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Problems and Proposed Solutions in Predicting Acid Drainage with Acid-Base Accounting

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ABSTRACT

Acid-Base Accounting (ABA) is reviewed from the perspectives of theory, application, and analytical procedures, highlighting the problems that can arise in the interpretation of Acid Potential and Neutralization Potential as well as practical problems involving the spatial and temporal dimensions. Problems include the conversion of acid-generating sulfur in a sample to Acid Potential through a conversion factor of 31.25 and the soaking of a sample in excess acid in order to determine the environmentally relevant content of Neutralization Potential. An enhanced version of ABA that would overcome many of these problems includes detailed geochemical evaluation of a sample and the interpretation of ABA results in light of the anticipated environmental conditions to which the sample will be exposed at a site.

1. INTRODUCTION

A relatively inexpensive and rapid test for acid-drainage potential is known as "acid-base accounting" (ABA) which measures the balance between acid-generating and acid-neutralizing components in a sample of rock or tailings (Coastech Research Inc., 1989; Everett, 1985; Caruccio, 1986). In its simple form, the test involves a measurement of total sulfur or sulfur forms which are

expected to generate acidity and a measurement of acid-neutralizing capability. When the measurements are converted to consistent units (e.g., tonnes of CaCO_3 /1000 tonnes of rock or tailings), the sample is declared to be potentially acid generating whenever the acid-generating content exceeds the acid-neutralizing capability.

Acid-base accounting has proven to be a valuable tool in acid-drainage assessments because dozens to hundreds of samples can be analyzed for relatively little cost compared to kinetic tests such as leach columns and humidity cells. However, ABA has been found to be inaccurate in that there are reports that samples predicted to generate net acidity by ABA have not done so and samples predicted to remain pH-neutral have eventually generated net acidity.

There are myriad explanations for why ABA has not accurately predicted the onset of acid drainage. These explanations, which are discussed in this paper, show that the failure of ABA can, in fact, be often attributed to improper interpretation and application of the results rather than the explicit failure of the test procedure itself. Interpretation of ABA analyses is best performed on a site-specific basis and may not necessarily be extrapolated to other sites.

Acid generation and neutralization are highly complex processes, affected by physical, geochemical, and biological reactions (e.g., Morin et al., 1990). In light of this complexity, the results of ABA analyses should not be interpreted through a simple comparison of acid-generating and acid-neutralizing values. The following sections will examine the factors of acid potential and neutralization potential by discussing the implications of theory, analytical procedures, and interpretation. This approach will provide guidelines for interpreting ABA analyses on a site-specific basis.

2. ACID POTENTIAL

One of the primary goals of acid-base accounting is to delineate the acid-generating potential of a sample. This is usually accomplished by an analysis for total solid-phase sulfur in the sample. In reality, a sample might contain several forms of sulfur (e.g., Tuttle et al., 1986; Goldhaber, 1983; Ammons and Shelton, 1984), including:

- o sulfide
- o elemental sulfur
- o polythionates, thiosulfate, sulfite, and other moderately reduced forms
- o sulfate
- o organically bound sulfur

Interpretation of ABA results is performed best on the basis of the relative acid potential of each form. Sulfate is often a product of acid generation but is not acid-generating in itself. For organically bound sulfur, published literature is contradictory on its acid-generating ability, perhaps because the oxidation of organic matter may generate acidic conditions as low as approximately pH 3 independent of sulfur oxidation. The remaining forms of sulfur may be considered acid-generating or acid-consuming as explained below. Consequently, separate analyses for at least sulfate followed by subtraction of sulfate from total sulfur will provide a more reliable estimate of acid-generating sulfur. The value of this refinement is obvious for samples which contain high concentrations of sulfate minerals such as gypsum (hydrated calcium sulfate). Separate analyses for the other forms may improve the accuracy of ABA results.

