

# The evaluation of various oxidants used in acid leaching of uranium

R. VENTER\* and M. BOYLETT\*

*\*GRD Minproc Pty (Ltd), Bryanston, South Africa*

Uranium in the tetravalent state has very low solubility in acid and alkaline solutions and needs to be oxidized to the hexavalent state to become soluble. In acid leaching of uranium the uranium is oxidized by ferric iron, which in turn is reduced to ferrous iron. The reduced iron needs to be oxidized to the ferric form to allow the uranium dissolution reaction to proceed.

This paper considers the known and proven oxidants used in acid leaching of uranium. These include manganese dioxide, sodium chlorate, hydrogen peroxide, ferric iron, oxygen and sulphur dioxide/air (oxygen). The reaction chemistry of each oxidant as well as the practical applications are considered and discussed.

All the oxidants discussed work adequately in acid leaching of uranium and it is possible to engineer solutions for the use of the oxidants in acid leaching on plant scale. The availability and supply of the oxidant as well as the cost and environmental impact of the oxidant plays a major role in the selection of a suitable oxidant. The location of the plant will also have an impact on these issues.

## Introduction

Dissolution of uranium is the first hydrometallurgical process in the extraction of uranium from the uranium containing minerals in the ore. Leaching procedures selected for dissolving uranium are dependant in part on the physical characteristics of the ore such as: type of uranium mineralization, ease of liberation, and the nature of other constituent minerals present.

In nature uranium occurs in the tetravalent and hexavalent oxidation states. Tetravalent uranium has a low solubility in both dilute acid and carbonate (alkali) solutions. To achieve economic recovery of uranium in the tetravalent state, oxidation to the hexavalent state is essential. It is therefore important to maintain proper oxidizing conditions during leaching of these minerals to achieve high uranium extraction.

A number of oxidants have historically been used and are currently used to oxidize tetravalent uranium to hexavalent uranium in acid and alkaline circuits. When selecting an oxidant to be used on commercial scale, consideration must be given to a number of factors. Apart from the effectiveness of the oxidant to maintain an oxidizing environment, the availability and cost of the oxidant must be considered. Part of this is the logistics of getting the oxidant to site and administering it to the leach. The latter could introduce a number of

engineering challenges to the flowsheet.

This paper considers different oxidants typically used in acid leach uranium circuits, namely:

- Manganese dioxide ( $\text{MnO}_2$ ) as milled pyrolusite
- Sodium chlorate ( $\text{NaClO}_3$ )
- Hydrogen peroxide (addition as  $\text{H}_2\text{O}_2$  or Caro's acid ( $\text{H}_2\text{SO}_5$ ))
- Ferric iron, frequently as roaster calcine from a pyrite burning acid plant
- Oxygen in pressure leach circuits and
- Sulphur dioxide/air(oxygen) mixture.

The reaction chemistry and application of each oxidant, as well as engineering issues on plant scale and the availability and logistics are discussed.

## Mineralogy

Uranium in typical uranium bearing minerals occur in the tetravalent ( $\text{U}^{4+}$ ) and hexavalent ( $\text{U}^{6+}$ ) oxidation states. Primary uranium ores are predominantly in veins or pegmatites but are also found in sedimentary and placer deposits. Primary ores include uraninite and pitchblende ( $\text{UO}_2$ ) (Kinnaird and Nex, 2008).

Secondary uranium ores are found in weathered zones of primary deposits and precipitated in sediments. These include autinite ( $\text{Ca}(\text{UO}_2)_2(\text{PO}_4)_2 \cdot 10\text{H}_2\text{O}$ ), torbernite ( $\text{Cu}(\text{UO}_2)_2(\text{PO}_4)_2 \cdot 10\text{H}_2\text{O}$ ), uranophane ( $\text{Ca}(\text{UO}_2)_2\text{SiO}_3(\text{OH})_2 \cdot 5(\text{H}_2\text{O})$ ), carnotite ( $\text{K}_2(\text{UO}_2)_2[\text{VO}_4]_2 \cdot 3\text{H}_2\text{O}$ ) and schoepite ( $(\text{UO}_2)_8\text{O}_2(\text{OH})_{12} \cdot 12(\text{H}_2\text{O})$ ) (Yan and Connely, 2008).

