

ACIDIFYING SOIL FOR CROP PRODUCTION WEST OF THE CASCADE MOUNTAINS (WESTERN OREGON AND WASHINGTON)

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Soil acidification sometimes is necessary for optimum plant growth west of the Cascade Mountain Range. Commercial producers of blueberries, azaleas, rhododendrons, and other ornamentals may need to reduce soil pH for optimum production.

Soil acidification is best performed prior to planting; it is much more difficult in established plantings. No routine soil test is available to determine soil acidification amendment rates. This publication is intended to provide guidelines for acidification of commercial fields. It is not intended for use with container-grown ornamentals.

Crop soil pH requirements

Table 1 lists optimum soil pH for selected crops grown in western Washington and Oregon. Soils in this region are naturally acidic. For most crops, liming to raise soil pH, rather than soil acidification, is needed. Blueberries and nursery crops such as azaleas and rhododendrons are exceptions. These “acid-loving” plants require a soil pH less than 6.0 and preferably below 5.5. Cultivation of these crops usually is the only reason to consider soil acidification west of the Cascade Mountain Range. Acidification of soil for production of these crops is common.

Table 1.—Optimum soil pH range for selected crops.*

Alfalfa	6.5–8.4
Vegetables	6.5–8.2
Garlic	6.5–7.5
Grass for seed or pastures	5.5–8.2
Fruit trees	6.0–8.0
Highbush blueberries and cranberries	4.5–5.5
Rabbiteye blueberries	4.2–5.0
Azaleas and rhododendrons	4.5–5.5
Field or silage corn	5.5–8.4
Wheat	5.5–8.4

*Soil pH determined in 1:2 soil:water ratio.

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Plant symptoms when soil pH is too high

Plants are excellent indicators of the need for soil acidification. Several symptoms are exhibited when soil pH is too high. A common symptom is yellowing (chlorosis) of leaves, with prominent, contrasting green veins (Figure 1). Leaves may be smaller than normal, and leaf edges may be brown (Figure 2). Symptoms are observed



Figure 1.—Chlorosis with contrasting green veins.



Figure 2.—Chlorosis with brown leaf edges.

in new growth (Figure 3). Chlorosis may affect a single branch or as much as the entire tree or shrub.

The yellowing associated with high soil pH usually is caused by iron deficiency, which can cause plants to perform poorly and eventually die. Nutrients such as zinc and iron have limited availability when the soil pH is above 7.5. (See “Micronutrient deficiencies,” page 4.)

Iron chlorosis is most likely to be observed in blueberries, rhododendrons, and azaleas. It sometimes is seen in silver maple, pin oak, sweet gum, some conifers, and fruit trees such as apples. Other ornamentals, such as dogwoods, also might express iron-induced chlorosis. For a more complete listing of susceptible plants, visit the University of Minnesota Web site listed in “References,” page 11.

Poor plant growth may result from high soil pH even without visible plant symptoms. A soil test will assist in assessing whether high soil pH is a problem.

A nitrogen or sulfur deficiency also causes leaves to turn yellow. As with iron deficiency, the new growth of sulfur-deficient plants first turns light green and then yellow. Nitrogen deficiency is seen first in older leaves. Yellow leaves resulting from nitrogen or sulfur deficiency do not have green veins, however. If applying nitrogen or sulfur does not change yellow plants to a healthy green color, an iron deficiency may be the problem. Note that blueberries rarely develop a sulfur deficiency.

Soil acidity

Acidity is a measure of the hydrogen ion concentration in soil solution (the water held between soil particles). Soil acidity is expressed as soil pH and is

measured on a scale from 0 to 14. Soil pH values below 7 indicate acidic soil, and values above 7 indicate basic (alkaline) soil. As pH decreases, soil acidity, or the hydrogen ion concentration, increases. Lime is added to acidic soil to reduce acidity and raise soil pH. Amendments such as elemental sulfur are added to basic soil to increase acidity and reduce pH.

Soil pH governs the overall environment for root growth. It determines the availability of some nutrients as well as elements toxic to root growth.

The importance of soil pH to plant growth can be explained using the following analogy.

Think of yourself in a swimming pool. If the water is too cold, too hot, or the chlorine content so high your eyes hurt, you might limit your time in the pool. Now think of roots growing in soil. If the soil pH is too high or too low, plant roots don't want to be there. Unfortunately, they don't have a choice. If roots are not “comfortable” because soil pH is too high or too low, they don't grow normally or can't take sufficient nutrients into the plant. When a plant shows severe deficiencies, it is not just “uncomfortable”; it has been injured from the lack of essential nutrients such as iron.

