Carnegie Mellon

NSF Grant DMR 079996

High School Teachers Summer Internship

Experiments in Steel

Linda Neumann
ChemistryTeacher
Taylor Allderdice High School

2003

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Introduction

A PURPORTED DEBATE AMONG MIDDLE AGE PHILOSOPHERS AND THEOLOGIANS REVOLVED AROUND, "HOW MANY ANGELS CAN DANCE ON THE HEAD OF A PIN?"

THE OBJECTIVE OF THIS PROJECT WAS MUNDANELY ANALOGOUS: "HOW MANY ACTIVITIES OR EXPERIMENTS RELEVANT TO THE CHEMISTRY CLASSROOM CAN BE ENVISIONED ON THE HEAD OF A STEEL NAIL?"
ANSWER MANY MORE THAN ARE INCLUDED HERE!

***No matter what conceptual real-life analogies we choose, what mathematics we apply, what media we use, which experiments we routinely conduct, or even what degree of sophistication our approach, as high school chemistry teachers we normally cover the same broad topics of study in our classrooms*.**
Why not try a fresh approach to your standard analogies and experiments using

Steel

an alloy of particular interest to students living in a city that calls its football team the "Steelers" ; a city with an economic and historical development that revolved around this very substance for many years!

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Objective

The overall purpose of this set of laboratory activities is to use a specific material, steel, in meaningful high school chemistry experiences. The corrosion of steel is an electrolytic process that can be studied easily within the parameters of a high school chemistry classroom. The experiments described herein provide some very tangible applications of; experimental design, phases of matter, crystal structures, oxidation/reduction reactions, Standard Reduction Potential, electrochemistry, LeChatlier's Principle, G, Free Energy and others as discussed throughout this study.

How many of theses concepts are drawn upon and taught with the following experiments, depends on the objectives which the teacher desires to fulfill. **From the simplest perspective**, the following experiment can function as no more than good practice in **experimental design**; practice observing steel corrosion, hypothesizing as to the underlying factors responsible for corrosion, setting up a multivariable experiment and gathering the data The **data is analyzed** for evidence of "trends" **without** interpretation of results in terms of any **underlying causal factors.**

A more complex treatment of the experiment would require that the student have a good grasp on **electrolytic processes** and the function of an electrochemical cell; mechanisms by which corrosion of steel occurs. This would enable the student to more effectively identify and "interpret" data trends in light of a **known mechanism**.

The most complex treatment of the experiment would include all of the above plus **knowledge of the composition** of the different steels being used (**SAE # 1008, 1045, 4140, 4340)** and some knowledge of the various **microstructures** that may occur in these steels. This would allow for interpretation of data trends in light of the microstructure of the steels involved (how the various alloying elements in the **crystal structure** or the %C phase mix might possibly **interact with the electrolytic mechanism** of steel corrosion.)

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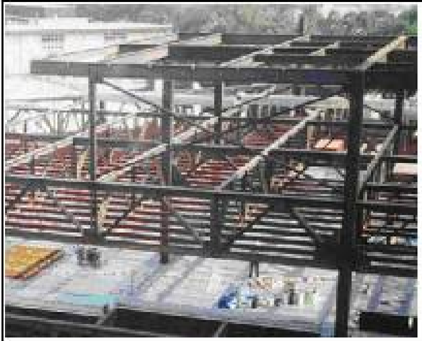

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Initial discussion with students will likely indicate that many know steel contains the element iron. Also, from everyday life, many will no doubt be familiar with the "corrosion" of iron appearing as a red, low-density substance commonly called "rust." A few students may even suggest that rust is the compound, iron (III)oxide (under most conditions rust is actually iron(III) hydroxide.) When it comes to hypothesizing what it is that causes steels to rust, a few students should know that moisture definitely plays a part (if they've ever left their bike out in the rain, etc.)

To hypothesize about corrosion, students must first be familiar with the simplest definition of steel as an iron-based carbon alloy (containing up to approx. 2% carbon.) Let them examine firsthand some partially corroded iron and steel objects; demo objects should be drawn from a variety of environmental conditions (rust on an "indoor" object such as scissors; rust on an object often left "outdoors," such as a garden tool; a picture of cars rusting in a junkyard etc.) After examining the objects, students should at least be able to hypothesize that rust might be related to the environmental conditions surrounding a steel object**.**

Experimenting with steel requires comprehension of the term alloy. According to the Metals Handbook, 8th ed., published by the American Society for Metals it is," A substance having metallic properties and being composed of two or more chemical elements of which at least one is an elemental metal." Have several examples of alloys available for students to examine, such as a brass ring. Steel is "iron-based" so that is its dominant metal; and carbon up to 2% is the other most common alloying element in steel.

