

**SEPTEMBER 1995**

**ARSENATE PRECIPITATION USING  
FERRIC IRON IN ACIDIC CONDITIONS**

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**WRRI Technical Completion Report No. 293**

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# **Arsenate Precipitation Using Ferric Iron in Acidic Conditions**

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TECHNICAL COMPLETION REPORT

Account Number 01-3-45631

September 1995

New Mexico Water Resources Research Institute  
in cooperation with  
Civil, Agricultural and Geological Engineering Department  
New Mexico State University

The research on which this report is based was financed in part by the U.S. Department of Interior, Geological Survey, through the New Mexico Water Resources Research Institute.

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## ABSTRACT

Arsenates (i.e., As(V)) can be removed from aqueous solution by precipitation with ferric iron (i.e., Fe(III)). The chemistry of arsenic acid describes the main properties of arsenates. This triprotic acid resembles the phosphoric acid system. For example, free arsenate ions (i.e.,  $\text{AsO}_4^{3-}$ ), like free phosphates, are present in significant concentration at pH values above  $\text{pK}_{a,3}$ . On the other hand, the concentration of free ferric iron in solution,  $\text{Fe}^{3+}$ , is limited by ferric hydroxide precipitation and hydroxy complexation under neutral or basic conditions.  $\text{Fe}^{3+}$  is the predominant iron form only under very acidic conditions. Therefore, the absence of either ferric ions or arsenate ligands prevents ferric arsenate ( $\text{FeAsO}_4$ ) precipitation in extreme pH conditions.

Precipitation studies using ferric chloride show that the formation of ferric arsenate in water containing 0.667 mM/L (50 mg/L as As) is favored in the pH range between 3 and 4. Ferric iron dose required to remove arsenic from solution increases with pH in the range of 3 to 10. Sludge production also increases with increasing pH conditions. Optimum ferric iron doses at pH 3 and 4 are 4.8 and 10.0 mM/L, respectively, where the arsenate is removed from solution by 98.72 and 99.68 percent. Corresponding iron requirement to arsenate ratios at these two pH conditions are 7.2 and 15.0.

Adverse effects on arsenic removal are observed at pH = 3, where the concentration of applied ferric iron exceeds the optimal dose. This effect is probably due to charge reversal on the surface of the precipitates. Overdosing above the optimal iron concentration at pH = 4 does not reduce treatment efficiency significantly. Presence of sodium chloride in solution at a concentration of 171 mM/L (10,000 mg/L as NaCl) does not impair system performance. However, sodium sulfate at a concentration of 104 mM/L (10,000 mg/L) affects adversely treatment performance.

Keywords: arsenic, arsenate, chemical precipitation, ferric iron, groundwater, water treatment

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## Introduction

Arsenic is ubiquitous in soils, rocks, water and biological food chains. Typical concentrations in the earth's crust vary from 2 to 5 ppm (Waldron 1980). Arsenic commonly is found in industrial and agricultural wastes. Approximately 70 percent of the arsenic is used as a pesticide in agricultural applications (Bodek, et al. 1988). Arsenic is also used in the ceramic, chemical, electronics, and pharmaceutical industries (Waldron 1980). This element may be found in produced waters (during extraction of petroleum) and in petrochemical refinery effluents. Most water supplies contain measurable concentrations of arsenic (Waldron 1980).

Currently, the maximum contaminant level (MCL) for arsenic in drinking water is 0.05 mg/L or 50 ppb (Pontius 1993). However, a court-ordered deadline was given to the U.S. EPA (USEPA) to propose revised regulations for arsenic no later than November 1992, but the agency is still undecided on what the new proposed regulations might be and has filed for an extension. The final rule is due no later than 24 months after the date of publication of the proposed rule, which is due sometime in late 1993 or early 1994 (Pontius 1993).

Several difficult decisions face the agency when assessing the health effects, which explains the delay for determining an appropriate MCL. Some scientists believe arsenic is an essential nutrient and suggest that an intake of 0.012 to 0.04 mg/L is safe for adults; however, no conclusive evidence exists on which to base the MCL of arsenic for humans (Pontius 1993). If the agency determines arsenic to be noncancerous, the maximum contaminant level goal (MCLG) would likely be proposed between 0.003 and 0.020 mg/L, based on preventing hyperkeratosis (thickening of the skin) (Pontius 1993). On the other hand, if it is determined that arsenic ingestion causes cancer (which some evidence in Taiwanese epidemiology studies suggests), then arsenic would be classified as a group A carcinogen with a MCLG of zero (Pontius 1993). These new regulatory trends are likely to affect both the drinking water and industrial wastewater treatment industries alike.

By far, the most common method for arsenic removal is chemical precipitation with lime and ferric salts, followed by coagulation and filtration processes (Harper and Kingham 1992; Huang and Fu 1984). Most processes, however, operate at a pH of around 7 to 12, and very little investigation has been done using this method under acidic conditions. Chemical precipitation using the lime and ferric salt method is very effective, particularly at high initial arsenic concentrations. However, this process produces large quantities of sludge (due to hardness and ferric hydroxide precipitation), which makes the process chemically intensive and inefficient from a waste minimization viewpoint. The technology presented in this report addresses the removal of arsenate ( $\text{AsO}_4^{3-}$ ) from industrial wastewaters containing relatively high concentrations of pentavalent arsenic (50 mg/L in this study). However, this methodology may present innovative alternatives to the water treatment industry, where this element typically is found at below the 1.0 mg/L level.

## Objectives

Ferric arsenate precipitation is affected by solution pH and by the dosage of ferric iron applied. To simplify study of these two variables on system performance, the authors selected to maintain the pH conditions constant, while varying the concentration of ferric iron in solution. Several pH conditions were chosen for this experimental approach. From these data, plots were constructed for each pH, and the minimum iron required for removal was determined. After the optimal pH and minimum iron requirement are determined, the effects of sodium chloride and sodium sulfate were investigated. Chloride and sulfate effects on treatment performance are important because these salts are common constituents of most industrial wastes. An experimental study was done using 10,000 mg/L sodium chloride to evaluate the potential effect of this salt on arsenic removal. A similar study using 10,000 mg/L  $\text{SO}_4^{2-}$  was conducted to determine the effect of this divalent anion on arsenic precipitation. These highly saline conditions were used to emulate an actual industrial waste from a produced water pit.

