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SPECIAL SECTION: ***Structural concepts***
Contribution from science education and history of science
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HELPING STUDENTS TO MAKE INFERENCES ABOUT THE ATOMIC REALM BY DELAYING THE PRESENTATION OF ATOMIC STRUCTURE

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ABSTRACT: The presentation in many introductory college chemistry courses jumps around in an unsystematic fashion among the domains of macroscopic observation, sub-microscopic particles, and symbolic representation, making it difficult for students to construct a coherent picture of the discipline. Such presentations usually pay little attention to the intellectual struggle that led to fundamental knowledge in chemistry. Topics like atomic and molecular structure are often introduced early in the course, even though they are far removed from direct experience, and historically were elucidated relatively late in the history of chemistry. This paper describes a freshman chemistry curriculum, in which the topic of atomic structure is delayed until the second semester. Concept development is linked to the observable behavior of matter, while the sub-microscopic and symbolic realms are introduced by engaging students in some of the detective work that established the relative atomic masses of the elements and formulas of simple compounds. In this way, students have an opportunity to become familiar with the relationships among facts, definitions, hypotheses, deductions, and predictions, which are central to the enterprise of science. A brief review of literature dealing with relevant history and educational research is included. [*Chem. Educ. Res. Pract. Eur.*: 2001, *2*, 183-202]

KEY WORDS: *Constructing concepts; element, compound & mixture; relative atomic mass; compound formulas; Avogadro's law; kinetic theory; active learning; historical approaches*

LESSONS FROM THE DISTANT PAST AND THE RECENT PAST

History

Practicing chemists, chemical educators, and chemistry students rely upon the data provided in the periodic table as they go about their business, but they may not be aware of the difficulties that were associated with acquiring this knowledge. In a recent article, Jensen (1998b) describes the revolution that replaced dualistic classifications of the elements with valence classifications, culminating in the discovery of the periodic law. This revolution was based in part on the availability of a correct set of relative atomic masses for the elements.

The struggle to determine a self-consistent set of atomic masses began in the early 1800's with the competing hypotheses of Dalton and Avogadro. Dalton maintained that the simplest compound formed by two elements consists of diatomic molecules, while Avogadro maintained that the composition of compounds must be consistent with the idea that equal volumes of gases at the same pressure and temperature possess equal numbers of particles.

The hypotheses lead to different sets of relative atomic masses, and compound formulas. Dalton's assertion, called the rule of simplicity, was eventually proven wrong, while Avogadro's assertion is now known as Avogadro's law. Avogadro's law is unlike the other gas laws, because it cannot be demonstrated directly by experiment. Its validity was inferred based on its ability to allow a body of experimental data to be interpreted in a self-consistent fashion.

Several authors have written about the problems associated with Dalton's rule of simplicity, and the assumptions associated with Avogadro's hypothesis (Bertanowicz, 1970; Brock, 1963; Feifer, 1966). Dalton's rule does not predict a unique set of formulas or atomic masses when there is more than one compound containing the same pair of elements. This was not recognized by some of Dalton's contemporaries, and Dalton himself apparently ignored this at times in his own work. Similarly, Avogadro's hypothesis does not predict a unique set of formulas for reactants and products in a given process, and the leading chemists of the day were not comfortable with the idea that elements could be composed of diatomic (or polyatomic) molecules, which the hypothesis required. Therefore, it is not surprising to learn that confusion reigned in the effort to obtain a correct set of relative atomic masses and compound formulas until the mid-1800's.

In 1858 Cannizzaro used Avogadro's hypothesis to predict that elemental gases were diatomic, and then used gas density measurements and gravimetric compositional data to determine a self-consistent set of relative atomic masses of some of the elements, and formulas of some compounds (Jensen, 1998b). Cannizzaro drew upon the work of Clausius as confirming Avogadro's hypothesis, raising it to the status of a law (Lund, 1968; Whitaker, 1979; Mendoza, 1990). A year earlier, Clausius argued that the common gases like hydrogen and oxygen were composed of diatomic molecules, when he put forth the kinetic theory of gases.

This theory proposed that what we call heat (thermal energy in modern terms) was due to the motion of submicroscopic particles. Predictions of kinetic theory, when compared to the observable gas laws, suggest that the quantity mv^2 is proportional to absolute temperature of the gas, where m is the mass of the individual gas particles, and v^2 is the average square of particle velocity. When the predictions of kinetic theory are combined with the idea of kinetic energy, one finds that a constant heat capacity for different gases is expected (Lund, 1968; Fitzgerel, 1960). Clausius determined that a gas composed of atoms undergoing only translational motion would have a heat capacity that was lower than those observed for the known elemental gases. In order to be consistent with existing heat capacity data, the known gases would have to be composed of diatomic molecules that would require energy to undergo other forms of motion, like rotation, in addition to translation. The idea of diatomic molecules was also consistent with Gay-Lussac's law of combining volumes, and gas density data. Thus Avogadro's hypothesis finally became Avogadro's law, after almost half a century. Although the confluence of these ideas was critical to the development of the periodic table and further work in structure determination, the complete acceptance of the particulate nature of matter did not come until the early 1900's (Jensen, 1998b).

Educational research

As one develops an appreciation for the struggles that scientists faced in piecing together the knowledge about atomic masses and molecular structure that today we take for granted, it becomes less surprising to find that today's chemistry students struggle when they are asked to draw inferences about the atomic realm based on macroscopic observations. Nakhleh (1992) summarizes a large body of evidence indicating that students of all ages seem to have trouble understanding and using the scientifically accepted model that matter is made of discrete particles that are in constant motion. She points to studies indicating that students do not have a clear picture of concepts like element, compound, mixture, or solution. Others indicate that students can solve traditional gas law problems, but may not understand the behavior of gases in terms of the underlying behavior of molecules. In addition, she notes that students may be able to balance chemical equations without being able to draw correct molecular diagrams associated with the equations. Nakhleh stresses that students need help in understanding the difference between atoms, molecules, and ions, and in visualizing chemical events as dynamic interactions, in order to avoid developing misconceptions that can hinder subsequent learning.

Tsaparlis (1997a) summarizes why students may have such difficulty with atomic and molecular structure in terms of different perspectives in educational research. In terms of Piagetian levels of development, the concepts of atomic and molecular structure require formal operational reasoning, since they cannot be learned about through direct concrete experience. Drawing inferences about the invisible world of atoms and molecules is difficult for students who are not at this level of reasoning. In terms of Ausubel's theory of meaningful learning, learners link new information to what they already know. Since structural concepts related to the atomic world must be built on new ground, it is difficult for students to learn about such concepts. In terms of information processing theory, students may be confounded and overwhelmed by presentations in which the instructor jumps back and forth among observations in the macroscopic world, the behavior of particles in the atomic world, and the symbolic representations used to link the two.

Various reports lend support to these ideas. Hesse & Anderson (1992) and Johnstone (2000, 1991a) indicate that students have great problems moving among the domains of macroscopic observation, atomic description, and symbolic representation. Other researchers (Haidar & Abraham, 1991; Abraham, et. al., 1994) have found that formal reasoning ability and preexisting knowledge play a role in development of students' conceptions, and their use of particulate theory. Students are more likely to use particulate theory when cued to do so. Lawson (1991 & 1998) has found that reasoning ability limits achievement in science courses more than prior knowledge. Finally, a study performed by Johnstone (1991b) indicates that only 25% of secondary school biology students can formulate hypotheses from observations, and that they must be taught to differentiate relevant from irrelevant information when evaluating a situation.

