).

#### 3.2 SATURATION SORPTION EXPERIMENTS

##### 3.2.1 Solution Analyses

All silicon analyses are shown in Table 2 and Figure 1, and they are discussed in detail in Section 3.2. Sodium analyses varied between 2270 and 2420  $\text{mg/dm}^3$ , compared with the expected value of 2299  $\text{mg/dm}^3$  ( $0.1 \text{ mol/dm}^3$ ). Chloride analyses were also essentially as expected: 3400 to 4200  $\text{mg/dm}^3$ , compared with an expected value of 3545  $\text{mg/dm}^3$ . Calcium, magnesium and sulfate analyses are given in Table 3. They indicate that these three species were present as impurities, and that they were leached from the surface of the iron oxide in the early stages of each run. Otherwise, no significant trends were detected in impurity levels in solution. Potassium analyses varied between 0.13 and 0.46  $\text{mg/dm}^3$ , with no obvious trends other than



slightly higher values in the second batch of solution (0.23 to 0.45 mg/dm<sup>3</sup>) than the first (0.13 to 0.18 mg/dm<sup>3</sup>). Aluminum and iron analyses were all below the detection limits of 0.1 mg/dm<sup>3</sup> and 0.01 mg/dm<sup>3</sup>, respectively.

### 3.2.2 Estimation of Sorption Capacity

The solution analyses for silicon (Table 2) were used to calculate the quantity of silicate adsorbed on the hematite surface after each sampling interval. In those cases where the solution was replenished after sampling (i.e.,  $V:m = 20:1$ ,  $60:1$  and  $120:1$ ), the initial silicon concentration was corrected for the residual solution that remained after decanting (see Section 2.2). The calculated quantities of adsorbed silicate are presented in Table 4 and Figure 2.

The silicon analyses in Figure 1 indicate that, for the two intermediate  $V:m$  values ( $60:1$  and  $120:1$ ), silicate adsorption occurred quite rapidly over the first 10-20 days, then more slowly up to 40 days, after which the solution analyses are only marginally different from the blanks, indicating that silicate adsorption was close to equilibrium with the starting solution. Adsorption was more prolonged at the lowest  $V:m$  value ( $20:1$ ), presumably because the smaller solution volume became depleted more rapidly than the larger volumes. In the experimental runs at the highest  $V:m$  value (initially  $240:1$ ), in which the solution was not replaced after sampling, silicate adsorption proceeded slowly throughout the 74-day experiment; the 30-day data point appears to be in error (Figure 1).

The calculated quantities of adsorbed silicate (Table 4 and Figure 2) show asymptotic approach to a value of about 0.64 mg Si/g Fe<sub>2</sub>O<sub>3</sub>, which is equivalent to 1.93 μmol/m<sup>2</sup> or one Si(O,OH)<sub>4</sub> moiety per 86 Å<sup>2</sup> (1 Å = 0.1 nm). This appears to represent an equilibrium surface concentration in contact with the starting silicate concentration of ~7.2 mg/dm<sup>3</sup> ( $2.56 \times 10^{-4}$  mol/dm<sup>3</sup>) at pH 7.

### 3.3 COMPARISON WITH LITERATURE DATA

The adsorption of silicate on iron oxides and oxyhydroxides has received much less attention than adsorption of the two isoelectronic and isostructural species, phosphate and sulfate, as indicated by the space accorded the three species in the review by Parfitt (1978). Silicate adsorption is known to reach a maximum at about pH 9; also, it is diminished by competition with phosphate, which appears to occupy similar surface sites (Parfitt 1978; Taylor 1995 and references therein). Thus, it is useful to compare our results with literature information on phosphate as well as silicate.

Our measured surface concentration of 1.93 μmol/m<sup>2</sup> of silicate on hematite, in equilibrium with  $2.56 \times 10^{-4}$  mol/dm<sup>3</sup> of dissolved silicate at pH 7, is similar to various reported values for the sorption capacities of iron oxides for phosphate and silicate. Colombo et al. (1994) noted that natural and synthetic goethites have a “relatively constant” phosphate adsorption capacity that approaches the theoretical value of 2.51 μmol/m<sup>2</sup> for the binuclear phosphate complex on the (110) crystallographic face, which usually makes up 90% of the goethite crystal surface. This surface complex has been described by Russell et al. (1974) and Parfitt et al. (1975), among others.