## Solution Chemistry

Arsenic can exist in four different oxidation states, which include: (-III), (0), (III), and (V) (Baes and Mesmer 1986; Merrill et al. 1986). The predominant aquatic forms are trivalent arsenite (III) and pentavalent arsenate (V) (Harper and Kingham, 1992). Arsenate is more common in aerobic surface waters, while arsenite is more common in anaerobic, or oxygen-free groundwater (Harper and Kingham 1992). Oxygen, chlorine ( $\text{NaOCl}$ ), or potassium permanganate can oxidize arsenite to arsenate (Gulledge and O'Connor 1973). The chemistry of arsenate resembles that of phosphorous (Baes and Mesmer 1986; Bodek et al. 1988). Arsenic acid,  $\text{H}_3\text{AsO}_4$ , like phosphoric acid,  $\text{H}_3\text{PO}_4$ , exists only in extremely acidic conditions.  $\text{H}_2\text{AsO}_4^-$  is the predominant arsenate form in slightly acidic conditions, while  $\text{HAsO}_4^{2-}$  predominates above neutrality. The deprotonated  $\text{AsO}_4^{3-}$  becomes the main arsenate form only at extremely high pH conditions. The following reactions and corresponding equilibrium constants are used to describe the chemistry of the arsenic acid system at 25°C (Baes and Mesmer 1986, Ramana and Sengupta 1992):

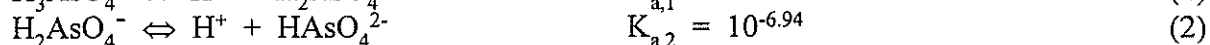
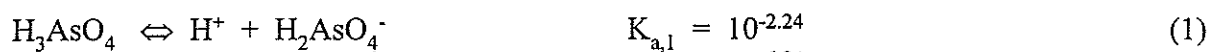


Figure 1, which is constructed from these equations, shows the pC-pH diagram for a total arsenate concentration of  $10^{-3.18}$  M As(V) (50 mg/L As(V)). This concentration also is used for the experiments in this report. This diagram resembles closely that of the phosphoric acid system, whose three deprotonation constants are  $10^{-2.15}$ ,  $10^{-7.20}$ , and  $10^{-12.38}$ .

The dissolution of ferric chloride in water results in the formation of various ferric oxohydroxide complexes (Baes and Mesmer 1986). The predominant complexes are determined by the availability of hydroxide ions in solution. Thus, the degree of complexation with hydroxide ions increases with the pH of the solution. Amorphous ferric hydroxide may precipitate from the solution when the solubility product for this solid is exceeded.



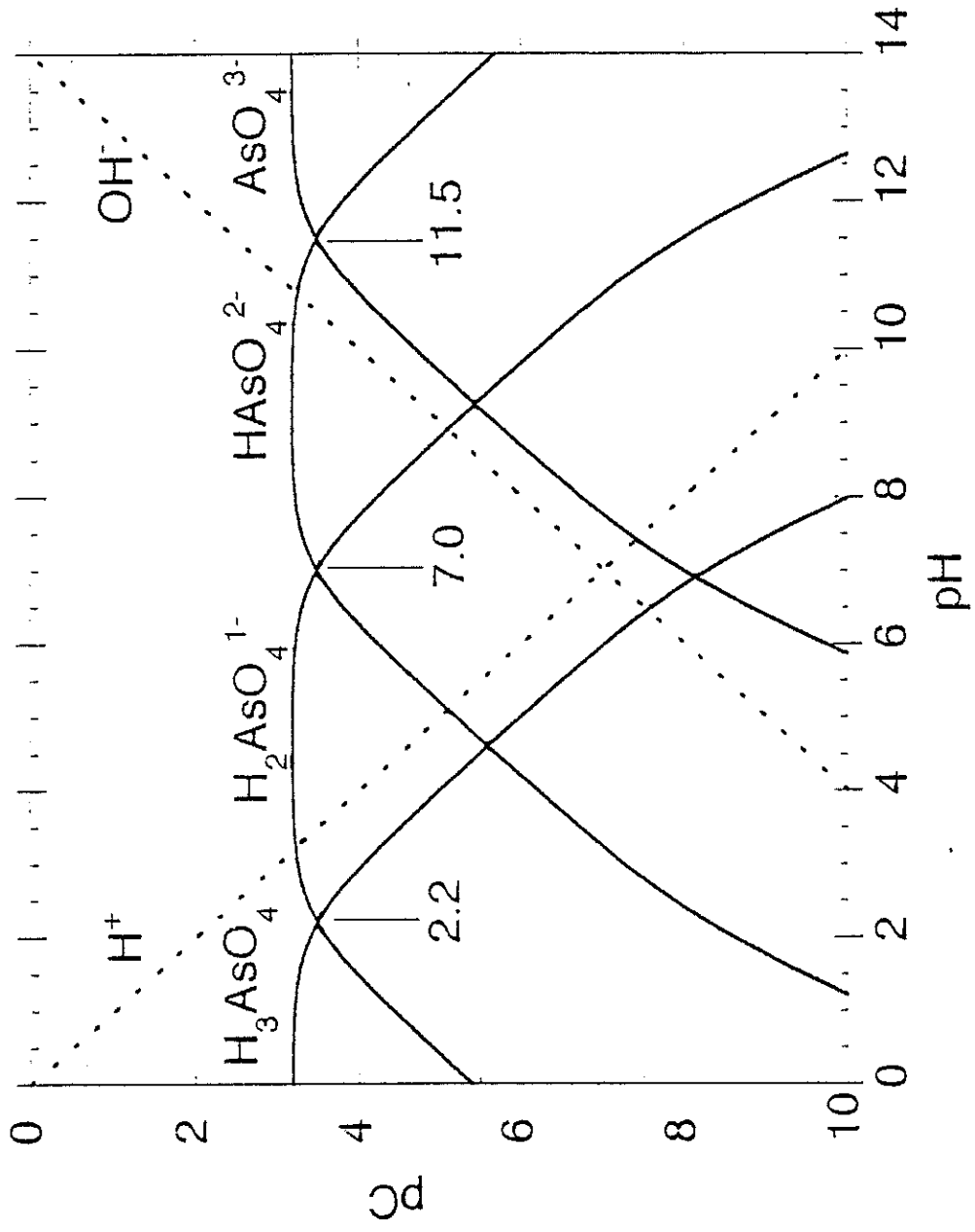


Figure 1. Arsenic acid diagram for  $pC_{T,As} = 3.18$  (50 mg/L As)

Thermodynamic information on the hydrolysis of Fe(III) has been compiled by Snoeyink and Jenkins, 1980, and Huang and Fu, 1984. Their reactions and corresponding equilibrium constants represent the chemical model for ferric iron solubility and hydroxy complexation in water solutions:

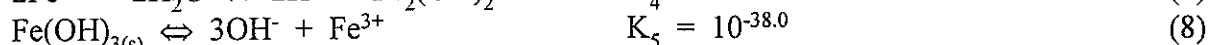
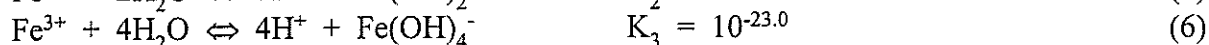


Figure 2 is a graphical representation of the solubility of Fe(OH)<sub>3</sub> based on Equations (4) through (9). Minimum solubility condition occurs at pH = 8.3 and equals 10<sup>-10.8</sup> M (8.85 x 10 mg/L as Fe(III)). This low solubility explains the excellent ability of Fe (III) to remove trace elements (dissolved and particulate) which are adsorbed onto and trapped within the iron hydroxide precipitate (Gulledge and O'Connor 1973; Pierce and Moore 1982; Merrill et al. 1986).

