

Planning and Teaching Science

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Course 15.4869 Theory of Science Teaching.

Senior 3 Chemistry: Conservation of Mass

Introduction

This presentation will cover the first three lessons of a Senior 3 Chemistry course. In preparing these lessons I first conducted a careful reading of the entire Chemistry curriculum in both Senior 1 and Senior 2. For the purpose of this assignment I focussed more on the analytical and experimental aspects of the courses rather than the societal and environmental issues, but otherwise I hope I was fairly thorough. On the basis of this review I am able to make the following observations.

In Senior 1, the student becomes familiar with the periodic table. He learns the names of some common compounds and their chemical formulas. He interprets the formulas by counting atoms; *but he does not deal with the concept of bonding*. Even though he classifies atoms with similar properties according to their place in the periodic table, and draws Bohr diagrams of the orbitals, he does not relate any of this to the theory of bonding.

The student learns in general terms about chemical and physical changes; he can identify combustion of wood as a chemical change and dissolution of sugar as a physical change; he is introduced to “exchange” reactions involving compounds that he will later learn (in Senior 2) to be “acids” and “bases”. In the lab, he observes the famous vinegar-and-baking-soda demonstration. But all this is done on a descriptive level. *The student does not write out and track the mechanics of specific chemical reactions.*

This background sets the stage for Senior 2. It is clear what the starting point ought to be; in fact, we will find that we revisit most of the topics in Senior 1, but we do it at a more sophisticated

level. The only question is where do we draw the line separating Senior 2 from Senior 3, and the Frameworks are quite specific on this, as we shall see when we go through the Specific Learning Objectives.

In Senior 2, we replace the descriptive Bohr orbitals with the Lewis diagrams (S2-2-01) which allows us to begin using the concept of valence to analyze chemical bonds. We generally limit ourselves to binary compounds, avoiding the polyatomic, but the analytical principal is otherwise fully developed. The classification of physical processes is sharpened from the previous year's "physical vs. chemical" to distinguish up to five types of chemical changes: synthesis, decomposition, single and double displacement, and combustion (S2-2-07). Note that the concepts of acid-base and oxidation-reduction reactions fall within the framework of these categories, but only in a descriptive way. The point is just that there is a further level of sophistication (role of electron transfer, concentration of OH ions, etc.) in this matter which is in fact deferred until Senior 3 or even Senior 4.

Most importantly, unlike Senior 1, in Senior 2 chemical equations are written out in full and balanced (S2-2-06). What is studiously avoided, however, is the quantified analysis of these equations: they are referred to obliquely as "skeleton" equations, and presumably it is to be hoped that the student does not ask what quantities are represented by the terms in the equation; because the concept of "moles" is, rather awkwardly in my opinion, not included as part of the Senior 2 curriculum.

We see most clearly the studious avoidance of this type of quantification in the very first lab experiment of the course, the Conservation of Mass (S2-2-05). The Lab Guide talks about dissolving an antacid pill in water, but doesn't say what the chemical formula of an antacid pill is. It talks about gas escaping, but it doesn't name the gas or describe the reaction responsible for producing it. There is indeed a measurement, a "massing" of the reactants and products as we are

taught to call it, but we don't go so far as to mass the reactants separately. (More on this point later!) All things considered, in this one experiment, we find a succinct picture of the philosophical differences between Senior 2 and Senior 3 Chemistry.

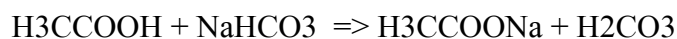
I find it most fitting and appropriate, therefore, that the best way to kick off the year in Senior 3 chemistry would be to have a learning cycle whose purpose is to re-do the Conservation of Mass Experiment of Senior 2, but from the more sophisticated, quantified and analytical perspective of Senior 3. And to that end, what could be a more natural starting point than the question of: "what is a mole."

I had originally divided my proposed lesson cycle into three lessons. The first lesson dealt with the concept of "moles". The second lesson was to take this concept and apply it to some reactions already familiar to the student from Senior 2: using the concept of moles, we can now take our stoichiometric equations (formerly "skeleton equations") and translate them into actual quantities of reactants and products, measurable in grams. In the third lesson, we would go to the lab and apply this knowledge to the familiar "volcano" experiment of baking soda and vinegar to see if we can now verify the conservation of mass more accurately than we were able to do with our limited knowledge in Senior 2.