Hingston et al. (1972) reported a peak capacity of about  $2.7 \mu\text{mol}/\text{m}^2$  for silicate adsorption on goethite at pH 9.5, declining gradually to about  $2.3 \mu\text{mol}/\text{m}^2$  at pH 7. Borggaard (1983) estimated the sorption capacity of various iron oxides and oxyhydroxides for phosphate, based on fitting experimental data to a Langmuir isotherm. With one exception, there was little variation among the various solid phases; capacities ranged from 1.8 to  $3.1 \mu\text{mol}/\text{m}^2$  at  $25^\circ\text{C}$ . The one exception was akaganéite ( $\beta\text{-FeOOH}$ ), which had an unusually high capacity of  $6.2 \mu\text{mol}/\text{m}^2$ , attributed to its microporous crystal structure (Borggaard 1983).

Colombo et al. (1994) conducted an exhaustive study of phosphate adsorption on thirty synthetic hematite specimens with different particle size and morphology. They used solutions of  $\text{KH}_2\text{PO}_4$  in  $0.1 \text{ mol}/\text{dm}^3$  KCl, adjusted to pH 6. They found that, in general, their data fit a Freundlich isotherm better than a Langmuir isotherm, both for fast (1 day) and slow (124 days) adsorption processes. They attributed this to the oxide having a heterogeneous sorbing surface, that is, different surface sites with varying affinity for phosphate. At an aqueous phosphate concentration of about  $1.3 \times 10^{-4} \text{ mol}/\text{dm}^3$ , they found that the phosphate capacity varied from 0.83 to  $3.12 \mu\text{mol}/\text{m}^2$ , with a mean value for the 30 hematite specimens of  $1.79 \mu\text{mol}/\text{m}^2$ . They concluded that "In comparison with goethite... hematite has more variability in phosphate-adsorption capacity, adsorbs on average less phosphate per unit surface area, has less affinity for phosphate, and exhibits a more important slow adsorption component".

Hansen et al. (1994a, 1994b) also noted the importance of slow adsorption of silicate on ferrihydrite (amorphous or marginally crystalline " $\text{Fe}(\text{OH})_3$ "), which necessitated at least 14 days of contact with aqueous silicate solutions to approach equilibrium.

### 3.4 SORPTION KINETICS

The sorption experiments were not designed to provide kinetic data, but rather were intended to provide information on saturation sorption capacity. However, the data (Tables 2 and 4) do give some indication of the rate of silicate adsorption on hematite. The rate varies with both surface coverage and aqueous silica concentration, and contact times of a month or more are required to approach equilibrium. At silica concentrations near  $7 \text{ mg}/\text{dm}^3$ , the sorption rates vary from  $\sim 10^{-13} \text{ mol}/\text{m}^2 \cdot \text{s}$  at 90% coverage to  $\sim 10^{-12} \text{ mol}/\text{m}^2 \cdot \text{s}$  at 50% coverage and  $> 5 \times 10^{-12} \text{ mol}/\text{m}^2 \cdot \text{s}$  at <25% coverage. These estimates are intended only to help in the design of possible future experiments to measure sorption rates accurately. The values for percentage coverage are based on the asymptotic limit of about  $0.64 \text{ mg Si}/\text{g Fe}_2\text{O}_3$ , or  $1.93 \mu\text{mol}/\text{m}^2$ , obtained in the present study.

### 3.5 SIGNIFICANCE AND IMPLICATIONS

Our results indicate that, in neutral solutions with silicate concentrations representative of granitic groundwaters, hematite (and probably other iron oxides) will adsorb silicate to a level approaching its sorption capacity. This implies that silicate adsorption may affect the ability of iron oxide minerals to adsorb groundwater contaminants such as radionuclides, that is, it may hinder the co-sorption of other ionic species. A previous literature review showed that silicate adsorption can reduce the affinity of iron oxide surfaces for other anions, such as chromate and selenite (Anderson and Benjamin 1985, Taylor 1995, Zachara et al. 1987). Adsorption of silicate, itself, is diminished by competition with phosphate (Taylor 1995 and references therein).