Do students think that "steel" always means the same substance? Go back to the term "alloy." An alloy is a "mixture" of substances; not a compound! Chemistry students should be able to differentiate a compound from a mixture. Calling on what they know about mixtures (constituents can be mixed in any proportion, separable by physical means), have them conjecture whether or not all steels are the same (of course, they are not.) There are many types of steel containing different proportions of the alloying elements; that is, steels vary widely in composition.

For students who have no difficulty with the above ideas, go on and build on their knowledge of mixtures, introducing them to the concept of the "microstructure" of steels. Students should be aware of two mixture types, heterogeneous and homogeneous, from their previous study of the Composition of Matter. They know that a heterogeneous mixture is "multi-phased" and that a homogeneous mixture is termed a "solution" which is the same phase throughout. Do students think these ideas also apply to the mixtures called alloys; particularly the alloy steel? Indeed, they do! Not only do steels vary in chemical composition; the same steel can have either single-phased or multi-phased forms depending on (among other things) its thermodynamic history. According to Noel F. Kennon in a recent book edited by Leonard E. Samuels called Light Microscopy of Carbon Steels , at normal temperatures many steels are heterogeneous two-phased mixtures of ferrite (iron atoms arranged in a body-centered cubic crystal structure) mixed with cementite (carbon in the form of the compound, iron carbide, Fe_3C). At higher temperatures, the same steel can transform into a homogeneous mixture, that is a solid solution, of carbon atoms dissolved in austenite (iron atoms arranged in a face-centered cubic crystal structure.) (6-9, 29)

After students realize that steels have many different compositions and microstructures, explain that steels can be classified according to their compositions by a Society of American Engineers (SAE) number such as 1008, 1045, etc. If teaching a more advanced course, if enrichment is desired, or if your students can handle the material, at this point the body-centered cubic crystal structure and the face-centered cubic crystal structure can be illustrated with models and associated with the phase of steel to which they apply. (Van Vlack 59 -84)

Building on these ideas; that steels are alloys containing iron, that iron "corrodes" into iron (III) hydroxide, that steel objects seem to corrode more when moisture (water) is present, and that not all steels are the same, have students "hypothesize" factors they think play a role in the corrosion of steel and to what degree (do some factors have a greater effect on steel corrosion than others?) Have them "brainstorm" an experimental set-up designed to test their ideas about steel corrosion. Any experimental design should include the above ideas. For mainstream chemistry classes, the foregoing ideas will be enough to handle and test. The discussion on steel microstructure and crystal structure should be eliminated.

Advanced Chemistry Classes and Gifted Chemistry Classes, however, should be required to understand the "mechanism" by which steel corrodes. They should call upon the equations for the corrosion of steel when designing any experiment on corrosion. Activities or discussions on these topics should be conducted in second and third lab sessions or during regular classes.

"The Selection of Mild Steel for Corrosion Service," (Metals Handbook 1:257) states, "the corrosion of iron and steel has become accepted as a phenomenon that is essentially electrolytic, especially where attack depends on the simultaneous presence of moisture and oxygen." This means that for corrosion to occur, moisture (water)containing charged particles, an electrolyte, must be present. Electrons are carried away from the neutral Fe atoms in steel, forming iron (II) ions, Fe^{2+} to a substance with a higher potential to gain electrons. In a sense, this forms a local "electrochemical cell" on the surface of the steel. (Corrosion in Action 15) Students who understand the function of and the nomenclature associated with the electrochemical cell will have a much better grasp on the corrosion process. Unless students are very comfortable with electrochemistry and the electrochemical cells, it would be good to spend several class periods diagramming electrochemical cells; building several simple cells in the lab. (Petrucci 471 - 490)

Students must at least know that oxidation, electron loss, occurs at the electrode in a cell called the anode (Fe functions as an anode when it gives up electrons from a steel surface and corrodes.) Reduction ,electron gain, occurs at the electrode in a cell called the cathode and normally O_2 functions in this capacity on the steel surface. Knowing the process is electrolytic, students hypothesize that water solutions containing ions, such as salt water, will have a greater effect on the speed of corrosion than will pure water. Solution type, as well as steel type, now becomes a factor to be considered in the corrosion experiment.