Soil acidification occurs naturally west of the Cascade Mountain Range. Winter rainfall leaches bases—primarily calcium and

magnesium—from the soil, making the soil acidic. The bases either are leached into the groundwater, causing “hard” well water, or are carried into surface streams and eventually to the ocean. The natural process of soil acidification is accelerated by the addition of nitrogen fertilizers, crop removal, and other farming practices.

Seasonal variation in soil pH

The season in which soil is sampled affects observed pH. West of the Cascades, seasonal variation of soil pH is associated with wetting and drying cycles. The highest soil pH (least acidity) is measured in late winter and early spring, prior to fertilizer application, increased microbial activity, and soil drying. Soil pH decreases



Figure 3.—Chlorosis caused by high soil pH appears in new growth.

in the late spring and summer as increasing microbial activity and the addition of fertilizer salts raise the salt content of the soil. The lowest soil pH (most acidity) occurs near the end of the growing season before winter rains begin.

Soil pH can change more than 1 unit from spring to fall in sandier soils. Typically, pH is from 0.3 to 0.5 unit higher in spring than during the growing season (Peterson 1972; Doerge and Gardner 1985).

Measuring soil pH

The first step in soil acidification is to determine soil pH. Without knowing your soil's pH, you can't know whether acidification is needed and, if needed, how much acidifying material to add.

In-field pH measurement

Many choices for in-field soil pH measurement are available. The values provided by in-field analyses should be used only as an estimation of soil pH. If an in-field measurement indicates soil pH might be above or below the range given for a crop, send a sample to a laboratory for confirmation. With any in-field method, soil pH should be confirmed at least periodically with a laboratory measurement.

In-field soil pH test methods have been evaluated by University of Arkansas Extension specialists. They compared soil pH of three soil samples from a national quality-control program with in-field soil

pH results (Stevens et al. 2001). Data from the study are presented in Table 2. The hand-held pH meter produced results closest to the laboratory average. The pH probe provided the same soil pH for all three samples. Results from the color kit were intermediate between the pH probe and hand-held meter.

The accuracy of hand-held meters is offset by their cost and complicated operation. The cost of a hand-held pH meter is approximately \$100 to \$300. The hand-held meter is a portable version of a laboratory meter and requires the same care and calibration with buffer solutions before each use. Accuracy depends on proper operation and on the condition of the electrode. A hand-held meter kept in a truck is likely to have a dried-out electrode and cannot be expected to perform adequately. The results for the hand-held meter presented in Table 2 were achieved with proper calibration, meter care, and operation, so they represent a best-case scenario.

The inability of the soil pH probe to distinguish between samples with more than a 2.0 pH unit difference makes it an unacceptable method of estimating soil pH in the field.

The color kit used to determine soil pH presented in Table 2 used a whole-unit (1.0) graduated scale. In many cropping situations, an 0.5-unit scale is necessary. Sometimes 0.1- to 0.2-unit differences are critical.

Color kits have several drawbacks. The solutions degrade with time and heat. Store them in a cool, dark place such as a refrigerator, and purchase new solution annually. Changing light conditions also can influence how color is perceived. In addition, matching the color of the sample mixture to the color chips is difficult for some people.

Table 2 does not include results from dye-impregnated paper strips, another in-field soil pH testing method. The accuracy of pH strips is similar to that of color kits (generally within 0.5 pH unit). The advantage of the strips is that solutions are not used.

Laboratory analysis

Soil pH is the most commonly performed soil test. Any soil testing laboratory will determine soil pH, generally for less than \$10. Soil samples should reflect differences in crop performance and soil type. See the publications listed under "For more information" (page 11) to learn more about taking a soil sample and choosing a laboratory.

Although soil pH laboratory measurements are recommended for decisions about liming and soil acidification, use them carefully. Laboratory results can be compared only if they are based on the same methodology. Be careful when interpreting soil pH data using unspecified methods.

Three common methods for soil pH determination are used by laboratories in the western United States—saturated paste, 1:1 ratio of soil to water, and 1:2 ratio of soil to water. Figure 2 shows that soil pH increases slightly for acidic soils as water-to-soil ratio increases. The difference in results between the two water methods is small. Results from the 1:1 method usually are 0.1 to 0.2 pH unit lower than those

Table 2.—Soil pH comparison of three in-field methods with the average soil pH from 82 laboratories participating in a quality-control program.

Method of pH measurement	Soil 1	Soil 2	Soil 3
Laboratory average	7.9	6.3	5.6
Hand-held pH meter	7.7	6.3	5.5
Color kit	8.0	6.5	5.5
pH probe	6.0	6.0	6.0

from the 1:2 method (Figure 4). Soil pH measurements based on the saturated paste method are slightly lower than those from the 1:1 method.