Multiple oxides are complex associations of uranium with rare earths, tin, tantalum, niobium and titanium with extensive crystal lattice substitutions that makes them extremely refractory. These include davidite ( $(\text{La,Ce,Ca})(\text{Y,U})(\text{Ti,Fe}^{3+})_{20}\text{O}_{38}$ ), brannerite ( $(\text{U,Ca,Ce})(\text{Ti,Fe})_2\text{O}_6$ ) and betafite ( $(\text{Ca,U})_2(\text{Ti,Nb,Ta})_2\text{O}_6(\text{OH})$ ) and are more difficult to process, and products can pose environmental issues (Yan and Connely, 2008.)

The silicates, uranothorite ( $(\text{Th,U})\text{SiO}_4$ ) and coffinite ( $\text{U}(\text{SiO}_4)_{1-x}(\text{OH})_{4x}$ ), having the zircon structure, would be expected to be refractory, but in practice readily dissolve in oxidizing acid or carbonate solutions (Lunt, *et. al*, 2007.)

Uraninite and coffinite occur in the tetravalent state (Lunt, *et al.*, 2007). In the Southern African context examples of uraninite are the Witwatersrand ores in South Africa (Ford, 1993) and Rössing in Namibia (Kinnaird and Nex, 2008). Kayelekera in Malawi is an example of coffinite (Chilumanga, 2008).

Minerals occurring in the hexavalent state typically include carnotite found at Langer Heinrich and Trekkopje in Namibia (Kinnaird and Nex, 2008). In phosphate deposits uranium is found in the hexavalent form in the apatite structure. Examples of the phosphate deposits are Bakouma in the Central African Republic, Florida and other commercial phosphoric acid producers from which uranium is a possible by-product (Merrit, 1971:22). There are also minerals described as refractory, such as brannerite, which require more aggressive leach conditions, but the basic chemistry is unchanged (Yan and Connely, 2008).

## Simple chemistry

Uranium is typically recovered in acid circuits using sulphuric acid, or in alkaline circuits using a sodium carbonate bicarbonate mix. Acid leach is the preferred method, but in cases where the acid consumption is very high, mostly caused by gangue minerals, it makes more sense to consider alkaline leach. The selection of acid or alkaline leach is based on a number

of key process drivers. These are discussed elsewhere (Lunt *et al.*, 2007).

In typical extraction circuits, after dissolution, uranium is usually extracted using strong base resins and/or alamine 336 in a solvent extraction plant. Solvent extraction is used only for acid circuits (Lunt *et al.*, 2007). In these circuits the uranium is solubilized in its hexavalent form. This is true for both acid and alkaline circuits:

Hexavalent uranium dissolves as the  $\text{UO}_2^{2+}$  ion and produces complex uranyl anions:

Uranyl sulphate  $[\text{UO}_2(\text{SO}_4)_3]^{4-}$  and

Uranyl carbonate  $[\text{UO}_2(\text{CO}_3)_3]^{4-}$

The focus of this paper is on oxidation in acid circuits and to leave the alkaline circuitry as a separate issue. A typical uranium acid leach flowsheet can be seen in Figure 1.

The typical flowsheet comprises comminution to reduce the particle size to a size acceptable to the leach. A pre-concentration step could also be included where the ore is upgraded. The material is then leached in sulphuric acid at atmospheric or elevated pressure and ambient or elevated temperature. Leaching is followed by liquid/solids separation in countercurrent decantation (CCD) thickeners or on belt filters, followed by solvent extraction (or ion

