

0. Green Analytical Chemistry⁴

In recent years, it has become a trend to develop items as “green,” with a perceived environmental friendliness. The field known as “green chemistry” can be traced back to the book *Green Chemistry: Theory and Practice*⁵ published in 1998 by Paul Anastas and John Warner. They defined green chemistry as “the use of a set of principles that reduces or eliminates the use or generation of hazardous substances in the design, manufacture, and application of chemical products.” A key item to notice with this definition is that green chemistry is a set of principles, that is, a way of thinking about chemistry. It is not a separate chemical discipline like analytical, organic, or physical. In fact, it is hoped that the current generation of chemistry students taking this course will have this mode of thinking become so natural to them that we will no longer have to think about “green chemistry,” but just “chemistry.”

How, then, does analytical chemistry fit into this way of thinking? What is green analytical chemistry? To answer these questions, we must first reconsider what is analytical chemistry and what is green chemistry.

During this course, you will learn a series of methods to qualitatively and quantitatively characterize a sample or chemical system. Knowledge of these methods and the chemical principles behind them is an asset valued by many industries. On the other hand, analytical chemists may be viewed as a necessary evil, for example, to analyze samples in order to comply with Food and Drug Administration or other regulatory agency requirements. Analytical chemistry is something that is tolerated, but the analysis costs money, rather than adding to a company's profits. Consider the following example: You work for a detergent manufacturer and a product development chemist comes into your lab with a bottle of detergent, asking you to determine the silicone level in this product. While you could easily do what you are told, upon questioning you find out that the product development chemist really doesn't want to know the silicone concentration. What he *really* wants to know is why the product isn't performing the way it should (maybe it foams too much or doesn't clean effectively) and he *thinks* it's related to the silicone concentration. A good analytical chemist does not merely provide data, but rather supplies information and knowledge upon which educated decisions can be based.

Analytical chemists are problem solvers. We use our knowledge to become partners with our customers to answer their questions. In the detergent analysis example, one analytical

chemist might have run the silicone analysis and then several other analyses in order to eventually solve the problem. But the analytical chemist who was a true scientific partner in the process understood the problem and used his knowledge of chemistry to develop a more refined hypothesis and streamline the analyses need to answer the ultimate question more easily. This role of the analytical chemist was summarized by Prof. Herb Laitinin who editorialized, "Analysis of a sample is not the true aim of analytical chemistry...the real purpose of the analysis is to solve a problem."⁶ Development of problem-solving skills comes with a combination of analytical knowledge and experience. *In summary, the analytical chemist is a problem solver and solving problems involves a means of thinking about chemistry.*

What is Green Chemistry?

We've defined green chemistry as something that is applied over the entire lifespan of a product. Hidden within this definition is a combination of environmental and economic concerns. That is, a product may have environmental benefits, but if it cannot compete in the marketplace, it is doomed to failure. Green chemistry is guided by the set of 12 principles developed by Anastas and Warner⁵ and shown in Table 1.

If green chemistry and concern for the environment are the right things to do, why haven't we been doing this all along? Think about organic chemistry. It isn't always easy to choose the right reagents to put the correct functional groups in the right places for a complex synthesis. Now consider that you are working as a product development chemist in industry. Not only must you synthesize the compound with properties desired by your company, but you must do it in a manner that can be passed on to a process engineer who can develop the process with high yield, speed, safety, and minimal cost. Now add environmental concerns on top of all of that and the lament of a certain amphibian character, "It's not easy being green," becomes more obvious. *In summary, green chemistry involves a way of thinking about chemistry in an environmentally and economically sound manner.*

Table 1. The 12 Principles of Green Chemistry⁵

1. Prevention

It is better to prevent waste than to treat or clean up waste after it has been created.

2. Atom Economy

Synthetic methods should be designed to maximize the incorporation of all materials used in the process into the final product.

3. Less Hazardous Chemical Syntheses

Wherever practicable, synthetic methods should be designed to use and generate substances that possess little or no toxicity to human health and the environment.

4. Designing Safer Chemicals

Chemical products should be designed to effect their desired function while minimizing their toxicity.

5. Safer Solvents and Auxiliaries

The use of auxiliary substances (e.g., solvents, separation agents, etc.) should be made unnecessary wherever possible and innocuous when used.