WHY DELAY THE TOPIC OF ATOMIC STRUCTURE?

The points raised in the previous section can be used to support several arguments for delaying the presentation of atomic structure in introductory science courses:

1) Chemistry entails linking observations in the see-touch world with inferences about the sub-microscopic world, and with symbolic representations. Since observations and inferences precede models and theories, an approach that delays the details of atomic models and theories is more consistent with the way in which science actually proceeds.
2) Knowledge of the behavior of substances, relative atomic masses, chemical formulas, and molecular geometry all preceded knowledge of atomic structure. Presenting topics roughly in the order in which they were historically elucidated can provide greater insight into how the science of chemistry developed, and greater appreciation of some of the uncertainties faced by the scientists who brought this knowledge to light.

3) Research indicates that students are easily overwhelmed by too much information, and that it takes time to develop reasoning skills. Therefore, an approach that delays topics like atomic structure can provide students with a better opportunity to develop the skills that are at the heart of connecting the see-touch world with the sub-microscopic world.

FRESHMAN CHEMISTRY AT THE MASSACHUSETTS COLLEGE OF PHARMACY AND HEALTH SCIENCES

Freshman chemistry at the Massachusetts College of Pharmacy and Health Sciences (MCPHS) comprises a sequence of two four-credit courses with a large classroom environment (100-130 students). Students attend four, 50-minute classroom periods, including an integrated prelab period, and one three-hour laboratory session each week. The instructor and one or two learning facilitators attend each class. Sections of approximately 40 students interact with the instructor and two or three assistants during laboratory sessions, which are coordinated with classroom activities as much as possible. Student majors include pharmacy, chemistry, pre-med, and various allied health programs. Almost all students have had high school chemistry and biology, and many have had high school physics. In recent years, average combined Scholastic Apptitude Test (SAT) scores have been in the vicinity of 1050, but based on a diagnostic examination, the class shows a wide range in their level of preparation.

The two-semester sequence is traditional in the sense that it comprises a survey of important topics. However, topic development is based on introducing experimental evidence before concepts and theories, and some attempt is made to follow the historical development of ideas. The corresponding author has been working in the area of curriculum development at MCPHS since the mid 1980's (Garafalo & LoPresti, 1986; Garafalo, LoPresti, & LaSala, 1988; Garafalo & LoPresti, 1993; Cohen, *et al*., 2000), and has been teaching freshman chemistry primarily from instructor-generated materials since academic year 1990-91. Efforts throughout the 1990's have centered on using constructivist learning theory (Ausubel, 1978; Resnick, 1983; Arons, 1990) to guide curriculum development. Important issues include choice and sequence of topics (including time spent on each topic and which topics to omit), nurturing an active learning environment, coordinating lab and classroom activites, and coupling content to the development of specific reasoning skills. These skills include distinguishing between observation and inference, hypothetico-deductive reasoning, and proportional reasoning (Arons, 1990). The work is based on an action research methodology (Nakhleh, 1996; Towns, 2000), which consists of planning and implementing specific classroom activities, observing and evaluating the results, and then using conclusions to revise the activities and perform another cycle of the process.

Instructor-generated handouts drive classroom discussion and think-aloud problem-solving sessions (Lochhead & Whimby, 1987), which provide students with rapid feedback. Students also experience active learning during the instructor's daily office hour, which is conducted in a classroom and usually attended by 10-15 students (Cohen *et al*., 2000). Presentations encourage students to link new information to what they already know, through Socratic lines of questioning. Progressively more complicated situations provide students with the opportunity to work through resolvable conflicts. Activities in the laboratory include observing properties of materials, interpreting experimental results in terms of a posed hypothesis, a known law, or a definition, and obtaining experimental data and using it to calculate quantitative properties of substances.

Table 1 indicates the current sequence of units in the first semester. The units are organized around themes that span and unify the sciences. Topics are presented in

TABLE 1: *Units comprising semester I of freshman chemistry.*

1. Mathematical foundations	6. Making inferences about the atomic realm
2. Introduction to measurement	7. Introduction to the periodic table
3. Observations about matter	8. The concept of energy
4. Ideas about motion	9. Gradients and equilibrium
5. The concept of force	10. Matter with a charge

approximate historical order within a given unit when possible (units 3, 6, 7, and 10), and the sequence of units is consistent with the historical development of concepts in chemistry. Units 3 and 6 are the focus of this report.

UNIT 3: OBSERVATIONS ABOUT MATTER

The topics comprising Unit 3 are listed in Table 2. Relevant sample questions for Units 3 and 6, cited in various places in the body of the paper, are found in Appendix I.

TABLE 2: *Topics comprising unit 3, Observations about matter.*

Properties of matter
Mixtures, compounds, and elements; identifying chemical processes
Laws of conservation of mass and constant composition
Evidence suggesting that matter is made of tiny particles
Law of multiple proportions / The atomic theory
Compounds vs mixtures: A closer look - Solutions and phase changes

Related Laboratories
Lab 3: Observations about matter

Total instructional hours: Five 50-minute classes; one three-hour lab period

The behavior of various samples of matter is first used to help students create a library of terms. Comparison of the properties of heterogeneous mixtures with those of compounds leads to the idea that compound formation involves substances combining in a special way. The term "chemical process" is introduced, but distinctions between chemical change and physical change are avoided (Gensler, 1970). Instead, focus is placed on identifying samples of matter as elements, compounds, or mixtures, based on their behavior (Appendix I, Q1).

The law of constant composition (definite proportions) is "discovered" by interpreting the results of chemical analysis data. Discussion is limited to mass ratios, avoiding atom and mole ratios, when considering limiting reagent problems (Cohen *et al*., 2000). Evidence is presented to support the atomic theory based on certain observations about gases, liquids, and solids. The laws of definite and multiple proportions are used to suggest that atoms may be bonding to one another to form molecules when a compound is formed. The definitions of element and compound are revisited in terms of the concept of atoms. All substances are referred to by name only. Identifying properties are stressed, and the hypothetical nature of any formulas that are introduced at this time is noted. An activity demonstrates that without more information, unique formulas for compounds are not possible (Appendix I, Q2, Q3). Finally, mixtures are revisited and phase changes are introduced. Changes in state of a sample of matter with heating suggest that phase changes involve different degrees of bonding, but because percent composition remains constant, such changes are not classified as chemical reactions. Based on their varying composition and their behavior during phase changes, solutions are identified as mixtures (Appendix I, Q4).

In the lab, several activities are used to reinforce important vocabulary. Students are shown a pair of processes, and asked to determine which is more likely a chemical reaction, based on observations like thermal energy or gas evolution. One activity entails observing a piece of zinc dropped into a solution of dilute hydrochloric acid, and another piece dropped into some water. Another activity entails mixing dilute acid with water vs. mixing dilute acid with dilute base. It is stressed that such limited investigations lead to conclusions that are suggestive, rather than definite.

