TREATMENT OF ACID MINE DRAINAGE USING MAGNESIUM HYDROXIDE

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ABSTRACT

Acid drainage from mining activities is of major environmental concern in South Africa. These effluents require neutralisation and metal removal prior to release to public watercourses. A novel process is proposed whereby magnesium hydroxide is being used for neutralization of free acid and subsequent raising of the mine wastewater pH to above 7 to facilitate rapid iron (II) oxidation and precipitation as ferric hydroxide. This is followed by lime treatment for removal of magnesium as $Mg(OH)_2$. By using magnesium hydroxide instead of $Ca(OH)_2$ or $CaCO_3$, gypsum precipitation is avoided and metal hydroxides can be precipitated separately from gypsum. Magnesium hydroxide can be recovered from the magnesium hydroxide /gypsum mixture through dissolution of magnesium hydroxide as magnesium bicarbonate, with carbon dioxide.

This study showed that magnesium hydroxide can be used for treatment of acid mine drainage rich in iron (II)-and magnesium hydroxide recovered from the sludge. Pilot-plant studies are in the planning stage to demonstrate the suitability of the magnesium process for full-scale application in the treatment of acid mine drainage.

1. INTRODUCTION

Acid mine drainage, AMD, from mining activities has become a major environmental concern. It has become a threat not only to the environment but also to human health. AMD is characterized by low pH (2-3), high salinity levels, containing a broad range of heavy metals ions and high concentrations of sulphate, iron, aluminium and manganese. Iron disulphide (FeS_2), commonly known as pyrite, is a major constituent of the strata being mined and large rock surfaces become exposed to air and water during mining activities (Sasowsky *et al.*, 2000). Pyrite is oxidized to soluble iron complexes and sulphuric acid, catalysed by sulphur oxidizing bacteria (Sawyer *et a.l*, 1994).

$$2FeS_2\ (s) + 7O_2\ (aq) + 2H_2O \rightarrow 2Fe^{2+} + 4SO_4^{2-}\ (aq) + 4H^+\ (aq) \qquad (1)$$

Traditionally, lime has been used for neutralization of acid mine water (Herrera *et al*., 2007; Sibrell *et al.*,2005, Watten *et al.*, 2005; Feng *et al.*, 2000; Semerjian *et al.*, 2003). Several new processes have been developed, based on the use of precipitated calcium carbonate or lime pre-treatment for neutralization of acid mine drainage and partial desalination (Maree *et al*., 1992; Maree *et al.*, 1994; Maree *et al*., 1996; Maree *et al.*, 1998; Maree *et al*., 2004; Maree *et al.*, 1994). The Council for Scientific and Industrial Researchers (CSIR) in South Africa has developed the fluidised-bed limestone neutralization process for the treatment of acid mine water (Maree & Clayton; 1991). Their studies showed that complete neutralization of discard leachate, containing 10 000 mg/L acid as $CaCO_3$ $_{and}$ 4 000mg/L Fe(II), can be achieved in a limestone neutralization, fluidized-bed reactor, provided that Fe(II) is oxidized beforehand. The process involves the neutralization of Fe(II)-rich water with lime to obtain iron(II) oxidation. This process can also speed up the settling process of ferric sludge (Semerjian *et al*., 2003).

The integrated limestone and Fe(II)-oxidation process was developed, which allows the oxidation of Fe(II) when limestone alone was used for neutralization (Maree, 1997). In this process powdered limestone was used for Fe(II)-oxidation at pH 5.5, neutralization of free acid, metal precipitation (e.g. Fe^{3+} and Al^{3+}) and gypsum crystallization, all in the same reactor. The novelty of this development lies in the fact that conditions were identified where Fe(II) can be oxidized at pH 5.5, by the addition of limestone. Previously, lime was used to raise the pH to 7.2 where the rate of Fe(II)-oxidation is rapid. A handling and dosing system was developed for using waste limestone containing 25% moisture, from the paper industry which (Maree, 2000)

The integrated limestone and lime process was developed for the treatment of acid and sulphate-rich effluents (Maree, 2003). The process consists of various stages. The bulk of the acid content is neutralized in first stage with limestone. CO_2 is produced and stripped off through aeration and transported to the third stage. In the second stage the water is treated with lime to allow precipitation of magnesium and other metals and the sulphate associated with these metals. The level to which sulphate is removed via gypsum crystallization is controlled by the solubility product of gypsum. In the third stage, the CO_2 that is produced in the first stage, is contacted with the high pH of the water from the second stage to adjust the pH to 8.3. This affords $CaCO_3$ precipitation. Due to its high purity, this $CaCO_3$ can be sold as a by-product or be recycled to the first stage, to supplement the limestone addition. This process offers the following benefits: (i) The treated water is under-saturated with respect to gypsum and, (ii) if the feed water contains aluminium, sulphate removal is not only achieved through gypsum crystallization, but also through ettringite ($3CaO.3CaSO_4.2Al_2O_3$) formation as it precipitates in the pH range 11.3 - 11.4.