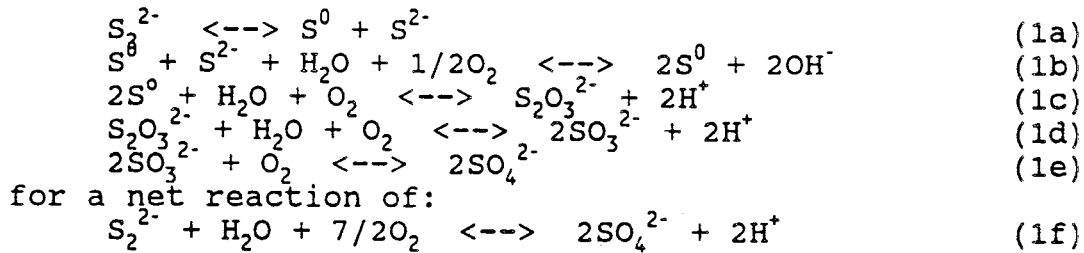
2.1 Stoichiometry and Mineralogy of Acid Generation

The reduced forms of sulfur can be considered acid-generating or acid-consuming depending on the original form and the extent of oxidation (Table 1). Analyses for all sulfur forms (Goldhaber, 1983) and delineations of oxidation pathways would provide an ultimate assessment of potential acidity, but would defeat the goal of a rapid and inexpensive test for acid drainage. Instead, an assumption is implicitly made in ABA that all sulfur exists as S_2^{2-} and oxidizes completely to sulfate (Equation 1f, Table 1).

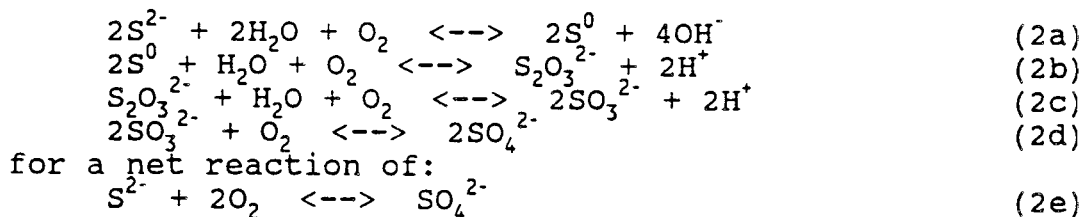
Based on the sulfur content judged to be potentially acid generating, the ABA procedure converts this sulfur to an equivalent $CaCO_3$ content for ease of comparison with neutralization potential. For conversion of sulfur as %S to tonnes $CaCO_3$ /1000 tonnes of sample (or kg $CaCO_3$ /tonne of sample or parts per thousand), the sulfur value is multiplied by 31.25. This conversion factor of 31.25 is rarely explained and justified.

TABLE 1--Oxidation sequence of sulfur originating as S_2^{2-} and S^{2-} , passing through selected intermediate inorganic species, to sulfate

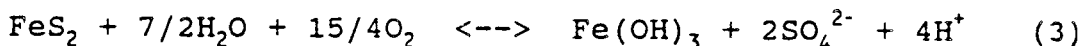
Oxidation of S_2^{2-}



Oxidation of S^{2-}



The conversion factor for sulfur to equivalent $CaCO_3$ is derived from the "standard" equation for pyrite oxidation:



Based on this equation, 4 moles of H^+ (acidity) are generated for each mole of pyrite, or 2 moles of H^+ for each mole of S (2:1). With atomic weights of sulfur and $CaCO_3$ (assuming at this point that one mole of H^+ is comparable to 1 mole of $CaCO_3$), the mole ratio of 2:1 becomes a weight ratio of 6.250:1 (6.250% $CaCO_3$ represents 1.000% S). The assumption is then made that each part of $CaCO_3$ is equivalent to 2 parts of H^+ , yielding a ratio of 3.125% $CaCO_3$:1.000% S. Finally, the conversion of $CaCO_3$ from percent (parts per hundred) to tonnes of $CaCO_3$ /1000 tonnes of sample or kg $CaCO_3$ /tonne of sample (parts per thousand) provides the ratio of 31.25:1.

The remainder of this section will focus on the implicit assumptions in Equation 3, which are so restrictive that values greater than and less than 31.25 may be more realistic in many cases. The additional assumption that 1 part of $CaCO_3$ is equivalent to 2 part of acidity will be discussed in Section 3.