Making hypotheses about observations associated with heating a metal in air encourages students to recognize the importance of making mass measurements. For example, students are presented with this sequence of activities:

> ***Activity 1:*** (Students are given the following to read.)
> When certain types of metallic substances are heated in an open container (called a crucible), they become a type of substance called **calx**. Calx is a powder, and does not exhibit metallic properties. Consider the following hypothesis to explain the experimental observations: Heating removes something from the metallic substance to make calx.
>
> *Answer the following questions:*
>
> - Is this a reasonable hypothesis?
> - How could you test the hypothesis?
> - Can you think of an alternative hypothesis that could also explain the observations?
> - If you have trouble, conduct the next activity first, then return to this one.
>
> ***Activity 2:*** You will now perform an experiment in which you heat a piece of metal in an open crucible, producing calx. Your instructor will tell you what size piece of metal to use. (Procedure and data sheets are omitted here.)
>
> *Answer the following questions:*
>
> - Did your sample of metal gain mass or lose mass?
> - What was the increase (or decrease) in mass of your sample in grams?
> - Does this result support the original hypothesis about what happens when calx is formed? If not, suggest a new hypothesis based on your experimental results.
> - Which is more likely an element, the piece of metal or the calx? Why?
> - Make the ratio: mass increase (or mass decrease) / original mass of metal. Check with other students who had different starting masses of metal. Did they get about the same **ratio** as you? Based on your findings, is calx more likely an element, a compound, or a mixture? Why?

A similar set of activities is used during this lab period to determine that an unknown liquid is most likely a mixture.

UNIT 6: MAKING INFERENCES ABOUT THE ATOMIC REALM

The topics comprising Unit 6 are listed in Table 3.

TABLE 3: *Topics comprising unit 6: Making inferences about the atomic realm.*

Measurement of pressure and temperature
Gas laws -The effect of temperature, volume, and mass of gas on gas pressure
Interpreting gas laws in terms of the motion of particles - (The kinetic theory)
Deducing relative atomic masses and compound formulas I - Dalton's rule of simplicity
Law of combining volumes
Deducing relative atomic masses and compound formulas II - Avogadro's hypothesis
Choosing the better hypothesis - help from kinetic theory / Avogadro's law
The concepts of mole, molar mass, and molar volume.
Deducing relative atomic masses and compound formulas III - Cannizzaro's method
Revisiting the gas laws - the ideal gas law

Related Laboratories
Lab 5: Observations and experiments regarding pressure
Lab 6: Determining relative atomic masses of two elements
Lab 7: Deductions using different hypotheses (dry lab)
Lab 8: Determining the molar mass of a componud by the Dumas method

Total instructional hours: Ten 50-minute classes; four, three-hour laboratory periods

Ideas about pressure

The concept of pressure is introduced immediately following short units that discuss basic ideas associated with motion and force. Emphasis is placed on distinguishing between pressure and force, differences associated with pressure exerted by solids, liquids, and gases, and using a column of liquid to measure gas pressure (Appendix I, Q5, Q6). These ideas are explored in the lab 5 using simple activities.

The gas laws are introduced next, with emphasis on the experimental relationship of pressure, P, to volume, V, temperature, T, and mass of gas, m: $P \propto mT / V$. Since for a given gas, greater mass suggests greater number of particles, N, the idea is introduced that

$$P \propto NT / V \tag{1}$$

Simple, but challenging activities are used to complement traditional gas law calculations in the lab. (Appendix I, Q7)

Using a model in which gas pressure arises from the collision of molecules with the container walls (the kinetic theory), the conclusion is reached that

$$P \propto Nmv^2 / V \tag{2}$$

Comparison of equations 1 and 2 leads to the idea that the measured property, temperature, may be related to the derived quantity, mv^2. Evidence from 20th-century experiments using rotating disks shows that absolute temperature is indeed proportional to average v^2 for a given gas. Careful use of proportions avoids the need for derivation of the more exact equation,

$p = (Nmv^2)/3V$, or for arbitrarily invoking other equations, when working with kinetic theory. Students are asked qualitative questions in which the behavior of a gas is interpreted in terms of the behavior of its moving particles (Appendix I, Q8).

Deducing atomic masses and molecular formulas

The focus of the presentation now returns to the early 1800's. The fixed, experimental hydrogen / oxygen mass ratio of 1/8 for water is presented, along with the hypothesis that water is composed of diatomic molecules with the formula, OH. The hypothesis is an example of Dalton's rule of simplicity. Using this information, students deduce that an oxygen atom should be eight times as massive as a hydrogen atom, and the concept of relative atomic mass units is introduced. Students then use the fixed mass ratio and various hypotheses to deduce other possible formulas for water, or other possible masses of the oxygen atom (Appendix I, Q9, Q10). Examples using data for other compounds are then introduced.

In the laboratory, students perform an experiment in which a solution of copper sulfate reacts with solid iron to produce a solution of iron sulfate and solid copper. They determine the copper / iron mass ratio and are then asked to determine possible relative atomic masses of the two elements based on possible reaction stoichiometries (Appendix I, Q11).

Law of combining volumes and Avogadro's hypothesis

Next, students are introduced to experimental observations about the volumes of gases that react with each other when the temperature and pressure have the same initial and final values. They are asked to use the fact that equal volumes of hydrogen and chlorine react completely at the same pressure and temperature, and the assumption that the reaction consists of hydrogen atoms combining with chlorine atoms, to deduce that equal volumes contain the same number of atoms. Next, students are asked to rationalize why one or the other of the reactant gases would be in excess when the temperature (or the pressure) of equal volumes of the two are different (Appendix I, Q12). Then the students are asked to consider other reactions like that of hydrogen with oxygen (where the combining volume ratio is 2/1). They are asked to assume that under these conditions, equal volumes of the reacting gases contain the same number of atoms. The idea that the product molecules would contain more than two atoms becomes evident, which reinforces the fact that Dalton's rule of simplicity is only a hypothesis.

The law of combining volumes is introduced next, and the fact that *two* volumes of the product gases are produced in various reactions. Students see that if Dalton's rule of simplicity is correct, then one must conclude that the number of particles per unit volume is not always equal for equal volumes of different gases under the same conditions of pressure and temperature.

Students are next introduced to Avogadro's hypothesis, which maintained that equal volumes of different gases do contain equal numbers of particles at the same temperature and pressure. At this point, students are asked to use Avogadro's hypothsis and the experimental facts about combining volumes to make various deductions (Appendix I, Q13). One sees that if diatomic element molecules are the reactants, the law of combining volumes and the observed density data can be made consistent with Avogadro's hypothesis. However, in this case Dalton's hypothesis cannot hold.

Dalton or Avogadro?

In the laboratory, students are given a work sheet and asked to perform several activities, analyzing data in an effort to determine which hypothesis is better. First they are asked to use Dalton's hypothesis and experimental mass ratios for water and nitrogen oxide to deduce the relative atomic masses of oxygen and nitrogen atoms (assuming that hydrogen is 1 amu.), and to predict the mass ratio in a new compound, ammonia. They are then asked to make various deductions using the law of combining volumes and Avogadro's hypothesis for the same compounds. One finds that the relative atomic masses of hydrogen, oxygen, and nitrogen atoms are in a ratio of 1/16/14 using Avogadro's hypothesis and certain assumptions, and 1/8/7 using Dalton's hypothesis. Also, a self-consistent set of reactant mass ratios is obtained using the relative atomic masses based on Avogadro's hypothesis, but not using those based on Dalton's.