The disadvantage of using lime or limestone is that metal hydroxides are precipitated together with gypsum. An attractive, alternative alkali for neutralization of acid mine water is magnesium hydroxide. Due to the high solubility of $MgSO_4$ (260 g/L) only metal hydroxides precipitate (Weast, 1971a). $Mg(OH)_2$ also has the capability of raising the pH to 10 which is sufficient for removal of metals like manganese and zinc that cannot be removed with limestone as the pH cannot be raised sufficiently. There are two mechanisms operating when metal containing acid water is neutralized with $Mg(OH)_2$. The first mechanism is where metals are precipitated at metal hydroxides. The second mechanism is where metals are absorbed onto the surface of $Mg(OH)_2$ particles. Teringo (1987) found that with $Mg(OH)_2$, metals can be removed from solution at one pH unit lower than when NaOH is used. $Mg(OH)_2$ produces a faster settling rate for metal hydroxide flocs as well as a denser sludge (Marshall & St. Armand, 1992), compared to lime and NaOH. This occurs because of the lower solubility of $Mg(OH)_2$, which results in a slower release of the OH-ion into solution, which in turn results in a more gradual increase in pH. The result is the formation of larger metal hydroxide particles and the formation of a more compact or denser sludge.

The purpose of this investigation was to demonstrate that $Mg(OH)_2$ treatment, in combination with lime treatment, offers an attractive solution to treatment of acid mine water that is rich in Fe(II) and other metals (Figure 1). In this approach, Fe(II) is oxidized rapidly and precipitated as $Fe(OH)_3$. The precipitated $Fe(OH)_3$, together with other metals hydroxides, are separated from the water. In the following stage, magnesium is precipitated with lime as $Mg(OH)_2$ and, together with gypsum, separated from the water. The $Mg(OH)_2$ can be separated from the gypsum by treating it with CO_2 to form $Mg(HCO_3)_2$ or with H_2SO_4 to form $MgSO_4$.

The overall objectives of the study were the following:

- Identify optimum conditions required for Fe(II)-oxidation when $Mg(OH)_2$ is used as the alkali (reported here).
- Determine conditions required for magnesium removal with lime (reported here).
- Determine the conditions required for rapid gypsum crystallization in the presence of $Mg(OH)_2$ (to be reported in a subsequent publication).
- Identify ways for the separation of $Mg(OH)_2$ from gypsum (to be reported in a subsequent publication).

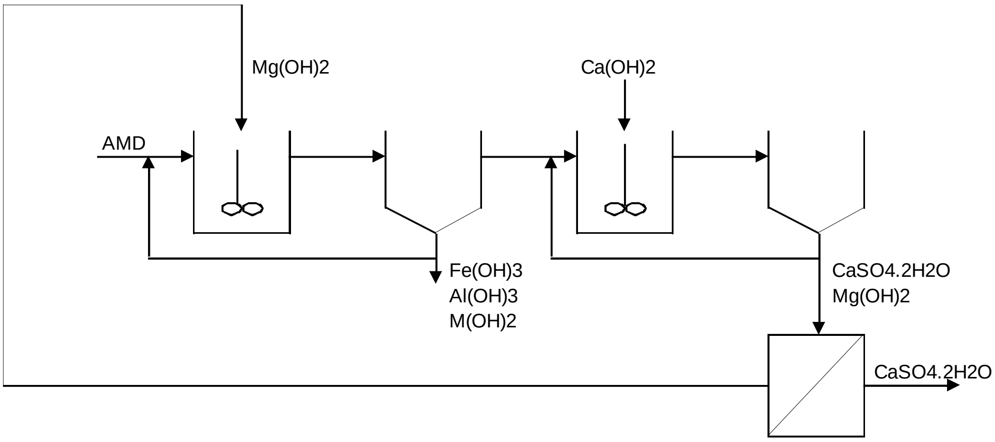

Figure 1. Process flow diagram of the $Mg(OH)_2$/Lime treatment process for acid mine water treatment

2. METHODS AND MATERIALS

Acidic mine water, rich in ferrous sulphate, was collected from a mine in Randfontein. Synthetic AMD containing ferrous sulphate and H_2SO_4 was prepared as aqeous solutions containing 100 - 900 mg/L Fe(II), some free acidity and sulphate. $Mg(OH)_2$ (CP grade) was used for neutralization. $Fe(OH)_3$, produced during the neutralization studies, was used for sludge recirculation. Air (HP compressed, Afrox, South Africa) was used as O_2 source.