Easy access to electronic data from Europe can complicate soil pH interpretations. In Europe, soil pH is measured in a dilute salt solution such as CaCl_2 . The resulting pH is 0.5 to 1.0 unit lower than that measured in distilled water. Dilute salt solutions also remove seasonal fluctuations.

Micronutrient deficiencies resulting from high soil pH

The solubility of metallic micronutrients such as iron (Fe), zinc (Zn), and manganese (Mn) in soil solution decreases rapidly as soil pH increases, making these elements less available to plants. Figures 5–7 illustrate how iron, manganese, and zinc concentration in soil solution changes with varying soil pH for Woodburn and Bashaw soils. Note that the concentration of these micronutrients in soil solution increases as soil pH decreases.

The decrease in metallic micronutrient availability is not necessarily

detectable by a soil test (DTPA extractable micronutrients). A plant may suffer from iron-deficiency chlorosis even when a soil test indicates that adequate amounts of iron are present. The chlorosis is caused not by a lack of iron in the soil; soil contains sufficient iron for plant growth, approximately 100,000 lb/acre. However, it is not in a plant-available form when soil pH is high. Soil acidification programs are designed to use iron present in the soil by decreasing soil pH and thus increasing the solubility of iron.

Although application of iron without soil acidification will increase the total Fe in the soil, it may not increase Fe in the soil solution and thus may not correct a deficiency. When iron is applied to

soil with a pH above 7.5, it does not remain soluble and becomes unavailable to the plant in a relatively short time, usually before the plant can use it. Repeated applications and high rates of application are needed to overcome an iron deficiency with soil treatments. Directions for an iron drench or soil application are given in the section titled “Iron drench,” page 6.

Zinc and manganese also may be deficient when soil pH is too high for a crop. Figures 6 and 7 show that the concentration of zinc and manganese decreases as soil pH increases. Blueberries are hyperaccumulators of manganese and do not exhibit manganese deficiencies. Red maple, on the other hand, frequently exhibits

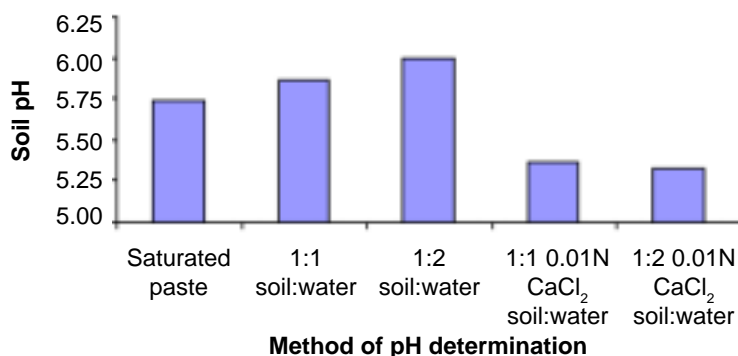


Figure 4.—Average soil pH by method of determination from 60 acidic soil samples used in the North American Proficiency Testing Program (1996–2002).

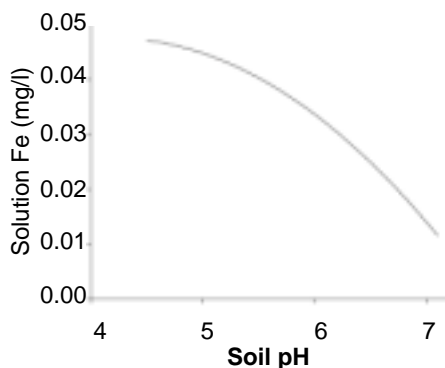


Figure 5.—The relationship of soil pH and iron (Fe) in solution extracted from a Woodburn soil.

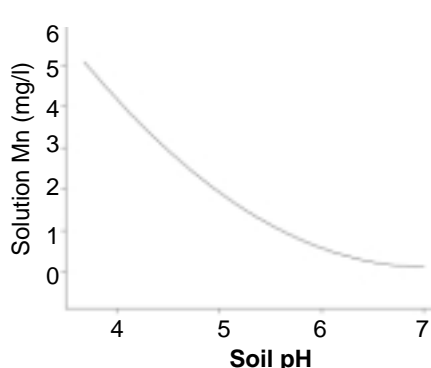


Figure 6.—The relationship of soil pH and manganese (Mn) in solution extracted from a Bashaw soil.

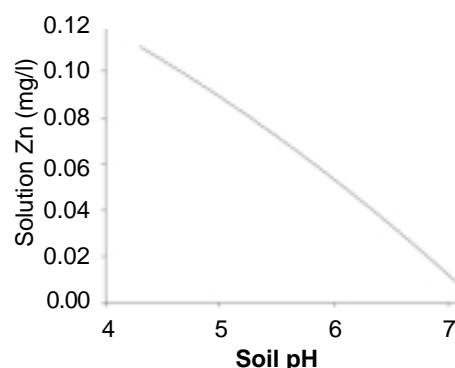


Figure 7.—The relationship of soil pH and zinc (Zn) in solution extracted from Bashaw and Woodburn soils.

manganese deficiency when soil pH is above 6.0 (Figure 8).