Unpublished, preliminary data obtained by Taylor and co-workers indicated that silicate adsorption can also diminish the adsorption of iodate by hematite. In general, the "adsorption edge" for anion adsorption by iron oxides appears to be lowered by 1 to 2 pH units when adsorbed silicate is present. This is consistent with the reported shift in the pH of zero charge of goethite from ~8.3 to ~6.8, when silicate is adsorbed on the surface (Hingston et al. 1972). Similar shifts for ferrihydrite, from ~8 to as low as 4, have been reported (Anderson and Benjamin 1985, Schwertmann and Fechter 1982). In contrast, cation adsorption properties of iron oxides appear to be virtually unaffected by co-sorption of silicate (Anderson and Benjamin 1985, 1990).

#### 4. CONCLUSIONS

Equilibrium adsorption of silicate by hematite in neutral aqueous solutions amounts to about  $1.93 \mu\text{mol}/\text{m}^2$  (one  $\text{Si}(\text{O},\text{OH})_4$  moiety per  $86 \text{ \AA}^2$ ). This is comparable to other reported values for silicate and phosphate adsorption on a variety of iron oxides and oxyhydroxides. It is important to take this adsorption into account when evaluating the ability of iron oxides to adsorb other species, especially anions. Silicate adsorption is known to diminish the ability of iron oxides to adsorb other anions; there appears to be minimal effect on cation adsorption. The presence of adsorbed silicate may explain some differences in sorption behaviour between natural and synthetic iron oxides (Ticknor and Cho 1990). Silicate adsorption is a relatively slow process, which may further complicate the interpretation of data on the adsorption properties of iron oxides in mixed-anion systems.

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\* Internal report available from SDDO, AECL, Chalk River Laboratories, Chalk River, Ontario K0J 1J0

TABLE 1

OBSERVED SILICA SORPTION DURING CONDITIONING OF SYNTHETIC  
GROUNDWATERS (SGW) WITH SYNTHETIC HEMATITE (Ticknor 1994a,b)

SGW <sup>1</sup>	Initial [Si] mg/dm <sup>3</sup>	[Si] mg/dm <sup>3</sup> After 7 d contact	[Si] mg/dm <sup>3</sup> After 14 d contact	[Si] mg/dm <sup>3</sup> After 28 d contact	[Si] mg/dm <sup>3</sup> After 56 d contact
A	7.26	0.54	1.15	1.74	2.26
B	6.90	0.16	0.47	0.58	1.00
1	3.57	nd	nd	0.06	nd
2	3.70	nd	nd	0.15	nd
3	3.70	nd	nd	0.14	nd
4	3.76	nd	nd	0.19	nd
5	3.53	0.08	0.14	0.24	nd
6	3.61	nd	0.24	0.34	nd
7	3.36	0.08	0.24	0.22	nd
8	3.64	0.10	0.30	0.32	nd

<sup>1</sup> SGWs A and B from Ticknor (1994a); SGWs 1-8 from Ticknor (1994b)

nd = not determined

TABLE 2  
SOLUTION ANALYSES FOR SILICON

TIME (Days)	SILICON ANALYSES (mg/dm <sup>3</sup> )					
	Blank NaCl-1	Blank NaCl-2	SOLUTION (cm <sup>3</sup> ) : SOLID (g)			
			20:1	60:1	120:1	240:1
0	7.31					
1						5.8
3	7.32		0.968	2.76	4.38	5.57
7	7.15		2.02	4.67	6.02	5.33
10	7.28		3.39	6.06	6.81	5.18
14	7.15		4.07	6.44	6.91	
17			4.87	6.7		
24	7.17		4.59	6.47		
30						3.5
31	7.18		5.14	6.51	6.56	
38	7.11	6.99	5.46	6.76		
45		6.96	5.8	6.78	6.84	
46						4.69
52		7.33	6.23	7.2	7.24	
59		7.29	6.27	7.11	7.13	
60						4.46
66		7.19	6.36	6.96	7.12	
74						4.34
80		7.14	6.23	6.9	7.02	
Average Blank		7.18				