Batch studies were carried out using 500 mL glass beakers and magnetic stirrers. As oxygen source, compressed air was bubbled through a diffuser.

For Fe^{2+}-oxidation, batch studies in beaker (500mL) were carried out by mixing reactants (H_2O, H_2SO_4, $FeSO_4$, $Fe(OH)_3$ and $Mg(OH)_2$) at time zero. The contents were stirred continuously at 240 rpm. Samples were taken at regular intervals and analyzed for pH, Fe(II), Mg, and alkalinity. $Fe(OH)_3$ produced during the experiments was kept for use in subsequent experiments. In cases where the $Fe(OH)_3$ was contaminated with $Mg(OH)_2$, due to excess dosages, the $Mg(OH)_2$ was dissolved by adjusting the pH to 6 with H_2SO_4.

The effects of the following parameters on the kinetics of Fe(II) oxidation were determined:

- Fe(II)-concentration (100, 200, 400 and 900 mg/L)
- $Mg(OH)_2$/Fe(II) (mole ratio 0, 1, 2, 4)
- $Fe(OH)_3$/Fe(II) (mole ratio 0, 1, 2, 4)
- O_2 concentration (various feed rates)
- Temperature (14, 25, 35 and 49 ºC)

Batch studies were also carried out to demonstrate:

- The removal of Mg^{2+} from solution with lime. The thin slurry formed was allowed to settle and decanted, whereafter the sludge formed was kept to be used in the following batch.
- Solubilisation of $Mg(OH)_2$ in the $Mg(OH)_2$/Gypsum sludge by contacting with CO_2 to produce soluble $Mg(HCO_3)_2$ at low temperature. The reaction mixture was allowed to settle, whereafter the filtrates were kept and used in the next step for $MgCO_3$ production.
- Precipitation of $MgCO_3$ from a $Mg(HCO_3)_2$-solution by heating. The $Mg(HCO_3)_2$ solution, formed in the previous step, was heated to 75 ^oC to decompose the $Mg(HCO_3)_2$ and allow CO_2 to escape and precipitate $MgCO_3$.

Water samples were collected at various stages in the treatment process and filtered (Whatman No 1 filter paper) for pH, Fe (II), acidity and magnesium determinations using standard procedures (APHA, 1989). Calcium was assayed using atomic absorption spectrophotometry. Acidity was determined by titration to pH 8.3 using a 0.1 N NaOH solution. The Fe(II) concentrations were determined by adding 10mL each of 1N H_2SO_4, 10mL of Zimmermann-Reinhard Reagent to 10ml of the sample pipetted into a 100mL Erlenmeyer flask. The solution was titrated with 0.1N $KMNO_4$ solution until pale pink , (Vogel, 1989).

3. RESULTS AND DISCUSSION

Water Quality and Chemical Reactions

$Mg(OH)_2$ and lime can be used for treatment of acid water. Table 1 shows the chemical composition of the water after treatment at the various stages. Overall the Total Dissolved Solids (TDS) content was lowered from 9 173 mg/l in the feed water to 5 075 mg/l after $Mg(OH)_2$ treatment and to 2 608 mg/l after lime treatment. During $Mg(OH)_2$ treatment, free acid and all the metals excluding calcium and magnesium, were removed to low levels. The lowering of the TDS in the case of $Mg(OH)_2$ treatment was mainly due to the oxidation of Fe^{2+} to Fe^{3+} (Reaction 2) and precipitation as $Fe(OH)_3$ (Reaction 3). Similarly, Mn^{2+} was oxidised to Mn^{4+} and precipitated as MnO_2. The other metals: Al^{3+}, Co^{2+}, Zn^{2+}, Pb^{2+}, Co^{2+} and Ni^{2+} precipitated as metal hydroxides. This was achieved due to the low solubility-products for $Fe(OH)_3(2.64 \times 10^{-39})$; $Al(OH)_3(8.5 \times 10^{-23})$; $Mn(OH)_2(2.06 \times 10^{-13})$; $Cu(OH)_2(2.20 \times 10^{-20})$; $Zn(OH)_2(7.71 \times 10^{-17})$; $Pb(OH)_2(1.42 \times 10^{-20})$; $Co(OH)_2(1.09 \times 10^{-15})$ and $Ni(OH)_2(5.47 \times 10^{-16})$ (Sillen & Martell, 1964, Lide, 1992). The Mg concentration increased from 234 mg/L to 840 mg/L during $Mg(OH)_2$ treatment (Reaction 4). The SO_4^{2-}-concentration remained constant during $Mg(OH)_2$ treatment due to the high solubility of $MgSO_4$ of 26 g/L (Weast, 1971a). During the follow-up, lime treatment , Mg^{2+} was removed completely as $Mg(OH)_2$ (Reaction 6) at pH 11.5. The SO_4^{2-} was lowered from 4 603 mg/l to 1 943 mg/l due to gypsum crystallization (Reaction 7).