For the factor of 31.25 to be valid at a particular

field site, specific geochemical and environmental assumptions must be met (Table 2). In the event that all of these assumptions are not met, other conversion factors may be warranted.

TABLE 2--Assumptions of geochemistry and environmental conditions that are required to validate the 31.25 ABA conversion factor

Assumptions:

- 1) sulfur occurs only as S_2^{2-} ,
 - 2) S_2^{2-} oxidizes completely to sulfate,
 - 3) pyrite is the only sulfide mineral,
 - 4) molecular oxygen and water are the only oxidants,
 - 5) all iron oxidizes to the ferric state, and
 - 6) all iron precipitates from solution in the form of $Fe(OH)_3$.
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In situations where only the first two assumptions of Table 2 are not met, the discussion surrounding Table 1 is relevant. As a result, the appropriate conversion factors range from 0.0 to 31.25 (Table 3). However, there is evidence that bacterial participation in sulfur oxidation may involve the consumption of the electrons in the redox reactions (Toran and Harris, 1989; Morin et al., 1990), which would change the stoichiometry of Equations 1 and 2. For example, if the electrons are removed by bacteria during the operation of Equation 1b, then the two hydroxide ions may not be formed. As a result, the net reaction (Equation 1f) would produce 4 moles of acidity for two moles of sulfur. Under experimental conditions, this might be misinterpreted as an increase in the rate of oxidation rather than a change in stoichiometry.

TABLE 3--Alternative conversion factors of sulfur to potential acidity (Assumptions 3-6 of Table 2 apply; 1 part of $CaCO_3$ is assumed to be equivalent to 2 parts of acidity; No bacterial effects on stoichiometry are considered)

<u>OXIDATION OF S_2^{2-} TO:</u>	<u>CONVERSION FACTOR BASED ON THE OXIDATION PATHWAYS OF TABLE 1</u>
S^0	0.0
$S_2O_3^{2-}$	15.625
SO_3^{2-}	31.25
SO_4^{2-}	31.25

At this point, the impacts of the other assumptions in Table 2 on the conversion factor will be examined. For simplification, the following discussion assumes that full sulfur oxidation to sulfate occurs in all cases. In contrast to Assumption 6, pyrite oxidation might not result in the precipitation of ferric iron if pre-existing conditions are already acidic (approximately pH < 2.5). In such situations, the appropriate conversion factor would be only 7.81 (Table 4). In geochemically reducing zones such as groundwater systems, there may even be negligible oxidation of the pyrite-derived iron (Assumption 5), leading to an appropriate conversion factor of 15.63.

In natural environments, there are often oxidants other than oxygen available (Swider and Mackin, 1989), in contradiction of Assumption 4. Oxidation by manganese, for example, would generate relatively large amounts of acidity, justifying a conversion factor of 125.00 if the resulting iron did not oxidize and precipitate (Table 4). Alternatively, oxidation of the iron without precipitation would reflect a factor of 117.19 whereas both oxidation and precipitation would indicate a factor of 140.63.

Oxidation of pyrite by ferric iron is a more complex case than by manganese because the total acidity generated by one mole of pyrite is dependent on the source of pyrite. If the ferric iron is derived from "standard" pyrite oxidation, then Equation 3 must operate 14 times for each operation of Equation 4 (Table 4). Consequently, the bulk acid generation is 30 H⁺: 30 S, representing a conversion factor of 15.63 (based on no oxidation or precipitation of the product iron). Alternatively, if the ferric iron is derived from a pre-existing ferric source, the relevant factor would be 125.00.