A common problem of field-grown red maple (*Acer rubrum*) is chlorosis that develops in late summer. Maples begin growth with no sign of chlorosis, and robust dark green growth continues until mid- to late August, when the chlorosis appears.



Figure 8.—Interveinal chlorosis and necrosis caused by manganese deficiency in red maple.

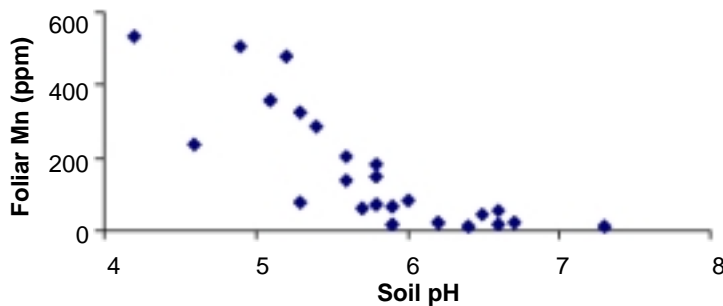


Figure 9.—Relationship between soil pH and foliar manganese in red maple.

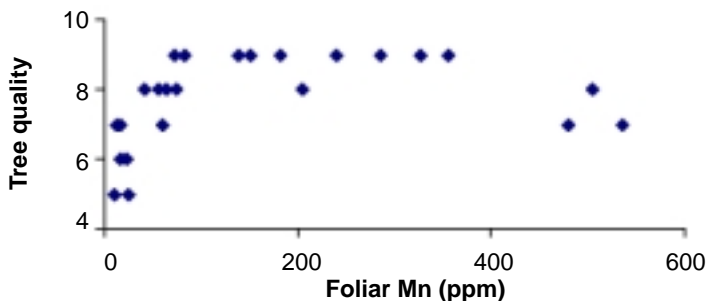


Figure 10.—Relationship between foliar manganese and tree quality in red maple.

Figure 9 shows the relationship between soil pH and foliar Mn in field-grown red maple. As soil pH increases from 4.5 to 6.0, foliar Mn decreases from several hundred ppm to less than 100 ppm. When the soil pH is above 6.0, uniformly low Mn was measured in the foliage. Trees grown in a soil with a pH above 6.0 contained less than

100 ppm Mn and produced stock with a lower quality rating than trees grown in soil with a pH below 6.0 (Figure 10). Tree quality was rated on a scale of 1 to 10, with 1 = extremely poor quality (foliar chlorosis and interveinal necrosis), 5 = poor quality (chlorosis evident), 7 = marginal quality (slight interveinal chlorosis), and 10 = good quality (dark green foliar color and no sign of chlorosis).

Generally, zinc and manganese do not need to be added to the soil. Acidification usually increases solubility and availability of existing nutrients and corrects the deficiency.

Lime as a cause of iron deficiency

In some cases, iron deficiency is not caused by high pH alone, but by the addition of carbonate (lime) or bicarbonate to the soil. Small lime additions can change the soil pH enough to reduce iron availability. These materials also inhibit a plant's ability to obtain iron from the soil solution. In addition, bicarbonate may interfere with a plant's ability to use iron within the plant (Marschner 1995). Lime-induced chlorosis has been observed on maple trees in the Willamette Valley even when soil pH is neutral (6.5 to 7.0 pH).

Lime-induced chlorosis can be caused by lime additions (or high free lime soils), some fertilizers, and irrigation water. Carbonates are an acid-neutralizing substance often dissolved in groundwater. Irrigating with water with a high carbonate concentration raises soil pH just as a lime application would. Surface water generally contains little carbonate or bicarbonate.

Lime-induced chlorosis is temporary and may last a few months to

Table 3.—Soybean yield and foliar iron concentration from iron-coated seed and foliar iron treatments from research in Minnesota during 2001.

	Iron chelate		Iron sulfate	
	Yield (bu/acre)	Iron uptake (oz/acre)	Yield (bu/acre)	Iron uptake (oz/acre)
No iron seed				
with or without foliar	0	0	0	0
Iron seed + no foliar	11	0.46	11	0.46
Iron seed + 1 foliar	17	1.13	19	1.27
Iron seed + 2 foliar	10	—	19	—

years until soil pH decreases. It usually is not detrimental for homeowners with perennials, but it may be a significant problem in the nursery industry, where appearance is critical.

Blueberries (Figure 11) and rhododendrons do not recover from lime-induced chlorosis as quickly as most ornamentals; therefore, corrective action may be needed. To avoid chlorosis from lime applications to these crops, add only small amounts of lime based on a soil test.