$2Fe^{2+} + ½O_2 + 2H^+ \rightarrow 2Fe^{3+} + H_2O$ (2)
$2Fe^{3+} + 6H_2O \rightarrow 2Fe(OH)_3 + 6H^+$ (3)
$6H^+ + 3Mg(OH)_2 \rightarrow 3Mg^{2+} + 6H_2O$ (4)
$2Fe^{2+}\ ½O_2 + 2H^+ + 3Mg(OH)_2 \rightarrow 2Fe(OH)_3 + 3Mg^{2+} + H_2O$ (5)
$Mg^{2+} + Ca(OH)_2 \rightarrow Mg(OH)_2 + Ca^{2+}$ (6)
$SO_4^{2-} + Ca^{2+} + 2H_2O \rightarrow CaSO_4.2H_2O$ (7)

The rate of sulphate removal by gypsum crystallization may be predicted from Equation 8 (Maree *et al.*, 2004).

$$d[CaSO_4.2H_2O]/dt = k[CaSO_4.2H_2O](S)[C-C_0]^2 \qquad (8)$$

where $d[CaSO_4.2H_2O]/dt$ represents the rate of crystallisation, k the reaction rate constant, $[CaSO_4.2H_2O](S)$ the surface area of the seed crystals, C the initial concentration of calcium sulphate in solution and C_0 the saturated concentration of calcium sulphate in solution.

A benefit of using $Mg(OH)_2$ for neutralization and metal removal is that the precipitated metal hydroxides are not mixed with gypsum. The sludge from the $Mg(OH)_2$ stage consists of 83.8% $Fe(OH)_3$, 8.5% $Al(OH)_3$, 6.7% MnO_2 and 1.0% $Zn(OH)_2$. The lime sludge consists of 70.3% gypsum and 29,7% $Mg(OH)_2$. Other benefits include:

Fe(Ii)-Oxidation Using $Mg(Oh)_2$ for Neutralization

Figure 2 shows the behaviour of the various parameters under batch conditions. It was noted that free acid was removed rapidly until only Fe(II) acidity was left in solution. Acidity associated with Fe^{2+} in solution was removed as Fe^{2+} was oxidised as indicated by the similarity of the Acidity and Fe^{2+} lines (both expressed in mmole/L). The pH of the water remained below 5.5 while Fe^{2+} was still in solution. The pH was raised to above 9 only after complete removal of Fe^{2+}. The behaviour of Fe^{2+} oxidation when $CaCO_3$ was used for neutralization, as reported by Maree *et al.* (2004), was similar to when $Mg(OH)_2$ was used for neutralization. Lime behaves differently. Fe^{2+}-oxidation takes place at higher pH values as controlled by the lime dosage. These differences can be explained by the solubilities of the various alkalis (Table 2).

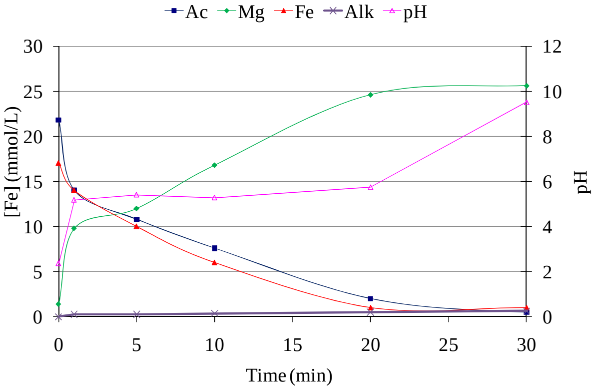

Figure 2. Behaviour of various parameters during Fe(II)-oxidation using $Mg(OH)_2$.

Alkali	Solubility (mg/l)		Mol mass	Ref
	mg/L	mmole/L	g	
$CaCO_3$	14	0.14	100.0	Weast, 1971b
$Mg(OH)_2$	9	0.15	58.3	Weast, 1971c
$Ca(OH)_2$	1850	25.00	74.0	Weast, 1971b

Table 2. Solubility of various alkalis.