Ferric oxides are capable of binding phosphates (Stumm and Morgan 1981). Maximum adsorption is reported under near-neutral pH conditions. Arvin and Petersen, 1980, report that phosphates can precipitate in natural waters in the presence of ferric iron. The main components in the precipitate are iron, phosphate, calcium, bicarbonate, and hydroxide, depending on the chemistry of the solution. According to the literature search by Arvin and Petersen, 1980, optimal pH removal conditions for phosphate removal by iron precipitation takes place in the 4.5 to 5.5 range; while their own experimental work favored near-neutral pH conditions. Merrill et al., 1986, also report excellent arsenate removal efficiencies at or near pH = 7.0. The similarities between phosphates and arsenates suggest that precipitation (and or adsorption) of arsenates by ferric iron may be favorable under acidic conditions. In fact, the theoretical model presented by Wagemann, 1978, predicts that arsenate is precipitated (or adsorbed) to a maximum extent at pH near 1.0. Interestingly, most experimental research on the removal of arsenate by iron salts has been conducted under basic conditions, leaving the acidic conditions practically unexplored. The works by Huang and Fu, 1984, and Huang and Vane, 1989 are exceptions to this limitation. These researchers investigated the removal of arsenate from water using activated carbon that was impregnated with ferrous iron. Their findings confirm that acidic conditions are favorable for the removal of this pollutant from water by adsorption onto the iron-carbon solid phase. Also, De Carlo and Thomas (1985) used an adsorptive bubble flotation method to remove arsenic from a high ionic strength geothermal fluid produced by a Hawaiian geothermal project well. Precipitation of ferric arsenate is described by the following chemical reaction (Huang and Fu, 1984):



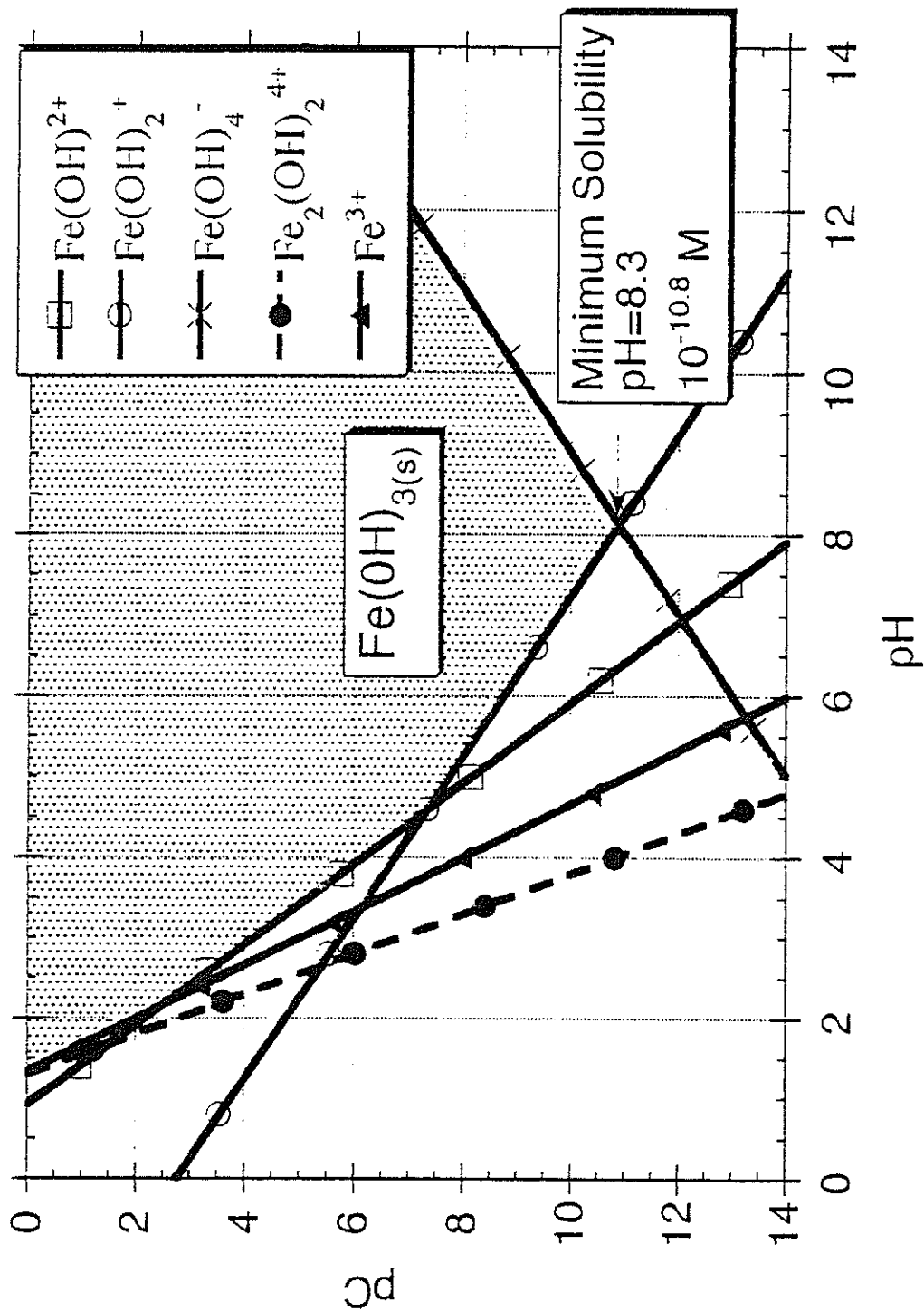


Figure 2. Ferric hydroxide solubility diagram

It is clear from this reaction that arsenate precipitation is favored when the arsenates are present in their most deprotonated form (i.e., as  $\text{AsO}_4^{3-}$ ) and when ferric iron is not hydrolyzed, but present in its free ionic state (i.e., as  $\text{Fe}^{3+}$ ). According to Figure 1, the concentration of  $\text{AsO}_4^{3-}$  is maximized only at extremely high pH conditions, when the pH of the solution is higher than  $\text{pK}_{a,3}$ . On the other hand, iron (III) tends to hydrolyze strongly (i.e., form hydroxy complexes) in basic or neutral solutions. Therefore, precipitation of arsenate by ferric ions is anticipated at a pH condition where both exist at significant concentrations. This rules out the possibility of precipitation at very high pH where free ferric iron is small and very low pH conditions where arsenate exists in a protonated state.