Consider the electrochemical equations associated with corrosion. Even in the absence of O_2, it has been demonstrated that ferrous metals will go into solution in water until a state of equilibrium is reached according to the following equation:

Fe	+	$2\,H_2O$	=	Fe^{2+}	+	$2OH^-$	+	H_2
iron		water		ferrous		hydroxide ions		hydrogen gas ions

The ferrous and hydroxide ions result from the single displacement reaction that reduces H^+ in the water to hydrogen gas; which is sometimes "plated out" on the surface of the metal. If the conditions are such that the equilibrium constant for this reaction is low enough and the vapor space small enough, the reaction will stop before there is saturation with respect to ferrous hydroxide. ("The Selection of Mild Steel for Corrosion Service" Metals Handbook 8th 1: 257) Chemistry students should be very familiar with single replacement reactions and recognize the ions involved. (However, be careful here! My students normally learn that Fe does not replace hydrogen ion from water at room temp...and indeed the kinetics are against it...but then it is not customary to use water devoid of oxygen in the high school chemistry lab.)

The corrosion of steel is an excellent application of LeChatelier's Principle. LeChatelier was a French industrial chemist and professor particularly interested in problems related to metallurgy. His principle states that a system in dynamic equilibrium subjected to a stress will shift in such a way as to remove that stress. (Wilbraham 502) The presence of iron and water will push the reaction to the right (in order to use up the reactants) and enhance corrosion; whereas the excess presence of either product, hydroxide ions, OH^-, or H_2 gas, should cause the above corrosion reaction to go into reverse, and thus slow down the corrosion. (The reaction shifts left to remove the excess OH^- ions or hydrogen gas.) Applying LeChatelier's principle, students can now hypothesize that "basic" conditions around the iron, such as the presence of $CaCO_3$ should retard corrosion. Likewise, if acid is present, the H+ will neutralize the OH- ion, effectively reducing it and driving the reaction to the right; restoring the OH- ion and enhancing corrosion. Hypotheses as to the same concerning basic and acidic solutions should be tested when designing any corrosion experiment.

If O_2 is in the water (as it most always is), then it normally enhances corrosion. Recall, if Fe acts as the "anode," giving up electrons in the local electrochemical cell that enables corrosion, then O_2 can function as the "cathode," taking on those electrons, because a potential difference exists between them.(Corrosion in Action 20) Oxygen is reduced according to the following equation; $O_2 + 2H_2O + 4e^- = 4OH^-$ (Petrucci 484)

Oxygen present in solution also can enhance corrosion in other ways. "Oxygen, if present, reacts with the hydrogen film (that may build up on the steel surface), removes it from the surface of the metal in the formation of water and also reacts with the ferrous ions to form ferric ions in such a concentration as to exceed the solubility product of ferric hydroxide, and hence a precipitate of ferric hydroxide is formed. Both reactions are involved in the corrosion process." ("The Selection of Mild Steel for Corrosion Service" 1:257) An additional equation shows why, as corrosion progresses, the water around the corrosion turns from yellow to orange to red. As mentioned, the O_2 goes on to convert the ferrous ions (yellow in solution) to the characteristic ferric ions (orange/red in solution) associated with "rust" :

$$4\,Fe^{2+} + O_2 + 2H_2O = 4\,Fe^{3+} + 4\,OH^-$$

(Metals Handbook 8th 1: 257)

Students apply LeChatelier's Principle and postulate that any O_2 present in solution will drive the reaction forward, enhancing corrosion. Oxygen level is indeed an important variable in any experiment on the corrosion of steel!

Finally, good experimental design often incorporates indicators as to the reactions being tested. Evidence that OH^- is a product of the reduction of oxygen during corrosion can be tested by the presence of phenolphthalein near the surface of the iron or steel; any place where the reduction of O_2 is occurring on the surface of ferrous metals should become "basic" and turn that indicator pink. Likewise, when corrosion initially begins, iron(II) or ferrous ions build up on the surface of the steel where the oxidation of iron atoms occurs; this site should turn the indicator for ferrous ions, hexacyanoferrate(III), blue. If the iron is held in a gel medium, as with the experiments found in Appendix B of this study, these colors develop nicely showing the electrolytic nature of corrosion (Corrosion in Action 17)