In contrast to Assumption 3 of Table 2, many other sulfide minerals occur in nature in combination with, or independent of, pyrite. These other minerals require values of the conversion factor that may differ from those of pyrite. For example, chalcopyrite oxidation with oxygen and water only requires a factor of 15.63 (Equation 8, Table 4) if oxidation were to occur in the pH range of approximately 2.5-5.0 so that the oxidized ferric iron precipitated and the copper remained in solution. In other ranges of pH and with other ions in the water, alternative factors would be warranted. As another example, galena is not expected to generate acidity under the sulfur-oxidation pathway of Table 1 if the lead remains in solution as the free ion (Equation 9, Table 4). On the other hand, precipitation of lead from solution as a simple hydroxide mineral would then

indicate acid generation should be expected with a factor of 31.25, equivalent to standard pyrite oxidation.

TABLE 4--Alternative conversion factors of sulfur to potential acidity (All sulfur is assumed to oxidize to sulfate and 1 part of CaCO₃ is equivalent to 2 parts of acidity)

		<u>CONVERSION FACTOR</u>
<u>"Standard" Pyrite Oxidation</u>		
$\text{FeS}_2 + 7/2\text{H}_2\text{O} + 15/4\text{O}_2 \leftrightarrow \text{Fe}(\text{OH})_3 + 2\text{SO}_4^{2-} + 4\text{H}^+$	[3]	31.25
<u>No Precipitation of Ferric Iron</u>		
$\text{FeS}_2 + 1/2\text{H}_2\text{O} + 15/4\text{O}_2 \leftrightarrow \text{Fe}^{3+} + 2\text{SO}_4^{2-} + \text{H}^+$	[4]	7.81
<u>No Oxidation and Precipitation of Iron</u>		
$\text{FeS}_2 + \text{H}_2\text{O} + 7/2\text{O}_2 \leftrightarrow \text{Fe}^{2+} + 2\text{SO}_4^{2-} + 2\text{H}^+$	[5]	15.63
<u>Manganese Replacing Oxygen as an Oxidant; No Oxidation and Precipitation of Iron</u>		
$\text{FeS}_2 + 8\text{H}_2\text{O} + 14\text{Mn}^{5+} \leftrightarrow 14\text{Mn}^{2+} + \text{Fe}^{2+} + 2\text{SO}_4^{2-} + 16\text{H}^+$	[6]	125.00
<u>Ferric Iron Replacing Oxygen; No Oxidation and Precipitation of Iron</u>		
$\text{FeS}_2 + 8\text{H}_2\text{O} + 14\text{Fe}^{3+} \leftrightarrow 15\text{Fe}^{2+} + 2\text{SO}_4^{2-} + 16\text{H}^+$	[7]	125.00 OR 15.63 ¹
<u>Chalcopyrite Replacing Pyrite; No Copper Precipitation</u>		
$\text{CuFeS}_2 + 11/2\text{H}_2\text{O} + 17/2\text{O}_2 \leftrightarrow \text{Cu}^{2+} + \text{Fe}(\text{OH})_3 + 2\text{SO}_4^{2-} + 2\text{H}^+$	[8]	15.63
<u>Galena Replacing Pyrite; No Lead Precipitation</u>		
$\text{PbS} + 2\text{O}_2 \leftrightarrow \text{Pb}^{2+} + \text{SO}_4^{2-}$	[9]	0.00
<u>Galena Replacing Pyrite; Lead Precipitation as Hydroxide</u>		
$\text{PbS} + 2\text{O}_2 + 2\text{H}_2\text{O} \leftrightarrow \text{Pb}(\text{OH})_2 + \text{SO}_4^{2-} + 2\text{H}^+$	[10]	31.25

¹ depending on the source of the ferric iron; see text

Based on this analysis of theory and stoichiometry, the determination of potential acidity as part of the ABA procedure is significantly more difficult than a simple analysis for total sulfur or a form of sulfur. A reliable determination requires attention to mineralogy and the environmental field conditions to which rock or tailings will be exposed. Nevertheless, an additional complication arises in the assumption that 1 part of CaCO₃ represents 2 parts of acidity. As discussed in Section 3, a ratio between 2:1 and 1:1 may be more appropriate so that conversion factors in this section would have to be multiplied by a value between 1 and 2, increasing the Acid Potential.