Returning to the kinetic theory, relative velocity data for different gases at the same temperature are compared, and their relative particle masses predicted using the equation $m_2 / m_1 = v_1^2 / v_2^2$. Students see that calculated mass ratios for elements agree with those predicted using Avogadro's hypothesis, but not with those based on Dalton's hypothesis. Experimental velocity ratios also suggest that non-hydrogen reactant molecules are more massive than product molecules (e. g. using experimental velocity data, oxygen particles are found to be more massive than water molecules, and nitrogen particles are more massive than ammonia molecules), which is also consistent with predictions of Avogadro, but not Dalton. While detailed discussion of the validity of the kinetic theory is not presented, it is stressed that the theory had gained acceptance because of its ability to predict certain observed phenomena (such as heat capacities). At last we refer to Avogadro's hypothesis as Avogadro's law, and the ideal gas law is introduced. The concepts of mole, molar mass, and molar volume are introduced, but the actual number of particles in a mole is determined several weeks later in the term. Finally, students are led through Cannizzaro's method of determining relative atomic masses of other elements using Avogadro's law, vapor density data for a series of compounds containing a given element, and percent composition data. In the lab, students determine the molar mass of a volitile compound by the Dumas method.

Unit six has set the stage for the discussion of the periodic classification of the elements based on increasing atomic mass (the three known discrepencies are avoided until semester 2), and stoichiometry based on moles. In unit 7, the relative atomic masses that appear in the modern periodic table are introduced and used throughout the rest of the academic year. The classification of some of the representative elements based on observed mole ratios in their compounds is used to reinforce the newly introduced concept of the mole, before it is used more extensively. No mention of valence electron configurations in the representative element families is made until unit 11 in the second semester, where the shift to element classification according to increasing atomic number is also introduced.

OBSERVATIONS AND DISCUSSION

Unit 3 material

Initially, it was challenging for instructors to avoid using the common terms "physical change" and "chemical change," since many students bring these terms with them from high school. However, such distinctions are often made based on arguments that deny one's experience (e.g. saying that water and steam are really the same substance), or that rely upon statements that must be taken upon faith (e.g. saying that sugar does not change when it is dissolved in water), and usually require more discussion of bonding than is warranted at this time in the course (Gensler, 1970; Jensen, 1998a). At this point, it is sufficient to note that different degrees of bonding appear to be involved in phase changes and the process of dissolving, and to rely upon other factors to exclude these two processes from what are called chemical reactions. These include comparison of the behavior of the gases produced when water is boiled and when it is electrolyzed, and recognition of the nonconstant composition and boiling point of solutions. This approach is consistent with that suggested by Bar & Travis (1991).

Limiting quantitative treatments in Unit 3 to discussion of mass ratios is important because it reduces the number of ideas that students must consider. Many students are extremely challenged by questions that ask them to interpret and use ratios in a fashion that goes beyond plugging numbers into proportions or cancelling units (Arons, 1990; Cohen *et al*., 2000). Given the additional demands of mastering vocabulary and hypothetico-deductive reasoning that are also placed upon them at this time, the more limited presentation of basic chemical calculations is justified. The mole concept is introduced in Unit 6, and not used extensively until Unit 7.

The laboratory activities dealing with the conversion of metal to calx challenge students to observe properties of matter, apply definitions, interpret change in terms of a posed hypothesis, and create an alternative hypothesis. As instructors circulate among the students, it is not unusual to find many individuals who have difficulty deciding whether or not the initial hypothesis is reasonable, and who cannot come up with an alternative hypothesis in which the metal is combining with something rather than releasing something. Some coaching is often necessary to help these students realize that mass measurements can be used to test the hypothesis.

When the students actually perform the transformation, they are given a piece of magnesium to heat, but the identity of the metal and the fact that the term “calx” is an archaic name for a metal oxide are omitted. After completing the procedure, students often say that they have proven the initial hypothesis is correct, because they have created calx. The fact that this procedure alone does not confirm whether the metal has lost or gained something seems to escape them. These difficulties in generating and testing hypotheses are consistent with the findings of Lawson *et al*. (1991).

The fact that some students recognize that they have performed this very manipulation in high school (usually as a demonstration of compound formation) does not apparently give away all of the inferences that they are expected to make. Several of these students have sought clarification on the meaning of their results. In the several years that these activities have been presented, instructor-student interactions have revealed no students who, prior to conducting the transformation, explicitly identify the fact that the metal is combining with oxygen. The initial hypothesis suggesting that the metal losses something, and the use of the term “calx” are apparently enough to divert at least the vast majority of them from drawing upon knowledge they might have acquired in high school. In addition, many students struggle when asked to apply the appropriate terms like element, compound and mixture during these activities. This is in spite of the fact that most students appear in lab with a set of written definitions, including these terms, which they were required to complete prior to arrival.

Unit 6 material

A diagnostic exam presented in the orientation week, prior to the start of classes, confirms that our students are challenged when they must interpret the gas laws in terms of the behavior of moving particles. For this reason, as much emphasis is placed on qualitative interpretation of macroscopic observations in terms of the behavior of particles as on quantitative calculations. While students often invoke mathematical relationships among relevant concepts in their deliberations, the questions posed of them demand an understanding of the relationships that goes beyond plugging numbers into formulas. This approach is supported by the findings of (DeBerg, 1995) who indicates that tasks drawing upon qualitative as well as quantitative knowledge have the potential to reduce dependence on algorithmic approaches to learning, particularly equation substitution and solution.

The study of history also provides insights into difficulties encountered by students. Recently, one student was not sure why temperature change should be related to change in molecular velocity. She maintained that there did not seem to be anything intuitive about it. It turns out that this student is in good company. Prior to kinetic theory, temperature change was interpreted as a change in amount of a substance called caloric that was hypothesized to surround molecules. Avogadro assumed that the volume occupied by a gas molecule was determined by the quantity of caloric attached to it, and that this would be the same for all molecules at a given pressure and temperature, regardless of their chemical nature (Goldstein & Goldstein, 1993). Increase of gas pressure with increasing temperature was interpreted in terms of increased **static** repulsion of molecules as they acquired more caloric, and not in terms of collisions with container walls. Isaac Newton was a proponent of this view (Niaz & Rodriquez, 2001). It is important to recognize that the dependence of temperature on v^2 was initially an inference drawn by comparison of the empirical gas laws with predictions based on kinetic theory (equations 1 and 2 above). The dependence was only later verified by experiment. The word inference is key, since some developments incorrectly identify the temperature-velocity link as a fundamental postulate of kinetic theory (Carpenter, 1966; Rhodes, 1992).

The introduction to hypothetico-deductive reasoning in which students use known mass ratios of elements in compounds, and either hypothetical formulas to deduce relative atomic masses of elements, or hypothetical atomic masses to deduce compound formulas is extremely challenging for many students. Students frequently resort to changing the experimentally determined mass ratios, and often mix up amu's with grams. When presented with different hypotheses about water, they frequently invoke the known formula for water, even though the hypothesis in the problem may not lead to that deduction. Sometimes students invoke facts learned in high school about electron configurations and the octet rule in order to support the correct formula for water. To keep things clear, mass ratios of substances use written names of the elements, instead of symbols (e.g. oxygen / hydrogen, *vs*. O/H), and atom ratios include the word "atom" (e.g. O atom/H atom).