As shown in Figure 2,  $\text{Fe}^{3+}$  is only found as the predominant iron form if the pH of the solution is kept significantly below the neutral point (Baes and Mesmer, 1986). Extreme pH conditions should favor the concentration of either  $\text{AsO}_4^{3-}$  or  $\text{Fe}^{3+}$ , but never favor the formation of both ions simultaneously. Therefore, an optimal pH condition must exist between both extreme pH conditions, where formation of ferric arsenate is enhanced. The authors have investigated a wide range of pH conditions in an attempt to identify optimal removal of arsenates by ferric salts. Whether the mechanism involved is precipitation, coprecipitation or adsorption is the subject of significant debate that falls beyond the scope of this project. For simplicity and consistency with most research in this area, the authors have adopted the concept that most arsenate is precipitated as ferric arsenate in conjunction with the excess iron, which is precipitated as a hydroxide. Kirk (1993) uses Equations (1) to (10) to develop the solubility diagram for  $\text{FeAsO}_4(s)$  that incorporates both the solubility of ferric hydroxide and the protonation states of the arsenic acid system. A similar procedure for aluminum phosphate is found in Snoeyink and Jenkins (1980) and is presented in more detail by Arvin and Petersen (1980) to incorporate several possible interferences. Kirk's diagram, which is reproduced in Figure 3, assumes a 1 to 1 molar ratio of Fe(III) to  $\text{AsO}_4^{3-}$  (i.e., the precipitate is composed of equimolar concentrations of the two components). This figure shows that the minimum arsenate solubility is approximately 0.0315 mM/L or  $10^{-4.5}$  M (2.36 mg/L As(V)) and occurs in acidic conditions at a pH of 5.5.

## Experimental Procedure

Arsenic solutions are made from Sigma brand  $\text{KH}_2\text{AsO}_4$  (FW=180.0 gm). A 13.35 mM/L (1,000 mg/L) arsenic stock solution (as As(V)) is prepared by adding 2.40 gm of  $\text{KH}_2\text{AsO}_4$  to 1 L of nanopure water from this stock solution. The required concentrations used in the experiments are made by dilution. For example, a 1.34 mM/L (100 mg/L) As(V) solution is prepared by diluting 100 mL of the stock solution to 1,000 mL. Iron solutions are made from Fisher brand  $\text{FeCl}_3 \cdot 6\text{H}_2\text{O}$  (FW=270.317 gm). The iron stock solution is 32.0 mM/L (1,787 mg/L) Fe(III), and is prepared by adding 8.65 gm of  $\text{FeCl}_3 \cdot 6\text{H}_2\text{O}$  to 1 L of nanopure water. Dilutions are made accordingly.

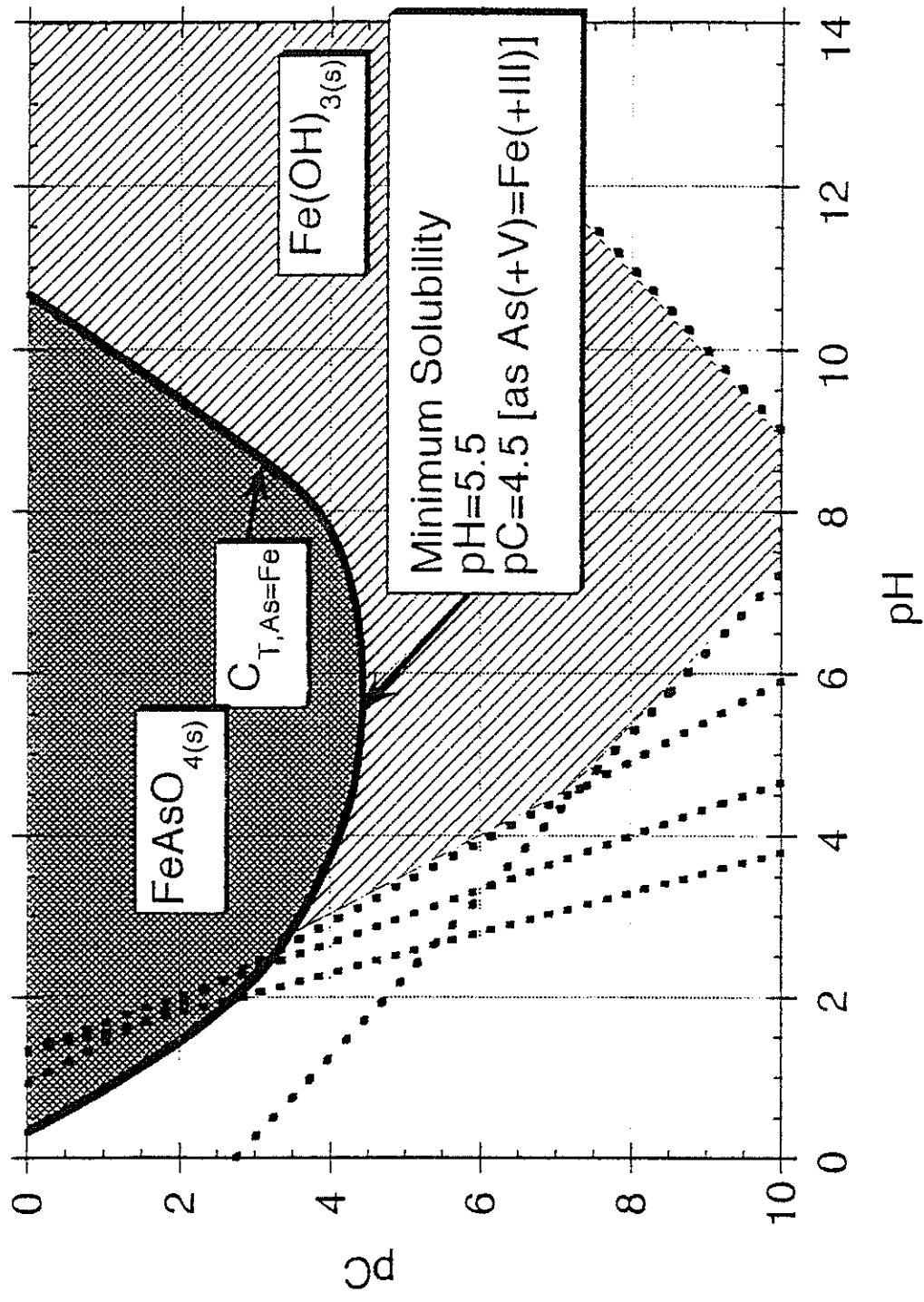


Figure 3. Ferric arsenate solubility diagram (Fe:As = 1.0,  $Fe^{3+}$  availability limited by ferric hydroxide) (from Kirk 1993)

Sodium chloride experiments are conducted using a NaCl solution. The 684 mM/L (40,000 mg/L) stock sodium chloride solution is prepared by adding 40.0 gm of Fisher brand NaCl (FW=58.45 gm) to 1 L of nanopure water. The 417 mM/L (40,000 mg/L)  $\text{SO}_4^{2-}$  stock sodium sulfate solution is prepared from the addition of 134.1 gm of Baker brand  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$  (FW=322.0 gm) to 1 L of water.