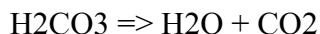
Since the material to be covered in the first two lessons of the cycle needs to be sufficient to do a complete analysis of the experiment, which is the ultimate goal of the cycle, I decided it would be natural to start at the end and work backwards, analyzing the experiment to see what knowledge will be needed. Then we can make sure we cover that knowledge in the first two lessons.

I set out to do just that, breaking down the lab into pieces and figuring out what I would need to teach the students theoretically. I soon found that my plan was flawed. My mistake was to

assume that the familiar volcano reaction was a suitable topic to include in Lesson 2 of the cycle, along with such old standards as $\text{HCl} + \text{NaOH} \Rightarrow \text{H}_2\text{O} + \text{NaCl}$. That reaction should be familiar from Senior 2, and in the idea would have been to use it, and similar reactions, to show the students how to convert molar equations into gram quantities. What I found was that the volcano reaction was just a bit too far outside the paradigm to be suitable for inclusion in a first lesson on this topic. The problem is the carbonate ion. It seems that what chemists do is tantamount to inventing a fictional compound called “carbonic acid” which occupies a place in chemical equations analogous to other familiar acids, with the small caveat that this “carbonic acid” doesn’t really exist in a stable form; that after writing the equation, you have to write a second line illustrating the breakdown of carbonic acid into CO_2 and water. In the case of the volcano reaction, it looks like this:



This in itself would be OK because although it is a bit longish we have still really just swapped an Na for an H. The problem is that the reaction isn’t complete until we add one more line:



We really must analyze this as a two-stage reaction; even with that, the reaction as written is a bit of an oversimplification, ignoring the role of the CO_3^{2-} and HCO_3^- ions; and for those reasons I couldn’t justify including it in a first lesson on stoichiometry. What to do?

The problem was that under the terms of the assignment, we are required to include at least one lab activity; and after two lessons we are simply not in a position to do the lab in the way I want to do it. I therefore made the hard decision that I would have to add an extra lesson to the cycle, in order to develop the theory to a sufficient level to justify the lab. (It is true that we were not strictly required to begin our lesson cycle with the first class of the term, but for reasons that

need not concern us I had a personal preference for doing so in this case.) The extra lesson is devoted almost entirely to reactions involving the carbonate ion. From the theoretical perspective it is perhaps a little out of the natural sequence, but I think it is justified by the practical necessity of getting to the lab activity. With that issue settled, we proceed to consideration of the lab.

I recently had the opportunity to assist one of the students in my practicum in writing up her results for the antacid pill experiment. (I was not present for the experiment itself.) She had recorded the mass of the reactants as something on the order of 180 gm. There were two trials: in the first trial, without a balloon, she lost 0.6 gm; in the second trial, capturing the escaping gas in a balloon, she lost 0.4 gm.

It would have been interesting to have recorded the mass of the pill, so we might have known how much of that was converted to gas. In Senior 3, with our new knowledge of stoichiometry, this is a wonderful example to compare the chemical equation to the measured quantity of gas production. Since I do not know the chemical formula of pepto-bismol, I will be substituting baking soda and vinegar for the reactants in my version, so that an actual equation can be written. Nevertheless I would think we are fairly safe in assuming that the unnamed gas in the Senior 2 experiment was indeed carbon dioxide, which brings us back to the question: what quantity of the antacid pill was converted into gas?

The Senior 2 experiment pointedly avoids this question by instructing the students to measure only the *combined* mass of the reactants and products. The “mass defect” in the experiment is therefore less than 1% of the total mass involved. Why did the designers of the lab seemingly go out of their way to obscure this most interesting question: how much of the *pill* was lost? Could it be that they didn’t *want* the students to ask this question?

That still leaves us in a bit of a pickle: how indeed to determine the mass of the CO₂ gas released in the experiment. In the Senior 2 experiment we trap it in a balloon and weigh it; but this is notoriously inaccurate. The Guide suggests that the weight of the gas is very small and hence hard to measure; it also suggests that some of the gas leaks out of the balloon. In the end, however, it says the students ought to conclude that mass is conserved.

I can hardly make a convincing case for this conclusion based on the data I saw last week. The mass defect with the balloon in place was only about 30% smaller than the mass defect when the gas was allowed to escape freely. It seems that the writers of the Guide overlooked the obvious explanation for this discrepancy: you can no more properly measure the weight of gas in a balloon than you can measure yourself on a scale if you are standing at the bottom of a lake. (Yes, you are indeed by definition measuring the *weight*; it's the *mass* that you're getting wrong!) In fact, if we take into account the greater density of CO₂ as compared to air, the aforementioned difference of 30% even starts to make sense. This is the direction I would like to go with the analysis of the experiment.