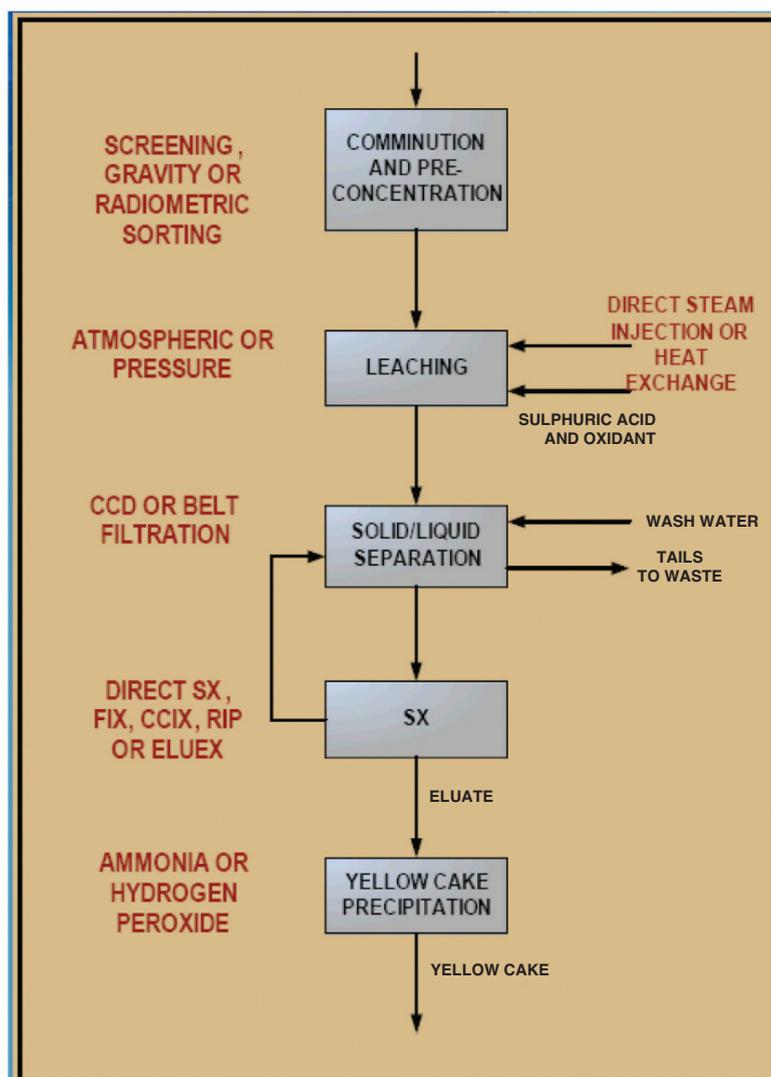


Figure 1. Typical acid leach flowsheet (Lunt and Holden, 2006)

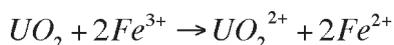
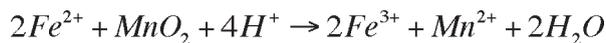
exchange or both) and finally product precipitation/recovery. Variations in ion exchange methodology include fixed bed ion exchange (FIX) and countercurrent ion exchange (CCIX). An alternative to liquid/solid separation is resin in pulp (RIP) the development of which will closely parallel the gold CIP designs.

### ***Uranium dissolution chemistry in acid***

Tetravalent uranium must be oxidized to the hexavalent form before dissolution occurs in sulphuric acid. Ferric iron ( $Fe^{3+}$ ) acts as the principal oxidant of tetravalent uranium in acid leaching circuits. Iron is generally present, either as a constituent in the ore, or introduced as metallic iron as a result of wear and abrasion in the grinding circuit. Iron can also be added on demand.

The mechanism for the dissolution of uranium in sulphuric acid is therefore based on the following reactions:

The  $Fe^{3+}$  oxidizes the tetravalent uranium ( $U^{4+}$ ) in the solution to the soluble hexavalent uranium ( $U^{6+}$ ). The  $Fe^{3+}$  is then reduced to  $Fe^{2+}$ . The oxidant is added to regenerate the  $Fe^{2+}$  to  $Fe^{3+}$  so that the leach reaction can continue. In practice it is found that the  $Fe^{3+}/Fe^{2+}$  mechanism is to be considered more than the uranium reaction. The mechanisms are described by the following reactions with  $MnO_2$  as oxidant (Merritt, 1971:64):



Thus, the  $Fe^{3+}/Fe^{2+}$  couple facilitates the leach process by acting as the electron carrier during oxidation.

The intention of this paper is to consider various ways of achieving the oxidation of ferrous to ferric iron in acidic conditions.

### ***Source of iron***

Iron needs to be present in the leach solution. If insufficient iron is present it can be added by adding ferrous or ferric sulphate or adding steel in some form. Iron can also be added via solution recycle streams from downstream processes.

### **Oxidants**

Following is a discussion of each of the oxidants typically used in acid leaching of uranium.

