Unit 3 – Matter I Standards

Chemistry 1. Atomic and Molecular Structure. Starting with atomic and molecular structure in the study of high school chemistry is important because this topic is a foundation of the discipline. However, because structural concepts are highly theoretical and deal with the quantum realm, they can be hard to relate to real-world experience. Ideally, from grades three through eight, students have been gradually introduced to the atomic theory; and by the end of the eighth grade, they should have covered the major concepts in the structure of atoms and molecules. By the time students reach high school, they should be familiar with basic aspects of this theory.

The study of structure can begin with the simplest element, hydrogen. Students can progress from a simple model of the atom to the historic Bohr model and, finally, to a quantum mechanical model, which is the picture of the atom that students should ultimately develop. Students should learn that the quantum mechanical model takes into account the particle and wave properties of the electron and uses mathematical equations to solve for electron energies and regions of electron density.

Students should understand that the energy carried by electrons either within an atom or as electricity can be transformed into light energy. Those who have completed high school physics will be familiar with the properties of electromagnetic waves. In chemistry students learn to apply the equations $E = h\nu$ and $c = \lambda\nu$. Students without significant training in physics will need to understand electromagnetic radiation as energy, frequency, and wavelength. The necessary mathematical background for the study of chemistry includes algebraic isolation of variables, use of conversion factors, and manipulation of exponents, all of which are covered in the mathematics standards for the middle grades.

1. The periodic table displays the elements in increasing atomic number and shows how periodicity of the physical and chemical proper-ties of the elements relates to atomic structure. As a basis for understanding this concept:

1a. ***Students know*** how to relate the position of an element in the periodic table to its atomic number and atomic mass. An atom consists of a nucleus made of protons and neutrons that is orbited by electrons. The number of protons, not electrons or neutrons, determines the unique properties of an element. This number of protons is called the element’s atomic number. Elements are arranged on the periodic table in order of increasing atomic number. Historically, elements were ordered by atomic mass, but now scientists know that this order would lead to misplaced elements (e.g., tellurium and iodine) because differences in the number of neutrons for isotopes of the same element affect the atomic mass but do not change the identity of the element.

1b. ***Students know*** how to use the periodic table to identify metals, semimetals, nonmetals, and halogens. Most periodic tables have a heavy stepped line running from boron to astatine. Elements to the immediate right and left of this line, excluding the metal aluminum, are semimetals and have properties that are intermediate between metals and nonmetals. Elements further to the left are metals. Those further to the right are nonmetals. Halogens, which are a well-known family of nonmetals, are found in Group 17 (formerly referred to as Group VIIA). A group, also sometimes called a “family,” is found in a vertical column in the periodic table.

1c. ***Students know*** how to use the periodic table to identify alkali metals, alkaline earth metals and transition metals, trends in ionization energy, electronegativity, and the relative sizes of ions and atoms. A few other groups are given family names. These include the alkali metals (Group 1), such as sodium and potassium, which are soft and white and extremely reactive chemically.

Alkaline earth metals (Group 2), such as magnesium and calcium, are found in the second column of the periodic table. The transition metals (Groups 3 through 12) are represented by some of the most common metals, such as iron, copper, gold, mercury, silver, and zinc. All these elements have electrons in their outer *d* orbitals.

Electronegativity is a measure of the ability of an atom of an element to attract electrons toward itself in a chemical bond. The values of electronegativity calculated for various elements range from one or less for the alkali metals to three and one-half for oxygen to about four for fluorine.

Ionization energy is the energy it takes to remove an electron from an atom. An element often has multiple ionization energies, which correspond to the energy needed to remove first, second, third, and so forth electrons from the atom. Generally in the periodic table, ionization energy and electronegativity increase from left to right because of increasing numbers of pro-tons and decrease from top to bottom owing to an increasing distance between electrons and the nucleus. Atomic and ionic sizes generally decrease from left to right and increase from top to bottom for the same reasons. Exceptions to these general trends in properties occur because of filled and half-filled sub shells of electrons.

1d. *Students know* how to use the periodic table to determine the number of electrons available for bonding. Only electrons in the outermost energy levels of the atom are available for bonding; this outermost bundle of energy levels is often referred to as the *valence shell* or *valence shell of orbital.* All the elements in a group have the same number of electrons in their outermost energy level. Therefore, alkali metals (Group 1) have one electron available for bonding, alkaline earth metals (Group 2) have two, and elements in Group 13 (once called Group III) have three. Unfilled energy levels are also available for bonding. For example, Group 16, the halogens, has room for two more electrons; and Group 17, the halogens, has room for one more electron to fill its outermost energy level.