6. Design for Energy Efficiency

Energy requirements of chemical processes should be recognized for their environmental and economic impacts and should be minimized. If possible, synthetic methods should be conducted at ambient temperature and pressure.

7. Use of Renewable Feedstocks

A raw material or feedstock should be renewable rather than depleting whenever technically and economically practicable.

8. Reduce Derivatives

Unnecessary derivatization (use of blocking groups, protection/ deprotection, temporary modification of physical/chemical processes) should be minimized or avoided if possible, because such steps require additional reagents and can generate waste.

9. Catalysis

Catalytic reagents (as selective as possible) are superior to stoichiometric reagents.

10. Design for Degradation

Chemical products should be designed so that at the end of their function they break down into innocuous degradation products and do not persist in the environment.

11. Real-time analysis for Pollution Prevention

Analytical methodologies need to be further developed to allow for real-time, in-process monitoring and control prior to the formation of hazardous substances.

12. Inherently Safer Chemistry for Accident Prevention

Substances and the form of a substance used in a chemical process should be chosen to minimize the potential for chemical accidents, including releases, explosions, and fires.

What is Green Analytical Chemistry?

Reviewing the previous summary statement, we can correctly guess that green analytical chemistry is a marriage of thought processes — a unique way of thinking about how to do chemistry. Performing chemical analyses in a green manner does not mean that we must make compromises to the requisite accuracy, precision, and other analytical demands. Inserting green values into the laboratory is about several factors related to the 12 green chemistry principles, including waste, energy, toxicity, and others. For example, the hazard associated with a given chemical is related to both our exposure to the chemical and the inherent risk possessed by the chemical. While we can wear personal protective equipment, like safety glasses, lab coats, and latex gloves, or work in a fume hood, we cannot change the inherent risk associated with a compound. So perhaps we can change the reagent. We've been doing this for years. Benzene and carbon tetrachloride were banished from the lab since the mid-1970s. But our practices must continue to evolve as our knowledge about chemicals and their inherent risk progresses.

At least half of the 12 principles of green chemistry apply to the analytical laboratory:

Principle 1: Prevent Waste. While the wastes generated in chemical production overshadow the amount of waste coming from a typical analytical laboratory, individual analytical procedures may involve as much as one liter of organic solvent to extract the analytes from the sample. When used in an analytical procedure, even water, which everyone would agree is an environmentally benign solvent, is converted to wastewater which must be treated appropriately. So the professional analyst must stay abreast of developments that can minimize waste generation.

Principle 5: Safer Solvents and Auxiliaries. The inherent risk associated with a chemical must be considered and alternatives explored. The risk may be toxicological, but may also be flammability, corrosivity, or other attributes. If alternatives do not exist, appropriate exposure controls are needed.

Principle 6: Energy Efficiency. Several common analytical procedures can be considered energy intensive. For example, not only are large amounts of solvents used in extractions, in most cases the solvent must then be evaporated to concentrate the solutes for analysis. Instrumental procedures can require high temperatures or have high power demands. Sometimes trade-offs must be considered. Does it require less energy to keep a drying oven turned on overnight and maintain a constant temperature or to turn off the oven and require a rapid heat up just prior to use? If water is used as a nontoxic solvent, do we pay a price in the energy needed to evaporate water compared with many organic solvents? Later, we will present a set of metrics to judge these compromises.

Principle 8: Reduce Derivatives. Many times chemical derivatives are employed to enhance the solubility or detectability of sample components. For example, the determination of fats in foods requires conversion of fatty acids into methyl esters for gas chromatography. However, technological advances can provide similar analytical results while avoiding the need for derivatization. For example, modern high-performance liquid chromatography can provide results equivalent to previous generations of gas chromatography.

Principle 11: Real-time Analysis. This principle is analytical chemistry!!! The field of process analytical chemistry, involving in-line, on-line, or at-line procedures, is a valuable tool in chemical processing. By monitoring the creation of a by-product, the poisoning of a catalyst, reaction pH, or other properties, feedback is provided to allow a reaction or process to stay in control, avoiding potential process disruptions.

Principle 12: Safer Chemistry. The flammability of solvents, oxidative capability of reagents, and similar factors should be considered in developing analytical methods.