#### ***Manganese dioxide ( $MnO_2$ )***

This is the traditional oxidant used for Witwatersrand ores and elsewhere, e.g. Ranger in Australia. The oxidation takes place according to the following equation:



From the equation it is clear that each mole of  $MnO_2$  requires 2 moles of acid. Commercially available pyrolusite typically contains in the order of 30% to 50%  $MnO_2$  with the remainder potentially being acid consuming gangue minerals. These gangue minerals could also be a source of iron needed for the leach reaction to take place.

Manganese dioxide (as pyrolusite) was the first choice in the Witwatersrand as it was freely available as surface outcrops near to the mines and cheap sulphuric acid was locally produced from pyrite in the ore. Vaal Reefs South Uranium Plant operates with this exact flowsheet and has done so for almost 30 years.

In the 1950s through to 1990s in excess of 20 Witwatersrand uranium plants used pyrolusite as oxidant in the leach in preference to the auto oxidation process, which introduced SO<sub>2</sub> and air into the leach to oxidize the ferrous iron to ferric iron, as a result of cheap sources of pyrolusite being available (Ford, 1993). It is easy to mill and is added as a slurry under redox control. No special agitation is therefore necessary and the MnO<sub>2</sub> does not deactivate or dissociate in the slurry if it does not react with the iron immediately, as is the case with some of the other oxidants (e.g hydrogen peroxide).

At Randfontein, Cooke Plant, the nearest supply was from a small pit at the top of the access road. Today most of this pit is surrounded by housing.

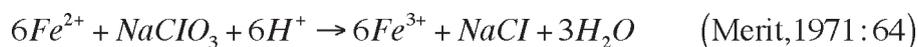
The apparent lack of availability and environmental considerations weigh against the use of pyrolusite today. South Africa is a major exporter of manganese ores from the Northern and Eastern Cape and a producer of electrolytic manganese dioxide. All of these products would be suitable but with increasing distances from site transport costs increase. Outside the Witwatersrand, for example in Namibia, the most likely source of supply is from Ghana. It is practical via the port of Walvis Bay, but also expensive.

The environmental aspect involves the production of Mn<sup>2+</sup> in solution. In theory this will be deposited on the slimes dam as the hydroxide. In practice, the pH to achieve this is high and must be maintained, which equates to high neutralization costs. Pyrite oxidation will lower the dam pH in the long term, thus soluble manganese may enter water tables and water courses. This is a greater problem on the Witwatersrand than in Namibia for instance. Namibian ores are essentially pyrite free and the climate is dryer, compared to the pyrite containing ores in a wetter climate on the Witwatersrand. The higher pH needed to precipitate the manganese still holds true for Namibia.

The relatively low MnO<sub>2</sub> concentrations in the pyrolusite cause relatively high volumes of pyrolusite to be shipped to site, which introduces logistical challenges. The acid consuming gangue minerals in the pyrolusite can also significantly increase acid consumption in the leach.

### ***Sodium chlorate***

Sodium chlorate was traditionally the first choice oxidant in North American uranium plants. The oxidation reaction takes place according to the following equation:



From the equation it is clear that each mole of sodium chlorate requires 3 moles of acid.

In Southern Africa sodium chlorate was not considered as an oxidant due to availability of cheaper and more convenient alternatives (i.e. pyrolusite).

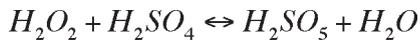
Sodium chlorate is added to the leach slurry as a solution. No special agitation is therefore necessary. During the reaction with sodium chlorate in the leach, chloride ions go into solution with the possible adverse effects on ion exchange loading and materials of construction.

Sodium chlorate is expensive. It is also used as oxidant in PGM refining. Partial reduction leads to intermediate products, such as chlorite and hypochlorite. Sodium chlorate powder has an auto-ignition risk when in contact with organic. It therefore needs procedures and proper engineering for safe handling.

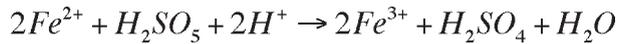
### ***Hydrogen peroxide (H<sub>2</sub>O<sub>2</sub>)***

Hydrogen peroxide was used as oxidant in acid leaching in the nineteen eighties. Lucas *et al.* (1983) describe plant trials performed in Queensland where pyrolusite was replaced by Caro's acid as oxidant. De Vries (1984) patented a process that utilizes hydrogen peroxide in an acid leach of uranium ore.