I want the students to take into account the buoyancy effect in their analysis. They will need to measure not just the *mass* of the gas produced, but also its *volume*. To calculate the buoyancy effect they will need to know the density of air and the density of CO₂. To make sense of these in a coherent way they will need to know, if not the full-blown Ideal Gas Law, at least the fact that 1 mole of an ideal gas occupies 22 litres at STP. They will need to write the stoichiometric equation for the reaction and interpret it in terms of masses of reactants and products. These are the skills they will need to do the lab, and that pretty much defines the scope of what we need to cover in the first three lessons of the cycle. Of course we can't just slam those ideas together one after another: each one of them has to be presented in its proper context. We will see the details of how this can be done when we get to the individual lesson plans. In the meantime I have gone

through the Frameworks and highlighted the SLO's that will be touched upon to a greater or lesser degree at some point in the cycle (see Appendix), and there are no less than nineteen of them. This doesn't mean I think I can fully cover 19 SLO's in four lessons: it just means that to some extent or another, this many topics will come into play in the course of the lesson cycle.

As part of this assignment, I am asked to give a constructivist overview of my lesson plan. I have in general a hard time with labels and definitions, but until this week I had little doubt that I was a staunch constructivist when it came to my teaching approach. I am generally Socratic in terms of prompting the students for the next step as much as possible, and I always try to find ways to incorporate new concepts within the framework of existing knowledge. I don't know how well these traits will be illustrated by a cold reading of my lesson plans, but I know how I teach, and this is how I teach. I had the definite impression that my style fell within the definition of constructivism; that is, I thought so until this week. I was in the library perusing a copy of Stinner's essay "How Students Understand Science" and I came across a passage that rang alarm bells in my head. Stinner is contrasting Piagetian theory with Constructivism and he says:

"Constructivists attribute poor comprehension of science conceptions to the presence of blocking misconceptions or ACF's acquired informally from previous experience. The constructivist approach to remediation would be to provide experiences that would facilitate reconstruction of the learner's reality with respect to the relevant science phenomena."

This sounded an awful lot like Arons, whom I had been required to read in my CIA Physics course. I thought I'd better verify this so I pulled a copy and this is what I found:

"Learner's difficulties in encompassing the law of inertia and the concept of force stem in large measure from the wealth of common sense preconceptions and experiential "rules" that most of us assimilate to our view of the behavior of massive bodies before we are introduced to Newtonian physics."

I'd say that was pretty much the way Stinner defined Constructivism. Now, I'm not very good at dealing in generalities, so maybe I'm misinterpreting the definitions; but I also did a pretty thorough reading of Arons in my other course, and I can say with no uncertainty that I did not like Arons' approach at all, not one little bit. So if being a Constructivist means I'm a follower of Arons, then I'm 100% sure that I'm not a Constructivist. (If you want to know why I don't like Arons, you might ask Professor Metz.)

Like I say, I thought I knew what Constructivism meant; and it seems that you, Professor Bush, also thought I knew what it meant, because you gave me full marks on the recent term test for my explanation of Constructivism. But if this puts me in cahoots with Arons, I want out!

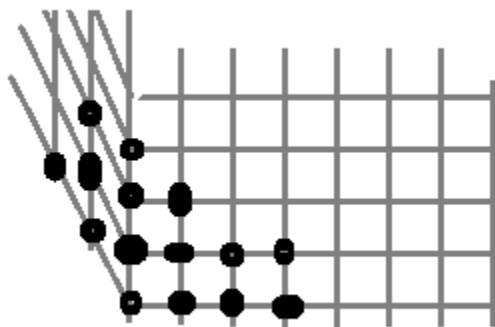
I am now ready to begin my lesson cycle. The topics do not follow in quite the natural order one might otherwise choose to teach them because I am working to the constraint of getting to a lab as soon as possible; but I think the sequencing is not too bad for all that. I have not specified a time for the lesson cycle, but it is probably closer to six hours than four; you will see that the lab itself must clearly span two class periods Either way, I don't believe in working to the clock. However long the lesson takes, that's how long it takes. I certainly don't believe in telling a student it would be a "waste of time" to deal with his questions. I have never believed in the idea of a one-to-one correspondence between lesson plans and SLO's (although I must admit that assignments like this one have gone a long way toward convincing me of the importance of knowing and tracking the SLO's). My idea is to go flat out and just teach *chemistry*; once I get near the end of the unit, I can always go back and see if there are any SLO's that haven't yet been fully dealt with, and take them up at that time. The lesson plans themselves will be rather brief, as I have covered the general direction pretty thoroughly in my introduction. Without further ado, let us procede.