Several of the experiments in this laboratory manual illustrate green chemistry. For example, Experiment 11 (Kjeldahl Nitrogen Analysis) uses selenium-coated boiling chips, rather than a mercury compound, as a catalyst (Principle 9); Experiment 15 (Preparation and Iodometric Analysis of a High-Temperature Semiconductor) reminds us that high-temperature superconductors could help us toward energy sustainability (Principle 6); and Experiments 28 (Analysis of Sulfur in Coal by Ion Chromatography) and 29 (Measuring Carbon Monoxide in Automobile Exhaust by Gas Chromatography) illustrate the role of analytical chemistry in pollution prevention (Principles 1 and 11). A new experiment 34, Green Chemistry: Liquid Carbon Dioxide Extraction of Lemon Peel Oil, demonstrates the potential of new technology (supercritical fluid extraction) to address solvent alternatives (Principles 1 and 5). We challenge students to discuss your experiments relative to the 12 green chemistry principles as you write your reports.

Now that we've gotten some idea of what green analytical chemistry is, let's discuss some specific green analytical considerations. Some review articles⁷⁻⁸⁻⁹ present more detailed discussions of advances in the field of analytical chemistry that exhibit green chemistry attributes. However, general considerations for the practice of green analytical chemistry can be summarized:

Planning. Proper planning allows the maximum amount of information to be obtained from the minimal number of analyses. A branch of statistics called Design of Experiments can be used to guide planning laboratory procedures. The developing field of chemometrics allows

us to uncover relationships between sample sets that might be hidden if conventional thinking is used.

Sampling. Attention to proper sampling is an often overlooked portion of chemical analysis, but perhaps the most important step in the entire analysis. A common misconception is that an analytical method can be made green simply by being performed at the micro-scale. While this is true, one must take care to ensure that the appropriate (minimum) sample size is used. The sample must be collected to be statistically representative of the system under study and homogenized to reduce error.

Direct Analysis. Methods which do not require sample work-up prior to the measurement step have several advantages, including greenness. Techniques employing ion-selective electrodes, reflectance spectroscopy, or surface analysis can often provide chemical information without using additional analytical reagents.

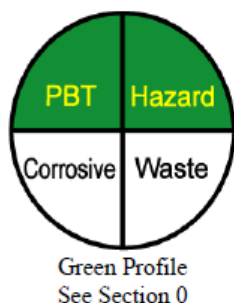
Sample Preparation. Organic extractions and acid digestions generate the greatest amount of waste (while using hazardous solvents or acids) in many procedures. This field has received considerable attention in modern times. Acid digestions can be made safer through the application of microwave irradiation. Alternative solvent processes using water, ionic liquids, or careful manipulation of heat and energy can greatly enhance the sample preparation process.

Chromatography. Separation of analyte from other sample components is necessary in most analytical procedures. Chromatography can involve copious amounts of solvents, sorbents, and other chemicals. Microcolumn chromatography reduces solvent consumption, minimizes the use of chromatographic sorbents, and provides superior performance.

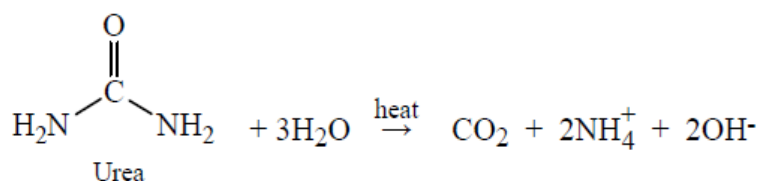
Data Reduction. Statistical manipulations can mine information from analytical results. Many times unambiguous compound identification is not needed; rather knowledge of trends is sufficient.

Field Analysis and Process Analysis. Taking the analysis to the sample generally provides several green advantages. Near instantaneous feedback can minimize the likelihood of a process upset. Similarly, the backhoe operator at an environmental clean-up site would like instantaneous notification that the site is clean so she can quit digging, rather than waiting for results to come back from the laboratory. One contribution of analytical chemistry to the environmental movement has been the identification of hazards within the environment. Careful execution of field and process analysis may advance analytical chemistry from an identification and compliance model to a hazard prevention model.