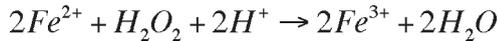
Hydrogen peroxide can be used as is or can be made up as Caro's acid or peroxymonosulphuric acid ( $H_2SO_5$ ) by mixing with concentrated sulphuric acid according to the following reaction:



The oxidant regenerates the ferric iron ( $Fe^{3+}$ ) from ferrous iron ( $Fe^{2+}$ ). In the case of Caro's acid the reaction is as follows:



In the case of hydrogen peroxide the reaction is as follows:



From these equations it is clear that one mole of acid is consumed for every one mole of Caro's acid or hydrogen peroxide.

Lucas *et al.* (1983) reported similar uranium extractions to using pyrolusite with a reduction in acid consumption and oxidant consumption. They also reported a reduction in lime consumption for neutralization of acid effluent but this could not be fully quantified. Mention was also made that excluding manganese from the leach liquors should reduce the neutralization requirement for the manganese and pyrolusite gangue minerals. Test work performed by GRD Minproc and Mintek supports the reduction in acid and oxidant consumption.

From the plant trial and other operations the main benefits of using hydrogen peroxide (or Caro's acid) as oxidant were claimed to be:

- Better process control
- Oxidant savings
- Sulphuric acid savings
- Lime savings
- Cleaner oxidant handling
- Simplified effluent treatment.

On the downside, hydrogen peroxide will dissociate if it does not react rapidly with the iron in solution. The dispersion of the hydrogen peroxide through the slurry is therefore very important. It is reported that difficulties with dispersion of the hydrogen peroxide in the leach slurry was experienced during scale-up to plant scale at Rössing and Western Areas (Boylett, 2008). It is also reported that using Caro's acid will reduce the hydrogen peroxide consumption as a result of more efficient use of the hydrogen peroxide (James, 2008).

Hydrogen peroxide can also be used for precipitation of uranium as uranyl peroxide ( $UO_4 \cdot xH_2O$ ) instead of the more traditional ammonium diuranate (ADU), which eliminates the use of ammonia as a precipitation reagent. Uranyl peroxide is also reported to be easier to thicken and filter or centrifuge and requires only drying and not calcining for dispatch. It is also reported to produce a purer product. This is the subject of another paper but it is sufficient to say that this product is the more common end product of an alkaline leach circuit than from an acid circuit. It is the end product from Langer Heinrich for example. As hydrogen peroxide becomes more popular for precipitation in acid circuits this might change.

Hydrogen peroxide was often discounted based on its relatively high price. It is very favourable from an environmental perspective as the reaction chemistry produces only water.

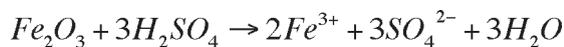
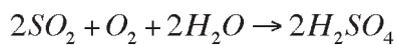
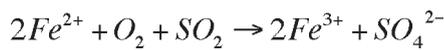
Production of Caro's acid was traditionally a challenge as a result of high heat generation by the reaction between the acid and the hydrogen peroxide. Manufacturers of hydrogen peroxide claim that they now have simple methods of producing 'Caro's acid' for use in the uranium

leach circuit (James, 2008). It is now made *in situ* by feeding sulphuric acid and hydrogen peroxide into a funnel device. The high heat of reaction is immediately dissipated into the slurry below.

For those gold plants using hydrogen peroxide for cyanide destruction the storage, use and availability of peroxide is an advantage. The greater use of peroxide by the gold plants enhances its applicability and availability in remote areas. Hydrogen peroxide is supplied in large volumes to a number of operations throughout Africa, for instance gold producers in Ghana. Suppliers maintain that availability and logistics of hydrogen peroxide is not a problem.

### ***Ferric iron from roaster calcine***

Calcine from the roasting of pyrite, which has been treated with concentrated sulphuric acid, is also used as oxidant in acid leaching of uranium. The resulting slurry is combined with the return filtrate and contacted with sulphur dioxide in flotation cells. The auto-oxidation reaction results in the conversion of ferrous iron to ferric iron with the formation of additional sulphuric acid. The process is described by the following reactions (Ford, 1993):



This process not only produces a concentrated source of ferric iron, but also liberates any uranium and gold present in the calcine for subsequent recovery.

