<u>Chapter 3</u> – <u>Elements and Compounds</u>

3.1 <u>Elements</u>

<u>Elements</u> are the fundamental materials of chemistry and are characterized by the fact that they cannot be broken down into simpler substances by chemical means. Only 88 occur naturally and only the first 92 are typically used in chemical laboratories. In chemistry classes, that number is usually further reduced in terms of laboratories and problems. All other materials encountered in the lab and everyday life are combinations of elements chemically bound to one another.

All elements exist as a collection of particles called atoms. These are the smallest units of an element. They retain the chemical properties (i.e. chemical reactivity) of the element, but not usually the physical properties. The reason is that physical properties are typically bulk properties (e.g. melting is a physical change and you can't melt one atom).

In nature, elements are typically found combined with other elements. A few, oxygen, nitrogen, gold, for example, are found pure, but that is uncommon. Ten elements comprise nearly all of the Earth's crust by mass and they are listed in Table 3.1 on p. 50. The top three: oxygen, silicon, and aluminum are the principal components of most rocks and the fourth, iron, is a major industrial material.

<u>Names of the Elements</u>

Each element has a name and associated symbol. The symbol is a one or two letter shorthand for the element, usually derived from the English name for the element, but there are some exceptions. For example, sodium is "Na" from the Latin for natruim and "W" is for tungsten from the German word for the element: wolfram. The first letter of the symbol is always capitalized. For elements with two letter symbols, the second letter is always lower case. Your book provides an example of why on p. 52. Another would be copper(II) oxide, CuO. If everything were capitalized, CUO, the compound then would be read as a carbon, uranium oxygen compound. For this reason, you should be careful when writing elemental symbols to make sure you have the capitalization correct.

The origin of the names of the elements is not important for this course, but many of them have interesting derivations. For example, cobalt is named after the German word for goblin (kobald) because in the purification process of the metal (smelting the ore) poison gases are emitted which *hundreds of years ago in poorly ventilated factories killed a high number of workers. Chromium is so named because the name derives from the Greek word for color. Chromium compounds are frequently brightly colored. Beryllium's name ironically derives from a word meaning "sweetness." It's ironic because beryllium compounds are highly toxic.*

3.2 <u>Introduction to the Periodic Table</u>

The periodic table is an essential tool for the chemist and using correctly is critical to successfully completing CHM 111, 211, and 212. You may use it less in upper division classes, but that is probably because by then it will be very familiar and you will have memorized much of the important material in it.

Most periodic tables have the same format. All contain at least 3 fundamental pieces of information. Most contain additional information. The periodic table on p. 53 of you textbook is the most basic. Each block represents one element and contains the atomic symbol and atomic number of an element. Each column has a label. Larger periodic tables usually include the atomic mass and elements full name. Less commonly, they will have even more information. For this course and CHM 211/212, you should have an 8.5" x 11" periodic table easily accessible. There is one you can download on the course website or you can find one via an internet search engine.

You must memorize the names and symbols for the first 30 elements, plus bromine, iodine, and uranium. The vast majority of problems will involve these 33 elements. You are not required to memorize their masses or atomic numbers (topics that will come up later in the semester). For example, you must know that that "S" represents sulfur, not sodium.

Column numbers (1-18 and IA/B – VIIIA/B) contain important information, but it's too early to go into the details on that. Looking at the periodic table, four distinct areas are visible. The first two columns, the ten in the middle, six on the right, and two rows underneath. The middle ten are known as the <u>transition metals</u> (3-12 or IB-VIIIB). The first two and last six columns are collectively called the <u>representative elements</u> (1, 2, 13-18, IA-VIIIA). The rows at the bottom are the lanthanides and actinides, respectively. You should know they are there, but except for uranium, they won't be a part of this course. The lanthanide elements are sometimes called "rare earth *elements," a term you may have seen in the news because of their importance in batteries of the kind in electric vehicles.*

About three-quarters of naturally occurring elements are metals, including all of the transition metals, lanthanides, and actinides. Another 20% are nonmetals, with the remain elements called metalloids. Metals and nonmetals are reasonably easy to differentiate because their properties are so different (vide infra), but metalloids can be a challenge. Inspection of a periodic table shows metalloids reside on a diagonal between the other two types of elements.

Metals generally possess the following physical properties. They are malleable (can be pounded into sheets) and ductile (can be drawn into wires). They are good conductors of both heat and electricity. Metals usually have high luster and are relatively dense. (Relatively is vague, so they are dense relative to water.) Metals are usually gray in color. Chemically, metals do not usually react with one another to form compounds, but they do react with nonmetals to form salts.

While many metals have all of these properties, there are a number of exceptions. For example, both lithium and sodium float on water (before they ignite the hydrogen gas produced from their reaction with the water). Of course, gold and copper aren't gray. Metals don't always have a shiny appearance, but that is usually because they react with oxygen and the metal oxide gives the metal surface a dull appearance. Zinc is neither malleable nor ductile. The point here isn't to say that these properties don't characterize metals, because they do, but realize there may be individual exceptions to the list of properties for some metals (and the nonmetals in the next section).