All tests are performed in 13 mL test tubes, which are washed in Citranox acid detergent and rinsed with nanopure water prior to the experiments. Three solutions are prepared and the pH of each solution is adjusted before mixing in the test tubes. These three solutions are: (1) 1.34 mM/L As(V) (100 mg/L) solution, (2) 32.0 mM/L Fe(III) (1788 mg/L) solution, and (3) nanopure water. The solution pH is adjusted by the addition of concentrated  $\text{HNO}_3$  or NaOH depending on the desired pH. The amount of concentrated  $\text{HNO}_3$  or NaOH added is typically below 1.0 mL into 500 mL, so the dilution due to concentrated  $\text{HNO}_3$  or NaOH addition is considered negligible. The pH is varied from 2 to 10, but is held constant for each run.

After pH adjustment, specific volumes of each of the three solutions are added to each test tube in the following order: (1) As(V) solution, (2) nanopure water, and (3) Fe(III) solution. There are 11 tubes total for each pH, and the total volume in each test tube after the addition of the three solutions is 10 mL. A 0.667 mM/L As(V) (50 mg/L) initial concentration is held constant and the Fe(III) initial concentration is varied from 0 to 16.0 mM (894.0 mg/L) Fe(III) at a given pH. The tubes are mixed for 10 minutes at 20 rpm in an end-over-end mixer, and then centrifuged at 3000 rpm for 10 minutes. The As(V) and Fe(III) concentrations in the supernatant are determined using Inductively Coupled Plasma (ICP) spectroscopy after centrifugation. An Jobin Yvon model JY70PLUS (argon gas) is used for this purpose. Arsenic solutions without iron are tested to ensure no arsenic is absorbed onto the glass test tube. Following ICP testing, the pH of each tube is then measured.

The sodium chloride experiments followed the same procedure and testing method as previously described, but with two exceptions. First, a 684 mM/L (40,000 mg/L) sodium chloride solution is introduced into the mixing scheme. Next, the concentration of the arsenic solution is changed from 1.34 mM/L As(V) (100 mg/L) to 2.68 mM/L (200 mg/L) As(V) to account for the sodium chloride solution addition. The Fe(III) initial concentration is varied the same as before, and the pH is adjusted in the same manner as before. Also, the solutions are made to ensure an initial As(V) concentration of 0.667 mM/L (50 mg/L), and a sodium chloride concentration of 171.0 mM/L (10,000 mg/L) in each test tube. The sodium sulfate experiments are performed following the same procedure and testing method as the sodium chloride tests, except that a 417 mM/L (40,000 mg/L) sodium sulfate solution (as  $\text{SO}_4^{2-}$ ) is used in place of the sodium chloride solution. After mixing, each test tube contains a 104.25 mM/L (10,000 mg/L) sulfate concentration.

## Results and Discussion

Experiments were done to determine the minimal Fe(III) concentration and optimal pH relationship for maximum As(V) removal. This task was accomplished by using a constant

As(V) concentration of 0.667 mM/L (50 mg/L) and varying the Fe(III) concentration from 0 to 16.0 mM/L (894.0 mg/L) for each pH condition. The pH range covered was 2 to 10, with increments of 0.5 over most of the pH range. Figure 4 shows the soluble arsenic remaining as a function of pH and Fe(III) concentration. It is clear from this figure that the lower pH conditions require less Fe(III) for equivalent arsenic removal for pH values below 3.0. Since Fe(III) can precipitate as either  $\text{Fe}(\text{OH})_3$  or  $\text{FeAsO}_4$ , it is likely that  $\text{Fe}(\text{OH})_3$  is the major precipitate at higher pH conditions (closer to minimum solubility point  $\text{pH} = 8.3$ , as predicted by Figure 2). The mass of waste (i.e., sludge) is obviously greater at higher pH levels for the same amount of As(V) removal than at the lower pH levels because both iron and arsenate are removed from solution. Figure 5 details the optimal removal zone from Figure 4. This figure shows that at  $\text{pH} = 3.0$ , arsenate removal depends heavily on the applied iron dose. The optimal iron concentration for As(V) removal at this pH is 4.8 mM/L (268.1 mg/L as Fe(III)). Iron addition beyond this point inhibits arsenic removal, possibly due to charge reversal on the surface of the arsenate and hydroxide precipitates. Secondly, the optimal pH range is between 3.0 and 4.0, depending upon what parameters are considered (lower final arsenic concentration or waste minimization).

Good arsenic removal is obtained using 4.8 mM/L (268.1 mg/L) as Fe(III) at a pH of 3.0, where the final arsenic concentration is 0.0086 mM/L (0.641 mg/L). Better arsenic removal efficiency is observed at a pH of 4.0, where the lowest attainable arsenic concentration is 0.002 mM/L (0.15 mg/L). Excepting one experimental point, over 10 mM/L (558 mg/L) of iron is required to operate at a pH of 4.0, which is more than twice the iron needed at a pH of 3.0. Theoretically, over twice as much sludge is produced for a relatively small improvement in arsenic removal efficiency.

Using the optimal iron dose at pH values of either 3.0 or 4.0, the percent arsenic reduction is 98.72 and 99.68 percent, respectively. Although final arsenic concentrations exceeds present drinking water standards ( $0.667 \times 10^{-3}$  mM/L or 0.05 mg/L), these are exceptional removal results considering the high initial As(V) concentration (50 mg/L).

It is likely that the present drinking water standard could be met with the addition of conventional water treatment technologies such as sand filtration, even for these highly contaminated water samples. The optimal pH range (3.0 to 4.0) is consistent with the results found in other articles which used iron coprecipitation for other applications. De Carlo and Thomas (1985) and Huang and Vane (1989) all show optimal As(V) removal in the pH range of 3-4.5 for their dissolved air flotation tests. Pierce and Moore (1982) showed increased As(V) removal with decreasing pH at a constant iron dose. However, the lowest pH they studied was 4.0.

Our experimental work shows that approximately 7.2 moles of iron are required to remove 1 mole of arsenic at a pH of 3.0. According to Equation (9), this should be a 1 to 1 molar ratio should the arsenic precipitate as pure ferric arsenate. This discrepancy may be explained by the hydroxide ion competition (against the  $\text{AsO}_4^{3-}$ ) for the Fe(III) ions.