Lesson 1: What is a Mole?

1. **Activation:** Teacher walks to the board and starts writing numbers: 6, 0, 2, 0, 0, 0 ...and keeps going for a long time. To class: “What have I just written?”

2. **Discussion:** How many atoms does this number represent? How much space would this many atoms occupy? This is a question we can deal with quite nicely.

3. **Analysis:** Draw a row of 10 dots, representing atoms. How many atoms would fill a cube this long?



Answer: 1000. Students must understand why. Then they can answer the question: what about a row of 100 atoms? How many atoms would fill a cube of that length? Answer: 1000000. Let the students make the generalization that for a row of 10^n atoms, the corresponding cube contains 10^{3n} atoms. (The generalization would be in the form of words, not a formula.)

Discussion: Point to the number on the board and ask, “what size of cube would contain *this* many atoms? Count the zeros and converted it to scientific notation. Let them struggle (not for too long!) with this question and then show them this trick: round the number up to 10^{24} . The answer should be immediately obvious: there are 10^8 atoms on the edge of this cube. How big is this cube? (Hint: the size of an atom is on the order of 10^{-8} cm).

Can anyone guess the significance of the number 6.02×10^{23} ? Hint: “the number of atoms in a mole” is *not* the “significance” of the number. Students may guess that it is the number of atoms in one cubic centimeter. Good guess, but wrong: it is the number of atoms such that the weight in grams is equal to the atomic mass: e.g. so many hydrogen atoms will weigh 1 gm and the atomic mass of hydrogen is 1; so many oxygen atoms will weigh 8 gm and the atomic mass of oxygen is 8. Do enough examples to clarify the working meaning of the definition.

What is the connection to the 1-cm cube we calculated previously? It is a natural fact that many real substances have a density on the general order of 1 gm/cm^3 , and now we have a way to explain this in terms of atoms. For the cube we calculated, the mass would be 1 if it was all hydrogen, 4 if it were all helium, etc; and for H_2O , which is mostly hydrogen, the very rough ballpark calculation we did happens (coincidentally) to give the *exact* value of the density.

SEGUE: Allow a free-ranging discussion for as long as possible, and then guide the class towards the question of the density of gasses. Ask the question: Does anybody know the density of air?

ANALYSIS: The density of air is 1.3 kg/m^3 . Round this off to 1 and compare to the density of water: students should see it is 1000 x less. Question: how far apart on average are the air molecules, compared to the water molecules? Let students understand that the factor of 1000 on density corresponds to a factor of 10 on linear separation. So the atoms are 10 times farther apart.

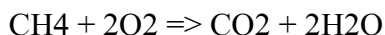
PROBLEM: If the average molecular weight of air is 29, and the density of air is 1.3 kg/m^3 , then what is the volume of air which weighs 29 gm? Students should be able to calculate that the volume is 22.4 litres. Go around the classroom and make sure they all know how to do this calculation; have students help each other until everyone can do it.

WRAP-UP: Look up the standard densities of different gasses, and have the students repeat the calculation for, e.g, a helium balloon: what size of helium balloon will weigh exactly 4 gm ($\text{He} = 4 \text{ AMU}$). Answer: 22.4 litres in all cases! *The same number of molecules occupies the same volume.* This is the mysterious generalization which we will return to later in the year when we do ideal gasses. For now, it lets us proceed to the next lesson.

Lesson 2: Stoichiometry

Sharing Objectives: In this lesson we will practise what we learned previously about the molar relations and apply it to balanced chemical equations. We will look at an equation and figure out how many grams of reactants and products are needed to balance the equation as written.