2.2 Rates of Acid Generation

The preceding section on stoichiometry was based on the implicit assumption that sulfide minerals would oxidize to another form of sulfur. In reality, this oxidation as well as any acid generation from metal precipitation takes place over a finite period of time. If sulfide minerals in a sample had a consistently negligible rate of oxidation under the relevant environmental conditions, then water-quality impacts would be minimal in spite of the findings of the ABA analysis.

Because the ABA procedure is a static test which is independent of time, rates of oxidation and acid generation cannot be measured and predicted by the test. This requires kinetic tests such as leach columns and humidity cells. Nevertheless, an understanding of rates and the factors affecting the rates (Table 5) will assist in the proper interpretation of ABA results. For example, a sample of coarse-grained massive chalcopyrite from a mine in British Columbia was predicted to be acid-generating on the basis of ABA (approximately -900 t CaCO₃/1000 t), but showed negligible acid generation in humidity cells probably due to crystallinity and grain size. On the other hand, if this sample were finely ground to reflect significant weathering, crushing, or processing in a mill circuit followed by discharge to a tailings impoundment, the rate of acid generation may be significantly enhanced. This illustrates the importance of delineating the handling of the rock and the environmental conditions under which the rock will be placed in order to properly interpret acid-generation tests.

TABLE 5--Factors affecting rates of oxidation and acid generation (from Morin et al., 1990)

FACTORS AFFECTING RATES OF OXIDATION:

Ratio of sulfide mineral to water	pH
Surface area (particle size)	Eh
Saturated/unsaturated condition	Temperature
Humidity in unsaturated zones	Surface roughness
Types and concentrations of oxidants	Crystallinity
Associated precipitation of minerals	Bacterial effects
Contact with other sulfide minerals	Catalysts such as Ag ⁺

TABLE 5 (CONTINUED)

FACTORS AFFECTING RATES OF ACID GENERATION:

All preceding factors, plus

- stoichiometry and extent of metal oxidation
 - stoichiometry and extent of metal precipitation
-

3. NEUTRALIZATION POTENTIAL

Because acid-base accounting defines the balance between acid-generating and acid-consuming materials, the second primary goal of the procedure is the delineation of Neutralization Potential (NP). As with Acid Potential, there are several implicit assumptions in the NP procedure and several difficulties in the proper interpretation of the NP data. The assumptions and difficulties appear to be less recognized than those of Acid Potential.

Unlike sulfur which is measured directly in the ABA procedure, Neutralization Potential is determined indirectly. The ground sample is soaked in excess acid usually at a pH less than 2 for a period of time and then the solution is titrated with a base usually to pH 7. This procedure introduces two major uncertainties which can affect the accuracy of the NP value and ability to interpret the value. These uncertainties are:

- o the impact of the large range of pH variation during the acid soak and the base titration, and
- o the impact of the length of time during which the sample soaked in excess acid.

The first uncertainty will be discussed in Section 3.1 and the second will be discussed in Section 3.2.

3.1 Stoichiometry and Mineralogy of Acid Neutralization

The neutralization of acidity is a complex series of reactions involving both solid-phase minerals and aqueous ions. In natural environments, neutralization is usually a sequential process consisting of stepwise increases in neutralization (Morin et al., 1988; Morin and Cherry, 1988; Morin, 1988). In fact, simulations of in-situ neutralization as a one-step process will usually lead to errors in predicted water quality and NP consumption (Morin and Cherry, 1988).

Neutralizing minerals typically contain some form of carbonate (CO_3) and hydroxide (OH). Although each mineral has a specified solubility, a range of solubility

occurs in natural waters due to the presence of other ions which enhance solubility and the variations in temperature and pressure. A simplified way to express the solubility of neutralizing minerals is through "equilibrium pH", which indicates the pH to which acidic water will be neutralized by a specific neutralizing mineral. Based on speciation and mineral-solubility calculations with data from six published field investigations and one laboratory study (Morin, 1988), the ranges of equilibrium pH for several neutralizing minerals were found to extend from 3.0 to 6.9 (Table 6).