Stumm and Lee (1961) determined the following relationship between the iron oxidation rate and pH in the absence of microorganisms for clear solutions (Equation 9).

$$-d[Fe^{2+}]/dt = k[Fe^{2+}][OH^-]^2P_{O2} \qquad (9)$$

where: $-d[Fe^{2+}]/dt$ = rate of iron oxidation; k = reaction rate constant; $[Fe^{2+}]$ = Fe(II) concentration (moles/L); $[OH^-]$ = hydroxide concentration (moles/L); and P_{O2} = partial pressure of oxygen (mm Hg).

Maree *et al.* (1997) showed that the rate of iron oxidation is also catalysed by suspended solids when $CaCO_3$ is used for neutralization. Figures 3 - 8 and Table 3 show the effects of the following parameters on the rate of Fe^{2+}-oxidation when pH is controlled by dosing $Mg(OH)_2$. The relative importance of various factors in terms of their influence on the rate of iron oxidation was determined by a series of controlled tests in which the dependence of the rate on one variable at a time was determined. In the pH range 5 - 9.5, which is of importance for $Mg(OH)_2$ neutralisation, the iron oxidation rate was assumed to have the following functional form:

$$-d[Fe^{2+}]/dt = k.[Fe^{2+}]^{n1}.[O_2]^{n2}.[OH^-]^{n3}.[SS]^{n4}.M^{n5} \qquad (10)$$

where: $-d[Fe^{2+}]/dt$ or R = rate of iron oxidation; k = reaction rate constant; $[Fe^{2+}]$ = Fe^{2+} concentration (moles/L); $[Mg(OH)_2]$ = magnesium hydroxide concentration (moles/L); $[O_2]$ = oxygen concentration (moles/L); $[Fe(OH)_3]$ = $Fe(OH)_3$-concentration and M = mixing intensity (rpm).

By varying the value of only one parameter in a series of experiments, say $[Fe^{2+}]$, equation 11 can be written as:

$$-\partial[Fe^{2+}]/\partial t = K.[Fe^{2+}]^{n1} \text{ or } \log(-\partial[Fe^{2+}]/\partial t) = \log K + n_1 \log[Fe^{2+}] \qquad (11)$$

where: $K = k.[O_2]^{n2}.[OH]^{n3}.[SS]^{n4}.M^{n5}$

Table 1. Chemical composition of feed and treated water when acid mine water was treated with $Mg(OH)_2$ and lime

Parameter	**Feed**	**$Mg(OH)_2$**	**Lime**
Dosage (mg/l)		1,487.8	2,171.8
Time (min)			180.0
Gypsum dosage (mg/l)			5,000.0
Water quality:			
pH	2.3	9.6	12.5
Alkalinity (mg/L $CaCO_3$)		60.0	2,180.0
Chloride (mg/L)			71.9
Fluoride (mg/L)			
Sulphate (mg/L)	4,635.6	4,641.9	2,210.2
Total acidity (mg/L $CaCO_3$)	2,320.0	0.0	0.0
Free acidity (mg/L CaCO3)	791.8	0.0	0.0
Sodium (mg/L)	133.6	123.5	146.7
Litium (mg/L)	0.6	0.5	0.6
Potassium (mg/L)	1.3	0.0	1.8
Magnesium (mg/L)	247.9	840.0	0.0
Calcium (mg/L)	465.7	460.0	1,696.0
Aluminium (mg/L)	36.2	1.3	1.9
Iron(II) (mg/L)	83.8	0.0	0.0
Total iron (mg/L)	438.2	2.2	0.9
Manganese (mg/L)	50.1	0.0	0.0
Copper (mg/L)	6.6	0.8	1.3
Zinc (mg/L)	7.6	0.0	1.0
Lead (mg/L)	5.2	1.7	0.5
Cobalt (mg/L)	3.4	1.1	0.1
Nickel (mg/L)	14.7	0.1	0.0
TDS (mg/L)	9,242.2	6,037.0	2,753.1
Anions (meq/L)	96.6	97.9	91.7
Cations (meq/L)	96.6	97.9	91.7
Sludge production:			
Gypsum (mg/L)			4,356.8
$Mg(OH)_2$ (mgL)			2,015.3
$Fe(OH)_3$ (mg/L)		994.4	
$Al(OH)_3$ (mg/L)		101.0	
MnO_2 (mg/L)		79.3	
Zn(OH)2 (mg/L)		11.6	
Sludge production:			
Gypsum (%)			68.4
$Mg(OH)_2$ (%)			31.6
$Fe(OH)_3$ (%)		83.8	
$Al(OH)_3$ (%)		8.5	
MnO_2 (%)		6.7	
Zn(OH)2 (%)		1.0	