***

The following experiment on the corrosion of steel was conducted during the summer months, and should be similar to the experiments generated by students who have undergone the previous discussions and activities. The experiment is best run by making each lab group responsible for setting up and monitoring one set of solutions throughout the course of the experiment. Joint sets of data could be pooled for a final evaluation of experimental results. The experiment can also be subdivided into a series of smaller experiments that limit the number of variables tested simultaneously. THIS EXPERIMENT DEFINITELY LENDS ITSELF TO AN "OPEN-ENDED APPROACH" IN WHICH EACH LAB GROUP RECEIVES 4 TYPES OF STEEL WITH WHICH TO DESIGN THEIR OWN EXPERIMENT TESTING THE VARIABLES THAT AFFECT STEEL CORROSION

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What will you Need?

Materials:

Steel Samples (cut in rectangular solids 1/2" x 1/2" x 1and 1/2")
13 samples of Low Carbon Steel SAE# 1008
13 samples of Medium Carbon Steel SAE# 1045
13 samples of Medium Alloy Steel SAE# 4140
13 samples of Medium Alloy Steel SAE# 4340
(Suggestion; Mark each block with Steel SAE# using magic marker)
13 1000mL beakers
26 glass rods (cut to fit bottom of above beakers)
2 Liters Distilled H_2O
2 Liters Saline Solution (1M NaCl, .01 M KBr, .01M KI)
2 Liters .005M H_2SO_4 (pH =2)
2 Liters 100ppm $CaCO_3$ solution
4 wash bottles (one for each solution above)
4 berol pipets (one for each solution above)

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Procedures:

**** See Pictures #1 & 2 Appendix* A**

1. Mark 3 beakers "Distilled Water" (A Freshwater Environment)
Additionally, label the first beaker "Minimum O_2", second "Dissolved O_2", and a third "Max O_2."

2. Mark 3 beakers "Saline Solution" (A Marine Environment)
Additionally, label the first beaker "Minimum O_2", second "Dissolved O_2", and a third "Max O_2."

3. Mark 3 beakers " H_2SO_4 solution, pH =2" (A Heavy Industrial Environment)
Additionally, label the first beaker "Minimum O_2", second "Dissolved O_2", and a third "Max O_2."

4. Mark 3 beakers " 100 ppm $CaCO_3$ solution" (A Hard water Area)
Additionally, label the first beaker "Minimum O_2", second "Dissolved O_2", and a third "Max O_2."

5. Mark the last beaker "Open to the Atmosphere/ Dry Conditions."

6. Arrange beakers on a table in groups by solution type (Distilled, Saline, H_2SO_4, 100ppm $CaCO_3$.)
Additionally arrange beakers within each group by increasing O_2 level: Minimum O_2, Dissolved O_2, Maximum O_2 ***The effect of both solution type and oxygen level will be tested.***

7. Sit the "Dry Conditions/ Atmospheric Exposure" beaker by itself.

8. Place two of the glass rods horizontally in the bottom of the "Dry Conditions" beaker and arrange the four types of steel in order of increasing **SAE # 1008, 1045, 4140, 4340.** **** See Picture #3***
These will serve as the experimental control. Any corrosion due to interaction between normal atmospheric conditions and specific steel type should show up here.

Also, make sure that the initial depth of each sample or the initial mass of each sample is marked on a folder assigned to each beaker prior to the experiment if you plan to use initial values in the calculation of extent corrosion. * See Measurement of Variables

9. Place two glass rods horizontally in the bottom of each remaining beaker.

10. As mentioned, prepare a manila **folder marked** "**Distilled Water/ Minimum O_2."** Remove the four steel samples from this folder. In the **beaker** marked **"Distilled Water/Minimum O_2"** arrange these samples in order of increasing SAE # 1008, 1045, 4140, 4340.

11. Proceed in the same manner as # 8 and #10 above for each of the remaining beakers, making sure to match the proper set of steel samples from the manila folders to the correct conditions under which they shall be tested as marked on the beaker.

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At the conclusion of these laboratory experiences (dependent on sophistication of approach) students should be able to do some or all of the following:

1. set up and conduct a multivariable experiment with appropriate control, systematically gather and classify data, and analyze data for any dominant trends therein.

2. describe steel as an iron-carbon alloy containing up to approx. 2% carbon (and always traces of manganese). It can also contain several other elements, called alloying elements, added to affect its properties. (Samuels 29-30).