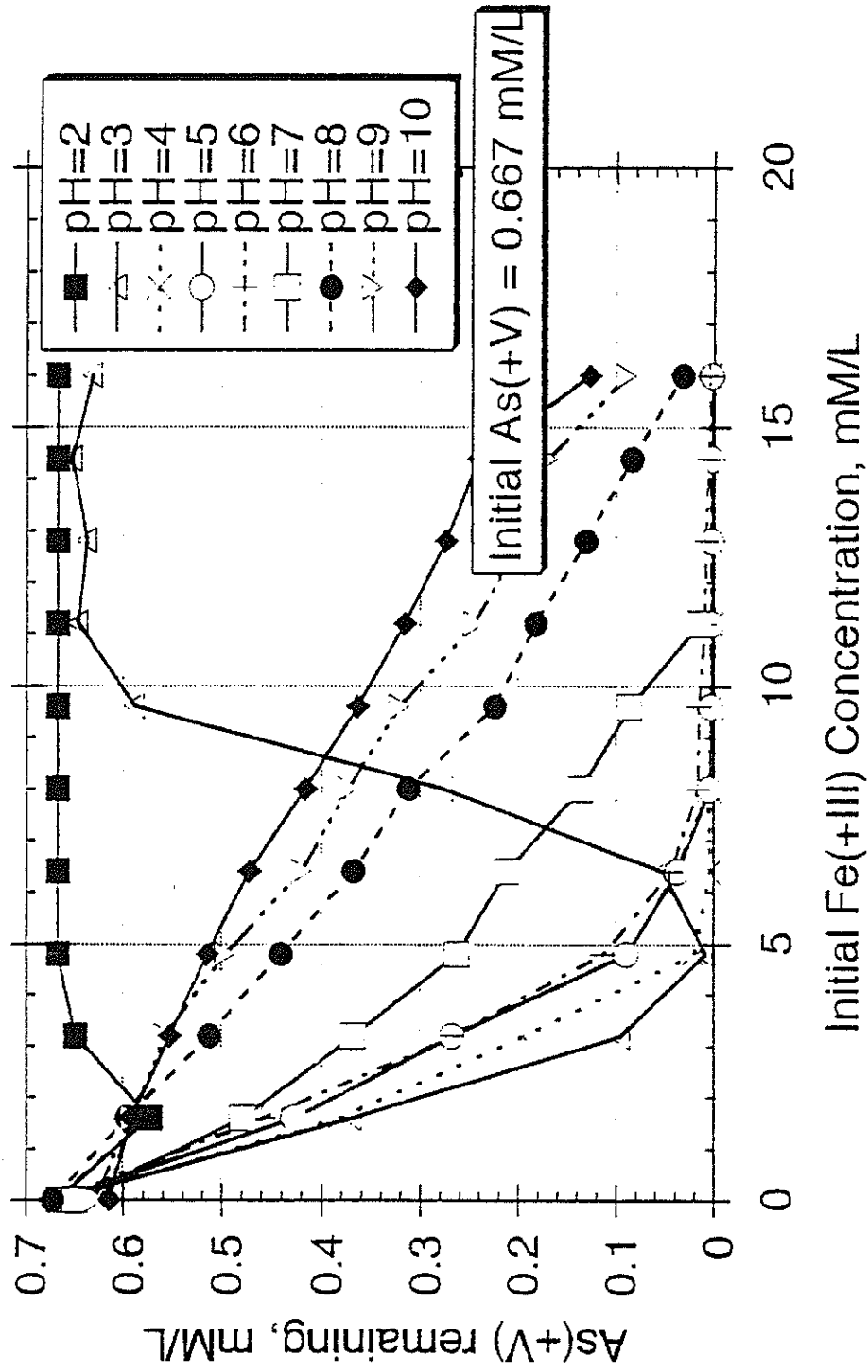


Figure 4. Effect of ferric iron dose on ferric arsenate solubility at various pH conditions



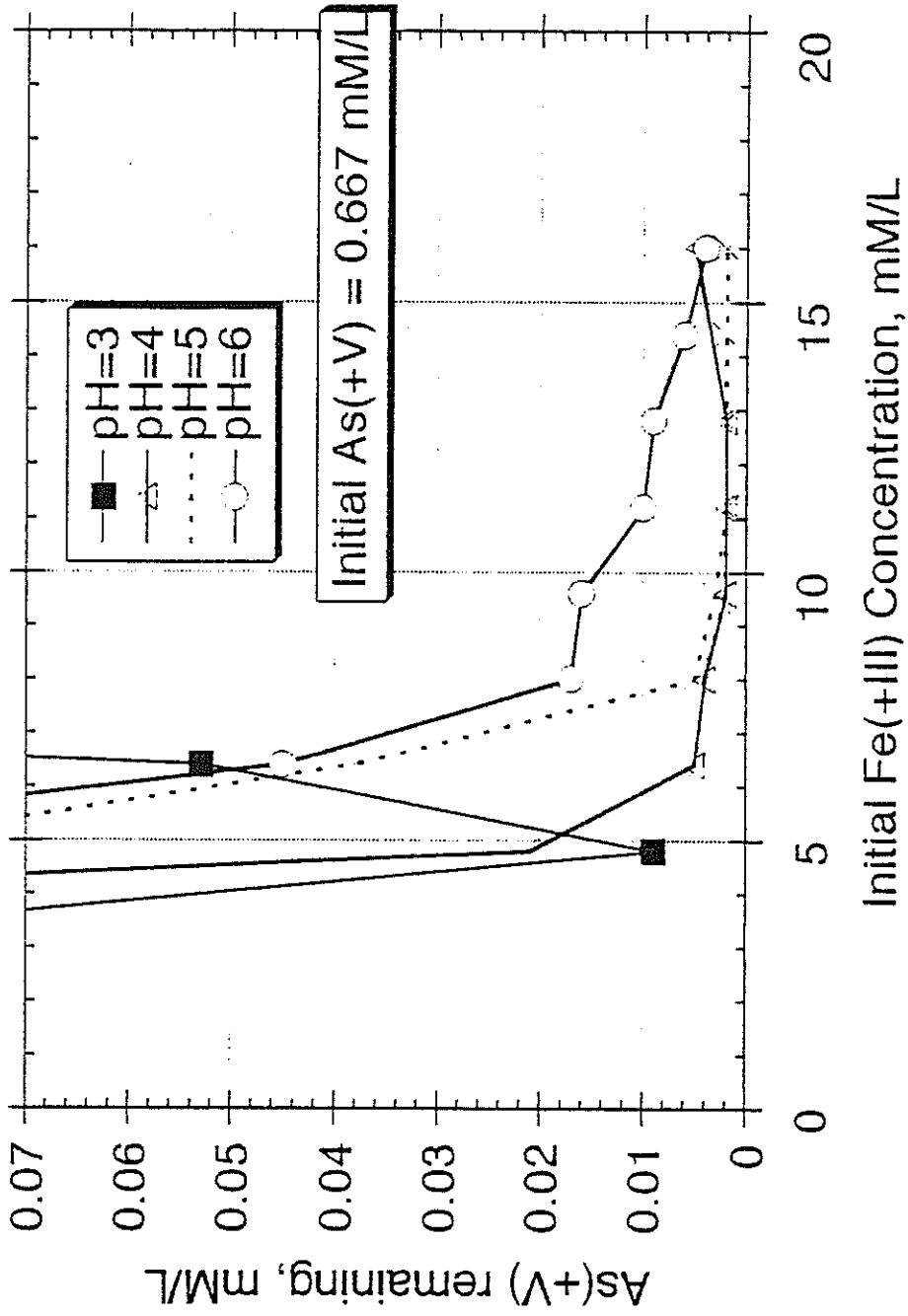


Figure 5. Detailed view of optimal arsenate removal zone.