The contribution, n_1, of Fe(II) to the overall reaction rate was determined from the slope of the graph obtained by plotting log R *versus* log $[Fe^{2+}]$. The data in Table 3 and Figures 3 - 5 showed that the rate of iron oxidation is of order 1.0 (≈ 1), 1.2 (≈ 1.0), 0.51 (≈ 0.5), to Fe^{2+}, $Mg(OH)_2$, and $Fe(OH)_3$ concentrations, respectively. Figure 6 shows that the reaction rate also increased with increased stirring rate and Figure 7 that the oxidation rate is higher with higher air feed rates. These findings suggested that the rate equation proposed by Stumm and Lee (1961) for clear solutions should be modified for suspensions to:

$-d[Fe^{2+}]/dt = k.[Fe^{2+}].[Mg(OH)_2] .[Fe(OH)_3]^{½}.[O_2]$ (12)

The Arrhenius equation, log k = log A - E/(2.303RT), can be used to estimate the value of the reaction rate k at other temperatures. The parameters, E, R, and log A have the values 7.065 kcal/mole (activation energy), 1.987 cal mole^{-1} degree^{-1} (gas constant) 7.464 (a constant), respectively (Figure 8). A linear relationship was obtained over the range 14 - 49 ^{0}C.

Separation Of Mg(Oh)$_2$ and Gypsum

Figure 9 showed that Mg(OH$_{)2}$ that is precipitated together with gypsum during lime treatment, could be converted, by contacting CO_2 with the $Mg(OH)_2$-rich sludge, to form soluble $Mg(HCO_3)_2$ (Reaction 13) It was noted that the solubility increased with decreasing temperatures. At 15˚C, 3 300 mg/L $Mg(OH)_2$ (as Mg), of the initial concentration of 8 300 mg/L, dissolved.

$Mg(OH)_{2(s)} + CO_{2(aq)} \rightarrow Mg(HCO_3)_{2(aq)}$ (13)

$MgCO_3$ can be precipitated from $Mg(HCO_3)_2$ by heating to drive off CO_2 (Reaction 14)

$Mg(HCO_3)_{2(aq)}$ + heat $\rightarrow MgCO_{3(aq)} + CO_2 + H_2O$ (14)

4. CONCLUSIONS

This study demonstrated that the integrated magnesium hydroxide and lime process is applicable for treating acid mine water effectively and that:

- $Mg(OH)_2$ can be used for pH control to allow Fe(II)-oxididation. Fe(II) with a concentration of 900 mg/L was completely oxidised within 10 min reaction time and precipitated as $Fe(OH)_3$ together with other metal hydroxides, separate from gypsum. The TDS concentration was lowered from 9 173 - 5 075 mg/L.
- Lime can be used to precipitate Mg^{2+} as $Mg(OH)_2$ and to partially remove sulphate by means of gypsum crystallization. The TDS concentration was lowerd from 5 075 2 608 mg/L.
- $Mg(OH)_2$ can be separated from the gypsum by reacting it with CO_2 to form $Mg(HCO_3)_2$.

5. ACKNOWLEDGEMENTS

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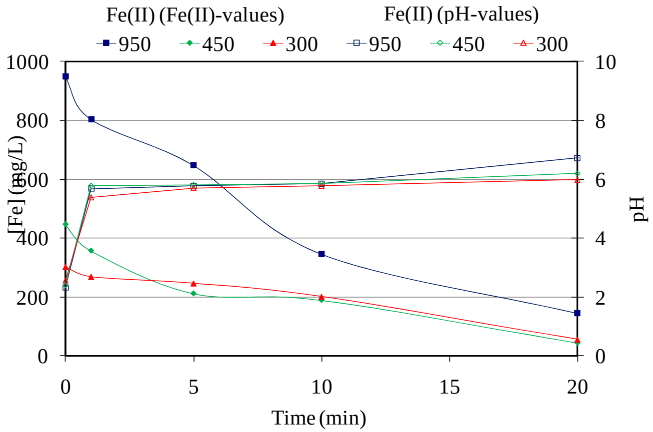

Figure 3. Effect of Fe(II)-concentration on the rate of Fe(II)-oxidation.