The lab in which students react iron with copper sulfate is very simple experimentally (Kiser, 1991), but posing a series of questions that requires students to deduce possible relative atomic masses, as opposed to telling them that they have determined the relative masses, makes it quite challenging. Getting them to draw pictures of atoms of different relative masses is often useful. Since the correct mass ratio is 1.1/1, students often mistake this as proof that the stoichiometry of the reaction is 1/1. These observations are consistent with those of Haidar & Abraham (1991), who maintain that hands-on experiences in chemistry are important but not sufficient for understanding concepts, and that students need instruction that will help them develop links between macroscopic observations in the laboratory and microscopic models that chemists use to explain them. Students are also challenged by the activities that center around the law of combining volumes and Avogadro's hypothesis to deduce formulas and relative atomic masses. The time spent comparing different hypotheses ensures that students have practice connecting a proposed rearrangement of particles with its appropriate symbolic representation in equations.

CURRICULUM EVOLUTION AND ASSESSMENT

The idea of a learning cycle, in which concept invention is preceded by student exploration of phenomena, and followed by application in new situations has been developed by Karplus (1977). Our classroom presentation relies upon a modified learning cycle, in which classroom discussion of an easily visualized situation (sometimes including a demonstration) occurs before the concept is introduced. This approach evolved when the curriculum was trimester based with no laboratory for the first ten weeks. The term "guided inquiry" is appropriate, because lines of questioning encourage students to reach certain conclusions.

The approach presented in the classroom is paralleled in the instructor-generated text. There, the description of phenomena is followed by feedback presented in a question and answer format. Students are encouraged to attempt answers before reading those provided. Applications are given in assigned homework problems. Answers are not provided until several days later, to encourage interaction with classmates, and attendance at help sessions.

The transition to a semester system in academic year 1996-97 has provided students with the opportunity to work more extensively with concepts introduced in Units 3 and 6 through laboratory-based learning activities. The activities in laboratories 3, 5, and 6 are based on what Domin (1999) describes as expository activities, but they have been modified. Simple activities that are usually used to "demonstrate" concepts have been turned into more challenging experiments by withholding some of the facts, and including a Socratic line of questioning, as described earlier. In this form, the labs are closer to what Domin (1999) describes as problem based, in which the concepts have been presented and must be used to answer specific questions.

While laboratory manipulations are relatively simple, most students fill the entire period with discussions among themselves, or Socratic dialogs with instructors. These are devoted to interpreting phenomena, repeating procedures to reinforce ideas, and writing up their findings and interpretations. This approach is consistent with the ideas of Nash (summarized by Pickering, 1988), who suggests that simple puzzle laboratory activities that do not bog students down in complicated procedures may be best for introductory courses. Students usually work in pairs, but hand in separate laboratory reports.

The most valuable source of feedback on how students are progressing in their understanding of material, and how intructors can improve future presentations, continues to be that from interactive learning sessions in the classroom, office hours, and the laboratory (Garafalo & LoPresti, 1993; Cohen *et al*. 2000)). Other sources include surveys and performance on examinations. The students' willingness to interact enables instructors to clarify points, and uncover misconceptions that might otherwise go undetected. Carefully listening to students and reflecting upon the phenomena within which concepts are rooted has allowed us to anticipate the disequilibrium that students face when they do not understand something or when an idea conflicts with the way in which they believe the world operates (Garafalo & LoPresti, 1993). Introducing material so that it produces only moderate disequilibrium is a major challenge to instructors, given the diverse population of learners. The construction of more logical Socratic sequences of questions for the classroom, text, and laboratory is an ongoing process.

In recent years, poor class performance on the first test in semester one has led to a steady decrease in the number of concepts that are introduced in the first few weeks of the curriculum. Some have been eliminated entirely, while others have been moved to more appropriate places. For example, prior to academic year 1997-98, the concept of energy was introduced before ideas about chemical processes and gases, but students complained that the course seemed more like a physics course than a chemistry course, and it was evident that they were being overwhelmed with too many ideas in too little time. In particular, not enough time was spent on evidence supporting the atomic theory, and how to distinguish between compounds, mixtures, and elements. The current presentation avoids any mention of the concept of energy until after the introduction of the periodic table, but discussion of the quantity mv^2 anticipates the idea. This approach necessitates resorting to velocity data from the 20th century instead of heat capacity data from the 19th century to support Avogadro's hypothesis.

Responding to student difficulties is producing a curriculum that is more in line with the historical evolution of chemistry. For example, in 1992-93 the presentation included much more about atomic structure and bonding in the first term, although ideas about structure were still delayed. Electrons were not introduced until week five, and atomic number was not introduced until week six in that presentation. Currently, electrons are not introduced until week 12, and atomic number is not introduced until week 1 of semester two. These changes reflect an increased awareness on the part of the instructors of which concepts are essential and which are not for students to explore a certain aspect of chemistry. For example, while the details of electron configurations may provide insight on stoichiometric relationships, knowledge of atomic masses and how to balance equations is sufficient to conduct lessons on calculations involving moles.

Presentations are also evolving so that experimental observations are better coordinated with relevant theory. Several years ago discussions of Dalton's rule of simplicity and Avogadro's hypothesis were not closely connected with the presentation of the gas laws, and the behavior of gases was discussed in terms of factors affecting gas volume. Now these topics are closely connected, and the emphasis is on factors affecting gas pressure, which makes for a better connection with the ideas of kinetic theory. Topics like quantum numbers and orbitals have been eliminated, and replaced in semester two by presentations that use comparison of ionization energies to suggest the existence of different energy levels in atoms. Tsaparlis (1997b) summarizes support for this controversial approach. Coverage of descriptive chemistry and colligative properties has become more restricted, and focused on supporting other topics like stoichiometry and equilibrium.

STUDENT PERFORMANCE AND ATTITUDE

It has not been possible to measure the effects on student performance of these interventions in a quantitative way. No opportunity exists for control groups, and several factors change from year to year. These include class size, average SAT scores, moving from a trimester system with 20 weeks of laboratory to a semester system with 30 weeks of laboratory, different learning facilitators of varying ability, different number of questions on exams, and pedagogic and content modifications that are made based on feedback from the previous academic year. The latter factor influences the nature of test questions each year. While instructors seek to interact with all students in the laboratory, the amount of help received by each student is different. In order to reach a conclusion, some students require more extensive interaction with an instructor than others. Some students actively engage in dialogue with instructors, while others prefer to keep this to a minimum. In addition, different instructors have different levels of expertise in guiding students in the right direction. In spite of all of these factors, the extensive feedback from students during the past ten years has driven and will continue to drive the evolution of a chemistry curriculum that is more consistent with the logical structure of the discipline, and that anticipates the needs of novice learners.