TABLE 6--Ranges of equilibrium pH for neutralizing minerals in natural waters (from Morin, 1988)

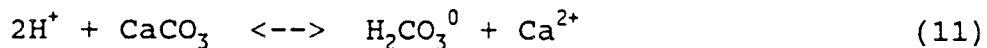
<u>MINERAL SUITE</u>	<u>RANGE OF EQUILIBRIUM pH</u>
Calcium-based carbonates	5.5-6.9
Iron-based carbonates	5.1-6.0
Aluminum-hydroxide-bearing minerals	4.3-5.0
Iron-hydroxide-bearing minerals	3.0-3.7

Because many water-quality guidelines call for aqueous pH to be greater than 6, most of the minerals in Table 6 are not important from a regulatory perspective, although their contribution to partial neutralization lessens the quantity of calcium-based carbonates needed to fully neutralize pH. From the perspective of ABA, a sample which is sometimes ground is placed in excess acid which causes all relevant neutralizing minerals to dissolve (reaction rates are critical, see Section 3.2). This could lead to the overestimation of the capability to neutralize pH to a value greater than 6. However, this situation becomes ambiguous when the titration with base is performed to about pH 7 in order to determine the quantity of excess acid that had been neutralized. During the titration, at least some of the metals liberated by the acid soak are re-precipitated. The acid soak causes much of the dissolved carbonate to leave the sample as CO₂ gas so that mineral precipitation is probably in the form of hydroxides. In this scenario, the precipitation of hydroxide minerals would consume the additional base, giving the false impression that less acid had actually been neutralized by the original minerals. In effect, this would result in an underestimate of NP. Because precipitation occurs over a finite period of time, the faster the base titration, the less the metal precipitation, and the more realistic the NP value.

As explained above, the standard measurement of NP reflects the content of all neutralizing minerals including those that might neutralize only to the environmentally insignificant level of pH 3. A refinement of the ABA procedure would include a direct measurement of CO₂ content and a mineralogical examination for the dominant forms of carbonate and hydroxide.

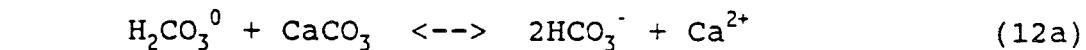
In addition to the aforementioned ambiguity in the NP value, there is additional ambiguity in the determination of acid consumption even when the content of calcium-based carbonates is known precisely. This source of ambiguity is centered on the number of moles of acidity that can be neutralized by one mole of CaCO₃.

In the simplistic situation where an acidic water with only aqueous H⁺ and pH less than 6 comes in contact with solid-phase CaCO₃, the appropriate equation describing the resulting neutralization up to a pH of approximately 6 is:

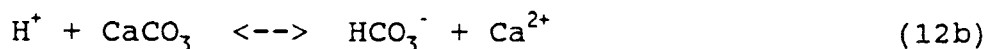


In this case, 2 moles of acidity are consumed by 1 mole of carbonate. The aqueous complex of H₂CO₃⁰ actually reflects the presence of dissolved CO₂ gas in the water (H₂O + CO₂ = H₂CO₃), which may be lost from the water under open-system conditions such as with surface waters or retained under closed-system conditions such as beneath a water table.

As carbonate continues to dissolve into this water, the aqueous pH changes little around pH 6.3 because of aqueous buffering. In other words, carbonate is being consumed while pH remains relatively constant due to the formation of bicarbonate:



OR



This neutralization above pH 6 is less efficient because only 1 mole of acidity is neutralized for each mole of CaCO₃. Full neutralization of acidic water to a pH greater than 6.5 is more complicated than indicated by Equations 11 and 12, but would still result in an overall ratio between 1:1 and 2:1. One laboratory study in British Columbia confirmed that the appropriate ratio was in the range of 2:1 and 1:1 (after alkalinity concentrations were converted from theoretical CaCO₃ reported by the laboratory to actual HCO₃ on the basis of another 1:2 molar ratio).