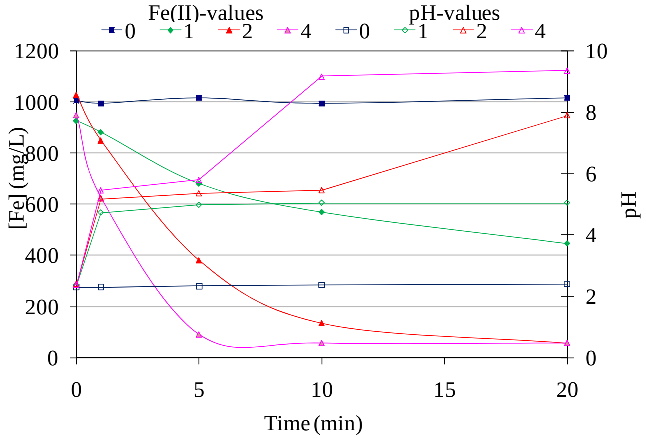

Figure 4. Effect of $Mg(OH)_2$/Fe(II) mole ratio on the rate of Fe(II) removal

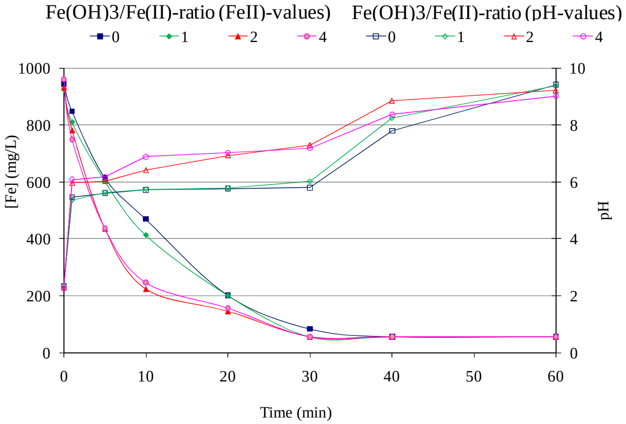

Figure 5. Effect of $Fe(OH)_3$/Fe(II) mole ratio on the rate of Fe(II) removal

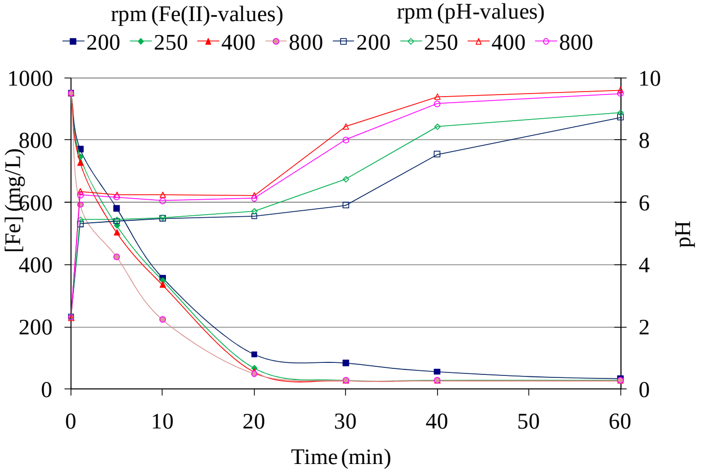

Figure 6. Effect of stirring rate

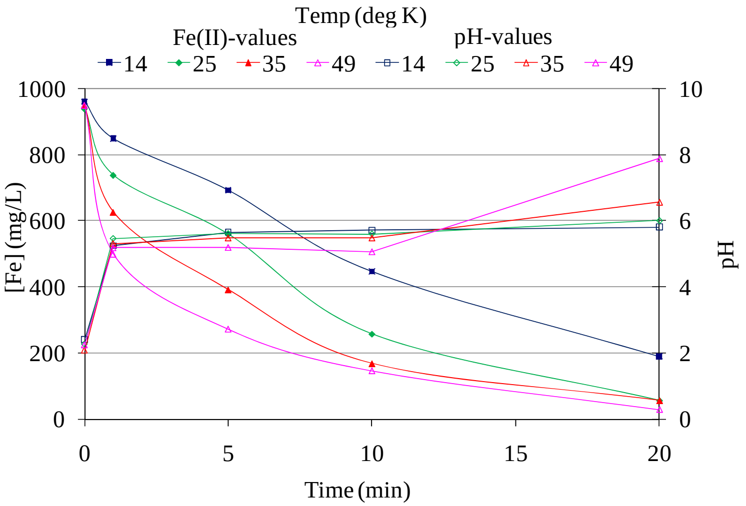

Figure 7. Effect of Temperature on the rate of Fe(II)-oxidation.