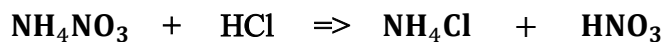
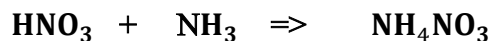
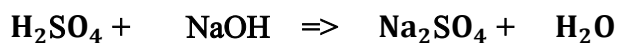
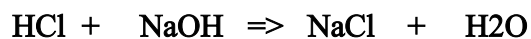
Discussion: Start by reviewing what we did the previous class and ask the students to come up with examples of chemical equations we can apply our knowledge to. They should come up with examples they have already seen in Senior 2, like:



Ask the students to come up with a suitable *question* to go along with this example. The goal is that they should come up with things like: how many grams of oxygen do you need to combust one gram of methane? Or, “how many pounds of CO_2 is produced when you burn one pound of methane (natural gas)? On a more sophisticated level, students might ask questions like: how many cubic meters of oxygen are needed to combust one cubic meter of methane? (Easy, but sophisticated!) Or to take it one step further: burning one cubic meter of natural gas (the metered quantity) produces how many pounds of CO_2 (the greenhouse gas)? How many cubic meters of methane does a household consume in one winter, and how many pounds of greenhouse gas are produced? Lead the students through calculations to answer these questions.

The teacher needs to turn the discussion at some point to reactions involving acids and bases, because we want to start introducing the ionic species which will figure in the lab activity of Lesson 4. For this purpose, we will do examples on the board involving things like NaOH, HCl, H₂SO₄, and especially acetic acid H₃CCOOH. (This one should be drawn on the board to show the bond layout.) We can hold back on the carbonate ion for now because it will get a whole lesson to itself next day.

WORKSHEET EXERCISE: Consider the following four chemical reactions:



1. If the equation is not balanced, balance it. Then calculate the number of grams corresponding to each of the molar quantities.
2. How many grams of lye (NaOH) are needed to react with 25 gm of hydrochloric acid?

3. How many grams of sulfuric acid do you need to make 100 gm of sodium sulfate?

4. BONUS QUESTION: If 1 lb = 454 gm, how many pounds of sulfuric acid do you need to make 100 pounds of sodium sulfate?

Lesson 3: The Carbonate Ion and its Friends

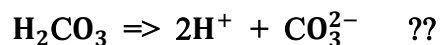
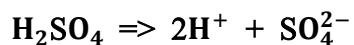
Activation: Do you remember the Seinfeld episode when Jerry told Kramer he'd have to start paying for the groceries he was mooching, and at the end of the week Kramer paid him for a bunch of stuff including "half a can of pop"? What did Jerry say?

Answer: "You know that sound a can of pop makes when you pull that little tab? That's the sound of you paying for a WHOLE can of pop."

Discussion: Why do they use CO₂ to put the fizz in soft drinks? Why not just compressed air?

Answer: Because CO₂ is far more soluble in water than ordinary gasses. The reason involves chemistry. CO₂ behaves like an acid in water, and there is something we call "carbonic acid" which is the result of combining CO₂ and H₂O to get H₂CO₃. We can analyze carbonic acid like other well known acids such as sulfuric acid, which has a similar formula. There is only one little thing to remember: H₂CO₃ doesn't exactly exist in normal conditions in any recognizable form. But it's a useful idea, so let's see how it analyzes.

Analysis: Using Sulfuric Acid as our "template", let's see what we should expect from carbonic acid. First let's just try dissolving it in water:



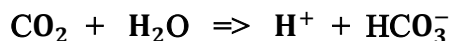
Is this how it works? Not exactly. In sulfuric acid, the two hydrogen ions come right off and go into solution. Carbonic acid doesn't work quite the same way. The reaction as written above does indeed proceed to a very slight extent, but to a much greater degree we get:



Even this doesn't give the whole story. Unlike sulfuric acid, carbonic acid is a "weak acid" which means it doesn't completely dissociate in water. So whereas in sulfuric acid, only the hydrogen ions and sulfate ions are present in solutions, in carbonic acid, all three species coexist in the solution: hydrogen carbonate, the hydrogen ion, and the carbonate ion. Oh...did I forget a species? Yes I did; the *bicarbonate* ion (HCO_3^-) is also there.

If that's not complicated enough, we are told (don't ask by whom) that the carbonic acid molecule (which we've never actually seen on its own) actually decomposes (partially!) to its constituent parts of water and CO_2 ; and that the dissolved CO_2 is a significant, if not the largest, component of the mixture. So if we include dissolved gaseous CO_2 , there are *five* species present in carbonic acid. (Six if you count good old H_2O .)

Don't be too intimidated. We can (and will) still write the simplified form of the equation like so:



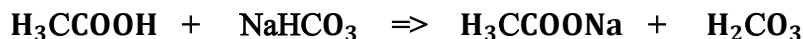
because that equation tells us most of the useful stuff that's going on. But we shouldn't forget that the reality may be quite a bit more complicated than that.