The administration of a relatively large number of tests during the semester is consistent with a correlation between frequent testing and student performance (Martin &

Srikameswaran, 1974). The five, hour-long tests currently comprise 20 multiple choice questions, while the two-hour final exam comprises 40 of these questions. The advantages of carefully constructed multiple choice questions over open-ended essay questions in focusing students on specific information and skills has been discussed (Statkiewicz & Allen, 1983). There are usually more qualitative than quantitative questions, with only a few of the recognition and recall variety. For the units under discussion, multiple-choice versions of questions like those in Appendix I are common. Such questions allow students to demonstrate that they are developing the ability to apply skills and knowledge to specific situations, but our observations suggest that most students do not come away from Unit 6 with a complete understanding of all of the interrelated ideas. In addition, the survey nature of the course leads to limited testing of each topic on the final exam, making it hard to judge how much students have retained throughout the semester.

The historical approach to relative atomic mass and compound formulas reinforces the construction of concepts from observation, but students' prior knowledge is often a stumbling block at first. For example, essentially all students know that the formula for water is H_2O. Since the fact-hypothesis-deduction activities in Unit 6 are stressful for many students, they resist immersing themselves in the activities, hoping that they can cling to their knowledge of the correct formula for water. It is only when they realize that the accepted formula cannot be concluded merely from hydrogen / oxygen mass ratios or combining volume ratios, and that various formulas are possible depending on the starting hypothesis, that they see the need to engage in the struggle to reason toward specific conclusions. Students frequently need to be reminded that experimentally determined mass and volume ratios are facts, that laws are statements of fact, and that at this point formulas and atomic masses are deductions based on hypotheses.

Many students seem relieved when we move on to Unit 7 and accept the relative atomic masses in the modern periodic table. On the positive side, since the unit exposes students to dilemmas faced by actual scientists, most are gratified when they realize that they are capable of engaging in the same type of reasoning as that of mature scientists. Students have commented that they could not appreciate what was happening when they were first exposed to the material, but later realized that the approach helped them to learn how to think. This has been reinforced by responses on surveys and student evaluations. Realistically it can be said that Unit 6 begins to make students aware of the fact that knowledge about the invisible atomic world is based upon inferences drawn from macroscopic observations, and that while experimental facts have an immutable nature, there may be room for several interpretations of those facts, and only when many facts can be consistently connected does a given interpretation gain wide acceptance.

CONCLUSIONS

Several authors have suggested the potential benefits of using history as more than just a source of anecdotes in the teaching of chemistry. Niaz & Rodriquez (2001) indicate that a historical perspective can facilitate students' conceptual understanding. According to Jensen (1998c), the study of the history of chemistry provides us with a blueprint of how to logically organize the current concepts and models while simultaneously revealing many of the underlying assumptions and relationships. Tsaparlis (1997a) points out that the history of scientific discoveries shows the natural route of human thinking and matches the cognitive development of the human mind, while Wicken (1976) stresses that describing the development of physical theories, instead of presenting them as proven, may encourage students to respond more creatively to new problems and empirical data.

Indeed, when the corresponding author sought to create a more logical presentation, based on experimental evidence, the result corresponded closely to the historical emergence of concepts. Studying the history of kinetic theory has resulted in deeper insights about the connection of observations with inferences drawn from the theory, and has provided insights on how to structure presentations at a more advanced level. The current authors are enthusiastic about a historic approach, but caution that it is easy to overwhelm students with too many ideas. Sanitizing the details is important. For example, the accuracy of the analytical data in the 19th century was not as good as it is today, and this added to the confusion in determining correct relative atomic masses. No mention is made of this in our development, nor is mention made of the caloric theory. Since concept development takes time, care must be taken in choosing content. Several years ago the historical concepts of combining weight and combining capacity were used, but this approach was abandoned, since it took too long for the students to construct these ideas, only to dispose of them shortly thereafter.

Formal operational reasoning has been identified as essential for success in science and mathematics courses (Bitner, 1991). Work by Lawson indicates that hypothetico-deductive reasoning is the essence of formal reasoning (Lawson, 1992), and that such reasoning limits acedemic performance in science courses more than prior, domain-specific knowledge (Lawson & Johnson, 1998). When an individual engages in this type of reasoning, intuitively generated ideas are proposed as hypotheses, the consequences are deduced, and evidence of some sort is compared with those deduced consequences to allow rejection or retention of the initial hypotheses (Lawson, *et. al*., 1991). Individuals who are unskilled in this type of reasoning often modify factual data in order to make it consistent with their hypothesis instead of using the data to test hypothesis validity (Lawson, 1992). In essence, the hypothesis is taken as the truth. When presented with correct reasoning, such individuals follow along easily, indicating that the problem is not one of comprehension, but one of generation.

Many MCPHS freshman students exhibit behavior consistent with these findings. Our approach attempts to bridge the gap by providing numerous examples where students are given experimental facts and hypotheses about the atomic realm, and must either test the hypotheses, or draw conclusions based on them. A few students with prior college degrees but no or very limited chemistry exposure have done very well with the material described here, which is also consistent with Lawson's findings. This observation is encouraging, suggesting that we are on the right track in constructing a logical presentation. It is also consistent with findings indicating that students perform better when presentations avoid careless mixing of material from the macroscopic, atomic, and symbolic realms (Georgiadou & Tsaparlis, 2000).

However, optimism must be tempered by the fact that many students continue to struggle. Topic sequence and time spent per unit in semester I have not changed for three years. The percent of the class performing poorly in semester I, and the grades on exams related to Units 3 and 6 during this three-year period also have been fairly constant. Last year 17% of the class (116 total students) received a grade below that of C-. While the sequence and presentation of topics appears reasonable, the amount of information presented may still be too great, particularly for our more poorly prepared students. This is supported by the fact that correct response rates on the more difficult test questions are often less than 50%. A likely contributing factor to poor performance for many students is resistance to active learning, and failure to attend help sessions.

In a survey to which 75% of last year's class responded (87 out of 116 total students), 10 students felt that the level of difficulty for this course was too high, while 20 students felt that it was too low. The rest thought it was appropriate. It is clear that the wide range of backgrounds our students bring to their freshman year continues to pose challenges. As a next step, a coordinated effort between instructors and Student Support Services is currently underway to help students embrace active learning strategies as early as possible. Further reduction of content in some of the other units will also be considered.

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APPENDIX I: SAMPLE QUESTIONS FOR UNITS 3* AND 6**

*** Observations about matter***
***** Making inferences about the atomic realm***

Q1. Substance x decomposes when heated to give substances y and z. Which statements are true?

a) x must be an element b) x may be an element c) y must be an element
d) y may be an element e) z must be an element f) x must be a compound
g) y must be a compound h) y may be a compound i) z must be a compound

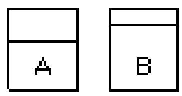
Q2. Two compounds are known, containing only the elements carbon and oxygen. The mass ratio of oxygen / carbon for compound A is 1.33/1 and for compound B it is 2.66/1. Remember that these ratios represent the number of grams of oxygen that combine with 1 g of carbon. Since the ratio is twice as large for compound B as it is for compound A, there is twice as much oxygen in compound B as in compound A. One way to interpret this is to assume that twice as many atoms of oxygen combine with carbon in compound B as in compound A. If the formula of compound A were CO, then what would be the formula for compound B?