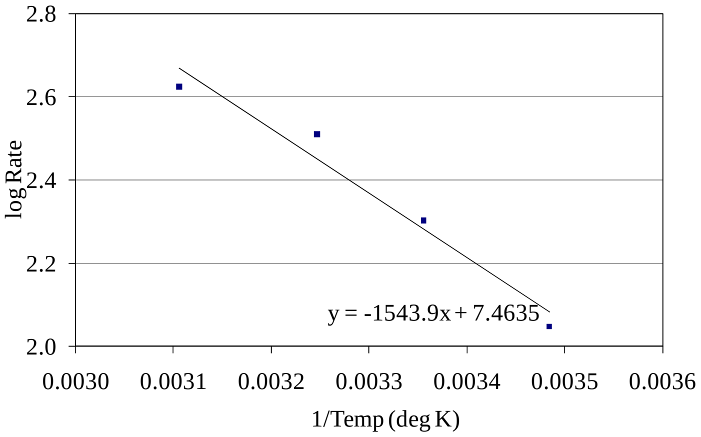

Figure 8. Activation enegy.

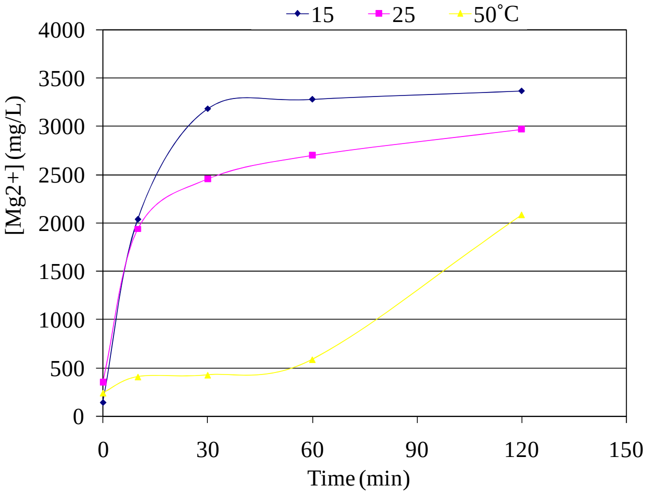

Figure 9. Effect of temperature on the solubility of $Mg(HCO_3)_2$

Table 3. Rate of Fe(II)-oxidation under various conditions

Rate	Order	Reaction conditions									
		Fe(II)	$Mg(OH)_2$/ Fe(II) mole ratio	$Fe(OH)_3$/ Fe(II) mole ratio	$Mg(OH)_2$	$Fe(OH)_3$	O_2	Mixing intensity	Temp	log Rate	log C
mg/L/min		mg/L	mmole/L	mmole/L	mg/L	mg/L	L/min	rpm	ºC		
Effect of Fe(II)											
39.1	1.1	**950**	1.67	0.00	1652	0		250	25	1.592	2.978
18.0		**450**	1.43	0.00	670	0		250	25	1.255	2.653
11.6		**300**	1.07	0.00	335	0		250	25	1.065	2.477
Effect of Mg(OH)2/Fe(II) mole ratio											
11.2		950	0.00	0.95	0	1720		250	25		
44.5	1.2	950	0.72	0.95	**713**	1720		250	25	1.648	2.853
178.7		950	1.67	0.95	**1652**	1720		250	25	2.252	3.218
323.9		950	3.56	0.95	**3531**	1720		250	25	2.510	3.548
Effect of Fe(OH)3/Fe(II) mole ratio											
96.6		950	1.67	0.00	1879	0		250	25		
113.1	0.5	950	1.67	0.95	1879	**1720**		250	25	2.053	3.236
150.7		950	1.67	1.89	1879	**3440**		250	25	2.178	3.537
212.2		950	1.67	3.79	1879	**6890**		250	25	2.327	3.838
Effect of air feed rate											
159.6		950	1.67	0.95	1879	1720	**0.39**	250	25	2.203	3.236
168.5		950	1.67	0.95	1879	1720	**0.68**	250	25	2.227	3.236
178.7		950	1.67	0.95	1879	1720	**1.23**	250	25	2.252	3.236
212.2		950	1.67	0.95	1879	1720	**4.92**	250	25	2.327	3.236
Effect of stirring rate											
178.7		950	1.67	0.95	1879	1720		**200**	25	2.252	3.236
204.2		950	1.67	0.95	1879	1720		**250**	25	2.310	3.236
323.9		950	1.67	0.95	1879	1720		**400**	25	2.510	3.236
363.0		950	1.67	0.95	1879	1720		**800**	25	2.560	3.236
Effect of Temperature											1/T
111.7		950	1.67	0.95	1879	1720			14	2.048	0.00348
201.1		950	1.67	0.95	1879	1720			25	2.303	0.00336
323.9		950	1.67	0.95	1879	1720			35	2.510	0.00325
421.0		950	1.67	0.95	1879	1720			49	2.624	0.00311

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