Correcting iron-deficiency chlorosis

Because yellow leaves can be caused by several factors, do not start a program to acidify your soil based on plant symptoms alone. If you suspect a pH problem, test your soil to determine the pH. Compare the results of the soil test with Table 1 to determine whether acidifying the soil is necessary.

Correcting irrigation water pH

After determining soil pH, test your irrigation water for carbonate, bicarbonate, and pH. If irrigation water has a pH above 8.2 and total carbonates ($\text{HCO}_3^- + \text{CO}_3^{2-}$) are 2 to more than 5 meq/l, monitor soil pH and plant performance. Irrigation water can be acidified to a pH of

6.0 to 6.5 to reduce or eliminate carbonate problems. (See Burt et al. in “References” for more information.)

Foliar iron application

Many people ask about using foliar sprays to alleviate iron deficiency. Foliar iron sprays are an option, but it is important to recognize that their effects usually are short lived. They must be applied at least every other week during the growing season, as iron is immobile within the plant and new growth must be treated. However, foliar iron sprays, in conjunction with a soil acidification program, may be the only way established plants can survive a severe iron deficiency.

An example of the need for more than one type of treatment in situations with severe iron-induced chlorosis is shown in Table 3. When soybeans were planted without a seed treatment of iron, no yield was produced, even when a foliar application of iron was made. The combination of a seed treatment and foliar spray produced the highest yield.

To prepare 100 gallons of spray, dissolve 50 ounces of ferrous sulfate in 100 gallons of water (1 oz/2 gallons) and add 2 cups of mild household detergent or an appropriate amount of another



Figure 11.—Blueberry with iron-deficiency chlorosis.

wetting agent. An alternative recipe can be found in *Iron Deficiency in Plants* (ARS 1976). Thoroughly wet the foliage with the spray.

Iron drench

Another common question about iron application is whether an iron solution can be added to the soil. Soil applications require high rates of iron and are expensive. Also, an iron drench is best applied during the dormant season. The following method is unlikely to be practical for commercial production, but would be appropriate in a landscape situation.

To apply an iron drench for dormant trees, vines, and shrubs, dissolve ferrous sulfate in water at the rate of 1 lb ferrous sulfate/gallon of water. The amount of solution needed is determined by the size of the plant. Measure the diameter of the drip line. Apply 1 gallon of solution for every foot of diameter.

For shrubs and small trees, dig a shallow trench around the drip line and add the solution to the trench. If the drip line is too large to conveniently excavate a trench, dig holes

large enough to accommodate a gallon of liquid around the drip line at 2- to 3-foot intervals. Apply the solution to the holes. After the solution has soaked into the soil, replace the soil in the holes or trench.

Acidifying soil

If the initial soil pH is below 6.5, acidification can be accomplished fairly easily by adding finely ground elemental sulfur (S). Elemental sulfur is oxidized and changed to the sulfate form by the bacteria *Thiobacillus*. The oxidation process creates sulfuric acid in the soil solution.

Soil pH reduction with elemental S is not rapid, but it is economical over the long term. Add sulfur no later than the fall prior to planting and preferably a year before planting. Application a year prior to planting gives elemental S the time needed to start to react. Elemental S usually takes at least 2 years to fully react. Before planting, test the soil to determine whether the desired soil pH has been attained.

Fertilizers such as ammonium sulfate acidify the soil through a different mechanism. It is the ammonium N, not the sulfate, that has the acidifying effect. Sulfate does not acidify soil. The ammonium form of nitrogen ($\text{NH}_4\text{-N}$) is transformed to the nitrate form ($\text{NO}_3\text{-N}$) by bacteria. This process provides the acidifying ion hydrogen (H^+) to the soil. Ammonium-containing fertilizer materials react in this manner, including urea (46-0-0), ammonium nitrate (34-0-0), urea-ammonium nitrate solutions (Solution 32), and ammonium phosphate sulfate (16-20-0-14). Ammonium sulfate is the most acid-forming N fertilizer.

Acidifying N fertilizers lower soil pH gradually. Data collected from

four Willamette Valley grass seed fields showed an annual reduction in soil pH of 0.1 unit per 100 lb ammoniac N/acre applied. This rate of pH reduction usually is not sufficient for blueberries or rhododendron nursery production.

Other materials, such as sulfuric acid (H_2SO_4), acidify soils more quickly than common nitrogen fertilizers or even elemental sulfur. However, these materials are difficult and hazardous to use.