Discussion: How much CO_2 can you dissolve in water? Get data on typical concentrations for carbonated beverages and breakdown into species. How does this depend on temperature and

pressure? Wikipedia gives the equilibrium concentration of CO₂ in water as 1.45 gm/litre at 100kPa (one atmosphere). Is that 1 atmosphere of total air pressure or is that 1 atmosphere of pure CO₂ pressure? Does it make a difference?

Carbonic acid comes in the form of a gas, CO₂; but there is also a base that comes in the form of a gas: Can anyone name it? Is the chemistry of ammonia complicated in the same way as the chemistry of carbonic acid?

More Analysis: We are taking up the question of the carbonate and related ions because we want to do the volcano experiment next class. So let's look at the actual volcano reaction of vinegar and baking soda and see what we can make of it. The reaction proceeds in two stages; first:



And then,
$$\text{H}_2\text{CO}_3 \Rightarrow \text{H}_2\text{O} + \text{CO}_2$$

The difficult question is: How much of the carbonic acid in Stage 2 actually ends up as CO₂ gas? How do we know some of it doesn't remain in solution in the form of carbonate ions, bicarbonate ions, and dissolved CO₂ gas?

The answer to this question is beyond the ordinary scope of Senior 3 Chemistry, but it ought to be noted here for the sake of completeness. Essentially, when we dissolved baking soda in water, we put carbonate ions into solution in much greater concentrations than we could normally achieve by cramming them into solution in the form of a dissolved gas. You can see this is true by comparing the solubilities of CO₂ and sodium bicarbonate. This is because when we dissolve CO₂ in water, every carbonate ion is accompanied by a hydrogen ion; and these two ions have a limited capacity to coexist with each other. The carbonate ion is much happier to coexist with the sodium ion, so the baking soda dissolves much more readily. What we are doing when we

introduce the vinegar is we are bringing in a large quantity of “hostile” hydrogen ions, which the carbonate is incompatible with. Therefore, the carbonate is almost completely driven out of solution in the form of CO₂ gas. (This can be illustrated by pouring some vinegar in a freshly opened can of coke.)

We could have just written the simple chemical equation which gives the same impression and left it at that, but now you know.

Lesson 4: Conservation of Mass (Lab Project)

Sharing Objectives: We are going to try and verify the conservation of mass as seen in the well-known volcano demonstration. A similar experiment was done in Senior 2 but it was lacking in certain quantitative details, and did not properly take into account the buoyancy effect when measuring the mass of the reaction products. We are going to collectively design an experiment to correct these deficiencies. We will spend one class designing the experiment, and another class executing it.

The group will divide itself into teams, each with a definite role in making the experiment work. The total project team will consist of ten students, or five teams of two. If some students want to do their own version of the experiment as a smaller independent team, they may also do so.

These are the project teams, which may be up to two students each:

- 1. Volume Measuring Group:** This team is responsible for devising a system to measure the volume of the escaping gas. The rest of the experiment will be designed around whatever system they come up with.

2. Analytical Group: Based on the system devised by the Volume Group, the Analytical team will calculate the suitable quantities of reactants to generate an appropriate volume of gas for the experiment.

3. Containment Group: This team is responsible for devising a suitable containment system so that the gas does not leak out before it is measured.

4. Fabrication Group: This team is responsible for fabricating and putting into practise whatever systems are designed by the Volume and Containment Groups.

5. Material Handling Group: This group is responsible for measuring and preparing the correct quantities of reactants for the experiment.

PLANNING STAGE: Review with the class the flawed experiment of Senior 1, emphasizing the need to eliminate two sources of error: the possibility of gas escaping from the containment, and the effect of buoyancy on tracking the mass changes. These are the responsibilities of Groups 1 and 3, who will need to collaborate with each other. Group 2 will need to work out the theoretical background regardless of the final configuration, so they will be occupied full time, needing only to plug in the final numbers once the experiment is designed. The work of the other groups will depend on the results of groups 1 and 3, so during class time they can also work on coming up with their solutions to these problems.

Presentations: As groups complete their tasks, they will present their results to the larger group for discussion and review. At the end of the class a list of materials will be made, and the instructor will ensure that these are on hand for the next day.

THE EXPERIMENT: The experiment will be carried out the following class. All students will record the data and write up an analysis showing whether the measurements are consistent with the balanced chemical equation.

This completes the lesson cycle.