3. indicate that there are many types of steel with great variations in their microstructures; these being dependant on the initial composition of the steel and how the steel was processed both thermodynamically and mechanically. ("Compositions of Standard Steels." and "Hot Finished Carbon Steel." Metals Handbook. 8th ed., Vol 1: 61-63)

4. show a general awareness of the range of microstructures that can occur in steels and be able to describe several of these.

5. describe, using a body-centered cubic crystal model, the most common microstructure of many steels at normal temperatures. This is essentially a two-phased heterogeneous system consisting of ferrite (iron atoms arranged in a bcc array) mixed with carbon In some from of an iron carbide compound, Fe_3C, called cementite. (Samuels 6-9)

6. describe, using a face-centered cubic crystal model, another microstructure that can occur in steels (at elevated

temperatures.) This is a one phase, homogenous mixture or a "solid solution" of carbon atoms "dissolved" in austenite (iron atoms arranged in a fcc array.) The carbon atoms are found in the interstitial spaces (spaces in the crystal structure between the iron atoms.) (6, 29)

7. explain that the corrosion of steel is an electrolytic process in which electrons are transferred away from the Fe in steel through a solution containing charged particles, called electrolytes. ("The Selection of Steel for Mild Corrosion Service." 1: 257)

8. define an oxidation half reaction as one in which electrons are lost

9. define reduction half reaction as one in which electrons are gained

10. describe an electrolytic process as one in which electrons will flow from an area of high charge density to an area of low charge density through an electrolyte; a moist interface containing charged particles.

11. describe the components of a "half-cell" as a metal strip called an electrode immersed in a solution of its own ions

12. construct a simple electrochemical cell (also called a voltaic or galvanic cell) by connecting two "half-cells" to a voltmeter (after placing a salt bridge between them.) Give evidence that electrons are flowing between the two half cells.

13. identify the parts of the electrochemical cell (anode where oxidation occurs) and cathode (where reduction occurs), electrolyte, salt bridge, etc.

14. use the Standard Reduction Potential Table to explain why electrons flow between the anode and the cathode in the electrochemical cell constructed. Connect this to emf.

15. explain corrosion on the surface of a piece of steel in terms of an electrochemical cell.

16. give the equations which define corrosion in steels and explain how each reactant and product affects the corrosion reaction in terms of LeChatelier's Principle.

17. explain the design of an experiment to test the environmental factors that affect the rate of corrosion in steels and tell why each environmental factor is included in the experiment.

18. in terms of what is known about steel corrosion, hypothesize as to the results of this corrosion experiment; both as to steel type and as to specific environmental conditions.

19. define "galvanic couple" in terms of the electrochemistry of metals.

20. explain what is meant by steel acting "anodic" and what is meant by steel acting "cathodic."

21. explain how steel can be made to act "cathodic" and what is meant by "cathodic" protection.

22. tell what indicator you would use to test for steel areas acting
"anodic" and forming Fe^{2+} ions; and what color you expect

23. tell what indicators you would use to show that steel is acting
"cathodic" and what color to expect

24. explain why the potential on stainless steel should differ from that on low carbon steel, and describe an experiment that would show this

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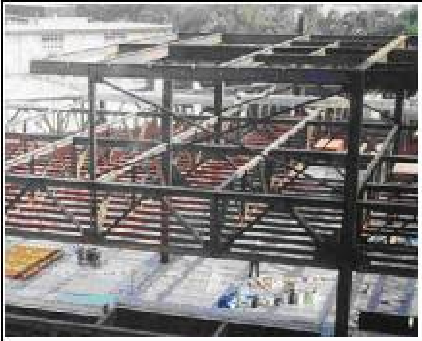