The properties of nonmetals are roughly the opposite of the metals. Therefore, they are brittle, dull, colored, have low densities (in fact several are gases), do react with each other (and metals), and are poor conductors of heat and electricity.

Metalloids do **NOT** *have properties that are the average of metals and nonmetals. Rather, metalloids have a mix of properties, some of which are those of metals, others of nonmetals. For example, crystalline silicon has a shiny mirror finish, grey color, and a density similar to aluminum, but is very brittle and is a poor conductor of heat and electricity (indeed, it is a semiconductor). Chemically, it reacts like a nonmetal. This mix of property types is typical of metalloids.*

Pure elements exist in a number of different forms. For example, metals are simply sheets of *atoms stacked on upon another. (Think of a box filled with tennis balls, but stuck together.) Others have small numbers of atoms discreetly bonded to each other. For example, oxygen gas is two oxygen atoms bound to each other, but no other atoms. Hydrogen, nitrogen, fluorine, chlorine, bromine, and iodine exist with the same structure, which is are called diatomic molecules. Other nonmetals have other numbers of atoms bound together. Finally, the Group 18 (VIIIA) elements, the noble gases, exist as single atoms. You should know the elements that form diatomic molecules and those that are free standing atoms.*

3.3 Compounds and Formulas

As presented earlier, when atoms of two or more different elements chemically combine, the result is a compound. Formally, a compound is a substance that contains two or more elements combined in a definite proportion by mass. The two major kinds of compounds, and the only two that we will work with in CHM 111, are molecular and ionic compounds. Both molecules and ionic compounds are neutral, but internally molecules are also electrically neutral, while the internal components of ionic compounds (aka salts) are charged particles. In salts, the total positive and negative charges are the same and cancel out leaving the bulk solid electrically neutral.

Ionic compounds are comprised of ions, which are simply charged particles. Ions with a positive charge are cations (cat ions), while those with a negative charge are anions (an ions). Charges are typically ±1, ±2, or ±3 (but can be higher in magnitude in rare cases). Salts almost always form lattices, with cations at some vertices and anions at the others. Salts are held together because opposite charges attract and the ions arrange themselves to maximize contact with oppositely charged ions and minimize contact with ions with the same charge.

Molecules are held together by a different kind of bonding, that we will discuss later in the semester. Basically, in an ionic compound neutral elements transfer an internal component (electrons), which create a charge imbalance leading to the charge attraction discussed in the previous paragraph. In a molecule, the atoms share the electrons and so neither atom develops a charge.

Writing Formulas of Compounds

There are two ways chemists refer to compounds: their names and their formulae. When *speaking the name is usually used, but when writing the formula is more common. For example, when referring to the liquid that flows out of the tap in your kitchen, I would call it water, but write down H_2O. Table salt is sodium chloride or NaCl. In general, names are less clunky to use when speaking and the listener identifies the compound more readily, but formulas are more compact and can be a shorthand if the reader knows what is being referred to. Finally, different compounds can have the same chemical formula, which can make using the formula confusing. For example, $C_6H_{12}O_6$ is the formula for glucose, but also 11 other sugars and 7 other non-sugar biomolecules. You can see how confusing it would be if someone referred to glucose as C-six-H-twelve-O-six. On the other hand, if the reader knows the only $C_6H_{12}O_6$ compound being discussed is glucose, using the formula is acceptable.*

There are different kinds of formulas that chemists use, the simplest of which is the compact formula. It is simply a listing of the number of atoms of each element in the compound as their atomic symbols. Thus, a glucose molecule is made up of six carbon atoms, twelve hydrogen atoms, and six oxygen atoms. It tells you ***<u>nothing</u>*** *about which atoms are bound to which other atoms. Thus, in water, H_2O, you may not infer that the hydrogen atoms are bound to each other (which, in fact, they are not).*

There is a variant of the condensed formula that is common. This formula is somewhat expanded and includes some structural information. As noted above, neither H_2O nor $C_6H_{12}O_6$, provides any information about the structure of the chemical. Consider octane. As the name suggests, there are eight of something in the molecule. The condensed formula is C_8H_{18}. This can be expanded a little to $CH_3(CH_2)_6CH_3$. In this formula, we can see there are two CH_3 units at the end of a chain of six CH_2 groups. Similarly on p. 59, your textbook shows that calcium nitrate is better written as $Ca(NO_3)_2$ than CaN_2O_6 because in the former you can see that there are two NO_3 groups of atoms, whereas in the latter formula you get no information.

Unfortunately, until we get to nomenclature (the method of naming compounds), you have no alternative but to read the formulas as described in the book, which is strings of letters and numbers. You'll find that chemists rarely use this method in practice, so you shouldn't get too attached to it.

<u>Composition of Compounds</u> - skip

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