To find the number of electrons available for bonding or the number of unfilled electron positions for a given element, students can examine the combining ratios of the element's compounds. For instance, one atom of an element from Group 2 will most often combine with two atoms of an element from Group 17 (e.g., $MgCl_2$) because Group 2 elements have two electrons available for bonding, and Group 17 elements have only one electron position open in the outermost energy level. (Note that some periodic tables indicate an element's electron configuration or preferred oxidation states. This information is useful in determining how many electrons are involved in bonding.)

1e. *Students know* the nucleus of the atom is much smaller than the atom yet contains most of its mass. The volume of the hydrogen nucleus is about one trillion times less than the volume of the hydrogen atom, yet the nucleus contains almost all the mass in the form of one proton. The diameter of an atom of any one of the elements is about 10,000 to 100,000 times greater than the diameter of the nucleus. The mass of the atom is densely packed in the nucleus.

The electrons occupy a large region of space centered around a tiny nucleus, and so it is this region that defines the volume of the atom. If the nucleus (proton) of a hydrogen atom were as large as the width of a human thumb, the electron would be on the average about one kilometer away in a great expanse of empty space. The electron is almost 2,000 times lighter than the proton; therefore, the large region of space occupied by the electron contains less than 0.1 percent of the mass of the atom.

Chemistry 2. Chemical Bonds. Standard Set 2 deals with two distinct topics: chemical bonds and intermolecular attractive forces, such as hydrogen bonds. A logical place to begin the study of this standard is with a discussion of the chemical bond. A key point to emphasize is that when atoms of two different elements join to form a covalent bond, energy is almost always released. Conversely, breaking bonds always requires the addition of energy. Students should understand that the sum of these two processes determines the net energy released or absorbed in a chemical reaction.

This standard set requires a basic knowledge of electrostatics and electronegativity and a thorough knowledge of the periodic table. After studying standards for chemistry for the elementary grades, students should know that matter is made of atoms and that atoms combine to form molecules. Students can also be expected to know that atoms consist of protons, neutrons, and electrons. Although knowledge of complex mathematics is not required for this standard, some background in three-dimensional geometry will be helpful.

2a. ***Students know*** atoms combine to form molecules by sharing electrons to form covalent or metallic bonds or by exchanging electrons to form ionic bonds.

In the localized electron model, a covalent bond appears as a shared pair of electrons contained in a region of overlap between two atomic orbitals. Atoms (usually nonmetals) of similar electronegativities can form covalent bonds to become molecules. In a covalent bond, therefore, bonding electron pairs are localized in the region between the bonded atoms. In metals valence electrons are not localized to individual atoms but are free to move to temporarily occupy vacant orbitals on adjacent metal atoms. For this reason metals conduct electricity well.

When an electron from an atom with low electronegativity (e.g., a metal) is removed by another atom with high electronegativity (e.g., a nonmetal), the two atoms become oppositely charged ions that attract each other, resulting in an ionic bond. Chemical bonds between atoms can be almost entirely covalent, almost entirely ionic, or in between these two extremes. The triple bond in nitrogen molecules (N_2) is nearly 100 percent covalent. A salt such as sodium chloride (NaCl) has bonds that are nearly completely ionic. However, the electrons in gaseous hydrogen chloride are shared somewhat unevenly between the two atoms. This kind of bond is called *polar covalent.* (Note that elements in groups 1, 2, 16, and 17 in the periodic table usually gain or lose electrons through the formation of either ionic or covalent bonds, resulting in eight outer shell electrons. This behavior is sometimes described as "the octet rule.")

2b. ***Students know*** chemical bonds between atoms in molecules such as H_2, CH_4, NH_3, H_2CCH_2, N_2, Cl_2 and many large biological molecules are covalent.

Organic and biological molecules consist primarily of carbon, oxygen, hydrogen, and nitrogen. These elements share valence electrons to form bonds so that the outer electron energy levels of each atom are filled and have electron configurations like those of the nearest noble gas element. (Noble gases, or inert gases, are in the last column on the right of the periodic table.) For example, nitrogen has one lone pair and three unpaired electrons and therefore can form covalent bonds with three hydrogen atoms to make four electron pairs around the nitrogen. Carbon has four unpaired electrons and combines with hydrogen, nitrogen, and oxygen to form covalent bonds sharing electron pairs.