Interestingly, Huang and Vane (1989) also discovered a 7 to 1 ratio between ferrous iron, Fe(II), and arsenic adsorbed with the activated carbon tests. It is possible that the Fe(II) might be oxidized to Fe(III) by the arsenate in their experimental work. This would lead to precipitation of both ferric arsenate and ferric hydroxide on the surface of the activated carbon. The experimental results, which show that maximum arsenic removal occurs around a pH of 3 to 4, also are in reasonable agreement with the theoretical predictions in Figure 3. The experimental results are certainly more compatible with this chemical model than with the theoretical model presented by Huang and Fu (1984) who propose a minimum solubility at a pH of 7.0 to 11.0. However, it is significant to point out that Huang and Fu (1984) assume a chemical model that requires a constant Fe(III) concentration. Such constraint is unlikely to occur in aquatic systems, where ferric hydroxide may control the solubility of ferric iron at concentrations much lower than the assumed value.

The minimum solubility predicted in the chemical model is rather conservative compared to the experimental results. The minimum solubility predicted in Figure 3 (about 0.03 M/L) is considerably higher than the experimental results at optimal conditions (0.002 mM/L = 0.15 mg/L as As). This discrepancy is probably due to arsenic coprecipitation with ferric hydroxide, which may improve arsenic removal. As stated before, the ferric iron to arsenic ratio should be 1 to 1 for pure ferric arsenate precipitation. However, 4.8 mM/L (268.1 mg/L) of iron are needed to remove 0.667 mM/L (50.0 mg/L) of arsenic at a pH of 3.0. Our measurements of soluble iron show that at this pH, 0.031 mM/L (1.73 mg/L) of iron remains in solution. This experimentally determined point corresponds well the theoretical solubility of 0.08 predicted in Figure 2. Using the mass balance principle, we conclude that the remaining 4.13 mM/L (230 mg/L) of the applied ferric iron precipitates as ferric hydroxide.

It is important to notice that according to Figure 3, the solubility of ferric arsenate increases by increasing the pH above neutrality. This may be explained by the limited amount of  $\text{Fe}^{3+}$  available in solution after ferric hydroxide precipitation. In fact total ferric iron solubility reaches a minimum condition at a pH of approximately 8.3 (Snoeyink and Jenkins 1980; see Figure 3). This condition coincides with the location of the zero point of charge (ZPC) for ferric hydroxide. Thus, positively charged iron species dominate at pH values below the ZPC point. Negatively charged iron compounds are predominant above 8.3. Since the arsenate species exist in anionic form (see Figure 1), the solubility of ferric arsenate should increase with increasing pH above 8.3 due to like charges repelling. It is also important to observe from Figure 4 that as the pH is increased, more iron is required for the same amount of arsenic removal. This suggests that ferric arsenate precipitation may not be the main removal mechanism at near-neutral pH conditions. Instead, the more dominant form of arsenic removal may be coprecipitation with the ferric hydroxide solid phase.

The sodium chloride and sodium sulfate experiments were conducted at a pH of 3.0 and 4.0 because of lower iron requirements and good arsenic removal in these conditions. The studies at a pH of 3.0 were performed using a 171.0 mM/L (10,000 mg/L) sodium chloride concentration. A similar experiment was performed using a 104.25 mM/L (10,000 mg/L) sulfate concentration. Figure 6 compares the results of these two experiments plotted against the findings in the

absence of these salts. Figure 6 shows that presence of high concentration sodium chloride enhances the removal of arsenates below the optimal iron dose at pH = 3.0. The presence of elevated sulfate levels decreases arsenic removal efficiency. However, this interference can be overcome by increasing the iron dose by a factor of 3. Similar patterns are noticed at pH = 4 as seen in Figure 7. The presence of NaCl in these conditions may hinder slightly the removal of arsenates from solution. It is also clear that the presence of sulfates present a significant interference to arsenate precipitation from solution. This interference can be overcome by addition of excess iron salt to the solution.

## **Conclusions**

Ferric arsenate precipitation plays an important role in arsenic removal in acidic conditions. However, arsenate coprecipitation with ferric hydroxide appears to be the main arsenic removal mechanism at near-neutral and acidic conditions. Arsenic removal by iron coprecipitation is very effective at a pH of 3.0, but removal at this pH is highly sensitive to iron concentration. The optimal iron concentration range is narrow and higher iron concentrations than the optimal amount hinders arsenic removal. At this pH, the initial 0.667 mM/L (50 mg/L) As(V) is reduced by 98.72 percent. Slightly better arsenic removal is obtained at a pH of 4.0, but more than twice the amount of iron is needed than at a pH of 3.0. The optimal iron concentration range is wider than at pH = 3.0, and arsenic removal is more tolerant of higher iron concentrations. At a pH of 4.0, 0.667 mM/L (50 mg/L) arsenic is reduced by 99.68 percent. The amount of iron needed for equivalent arsenic removal increases with pH. This holds true for the pH range of 3.0 to 10.0. The presence of elevated concentrations of sodium chloride has little effect on arsenic removal and actually widens the iron concentration tolerance range at a pH of 3.0. However, the presence of elevated sodium sulfate concentrations hinders significantly arsenic removal at both pH = 3.0 and 4.0.