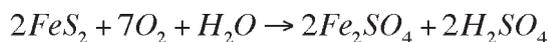
A number of plants practise milder forms of calcine digestion by adding sulphuric acid to the calcine slurry in air-agitated pachucas at between 70°C and 90°C. Pyrolusite is also added to maximize the concentration of ferric iron. This practice is followed by Rössing and was followed by Hartebeesfontein uranium plant amongst others.

### ***Oxygen***

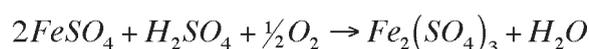
Uranium ores containing sulphidic minerals can be leached by the addition of oxygen only at elevated temperature and pressure. The addition of sulphides to ores that do not contain sulphidic minerals has also been proposed.

Sulphuric acid and ferric sulphate are generated *in situ* during pressure leaching by the reaction of oxygen with the sulphides. If pyrite and uraninite are present the following reactions might be expected:

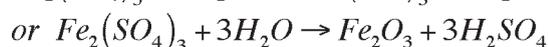
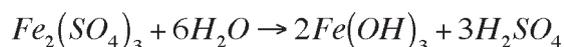
Pyrite is oxidized to produce soluble iron and acid:



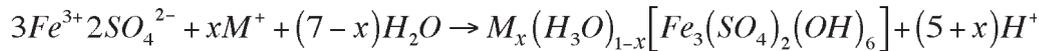
Ferrous iron is oxidized to ferric iron by oxygen:



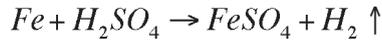
As the temperature increases above 170°C ferric iron is removed from the solution through iron hydrolysis reactions:



Jarosite also forms during the hydrolysis reaction as described by the following general reaction (Guise and Castro). *M* indicates the cation.

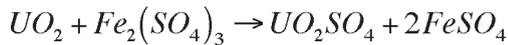


An additional source of dissolved iron is from steel from the crushing and grinding plants:

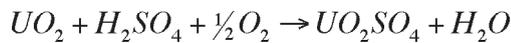


The following reactions represent the dissolution of uranium:

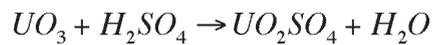
Tetravalent uranium is oxidized by ferric iron:



The dissolution of tetravalent uranium by oxygen:



Hexavalent uranium will dissolve in the presence of sulphuric acid alone:



Acid consuming gangue, including carbonates, is leached by acid:



These reactions proceed more rapidly to completion at high temperature and with the increased solubility of oxygen in solution at increased pressure. Temperatures of above 170°C are suggested. Oxygen consumption is in the order of 2 tons of oxygen per ton of sulphur.

Pressure leaching is utilized in the treatment of refractory ores and is performed in autoclaves at elevated temperature and pressure. Autoclaves were trialled at the Western Deep Levels plant in the late '70s (Bovey and Stewart, 1978). Acid leaching in autoclaves, based on these plant trials, were used most recently in the design of the autoclaves for the Dominion Reefs plant owned by Uranium One.

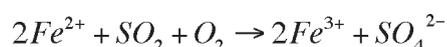
Indicated advantages of the pressure leaching process include: improved extraction, decreased operating costs, a decreased amount of impurities and free acid in the leach solutions and improved slurry filtration properties. Pressure leaching also results in increased recovery of gold and uranium from the pyrite.

Disadvantages include increased corrosion and maintenance. Another disadvantage is the production of soluble SiO<sub>2</sub> that contaminates resin in ion exchange circuits and forms crud in solvent extraction circuits.

### ***Sulphur dioxide/air (SO<sub>2</sub>/air)***

It is a well known fact that a mixture of SO<sub>2</sub> gas and air acts as an oxidant in hydrometallurgy (Ho and Quan, 2003). It is used in iron and manganese removal in cobalt circuits, although it has not been proven commercially yet. It was first suggested by workers in the USA and the operating conditions and detailed chemistry was determined by the Government Metallurgical Laboratories (GML) in South Africa in the 1950s (Ford, 1993). It has also been suggested for use in leaching of uranium in an oxidizing sulphuric acid solution as an alternative to MnO<sub>2</sub>. (Ho and Quan, 2007).