Broadcast application

To acidify soils west of the Cascade Mountains, broadcast elemental S and mix thoroughly to a depth of 6 to 8 inches. Tables 4–7 (page 8) provide recommended S rates based on soil pH and soil type. These tables require you to know your soil series and to perform an initial soil test to determine starting pH. Determine the soil series from a Soil Survey Report or enlist the aid of your local Extension or Natural Resources Conservation Service personnel. Do not guess soil types in a field. You could severely overestimate the rate of S to apply.

Each table gives recommendations for soil series representing one of the major landscape groups in the Willamette Valley:

- Sandy “river bottom” soils such as Newberg that are low in organic matter and have a CEC (cation exchange capacity) less than 14 meq/100g—*Table 4*
- Silt loam soils of the valley floor, such as Dayton or Woodburn (CEC 14 to 25 meq/100g)—*Table 5*
- “Red” upland soils, such as Jory and Nekia—*Table 6*
- Dark, clayey, high organic matter soils such as Bashaw (CEC greater than 25)—*Table 7*

Estimated elemental S rates are given as ranges in Tables 4–7 since the tables are created for groups of soil types. A single rate for a group of soils or even a single soil is not appropriate because of soil and seasonal variability. Adjust elemental S application rates based on experience with the field (for example, if soil pH changes are greater or less than expected).

Although Tables 4–7 provide estimated elemental S rates of as much as 5,200 lb/acre, the maximum for a single application should be 3,000 lb/acre. Higher rates may lower soil pH more than desired. If you overacidify soil, plant several thousand dollars of stock, and have to raise soil pH after planting, you risk losing your investment in planting stock. At a minimum, you will spend money on lime to neutralize the excess elemental S.

Where the tables estimate more than 3,000 lb/acre, apply only the initial 3,000 lb/acre and mix it into the soil thoroughly. Allow the elemental S to react in moist soil for 6 months to a year. Then measure soil pH to determine whether additional elemental S is needed.

The estimated elemental S application rates in Tables 4–7 are based on data used to determine the amount of lime needed to raise soil pH. Lime rate to increase soil pH was transformed into elemental S rate to lower soil pH. See the appendix (page 10) for further explanation of the process and for equations to estimate elemental S application rates for specific western Oregon soils or for soils outside the scope of Tables 4–7.

Research conducted near Wilsonville, Oregon, provides some verification of the assumptions used to create Tables 4–7. Elemental S was applied at a rate of 1 ton/acre to

continued on page 9

Table 4.—Estimated elemental sulfur (S) required for a broadcast application required to reduce soil pH of Newberg or similar soil. (See notes.)

Desired soil pH	Current soil pH				
	6.5	6.3	6.1	5.9	5.7
	lb/acre elemental S to apply				
5.4	1,100–1,500	800–1,100	500–800	200–500	0–200
5.2	1,500–1,800	1,100–1,500	800–1,100	500–800	200–500
5.0	1,800–2,100	1,500–1,800	1,100–1,500	800–1,100	500–800
4.8	2,100–2,400	1,800–2,100	1,500–1,800	1,100–1,500	800–1,100

Table 5.—Estimated elemental sulfur (S) required for a broadcast application required to reduce soil pH of Chehalis, Dayton, Laurelwood, Malabon, Powell, Sauvie, Willakenzie, Woodburn, or similar soil. (See notes.)

Desired soil pH	Current soil pH				
	6.5	6.3	6.1	5.9	5.7
	lb/acre elemental S to apply				
5.4	1,600–2,000	1,200–1,600	800–1,200	400–800	0–400
5.2	2,000–2,400	1,600–2,000	1,200–1,600	800–1,200	400–800
5.0	2,400–2,800	2,000–2,400	1,600–2,000	1,200–1,600	800–1,200
4.8	2,800–3,200	2,400–2,800	2,000–2,400	1,600–2,000	1,200–1,600

Table 6.—Estimated elemental sulfur (S) required for a broadcast application required to reduce soil pH of Jory, Nekia, McBee, or similar soil. (See notes.)

Desired soil pH	Current soil pH				
	6.5	6.3	6.1	5.9	5.7
	lb/acre elemental S to apply				
5.4	2,300–2,800	1,800–2,300	1,200–1,800	800–1,200	0–800
5.2	2,800–3,400	2,300–2,800	1,800–2,300	1,200–1,800	800–1,200
5.0	3,400–3,800	2,800–3,400	2,300–2,800	1,800–2,300	1,200–1,800
4.8	3,800–4,300	3,400–3,800	2,800–3,400	2,300–2,800	1,800–2,300

Table 7.—Estimated elemental sulfur (S) required for a broadcast application required to reduce soil pH of Bashaw, Salem, Amity, or similar soil. (See notes.)