**TABLE 3**  
**SOLUTION ANALYSES FOR CALCIUM, MAGNESIUM AND SULFATE**

Days	Sample	Ca	Mg	SO4	Days	Sample	Ca	Mg	SO4
0	NaCl-1	0.163	0.029	nd	3	60:1	2.69	1.29	nd
3	NaCl-1	0.144	0.007	<0.1	7	60:1	0.333	0.093	nd
7	NaCl-1	0.141	<0.005	<0.1	10	60:1	0.183	0.022	nd
10	NaCl-1	0.156	0.007	<0.1	14	60:1	0.176	0.013	nd
14	NaCl-1	0.139	0.007	<0.1	17	60:1	0.16	0.01	nd
24	NaCl-1	0.14	0.006	nd	24	60:1	0.17	0.012	nd
31	NaCl-1	0.138	<0.005	<0.1	31	60:1	0.159	<0.005	nd
38	NaCl-1	0.159	<0.005	0.18	38	60:1	0.167	<0.005	nd
					45	60:1	0.184	<0.005	nd
					52	60:1	0.189	0.011	nd
					59	60:1	0.168	0.01	nd
					66	60:1	0.195	0.017	nd
38	NaCl-2	0.152	<0.005	<0.1	80	60:1	0.173	0.011	nd
45	NaCl-2	0.15	<0.005	<0.2					
52	NaCl-2	0.137	0.008	<0.2	3	120:1	1.3	0.627	7.4
59	NaCl-2	0.128	0.008	<0.2	7	120:1	0.17	0.027	0.7
66	NaCl-2	0.129	0.008	<0.2	10	120:1	0.321	0.014	<0.1
80	NaCl-2	0.131	0.008	<0.2	14	120:1	0.138	0.009	<0.1
					17	120:1	nd	nd	nd
					24	120:1	nd	nd	nd
					31	120:1	0.142	0.01	0.29
					38	120:1	nd	nd	nd
					45	120:1	0.151	<0.005	<0.2
3	20:1	6.74	3.42	nd	52	120:1	0.147	0.009	<0.2
7	20:1	1.15	0.561	nd	59	120:1	0.149	0.009	<0.2
10	20:1	0.351	0.12	nd	66	120:1	0.138	0.009	<0.2
14	20:1	0.204	0.042	nd	80	120:1	0.143	0.01	<0.2
17	20:1	0.148	0.021	nd					
24	20:1	0.171	0.023	nd	1	240:1	0.702	0.293	nd
31	20:1	0.167	0.01	nd	3	240:1	0.812	0.327	nd
38	20:1	0.142	0.008	nd	7	240:1	0.791	0.329	nd
45	20:1	0.157	0.008	nd	10	240:1	0.794	0.329	nd
52	20:1	0.172	0.015	nd	30	240:1	0.788	0.315	nd
59	20:1	0.181	0.015	nd	46	240:1	0.768	0.331	nd
66	20:1	0.17	0.014	nd	60	240:1	0.792	0.332	nd
80	20:1	0.157	0.014	nd	74	240:1	0.761	0.331	nd

nd = not determined; sample designations correspond to the solutions in Table 2.



**TABLE 4**  
**CALCULATED QUANTITIES OF ADSORBED SILICATE**  
**(mg Si per g Fe<sub>2</sub>O<sub>3</sub>)**

Time (Days)	Solution Volume to Solid Mass Ratio (cm <sup>3</sup> /g)			
	20:1	60:1	120:1	240:1
1				0.336
3	0.125	0.266	0.337	0.386
7	0.226	0.413	0.475	0.437
10	0.301	0.478	0.520	0.468
14	0.362	0.519	0.555	
17	0.407	0.548		
24	0.459	0.592		
30				0.574
31	0.499	0.605	0.632	
38	0.532	0.613		
45	0.555	0.616	0.667	
46				0.573
52	0.575	0.634	0.654	
59	0.592	0.635	0.655	
60				0.608
66	0.608	0.639	0.663	
74				0.625
80	0.627	0.644	0.679	