Q3. Let us look again at the two compounds that contain only oxygen and carbon. When we make the reverse ratios, we get: carbon/oxygen = 1/1.33 = 0.7519 for compound A, and carbon / oxygen = 0.3759 for compound B. Compare the ratios. How much larger is the ratio for compound A compared to that for compound B? If the formula for compound B were CO, what would we have to conclude about the formula for compound A?

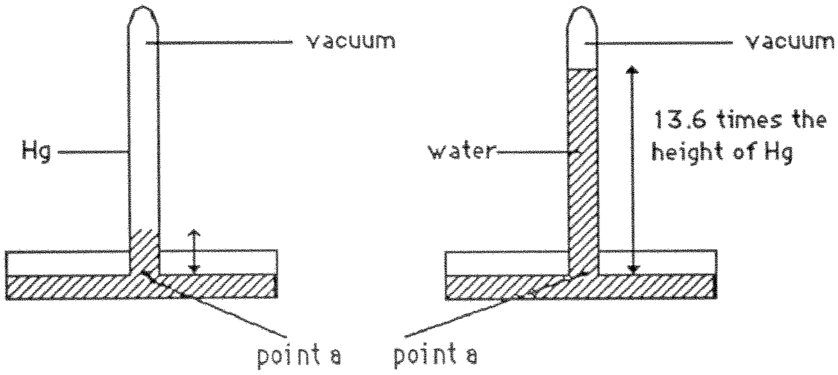
Q4. A sample of a certain liquid appears homogeneous, but when it boils, the remaining liquid exhibits properties that change as the boiling continues. For example, the boiling point and freezing point of the liquid change. The liquid is most likely

a) an element b) a compound c) a solution d) a heterogeneous mixture

Q5. Consider two identical beakers sitting on a table (to the right), A containing 50 g of liquid X, and B containing 50 g of water. Which exerts the greater force on the table? Do they exert the same pressure on the table?

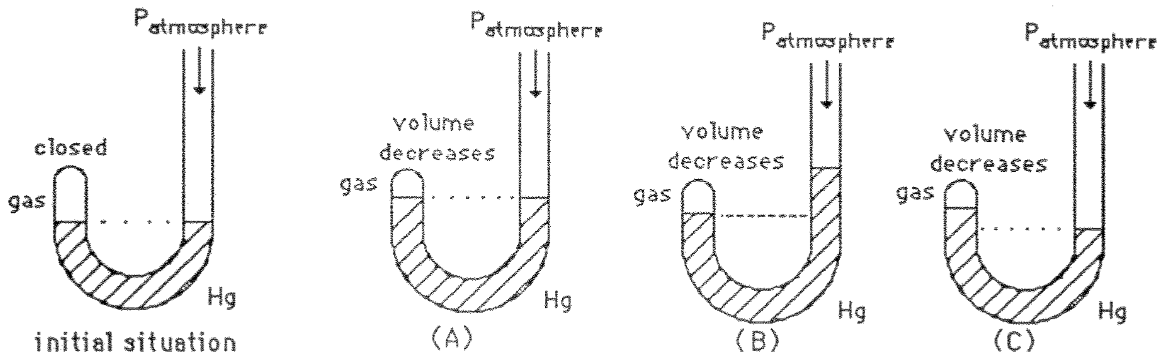


Q6. Consider the two barometers pictured below, side by side on a lab bench.



For each barometer, at the bottom of the liquid column (marked as point **a** in the pictures), what is the pressure due to? If there is no motion in the column, what can we conclude about the pressure at this point, compared to the pressure exerted by the atmosphere? What can you conclude about the weight of the Hg column in the tube on the left compared to the weight of the water column in the tube on the right? What about the masses? the densities?

Q7. Consider the J tube pictured below, to the left, which is labeled initial situation. A sample of gas is in the space above the mercury in the left arm. The right arm is open to the atmosphere. Make a hypothesis about which picture, A, B, or C represents what the tube will look like after more mercury is added to the right arm of the tube. The temperature remains constant.



You will be given a piece of clear plastic tubing, to which you can add water, to simulate a J-tube. Conduct some experiments with the tube and water to help you answer the question. Interpret what happens in terms of Boyle's law. Explain why each choice above is right or wrong based on Boyle's law.

Q8. For each of the following, describe your answer in terms of what would be happening **at the atomic level.**

a. When the volume of a gas increases, why must temperature increase to keep pressure constant?
b. If you add more gas to a container without changing the volume, what must you do to keep the pressure constant? Why?

Q9. For the water formula OH, we concluded that the ratio of the mass of a hydrogen atom to that of an oxygen atom was 1 amu/8 amus, if we assign an H atom a mass of 1 amu. What would be the formula for water if:

a) an oxygen atom were only 4 times as massive as a hydrogen atom?
b) an oxygen atom were 16 times as massive as a hydrogen atom?

Q10. Given the fact that the hydrogen / oxygen mass ratio is 1/8 in water, if the formula for water is

a) H_2O, what is the mass of an O atom compared to that of an H atom?

b) HO_2, what is the mass of an O atom compared to that of an H atom?

Q11. For this reaction, the Law of Definite Proportions allows us to write:

$$\frac{\text{mass of copper produced}}{\text{mass of iron dissolved}} = \text{fixed value} = \frac{\text{total mass of copper atoms produced}}{\text{total mass of iron atoms dissolved}}$$

a) Suppose the atoms react 1 to 1: Fe (s) + Cu (aq) $\rightarrow$ Cu (s) + Fe (aq)
Use the experimentally determined copper/iron mass ratio to determine the mass of a copper atom compared to the mass of an iron atom. (assume an Fe atom = 1.00 amu)

b) Suppose 2 copper atoms react for each iron atom: Fe (s) + 2Cu (aq) $\rightarrow$ 2 Cu (s) + Fe (aq).
Use the experimentally determined copper/ iron mass ratio to determine the (copper atom)/(iron atom) mass ratio. (assume an Fe atom = 1.00 amu)

Q12. It is an experimental fact that one liter of hydrogen gas reacts completely with one liter of chlorine gas, when they are both at the same temperature and pressure, to form the gas hydrogen chloride. If the one liter of hydrogen were at a higher temperature but still the same pressure as the one liter of chlorine, would reaction still be complete, or would one of the gases now be in excess? Explain your reasoning.

Q13. At a given pressure and temperature, one volume of nitrogen reacts completely with one volume of oxygen to produce two volumes of nitrogen oxide at this pressure and temperature. To summarize: 1 vol nitrogen + 1 vol oxygen $\rightarrow$ 2 vol nitrogen oxide Avogadro's Hypothesis: Equal volumes of any two gases, each at the same temperature and pressure, contain an equal number of particles. Note that a particle could be either an atom, or a molecule (a molecule remember contains two or more atoms bound together). If the formula for nitrogen oxide is NO, deduce what the particles are in the nitrogen container and in the oxygen container that combine to form NO, if Avogadro's hypothesis is assumed. Write an equation showing the formation of NO from these species.

REFERENCES

Abraham, M.; Williamson, V. & Westbrook, S. (1994). A cross-age study of the understanding of five chemistry concepts. *Journal of Research in Science Teaching, 31,* 147-165.

Arons, A. (1990). *A Guide to introductory physics teaching*, Wiley: New York.