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EXPERIMENTAL RESULTS

Qualitative

Data Grid (*Observations)*

Corrosion of Steels **SAE# 1008, 1045, 4140, 4340**

Observations

Date *7/29/03 Day 14* **Time** *3:00 PM*

Solution Type	**Minimum O_2 (sealed)**	**Dissolved O_2**	**Max O_2 (wet/dry)**
Saline	orange/red deposit on bottom (1mm) thick(3mm) dep on top of all blks, furry red/orange on sides solution: yell/or particles	thick red deposit on bottom (5-6mm) thick, furry ,drk red corr 1008(most)-1045 4340-4140(least) solution:org/ ppt red	film of red/org susp on bott 1008(blk pts 4140-1045-4340(least)
Acidic pH=2 H_2SO_4	bottom clear 4140 bubbles on surface no corr-4340 few bub slight corr-1045(no bub top (sl corr)-1008 lght yell top (most corr) solution: clear/yell tinge	orange deposit bottom (1mm) blks covered thin or film 1008 (most) 4340-1045-4140(least) solution: clear/or fine or deposit on top	slight org chnk under 1008 Orang cor top 1008-4140-4340 -1045 (no)
Basic 100ppm $CaCO_3$	white/or deposit bottom thk org "icing" all blk, esp 4140 clear yellow sol	deep or/red depost bottom thk corr 1008-4140-1045-4340 clear or sol (suspended org part top)	white/org under 1008 4140 most cor then 1008 white powder top1045,4340

Distilled Water	fine or powder mold-like furry 1008,4140 most	thick red deposit furry or needles 4140 most, oth	flecks of 4140 most, then1008 medium

Quantitative

Ultrasound Measurements

Index to Corrosion:

Sum of depth lost on a 3-Point Test Instrumental Uncertainty

Values x 10^{-4} in - 3 x 10^{-4} in

Minimum Oxygen Level (Immediately Sealed)

	1008	1045	4140	4340
Saline	13	8	22	7
pH =2 Sulfuric Acid	4	11	4	4
Basic ppm 100 Calcium Carbonate	9	9	6	17
Distilled	12	28	0	7

Ultrasound Measurements

Index to Corrosion:

Sum of depth lost on a 3-Point Test

InstrumentalUncertainty Values x 10^{-}

^{4} in +/- 3 x 10^{-4} in

Dissolved Oxygen Level (Liquid Maintained at 400mL Level)
Oxygenated & Stirred Daily)

	1008	**1045**	**4140**	**4340**
Saline	6	23	0	0
pH =2 **Sulfuric Acid**	3	16	9	10
Basic ppm 100 **Calcium** **Carbonate**	3	24	0	6
Distilled	8	4	0	2

Ultrasound Measurements

Index to Corrosion:

Sum of depth lost on a 3-Point Test Instrumental Uncertainty

Values x 10^{-4} in +/- 3 x 10^{-4} in

Maximum Oxygen Level (Bathed Daily in Given Solution)
Wet/Dry Interface

	1008	**1045**	**4140**	**4340**
Saline	38	12	0	13
pH =2	22	16	2	1

Sulfuric Acid					
Basic Calcium Carbonate	**ppm 100**	0	13	1	0
	Distilled	21	6	0	6

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INTERPRETATION OF EXPERIMENTAL RESULTS (CONCLUSIONS)
FEEDBACK FOR FUTURE EXPERIMENTS

After conducting this experiment, several conclusions were immediately apparent. **First off, from a "qualitative" standpoint, the experiment was indeed a success. More viable data can be gleaned "visually" from the observations made during this experiment than would at first appear to be the case.** For example, steel type 4140 began to corrode in the distilled water, max oxygen (wet/dry interface) environment on the second day of the experiment; progressing rapidly from the yellow Fe(II) initial corrosion stage, through the "intermediate orange" Fe(II) and Fe(III) mix stage, to the reddish Fe(III) hydroxide stage within the duration of this two week experiment. On the other hand, steel type 4340, in the same environment, took a week before showing any signs of corrosion and hardly progressed through the light yellow Fe(II) stage in two weeks. High school students could make such observations and record as quantitative data the number of days it takes for a steel type in a given environment to show any corrosion at all; the number of days only yellow corrosion is present; the day on which orange corrosion first appears (such as 7th or 8th) etc. Looking up the compositions of steels 4140 and 4340 and relating these compositions to the recorded experimental observations is an exercise in connecting the external properties of materials with their internal compositions. The compositions of these steels are as follows:

(Ladle chemical composition limits %) (**Metals Handbook, 8th ed 61)**

SAE
4140 .38-.43%C .75-1.00%Mn .040%P .040%S .20-.35% Si **0% Ni**
.80-1.10%Cr .15-.25% Mo 0%V

4340 .38-.43%C .60-.80% Mn .040%P .040%S .20-.35% Si **1.65-2.00%** .70-.90% Cr .20-.30%Mo 0%V

Ni

Both steels are considered to be medium-carbon alloy steels and their compositions are quite similar. The one component noticeably different between 4140 and 4340 is the nickel content; 4140 has none! Analyzing the data drawn from experimental observations that in distilled water (wet/dry interface) steel 4140 corrodes much faster and to a greater extent than steel 4340, it would not be unreasonable to conclude that Ni as an alloying element definitely must play some role in retarding the corrosion of steel; at least in that specific environment.