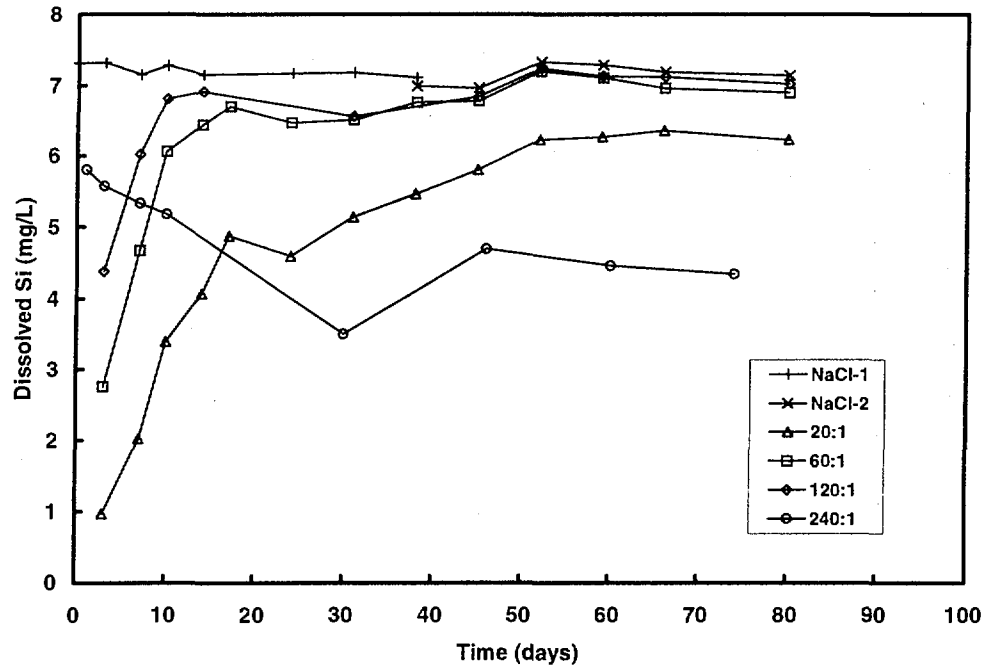


FIGURE 1: Analysed Silicate Concentrations as a Function of Time for Blank Solutions (NaCl-1 and NaCl-2) and for Solutions Contacted with Hematite with Various  $V:m$  Ratios.

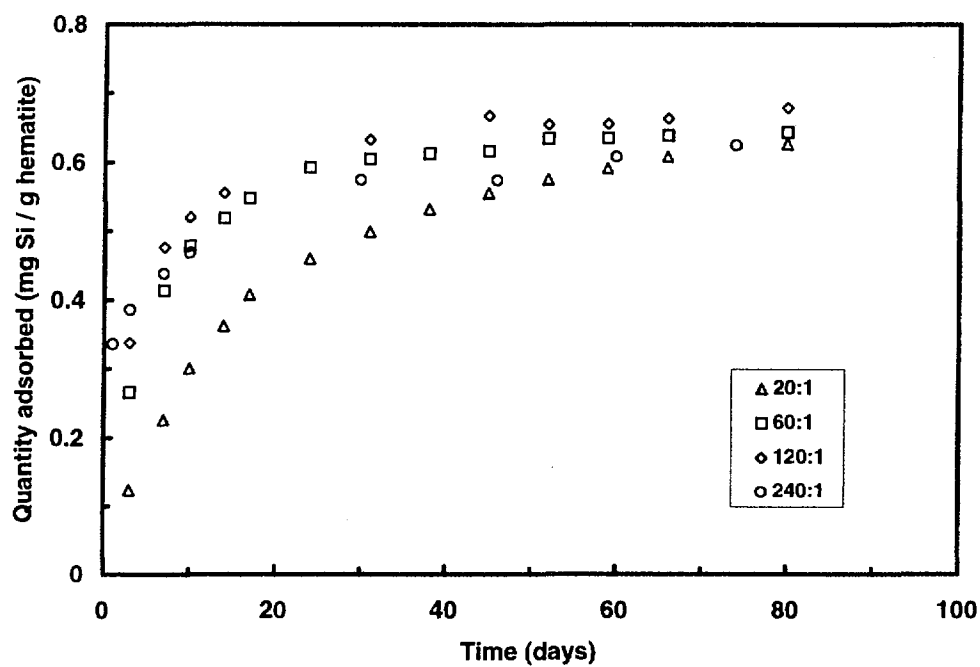


FIGURE 2: Quantity of Silicate (Analyzed as Si) Adsorbed on Hematite Powder as a Function of Time, for Various  $V:m$  Ratios.

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