Desired soil pH	Current soil pH				
	6.5	6.3	6.1	5.9	5.7
	lb/acre elemental S to apply				
5.4	2,500–3,000	1,500–2,500	1,200–1,500	800–1,200	0–800
5.2	3,000–3,600	2,500–3,000	1,500–2,500	1,200–1,500	800–1,200
5.0	3,600–4,400	3,000–3,600	2,500–3,000	1,500–2,500	1,200–1,500
4.8	4,400–5,200	3,600–4,400	3,000–3,600	2,500–3,000	1,500–2,500

Notes

- Two years are needed for complete reaction.
- Apply a maximum of 3,000 lb/acre/application. Check soil pH in 12 months and apply additional material if needed.
- See appendix (page 10) for sources of recommendations and individual soil equations.

a loam soil with a pH of 5.7 and 3.5 percent organic matter. Soil pH was reduced to and stabilized at 5.1 after 2 years (Hemphill and Jackson 1982). Table 5 recommends up to 1,200 lb/acre of elemental S to obtain a similar result. This example indicates that, for some soils, more than one application might be needed to achieve a desired pH reduction. Remember, when applying elemental S for soil acidification, start conservatively and then apply additional S a year later if needed.

Partial root zone or band application

Lowering soil pH of the complete plant root zone or an entire field may not be practical or necessary. A band or other concentrated treatment may be sufficient to overcome

an iron deficiency. For established trees and shrubs, you can acidify a small zone from which the plant can obtain iron (Figure 12). Elemental S, N-pHuric, phosphoric acid, or sulfuric acid can be used. Elemental S is the easiest, longest lasting, and safest.

To lower soil pH for established shrubs, ornamental trees, or fruit trees, dig at least 4 holes, preferably 8 to 12, per tree or shrub. Holes should be at least 1 foot deep. Mix an acidifying material with the soil. For each hole, use 2 oz elemental S, 4 oz N-pHuric, or 4 oz phosphoric acid. A small amount of iron, zinc, and manganese sulfate can be added to supply these nutrients to the crop. Refill the holes with the mixture.

Banded elemental S applications to established blueberries can be

effective in improving plantation performance. Apply 300 to 500 lb/acre elemental S in a 5-foot-wide surface band on either side of the row. Apply material in the fall with some incorporation if possible. Increasing the S-to-soil contact increases the acidification rate.

In the spring, measure soil pH in a 1- to 6-inch sample from the banded area. If the desired soil pH has not been reached, make a second application at the same rate.

In very poorly drained soils where ponding occurs, hydrogen sulfide (a rotten egg smell) may be generated. Avoid elemental S applications to these areas.

Some nurseries and plantations also apply large amounts of sawdust or bark dust to reduce soil pH as well as increase aeration.



Figure 12.—Individual trees can be treated with a partial root zone application of elemental S.

Summary

Acidification of soil sometimes is necessary for optimum growth of blueberries, azaleas, rhododendrons, and other ornamentals. Plants with yellowing of leaves in new growth, similar in color to the symptoms for nitrogen and sulfur deficiencies, may be suffering from an iron deficiency. Iron deficiency is caused by the soil pH being higher than tolerated by the plant being grown. The solution is to acidify the soil and reduce soil pH. When attempting to acidify soil:

- Proceed cautiously, as soil pH can be lowered beyond the desired level, especially in sandy soils.
- Monitor the change through annual soil testing.
- Keep complete records of the amount of material added, the amount of mixing, and time of year the acidifying material was added.

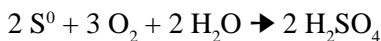
Appendix. Derivation of Tables 4–7

This appendix explains how research data were used to derive the rates in Tables 4–7. Data from Peterson’s thesis have long been used to create the SMP buffer tables for lime requirement in Oregon. These data were reversed via calculation to produce the estimated elemental S rates in Tables 4–7.

For each soil series, the increase in soil pH was regressed against the amount of lime (milliequivalents per 100 grams soil) applied (see Figure 13 for the Newburg soil series). Soil series were grouped by landscape position, and the slope of the regression line was combined and fitted for a linear relationship. The R^2 for the combined data was 0.84, which means that by knowing the amount of lime added you can predict the resulting change in soil pH.

Elemental S chemistry

The following equations show that one molecule of elemental S is required to neutralize one molecule of calcium carbonate. The substitution of elemental S for calcium carbonate using the linear response in Figure 13 allows calculation of the elemental S needed for acidification from the liming regression equations derived from Peterson’s data.