The great variety of combinations of carbon, nitrogen, oxygen, and hydrogen make it possible, through covalent bond formation, to have many compounds from just these few elements. Teachers can use ball and stick or gumdrop and toothpick models to explore possible bonding combinations.

2c. ***Students know*** salt crystals, such as NaCl, are repeating patterns of positive and negative ions held together by electrostatic attraction. The energy that holds ionic compounds together, called *lattice energy,* is caused by the electrostatic attraction of *cations,* which are positive ions, with *anions,* which are negative ions. To minimize their energy state, the ions form repeating patterns that reduce the distance between positive and negative ions and maximize the distance between ions of like charges.

Earth 7. Biogeochemical Cycles. Students who complete high school biology/life sciences before they take earth sciences will already have learned about biogeochemical cycles. Through standards presented in lower grade levels, other students should have been ex-posed to life cycles, food chains, and the movement of chemical elements through biological and physical systems. Students should also have studied chemical changes in organisms and should know that through photosynthesis solar energy is used to create the molecules needed by plants. In this standard set students will learn that within the biogeochemical cycles, matter is transferred between organisms through food webs or chains. Matter can also be transferred from these cycles into physical environments where the cycling elements are held in reservoirs. Matter can be transferred back into biological cycles through physical processes, such as volcanic eruptions and products of the rock cycle, particularly those from weathering.

7. Each element on Earth moves among reservoirs, which exist in the solid earth, in oceans, in the atmosphere, and within and among organisms as part of biogeochemical cycles. As a basis for under-standing this concept:

7a. ***Students know*** the carbon cycle of photosynthesis and respiration and the nitrogen cycle. Carbon and nitrogen move through biogeochemical cycles. The recycling of these components in the environment is crucial to the maintenance of life. Through photosynthesis, carbon is incorporated into the biosphere from the atmosphere. It is then released back into the atmosphere through respiration. Carbon dioxide in the atmosphere is dissolved and stored in the ocean as carbonate and bicarbonate ions, which organisms take in to make their shells. When these organisms die, their shells rain down to the ocean floor, where they may be dissolved if the water is not saturated in carbonate. Otherwise, the shells are deposited on the ocean floor and become incorporated into the sediment, eventually turning into a bed of carbonate rock, such as limestone. Uplifted limestone may dissolve in acidic rain to return carbon to the atmosphere as carbon dioxide, sending calcium ions back into the ocean where they will precipitate with carbon dioxide to form new carbonate material. Carbonate rocks may also be subducted, heated to high temperatures, and decomposed, returning carbon to the atmosphere as volcanic carbon dioxide gas. Carbon is also stored in the solid earth as graphite, methane gas, petroleum, or coal. Nitrogen, another element important to life, also cycles through the biosphere and environment. Nitrogen gas makes up most of the atmosphere, but elemental nitrogen is relatively inert, and multicellular plants and animals cannot use it directly. Nitrogen must be “fixed,” or converted to ammonia, by specialized bacteria. Other bacteria change ammonia to nitrite and then to nitrate, which plants can use as a nutrient. Eventually, decomposer bacteria return nitrogen to the atmosphere by reversing this process.

7b. ***Students know*** the global carbon cycle: the different physical and chemical forms of carbon in the atmosphere, oceans, biomass, fossil fuels, and the movement of carbon among these reservoirs.The global carbon cycle extends across physical and biological Earth systems. Carbon is held temporarily in a number of reservoirs, such as in biomass, the atmosphere, oceans, and in fossil fuels. Carbon appears primarily as carbon dioxide in the atmosphere. In oceans carbon takes the form of dissolved carbon dioxide and of bicarbonate and carbonate ions. In the biosphere carbon takes the form of sugar and of many other organic molecules in living organisms. Some movement of carbon between reservoirs takes place through biological means, such as respiration and photosynthesis, or through physical means, such as those related to plate tectonics and the geologic cycle. Carbon fixed into the biosphere and then trans-formed into coal, oil, and gas deposits within the solid earth has in recent years been returning to the atmosphere through the burning of fossil fuels to generate energy. This release of carbon has increased the concentration of carbon dioxide in the atmosphere. Carbon dioxide is a primary greenhouse gas, and its concentration in the atmosphere is tied to climatic conditions.