Nor are such conclusions the extent to which the "qualitative" approach can be taken. Students have ample experimental data to go on and access observations about the corrosion of 4140 and 4340 in all the other solutions tested (keeping the wet/dry interface constant.) The data shows that in basic solution and in acid solution, the same is the case; 4140 corrodes earlier and to a greater extent than 4340. But just when students are sure that the Ni content difference provides corrosion protection in every solution, observations show this is not the case in saline solution! However, the fact that 4340 seems corrosion protected in the three or the four solution types, must indicate that Ni alloying at least often retards corrosion in the wet/dry scenario.

Because this experiment has so many aspects to it, students can analyze the data further and check on the corrosion behavior of 4140 and 4340 at oxygen levels other than max (wet/dry.) No matter whether the solution is sealed from the atmosphere (low ox) or open to the atmosphere at 400mL line, in basic solutions or distilled water, 4140 is still the most corroded and 4340 the least or close to it. The conclusion that the effect of alloying with Ni retards corrosion still holds credence.
However, when the acid and saline solutions are checked, a remarkable trend surfaces...4140 is the least corroded at any oxygen level in the harsh saline and acid environments! Do these environments actually interact negatively with the presence of Ni so as to overcome its corrosion retarding effects; or does Ni do something to the microstructure of the steel so that it is actually "softer" than 4140 and therefore more susceptible to attack in the harsher environments of salt and acid? Students can investigate all the topics and trends generated by this one experiment alone!

Qualitatively, a few pleasant surprises even occurred. The most striking of these was observing all four types of steel sitting in pH 2 sulfuric acid (*sealed)* for ten whole days before a sign of corrosion appeared on any of them...not a tinge of yellow! These steel samples were covered with gas bubbles from the onset of the experiment. Even if not a shred of "quantitative" data was derived from this experiment (and that is just about the case as will be later discussed), a high school class could be given the corrosion equation under "low oxygen" conditions (see page 7 of this paper) and asked to explain the observed phenomenon. Students could be asked to think in terms of the H+ ion abundant in acid conditions and directed to what they know about the Reduction Potentials between Fe and H+. The potential to be reduced is 0.000ev for H+ and - .440 ev for Fe (Petrucci 476) showing that Fe will give up electrons to H+. Students could be asked to explain the origin of the bubbles surrounding the steel samples as well as the identity of the gas in the bubbles (H+ ion is reduced to hydrogen gas which appears as bubbles on the steel when the Fe in steel is oxidized.) Using the corrosion equation under "low oxygen" conditions, students should explain, via LeChatelier's principle, why hydrogen bubbles clinging to the surface of steel will retard corrosion. Specifically, "Where there is no electrolytic coupling, the film of hydrogen will be on the surface of the iron or steel and, in the absence of oxygen, will hinder further corrosion by insulating the metal from the water and by its own tendency toward a back reaction. The formation of such a gas film on metal is called 'polarization'." (" The Selection of Mild Steel for Corrosion Service" Metals Handbook, 8ed. 1: 257) ***See Picture #6 (Compare saline/front with acid (hydrogen bubbles)/ back)*** It is of interest to note the connection, again, between steel composition and the formation of hydrogen bubbles on the surfaces of the steel types in the sealed acid solution. Bubbles formed fastest and in greatest amount around steel type 4140! This was the only steel type to retain all of the bubbles for the duration of the experiment and the only steel type to show no corrosion. Students can go on to investigate connection between resistance to corrosion in acid and the 4140 composition.