2. Gravimetric Determination of Calcium as $\text{CaC}_2\text{O}_4 \cdot \text{H}_2\text{O}$ ¹³



Calcium ion can be analyzed by precipitation with oxalate in basic solution to form $\text{CaC}_2\text{O}_4 \cdot \text{H}_2\text{O}$. The precipitate is soluble in acidic solution because the oxalate anion is a weak base. Large, easily filtered, relatively pure crystals of product will be obtained if the precipitation is carried out slowly. Slow precipitation is achieved by dissolving Ca^{2+} and $\text{C}_2\text{O}_4^{2-}$ in acidic solution and gradually raising the pH by thermal decomposition of urea:



Reagents

Ammonium oxalate solution: Make 1 L of solution containing 40 g of $(\text{NH}_4)_2\text{C}_2\text{O}_4$ plus 25 mL of 12 M HCl. Each student will need 80 mL of this solution.

Methyl red indicator: Dissolve 20 mg of the indicator in 60 mL of ethanol and add 40 mL H_2O .

0.1 M HCl: (225 mL/student) Dilute 8.3 mL of 37% HCl up to 1 L.

Urea: 45 g/student.

Unknowns: Prepare 1 L of solution containing 15–18 g of CaCO_3 plus 38 mL of 12 M HCl.

Each student will need 100 mL of this solution. Alternatively, solid unknowns are available from Thorn Smith.²

Procedure

1. Dry three medium-porosity, sintered-glass funnels for 1–2 h at 105°C. Cool them in a desiccator for 30 min and weigh them. Repeat the procedure with 30-min heating periods until successive weighings agree to within 0.3 mg. Use a paper towel or tongs, not your fingers, to handle the funnels. Alternatively, a 900-W kitchen microwave oven dries the crucible to constant mass in two heating periods of 4 min and 2 min (with 15 min allowed for

cooldown after each cycle).¹⁴ You will need to experiment with your oven to find appropriate heating times.

2. Use a few small portions of unknown to rinse a 25-mL transfer pipet, and discard the washings. *Use a rubber bulb, not your mouth, to provide suction.* Transfer exactly 25 mL of unknown to each of three 250- to 400-mL beakers, and dilute each with ~75 mL of 0.1 M HCl. Add 5 drops of methyl red indicator solution to each beaker. This indicator is red below pH 4.8 and yellow above pH 6.0.
3. Add ~25 mL of ammonium oxalate solution to each beaker while stirring with a glass rod. Remove the rod and rinse it into the beaker with small portions of distilled water. Add ~15 g of solid urea to each sample, cover it with a watchglass, and boil gently for ~30 min until the indicator turns yellow.
4. Filter each hot solution through a weighed funnel, using suction (Figure 2-17 in the textbook). Add ~3 mL of ice-cold water to the beaker, and use a rubber policeman to help transfer the remaining solid to the funnel. Repeat this procedure with small portions of ice-cold water until all of the precipitate has been transferred to the funnel. Finally, use two 10-mL portions of ice-cold water to rinse each beaker, and pour the washings over the precipitate.
5. Dry the precipitate, first with aspirator suction for 1 min, then in an oven at 105°C for 1–2 h. Bring each filter to constant mass. The product is somewhat hygroscopic, so only one filter at a time should be removed from the desiccator, and weighings should be done rapidly. Alternatively, the precipitate can be dried in a microwave oven once for 4 min, followed by several 2-min periods, with cooling for 15 min before weighing. This drying procedure does not remove the water of crystallization.
6. Calculate the average molarity of Ca^{2+} in the unknown solution or the average weight percent of Ca in the unknown solid. Report the standard deviation and relative standard deviation (s/\bar{x} = standard deviation/average).

14. Iodimetric Titration of Vitamin C²³



Green Profile
See Section 0

Ascorbic acid (vitamin C) is a mild reducing agent that reacts rapidly with triiodide (See Section 16-7). In this experiment, we will generate a known excess of I_3^- by the reaction of iodate with iodide (Reaction 16-18), allow the reaction with ascorbic acid to proceed, and then back titrate the excess I_3^- with thiosulfate (Reaction 16-19 and Color Plate 11).

Reagents

Starch indicator: Make a paste of 5 g of soluble starch and 5 mg of Hg_2I_2 in 50 mL of distilled water. Pour the paste into 500 mL of boiling distilled water and boil until it is clear.

Sodium thiosulfate: 9 g $Na_2S_2O_3 \cdot 5H_2O$ /student.

Sodium carbonate: 50 mg Na_2CO_3 /student.