Ausubel, D., Novak, J. & Hanesian, H. (1978). *Educational psychology: A cognitive view*. New York: Holt, Rinehart , & Winston

Bar, V. & Travis, A. (1991). Children's views concerning phase changes. *Journal of Research in Science Teaching, 28,* 363-382.

Bernatowicz, A. (1970). Dalton's rule of simplicity. *Journal of Chemical Education, 47,* 577-579.

Bitner, B. (1991). Formal operational reasoning modes: Predictors of critical thinking abilities and grades assigned by teachers in science and mathematics for students in grades nine through twelve. *Journal of Research in Science Teaching, 28,* 265-274.

Brock, W. (1963). Prout's chemical bridgewater treatise. *Journal of Chemical Education, 40,* 652-655.

Carpenter, D. (1966). Kinetic theory, temperature, and equilibrium. *Journal of Chemical Education, 43*, 332-334.

Cohen, J., Kennedy-Justice, M., Pai, S., Torres, C., Toomey, R., DePierro, E., & Garafalo, F. (2000). Encouraging meaningful quantitative problem solving *Journal of Chemical Education, 77*, 1166-1173.

Deberg, K. (1995). Student understanding of volume, mass, & pressure of air within a sealed syringe in different states of compression. *Journal of Research in Science Teaching, 32*, 871-884.

Domin, D. (1999). A review of laboratory instruction styles. *Journal of Chemical Education, 76,* 543-547.

Feifer, N. (1966). The relationship between Avogadro's principle and the law of Gay-Lussac. *Journal of Chemical Education, 43,* 411-412.

Fitzgerel, R. & Verhoek, F. (1960). The law of Dulong and Petit. *Journal of Chemical Education, 37*, 545-549.

Garafalo, F.& LoPresti, V. (1986). An integrated college freshman natural science curriculum. *Journal of Chemical Education, 63,* 854-858.

Garafalo, F., LoPresti, V., & LaSala, E. (1988). Student evaluation of an integrated college freshman natural science curriculum *Journal of Chemical Education, 65,* 890-891.

Garafalo, F.& LoPresti, V. (1993). Evolution of an integrated college freshman curriculum. *Journal of Chemical Education, 70*, 352-359.

Gensler, W. (1970). Physical change versus chemical change. *Journal of Chemical Education, 47,* 154-155.

Georgiadou, A. & Tsaparlis, G. (2000). Chemistry teaching in lower secondary school with methods based on: A) psychological theories; B) the macro, representational, and submicro levels of chemistry. *Chemistry Education: Research and Practice in Europe* (*CERAPIE*), *1*, No. 2, 217-226. [http://www.uoi.gr/conf_sem/cerapie]

Goldstein, M. & Goldstein, I.(1993). *The refrigerator and the universe*, Harvard University Press, Cambridge, MA.

Haider, A. & Abraham, M. (1991). A comparison of applied and theoretical knowledge of concepts based on the particulate nature of matter. *Journal of Research in Science Teaching, 28,* 919-938.

Hesse, J. & Anderson, C. (1992). Students' conceptions of chemical change. *Journal of Research in Science Teaching, 29*, 277-299.

Jensen, W. (1998a). Logic, history, and the chemistry textbook (II). *Journal of Chemical Education, 75*, 817-828.

Jensen, W. (1998b). Logic, history, and the chemistry textbook (III). *Journal of Chemical Education, 75*, 961-970.

Jensen, W. (1998c). Logic, history, and the chemistry textbook (I). *Journal of Chemical Education, 75*, 679-687.

Johnstone, A. (1991a). Thinking about thinking. *International Newsletter on Chemical Education*, No. 6, 7-11.

Johnstone, A. (1991b). Cognitive style and analytical ability and their relationship to competence in the biological sciences. *Journal of Biological Education, 25*, 135-140.

Johnstone, A. (2000). The presentation of chemistry - Logical or psychological? *Chemistry Education: Research and Practice in Europe* (*CERAPIE*), *1*, 9-15. [http://www.uoi.gr/conf_sem/cerapie]

Karplus, R. (1977). Science teaching and the development of reasoning. *Journal of Research in Science Teaching, 14,* 169-175.

Kiser, B. (1991). Determining the relative masses of copper and iron atoms. *Journal of Chemical Education, 68*, 250.

Lawson, A.E., McElrath, C.B., Burton, M.S., James, B.D., Doyle, R.D., Woodward, S.L., Kellerman, L., & Snyder, J.D. (1991). Hypothetico-deductive reasoning and concept acquisition: Testing a constructivist hypothesis. *Journal of Research in Science Teaching, 28*, 953-970.

Lawson, A. (1992). What do tests of formal reasoning actually measure? *Journal of Research in Science Teaching, 29,* 965-983.

Lawson, A. & Johnson, M. (1998). What are the relative effects of reasoning ability and prior knowledge on biology achievement in expository and inquiry classes? *Journal of Research in Science Teaching, 35,* 89-103.

Lochhead, J. & Whimby, A. (1987). In *New Directions in Teaching and Learning, No. 30*; E. Stice (ed.); Jossey-Bass: San Francisco, 73-92.

Lund, E. (1968). _Activated complex - A centenarian? *Journal of Chemical Education, 45*, 125-128.

Martin, R. & Srikameswaran, K. (1974). Correlation between frequent testing and student performance. *Journal of Chemical Education, 51*, 485-486.

Mendoza, E. (1990). The lattice theory of gases: A neglected episode in the history of chemistry. *Journal of Chemical Education, 67*, 1040-1042.

Nakhleh, M. (1992). Why some students don't learn chemistry. *Journal of Chemical Education, 69,* 191-196.

Nakhleh, M., Lowery, K., & Mitchell, R. (1996). Narrowing the gap between concepts and algorithms in freshman chemistry. *Journal of Chemical Education, 73*, 758-762.

Niaz, M. & Rodriquez, A. (2001). Do we have to introduce history and philosophy of science or is it already inside chemistry? *Chemistry Education: Research and Practice in Europe* (*CERAPIE*), *2*, No. 2, 159-164. [http://www.uoi.gr/conf_sem/cerapie]

Pickering, M. (1988). Report of the NEACT conference: The chemistry lab and its future. *Journal of Chemical Education, 65*, 449-450.

Resnick, L. (1983). Mathematics and science learning: A new conception. *Science, 220*, 477-478.

Rhodes, G. (1992). Does a one-molecule gas obey Boyle's law? *Journal of Chemical Education, 69,* 16-18.

Statkiewicz, W. & Allen, R. (1983). Practice exercises to develop critical thinking skills. *Journal of Research in College Science Teaching, 12*, No. 5, 5-9.

Towns, M; Kreke, K.; & Fields, A. (2000). An action research project. *Journal of Chemical Education, 77*, 111-115.

Tsaparlis, G. (1997a). Atomic and molecular structure in chemical education – A critical analysis from various perspectives of science education. *Journal of Chemical Education, 74,* 922-926.

Tsaparlis, G. (1997b). Atomic orbitals, molecular orbitals, and related concepts: Conceptual difficulties among chemistry students. *Research in Science Education, 27*, 271-287.

Whitaker, R. (1979). The early development of kinetic theory. *Journal of Chemical Education, 56,* 315-318.

Wicken, J. (1976). The value of historical concepts in science education. *Journal of Chemical Education, 53*, 96-97.