(Elemental S + oxygen + water + *Thiobacillus* bacteria + time → sulfuric acid)



(Sulfuric acid + lime → gypsum + carbon dioxide + water)

The next step in estimating the amount of S required for soil acidification is to adjust for atomic weight differences between elemental S and $CaCO_3$. The gram molecular weight of calcium carbonate is 100, and that of sulfur 32. Therefore, approximately one-third the

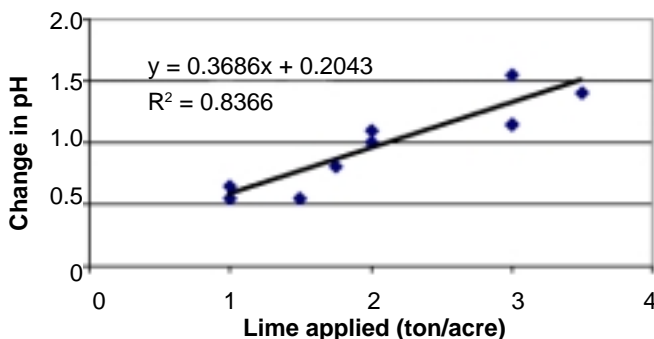


Figure 13.—Increase in soil pH from the application of lime to Newburg soil.

weight of elemental S (32 vs. 100) is required to acidify the same weight of calcium carbonate.

Realignment of lime equations

The equation displayed in Figure 13 ($Y = 0.3686x + 0.2043$) for the Newburg soil can be restated.

Changing liming to acidification

To apply lime and increase soil pH:

$$\text{Change in soil pH} = 0.3686 \times \text{tons lime} + 0.2043$$

To apply elemental S and decrease soil pH:

$$\text{Change in soil pH} = 0.3686 \times \text{tons S} \times 0.32 + 0.2043$$

Remember that one S reacts with one $CaCO_3$, but the weights are different, 100 vs. 32. Therefore, 100 pounds lime reacts with 32 pounds elemental S.

We want to solve for tons S (X) to get the rate of elemental S to apply for a desired change in soil pH.

Solving for elemental S rate

$$Y = 0.3686x + 0.2043$$

$$Y - 0.2043 = 0.3686x$$

$$(Y - 0.2043) \div 0.3686 = \text{tons S} \div 0.32$$

$$\text{Tons S to apply} = (Y - 0.2043) \div 0.3686 \times 0.32$$

$$\text{Pounds elemental S to apply} = (Y - 0.2043) \div 0.3686 \times 0.32 \times 2,000$$

where Y is the desired pH change

The previous equation can be expressed as:

$$\text{Pounds of elemental S needed for desired decrease in soil pH} = (\text{desired pH reduction} - 0.2043) \div 0.3686 \times 0.32 \times 2,000 \text{ (pounds in a ton).}$$

Using the equation

To reduce the Newburg soil pH from 6.5 to 5.4, or 1.1 pH units:

$$\text{pounds/acre of elemental S required to reduce soil pH 1.1 units} = (1.1 \text{ (the desired reduction)} - 0.2043) \div 0.3686 \times 0.32 \times 2,000 = 1,560 \text{ lb/acre}$$

$$1,560 \text{ lb/a S} = (1.1 - 0.2043) \div 0.3686 \times 0.32 \times 2,000$$

To calculate the estimated amount of S to apply for a soil for any target pH, you can use the regression equations for the soil series listed in Table 8, page 11.

Table 8.—Elemental S needed (Y) in tons per acre to achieve desired pH change (X).

Soil series	Regression equation to estimate elemental S to apply in lb/acre	Coefficient of determination (R ²)
Bashaw	$Y = (X - 0.2158) \div 0.1672 \times 0.32 \times 2000$	0.89
Chehalis	$Y = (X - 0.1219) \div 0.3052 \times 0.32 \times 2000$	0.89
Dayton	$Y = (X - 0.0784) \div 0.2662 \times 0.32 \times 2000$	0.71
Jory	$Y = (X - 0.0388) \div 0.2278 \times 0.32 \times 2000$	0.89
Laurelwood	$Y = (X - 0.0201) \div 0.2994 \times 0.32 \times 2000$	0.91
Malabon	$Y = (X - 0.1248) \div 0.2852 \times 0.32 \times 2000$	0.89
Nekia	$Y = (X - 0.1000) \div 0.2458 \times 0.32 \times 2000$	0.98
Newberg	$Y = (X - 0.2043) \div 0.3686 \times 0.32 \times 2000$	0.84
Powell	$Y = (X - 0.1667) \div 0.3166 \times 0.32 \times 2000$	0.87
Willakenzie	$Y = (X - 0.2250) \div 0.2572 \times 0.32 \times 2000$	0.94
Woodburn	$Y = (X - 0.1241) \div 0.3226 \times 0.32 \times 2000$	0.91

For more information

OSU Extension publications

Soil Sampling for Home Gardens and Small Acreages,
EC 628.

A List of Analytical Laboratories Serving Oregon,
EM 8677.

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University of Minnesota. Soil pH Preferences for Selected Landscape Plants (<http://www.extension.umn.edu/distribution/horticulture/components/1731-29.html>).

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