Potassium iodate: 1 g KIO_3 /student.

Potassium iodide: 12 g KI /student.

0.5 M H_2SO_4 : 30 mL/student.

Vitamin C: Dietary supplement containing 100 mg of vitamin C per tablet is suitable. Each student needs six tablets.

0.3 M H_2SO_4 : 180 mL/student.

Preparation and Standardization of Thiosulfate Solution

1. Prepare starch indicator by making a paste of 5 g of soluble starch and 5 mg of Hg_2I_2 in 50 mL of water. Pour the paste into 500 mL of boiling water and boil until it is clear.
2. Prepare 0.07 M $Na_2S_2O_3$ ²⁴ by dissolving ~8.7 g of $Na_2S_2O_3 \cdot 5H_2O$ in 500 mL of freshly boiled water containing 0.05 g of Na_2CO_3 . Store this solution in a tightly capped amber bottle. Prepare ~0.01 M KIO_3 by accurately weighing ~1g of solid reagent and dissolving it in a 500-mL volumetric flask. From the mass of KIO_3 (FM 214.00), compute the molarity of the solution.

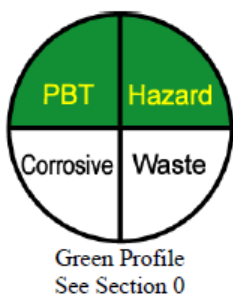
3. Standardize the thiosulfate solution as follows: Pipet 50.00 mL of KIO_3 solution into a flask. Add 2 g of solid KI and 10 mL of 0.5 M H_2SO_4 . Immediately titrate with thiosulfate until the solution has lost almost all its color (pale yellow). Then add 2 mL of starch indicator and complete the titration. Repeat the titration with two additional 50.00-mL volumes of KIO_3 solution. From the stoichiometries of Reactions 16-18 and 16-19, compute the average molarity of thiosulfate and the relative standard deviation.

Analysis of Vitamin C

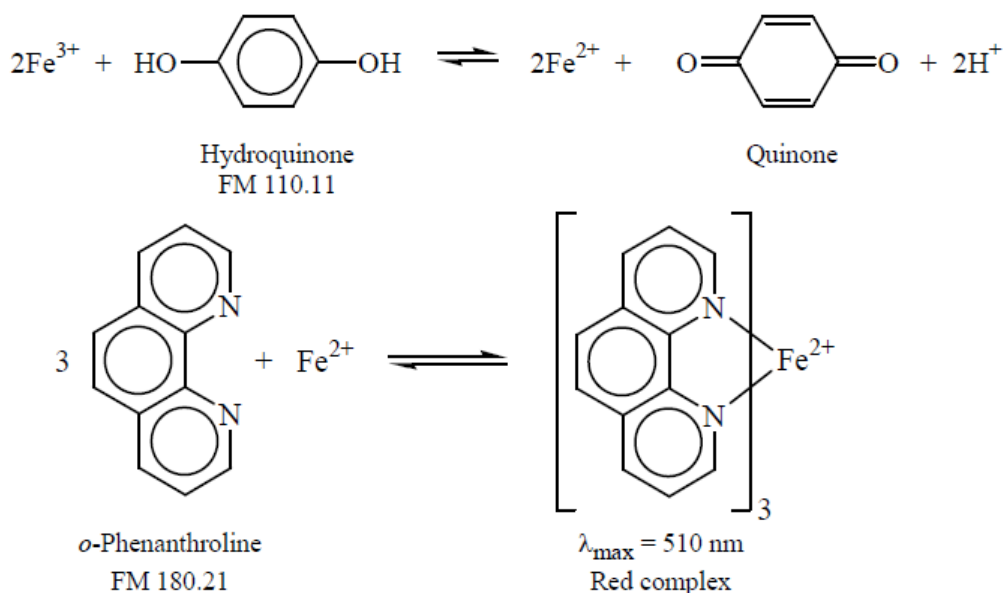
Commercial vitamin C containing 100 mg per tablet can be used. Perform the following analysis three times, and find the mean value (and relative standard deviation) for the number of milligrams of vitamin C per tablet.

1. Dissolve two tablets in 60 mL of 0.3 M H_2SO_4 , using a glass rod to help break the solid. (Some solid binding material will not dissolve.)
2. Add 2 g of solid KI and 50.00 mL of standard KIO_3 . Then titrate with standard thiosulfate as above. Add 2 mL of starch indicator just before the end point.