In the case of SO<sub>2</sub>/air mixture as oxidant, oxygen in the air together with the SO<sub>2</sub> is responsible for oxidizing the Fe<sup>2+</sup> to Fe<sup>3+</sup> according to the following reaction:



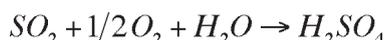
Both oxygen and SO<sub>2</sub> are required in solution for the reaction to proceed, but due to the lower solubility of oxygen compared to SO<sub>2</sub>/O<sub>2</sub> mass transfer limits the maximum reaction rate. The Fe<sup>2+</sup> oxidation rate is controlled by the SO<sub>2</sub>/O<sub>2</sub> ratio and the rate of oxygen mass transfer. It is independent of Fe<sup>2+</sup> concentration.

Higher oxidation rates can be obtained in solutions compared to slurries, as the solids in the slurries decreases the rate of oxygen transfer. Ho and Quan (2007) focused on producing a ferric solution that can be added to the leach slurry rather than *in situ* addition of SO<sub>2</sub>/air to the leach.

Earlier work by Ho and Quan used sodium sulphite as the source of SO<sub>2</sub>; however, at a metallurgical plant direct sparging of SO<sub>2</sub> would be preferred. The measured Fe<sup>2+</sup> oxidation rates using SO<sub>2</sub> were very similar when sodium sulphite was used compared to SO<sub>2</sub> gas.

Ho and Quan found that the SO<sub>2</sub>/O<sub>2</sub> volumetric ratio plays a major role in the oxidation rate. At relatively low SO<sub>2</sub>/air ratios where there is an excess of oxygen in the solution, the reaction rate is proportional to the SO<sub>2</sub> dissolution rate into the solution. However, once the SO<sub>2</sub> flow rate exceeds the maximum rate of oxygen mass transfer, the Fe<sup>2+</sup> oxidation rate decreases as a result of reducing conditions caused by SO<sub>2</sub>.

The main contributor to the reduced SO<sub>2</sub> efficiency is the side reaction of SO<sub>2</sub> to produce sulphuric acid, according to the following reaction:



A SO<sub>2</sub>/O<sub>2</sub> ratio of greater than 0.36 caused the Fe<sup>2+</sup> oxidation rate to become variable. Unlike most of the other oxidants, SO<sub>2</sub>/air is not an acid consumer.

From a large-scale implementation point of view, SO<sub>2</sub>/air has a number of challenges. For slurries dispersion agitation is necessary with more installed and utilized power. If an acid plant is running on site, SO<sub>2</sub> can be sourced from the acid plant. Care must, however, be taken to ensure that the SO<sub>2</sub>/air ratio from the burning sulphur is satisfactory, as this gas stream might have a higher nitrogen content compared to when pure SO<sub>2</sub> gas is mixed with air. Any SO<sub>2</sub> escaping from the leach tanks needs to be collected and passed through a scrubbing system. The relatively low SO<sub>2</sub> concentration in the gas stream can also give rise to large gas flow rates into the leach.

Introduction of the SO<sub>2</sub>/air mixture into the leach tanks also presents a challenge as it will probably need to be introduced against high slurry heads. This can influence the partial pressures of the different gases in the gas mixture. The solubility of SO<sub>2</sub> is higher than oxygen and if the SO<sub>2</sub>/oxygen ratio in the solution gets too high it will cause reducing conditions and affect the reaction kinetics.

Oxidizing iron in a solution stream added to the leach will be easier to achieve. Care should, however, be taken to ensure that sufficient iron is oxidized to the ferric state to oxidize the uranium.

As far as is known to the authors, at the time of writing this paper, no commercial operation exist that uses SO<sub>2</sub>/air as oxidant in acid leaching of uranium.

## Conclusion

Traditionally the tendency in Southern Africa was to use pyrolusite as oxidant in acid leach of uranium. Availability and logistical issues as well as environmental issues seem to be changing this tendency.

SO<sub>2</sub>/air has been tested on many occasions and has proved to work satisfactorily; however, administering the SO<sub>2</sub>/air gas mixture to the leach on large scale presents a number of challenges. SO<sub>2</sub> is an attractive option if acid is produced on site and if sulphur prices are low. It also reduces the number of reagents required to be shipped to site.

Hydrogen peroxide is also an alternative. It has no environmental effects and leads to a reduction in acid and oxidant consumption compared to pyrolusite. Handling and transport of hydrogen peroxide, as well as availability, may still be an issue. Hydrogen peroxide might also become costly compared to the other oxidants.