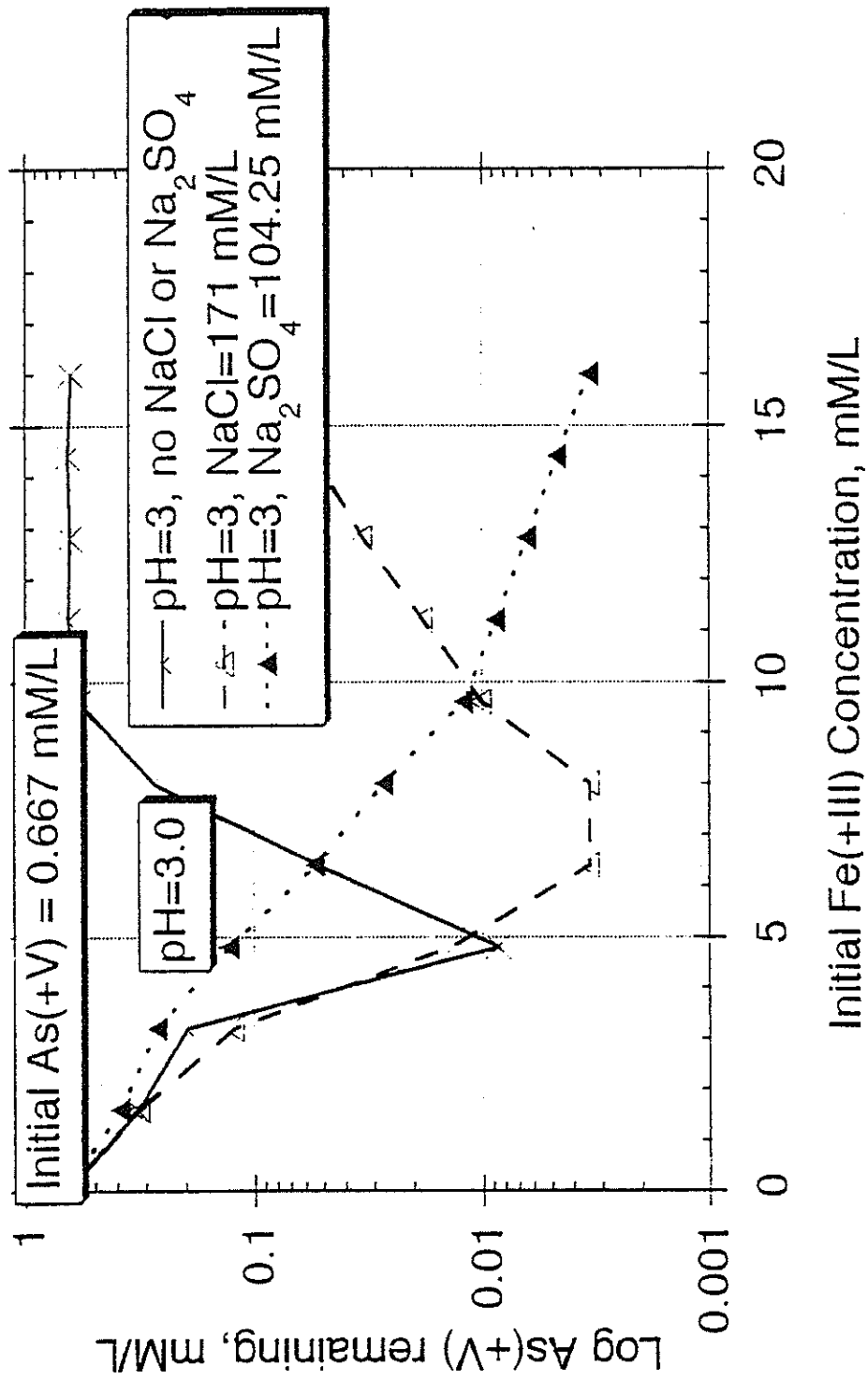


Figure 6. Effects of sodium chloride and sodium sulfate on ferric arsenate solubility at pH = 3.0

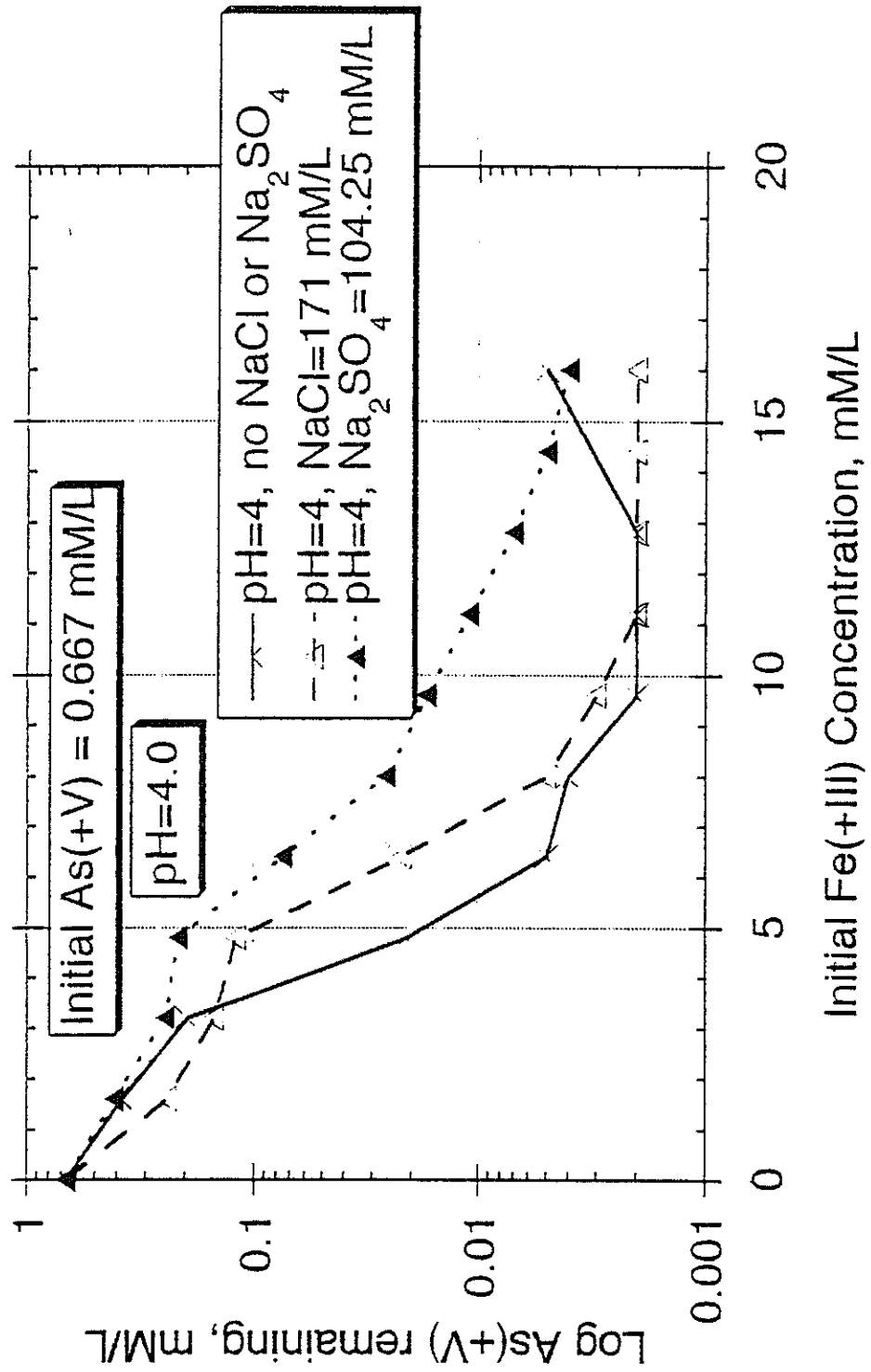


Figure 7. Effects of sodium chloride and sodium sulfate on ferric arsenate solubility at pH = 4.0

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