20. Spectrophotometric Determination of Iron in Vitamin Tablets³⁶



In this procedure, iron from a vitamin supplement tablet is dissolved in acid, reduced to Fe^{2+} with hydroquinone, and complexed with *o*-phenanthroline to form an intensely colored complex (Color Plate 15 in the textbook).



Reagents

Hydroquinone: (20 mL/student) Freshly prepared solution containing 10 g/L in distilled water.

Store in an amber bottle.

Trisodium citrate: (20 mL/student) 25 g/L $\text{Na}_2\text{citrate} \cdot 2\text{H}_2\text{O}$ (FM 294.10) in distilled water.

***o*-Phenanthroline:** (25 mL/student) Dissolve 2.5 g in 100 mL of ethanol and add 900 mL of distilled water. Store in an amber bottle.

6 M HCl: (25 mL/student) Dilute 124 mL of 37 wt% HCl up to 250 mL with distilled water.

Standard Fe (40 $\mu\text{g Fe/mL}$): (35 mL/student) Dissolve 0.281 g of reagent-grade

$\text{Fe}(\text{NH}_4)_2(\text{SO}_4)_2 \cdot 6\text{H}_2\text{O}$ (FM 392.14) in distilled water in a 1-L volumetric flask containing 1 mL of 98 wt% H_2SO_4 .

Procedure

1. Place one tablet of the iron-containing vitamin in a 125-mL flask or 100-mL beaker and boil gently (*in a fume hood*) with 25 mL of 6 M HCl for 15 min. Filter the solution directly into a

- 100-mL volumetric flask. Wash the beaker and filter several times with small portions of water to complete a quantitative transfer. Allow the solution to cool, dilute to the mark and mix well. Dilute 5.00 mL of this solution to 100.0 mL in a fresh volumetric flask. If the label indicates that the tablet contains <15 mg of Fe, use 10.00 mL instead of 5.00 mL.
2. Pipet 10.00 mL of standard Fe solution into a beaker and measure the pH (with pH paper or a glass electrode). Add sodium citrate solution 1 drop at a time until a pH of ~3.5 is reached. Count the drops needed. (It will require about 30 drops.)
 3. Pipet a fresh 10.00-mL aliquot of Fe standard into a 100-mL volumetric flask and add the same number of drops of citrate solution as required in step 2. Add 2.00 mL of hydroquinone solution and 3.00 mL of *o*-phenanthroline solution, dilute to the mark with water, and mix well.
 4. Prepare three more solutions from 5.00, 2.00, and 1.00 mL of Fe standard and prepare a blank containing no Fe. Use sodium citrate solution in proportion to the volume of Fe solution. (If 10 mL of Fe requires 30 drops of citrate solution, 5 mL of Fe requires 15 drops of citrate solution.)
 5. Determine how many drops of citrate solution are needed to bring 10.00 mL of the iron tablet solution from step 1 to pH 3.5. This will require about 3.5 or 7 mL of citrate, depending on whether 5 or 10 mL of unknown was diluted in the second part of step 1.
 6. Transfer 10.00 mL of solution from step 1 to a 100-mL volumetric flask. Add the required amount of citrate solution determined in step 5. Then add 2.00 mL of hydroquinone solution and 3.0 mL of *o*-phenanthroline solution. Dilute to the mark and mix well.
 7. Allow the solutions to stand for at least 10 min. Then measure the absorbance of each solution at 510 nm in a 1-cm cell. (The color is stable, so all solutions may be prepared and all the absorbances measured at once.) Use distilled water in the reference cuvet and subtract the absorbance of the blank from the absorbance of the Fe standards.
 8. Make a graph of absorbance versus micrograms of Fe in the standards. Find the slope and intercept (and standard deviations) by the method of least squares. Calculate the molarity of $\text{Fe}(\text{o-phenanthroline})_3^{2+}$ in each solution and find the average molar absorptivity (ϵ in Beer's law) from the four absorbances. (Remember that all the iron has been converted to the phenanthroline complex.)
 9. Using the calibration curve, find the number of milligrams of Fe in the tablet. Use Equation 4-27 in the textbook to find the uncertainty in the number of milligrams of Fe.