Sodium chlorate is not really considered as a result of introduction of chlorides into the system and the adverse effect it has on ion exchange resins and corrosion of materials of construction. Sodium chlorate also has an explosion risk.

All the oxidants described work adequately in acid leaching of uranium. It is also possible to engineer solutions to introduce the oxidant into the leach slurry on plant scale. The factors that play a significant role in selecting a suitable oxidant will include the availability and supply of the oxidant, the cost of the oxidant and the environmental impact of the oxidant. All three of these issues will be affected by the location of the plant and possible sources of oxidant close to the operation.

### Acknowledgements

The authors would like to thank the management of GRD Minproc (Pty)Ltd for permission to publish this paper and to acknowledge the input of their colleagues in undertaking uranium projects and uranium feasibility studies.

### References

- BOVEY, H.J. and STEWART, L.N. Pressure Leaching of Uranium Bearing Witwatersrand Ores, SAIMM, 22 March 1978.
- BOYLETT, M. Personal communication. 2008.
- CHILUMANGA, P. An Overview of Uranium Mining Development in Malawi, Implementation of Global Best Practice in Uranium & Processing IAEA HQs – VIENNA. 2008.
- DEVRIES, F.W. Hydrogen Peroxide in Sulfuric Acid Extraction of Uranium Ores, United States Patent 4,425,307, January, 10, 1984.
- FORD, M.A. Uranium in South Africa, *Journal of the South African Institute of Mining and Metallurgy*, vol. 93, no. 2, 1993. pp. 37–58.
- GUISE, L. and CASTRO, F. Purification of Sulphuric Acid Solutions from the Leaching of Nickelliferous Laterites, University of Minho, Department of Mechanical Engineering, Azurem – 4800 Guimaraes - Portugal.
- HO, E.M. and QUAN, C.H. Iron(II) Oxidation by SO<sub>2</sub>/O<sub>2</sub> in Uranium Leach Solutions, ANSTO, Australia. *Hydrometallurgy* 2003.
- HO, E.M. and QUAN, C.H. Iron(II) oxidation by SO<sub>2</sub>/O<sub>2</sub> for use in uranium leaching, *Hydrometallurgy*, vol. 85, 2007. pp. 183–192.
- JAMES, A. Sales and Technical Marketing Manager – Exports, Solvay SBU H<sub>2</sub>O<sub>2</sub>, Email communication. 2008
- KINNAIRD, J.A. and NEX, P. Assessing the Geological Occurrences of Uranium in Africa Implications for Mining and Processing, *Uranium Mining and Exploration Conference*, Johannesburg, March 2008.
- LUNT, D., BOSHOFF, P., BOYLETT, M., and EL-ANSARY, Z. Uranium Extraction The Key Process Drivers, The Southern African Institute of Mining and Metallurgy – Namibian Branch – *Uranium in Namibia*. 2007.

LUNT, D. and HOLDEN, A. Uranium Extraction The Key Issues for Extraction, *ALTA Uranium Conference*, Perth, Western Australia, May 2006,

LUCAS, G.C., FULTON, E.J., VAUTIER, F.E., WATERS, D.J., and RING, R.J. Queensland Mines Plant Trial with Caro's Acid, *Proc. Australas. Inst. Min. Metall.*, 287, September 1983.

MERRITT, R.C. The Extractive Metallurgy of Uranium, Colorado School of Mines. 1971.

YAN, D. and CONNELLY, D. Implications of Mineralogy on Uranium Ore Processing. *ALTA 2008 Uranium Conference*, Perth, Western Australia, June 2008.

---



**Riaan Venter**

*Senior Process Engineer, GRD Minproc, South Africa*

Riaan has 10 years of metallurgical and process control experience in the minerals processing industry. His experience includes various positions at Anglo Platinum where he started his career in 1998 where he gained production and technical experience on precious metals refining, and research, design and optimization experience on various concentrator plants, smelters and refineries. He moved to GRD Minproc in 2007 as a senior process engineer and has been involved in projects and feasibility studies for platinum, gold and uranium extraction facilities as well as techno economical evaluations of projects. Riaan is registered as a Professional Engineer at ECSA.

---