Hydrogen bubbles initially formed around all the steel samples in the acid solution open to the atmosphere. They did not start to corrode until the 3rd day of the experiment; even sitting in acid. Corrosion started on some of the blocks when the bubbles disappeared. Students can be asked to explain the observation that bubbles on the samples in the solution open to the atmosphere finally disappeared; whereas the bubbles stayed on the steel samples for two weeks in the acid sealed from the atmosphere. Here again, a simple qualitative observation leads to another avenue of investigation. Most likely, the oxygen initially dissolved in the sealed acid container was expended; the acid solution open to the atmosphere and stirred continued to bring oxygen to the surface of the steel. The oxygen removed the hydrogen film and corrosion proceeded. "Oxygen, if present, reacts with the hydrogen film, removes it from the surface of the metal in the formation of water and also reacts with the ferrous ions to form ferric ions......Both reactions are involved in the corrosion process;" (1: 257) See picture #7

Finally, when the parafilm was removed from the sealed acid beaker, little bubbles were observed all over the surface of the water and clinging to the under surface of the parafilm covering. Could the vapor pressure of the hydrogen gas being released above the acid solution in the sealed container have increased to the point that it created a back pressure of hydrogen on the liquid surface and therefore held the remaining hydrogen bubbles in solution? High school chemistry students could make that observation from this experiment and be asked to explain it in terms of Henry's Law (the solubility of a gas in a liquid is directly proportional to the pressure of that gas above a liquid.) **Within the analytical treatment of the qualitative observations generated by this experiment, lie many more investigations and experiments. Not a few concepts learned in high school chemistry can be drawn upon or reinforced with analysis of the qualitative data alone. The goal of this project has been reached; using the study of steel and its composition to generate meaningful activities and laboratory experiences for the high school chemistry classroom.**

As for the "quantitative" data generated by the **ultrasound measurements** in this experiment; it was **basically meaningless.** The limited time frame in which the experiment had to be conducted was from the outset a major concern. The experiment was run during the summer for little more than two weeks. This was sufficient to generate a wealth of qualitative data and to illustrate many chemistry concepts as discussed above; but it was not enough time for corrosion to remove a measurable amount of material; at least in terms of the ultrasound equipment being used. Any mass loss would have no doubt been equally, if not more so, imperceptible. It should be noted, however, that the ultrasound equipment being used was not designed to do "corrosion mapping"; this type of equipment would have given a digital map of the surface of each steel sample. However, that level of instrumental sophistication is not within the range of most high school budgets; and this was an attempt to simulate an experiment for the high school chemistry class.

Three- point test depth readings were taken on each of the 52 steel samples prior to and after conducting the experiment. The ultrasound gage was calibrated against the four steel samples that served as the control; determining the velocity of sound in each type of steel. All steel samples of a given steel type were measured before moving to a different steel type and depth readings were recorded to the ten- thousandths of an inch. The initial and final depths of each sample are not recorded herein, but only the sum of the depth loss over the 3-point test for each sample; this was considered to be an index to the extent of corrosion on each block. The total depth lost for each sample in each solution and at each level of oxygen is shown on pages 17 - 19. The following is an analysis of that data.

The "uncertainty" on any one measurement taken with the ultrasound equipment being used was +/- 3 ten -thousandths of an inch. This means that the uncertainty on one depth measurement on each block (final - initial depth measurement) was +/- 6 ten-thousandths of an inch as uncertainties are additive when subtracting. The "index to corrosion" was the sum of the depths lost at three places on the block; and again, since uncertainties are additive, **the uncertainty on each answer** was the sum of the three uncertainties for each depth measurement or **+/- 18 ten-thousandths of an inch!**

This means that if there was no change in depth, i.e., the depth change was 0 (no corrosion) it could conceivably show up as a loss of 18 ten-thousandths of an inch; and if a block actually lost 18 ten-thousandths of an inch (by corrosion), it could conceivably show up as losing nothing! So unless the sum of the depth lost is considerably more than the uncertainty, the value in this case is basically an index to nothing! And when only 7 data pieces out of 52 are greater than the uncertainty on their values, it appears that an instrument of much more precision must be used before any valid conclusion can be drawn as to depth loss due to corrosion.

Now the data is beguiling; it at first appears to have some plausible trends in it. It is true that some of the highest values of total depth lost appear in the max oxygen (wet/dry) set of data especially for steel type 1008. Yet at this same oxygen level in basic conditions, 4140 is the only steel sample that developed dark orange corrosion and it is shown as having lost only 1 ten-thousandths of an inch, while a sample of 1045 steel, coated with calcium carbonate deposit and still shiny when this was removed after two weeks and showed only a tinge of corrosion, supposedly lost 13 ten-thousandths of an inch! The ultrasound data just is not valid for this particular experiment.

The upshot of this is that the corrosion experiment as such should either be run from six months to a year in order to generate measurable quantitative data or run for the two-week duration and used qualitatively. In the high school chemistry curriculum, qualitatively is no doubt the more reasonable and practical choice!