Many laboratories implicitly assume a ratio for acid consumption by CaCO_3 during the NP procedure. The selection of the acid, the choice to express results as normality or molality, the equation for calculating acid consumption from the titration data, and the expression of the results as CaCO_3 , all involve some assumption on the ratio. Consequently, the details of the procedure followed by a particular laboratory should be known prior to interpretation of NP data. This chosen ratio must be consistent with that used for the calculation of Acid Potential.

3.2 Rates of Acid Neutralization

In contact with acidic water, each neutralizing mineral will dissolve at a specific rate. The rate will reflect site-specific factors such as mineral composition, aqueous pH, temperature, and grain size.

For the NP procedure, a sample which may be ground is allowed to soak for some period of time, usually for several minutes to a few hours, prior to base titration. The acid may be heated, which would accelerate rates of dissolution. As a result, the NP will reflect the content of rapidly-dissolving minerals. Because of the presence of excess acid which enhances rates of dissolution, most environmentally important neutralizing minerals will be dissolved. Slower-dissolving neutralizing minerals such as aged oxides and feldspar minerals may represent a significant source of long-term neutralization in relatively stagnant flow systems and, if these environmental conditions are expected, these minerals should be added to the NP results.

4. PRACTICAL PROBLEMS IN ABA INTERPRETATION

A significant issue in acid neutralization is the location of the neutralizing minerals relative to the sources of acid generation. If neutralizing minerals are located upgradient of the acid generation, then the minerals will only dissolve to a relatively minor extent in response to inflowing water (e.g., rainfall). When this water encounters strong acidity, the aqueous alkalinity may be overcome and acid drainage may result even if ABA analyses indicated an excess of NP in the sample. On the other hand, available NP in the vicinity, or downgradient, of the source will dissolve in response to the acidity and may then provide full neutralization. The dissolution of the NP would continue until exhausted and, if Acid Potential exceeds NP, acid drainage may then

flow freely from the rock.

An extension to the preceding problem involves a smaller scale, on the order of one rock grain or block. Sulfide minerals and NP are not usually distributed in the same pattern throughout a rock grain (Sorenson et al., 1980; Ferguson and Erickson, 1987). If NP is preferentially located within a grain whereas the sulfide is located on the exposed surface, then the ABA results have little validity and acid drainage may results despite the ABA balance. In contrast, the preferential encapsulation of sulfide within a grain may prevent the onset of acid drainage even in the presence of a bulk excess of sulfide. These scenarios can be significantly altered by weathering of grains.

5. CONCLUSION

This paper has reviewed the theory, application, and analytical procedures involved in the acid-base accounting. This review has highlighted problems that can arise in the interpretation of Acid Potential and Neutralization Potential as well as practical problems involving the spatial and temporal dimensions. These problems can affect the validity of the resulting Net Neutralization Potential and, thus, affect the predictive validity of the ABA procedure.

Problems with Acid Potential primarily center on the factor for converting sulfur to Acid Potential. The assumptions leading to the current ABA factor of 31.25 were examined and other factors ranging from 0 to greater than 100 were found to be appropriate under other sets of assumptions. These assumptions refer to the geochemistry of the rock or tailings and the environmental conditions to which rock or tailings will be exposed.

Problems with Neutralization Potential were attributed to the complexity of acid neutralization in the environment, the unique effects of each suite of neutralizing minerals, and the analytical procedure itself. Because of the ambiguities in the determination of Neutralization Potential, a more reliable test may often include the determination of carbonate content and the assumption that one mole of CaCO_3 will neutralize only one mole of acidity.

Practical problems that affect the validity of ABA results primarily involve spatial and temporal dimensions on various scales. On the relatively large scale of an

entire tailings impoundment or waste-rock pile, the sequence in which acid-generating and acid-neutralizing minerals are encountered by the flowing groundwater will determine whether acid drainage will leave the system at some point in time. On the smaller scale of a rock particle, the relative distribution of acid-generating and acid-neutralizing minerals across the surface of, and through a grain, will determine whether acid drainage might appear at some point in time.

An enhanced version of ABA includes detailed geochemical evaluations and the interpretation of ABA results in light of the anticipated environmental conditions at a site. This enhanced testwork will improve the accuracy of predictions based on acid-base accounting.

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