Chapter 3 – Chemical Foundations: Elements, Atoms and Ions

1) Sec 3.1 – The Elements – Ultimately all substances in the universe can be broken down chemically into elements. Nature uses a small number of these elements to make up all of matter. They are pure substances although they do not always appear in nature in their elemental form.

2) Sec 3.2 – Symbols for the Elements – Each element is represented by a symbol that consists of one or two letters of the name of the element. The first letter is always capitalized. If a second letter is needed, it must be lower case. Often the letters chosen are from the first and second letters of the name. Some element symbols are derived from older Latin names. (iron = Fe, from Latin name of ferrum)

3) Basic Laws of Chemistry (Experimental Basis of the Atomic Theory)

   a) Law of Conservation of Mass – The is no detectable change in mass during a chemical reaction.

      Example – If 100 g of calcium carbonate, $CaCO_3$, when heated decomposes to produce 56 g of calcium oxide, CaO. How many grams of carbon dioxide, CO_2, escape into the air? $CaCO_3 \longrightarrow CaO + CO_2$

   b) Law of Definite Composition – A compound always contains the same elements in the same proportion by mass.

      Example – Calcium oxide, CaO, is composed of 71.4% calcium and 28.6% oxygen. How many grams of oxygen must be combined with 40 g of calcium to make calcium oxide?

4) Sec 3.3 – Atoms – The smallest particle of an element that retains the properties of the element.

   Dalton's Atomic Theory

   a) All elements are composed of tiny indivisible particles called atoms.

   b) Atoms from the same element are identical. The atoms of any one element are different from those of any other element.

   c) Atoms of different elements can combine with one another in simple whole number ratios to form compounds. (Explains Law of Definite Composition.)

      Example – One atom of calcium combines with one atom of oxygen to form calcium oxide, CaO. Calcium atoms have a relative mass of 40 and oxygen atoms have a relative mass of 16. What percentage of calcium oxide's mass is due to calcium and what percentage is due to oxygen?

d) Chemical reactions occur when atoms are separated, joined, or rearranged. However, atoms of one element are not changed into atoms of another by a chemical reaction. (Explains Law of Conservation of Mass.)

$$C_3H_8 + 5\ O_2 \longrightarrow 3\ CO_2 + 4\ H_2O$$

5) Sec 3.4 – Formulas of Compounds

A compound is a pure substance that is made up of two or more different elements chemically combined. Chemists use the symbols of the elements to express which elements are present in the form of a chemical formula. If an element is present, then the symbol is there. If the element is not present, then it is not there. If a symbol is in the formula, we assume that there is one atom of that element unless we are told differently by using a number as a subscript to the lower right of the symbol.

Examples:

a) $C_6H_{12}O_6$ represents the formula for a compound that contains 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms.

b) $CaCl_2$ represents the formula for a compound that contains 1 calcium atom and 2 chlorine atoms.

5) Sec 3.5 – The Structure of the Atom (Electrons, Protons, and Neutrons)

a) Electrons are negatively charged subatomic particles. They were discovered by scientists whose main interests were in electricity rather than chemistry. These scientists studied the flow of electric current through gases at low pressures. They contained the gases using a closed gas tube with metal disks called electrodes at each end. When connected to a high voltage, the tube glows. One electrode, the anode, becomes positively charged. The other electrode, the cathode, becomes negatively charged. The glowing beam which travels from the cathode to the anode is called a cathode ray. J. J. Thomson in 1897 showed that the cathode ray could be deflected by either magnets or by electrically charged plates and was a collection of very small negatively charged particles, all alike, moving at high speed. He named these particles electrons. He found that the electron was almost 2000 times lighter than a hydrogen atom (the lightest atom known).

b) The Proton – Shortly after electrons were discovered, scientists began to think about the particles left over when a hydrogen atom loses and electron. Since atoms are electrically neutral, researchers reasoned that the leftover particle should have a positive charge. Experimental evidence for such particles, protons, was soon found. The proton carries a single unit of positive charge and is 1840 times heavier than an electron. A proton is what remains when a hydrogen atom is stripped of an electron.

c) The Neutron – To account for the mass of most atoms, it was necessary to assume the existence of a third particle. In 1932, James Chadwick confirmed the existence of this third particle, the neutron. He proved that neutrons are subatomic particles with no charge, but their mass nearly equals that of the proton.

d) Summary [You will need to know!]

Particle	Relative Charge	Relative Mass	Actual Mass	Location
electron	-1	0.00055 (0)	9.11×10^{-28} g	revolves around nucleus
proton	+1	1.0	1.67×10^{-24} g	in nucleus
neutron	none (0)	1.0	1.67×10^{-24} g	in nucleus

6) The Structure of the Atom (continued)

a) Even before neutrons were discovered, scientists were wondering how electrons and protons were positioned within an atom. This was difficult to determine since atoms are such small particles. A series of discoveries around 1900 provided new methods for probing into the atom.

1. William Roentgen found that very penetrating rays that were not deflected by a magnetic field were emitted when a cathode ray tube was operating. These unknown rays were called X-rays.

2. Henri Becquerel, while studying the nature of x-rays, found that some substances spontaneously emit three kinds of rays (alpha, beta, and gamma). Some substances (such as ZnS) when exposed to radiation will fluoresce (glow – give off light).

b) In 1911, Rutherford used alpha particles (a helium nucleus with a charge twice that of the proton and a mass four times that of the proton) to show that atoms have a nucleus. Most of the alpha particles went straight through the gold foil but about 1 out of 8000 were deflected by more than 90°. Since no individual particle (proton, neutron, or electron) is massive enough to cause such a deflection, Rutherford concluded that all of the protons (and, later to include neutrons) [the two types of particles that account for over 99.9% of the mass of the atom] were concentrated in a small region of the atom, which he called the NUCLEUS. The deflection of the alpha particles was due to their collision with the nucleus of the atom. By the number that were deflected (1 out of 8000), it was obvious that the nucleus occupies only a very small amount of space in an atom. By contrast, the negatively charged electrons occupy most of the volume of the atom.

7) Atomic Number – In 1914 Moseley showed that the wavelength of the x-rays produced by a metal when it was used as the cathode in a cathode ray tube was related to the number of protons in the nucleus of the atom. He also showed that each element had a unique number of protons in its nucleus. The number of protons each element contains is designated by its ATOMIC NUMBER.

ATOMIC NUMBER = number of protons in the nucleus of an atom of that element

8) Mass Number

a) MASS NUMBER = number of protons plus number of neutrons in the nucleus.

b) NEUTRAL ATOMS must have an equal number of protons and electrons.

c) More on chemical symbols.

The chemical symbols are used to tell a lot more about an atom.

i) A number in the lower right gives the number of atoms of that element in a formula.

ii) A number in the lower left tells us the atomic number of the element.

iii) A number in the upper left is the mass number of that atom (the number or protons plus neutrons)

iv) A number in the upper right (with either a + or a –) represents the charge of the ion.

$^{A}_{Z}\mathbf{Ca}^{+2}_{n}$

Ca = symbol for the element
A = mass number (protons + neutrons)
Z = atomic number (protons)
+2 = charge of the ion
n = number of atoms in a formula

<u>Examples</u>:

a) How many protons, neutrons and electrons are in one atom of ^{52}Cr?

b) How many protons, neutrons and electrons are in one atom of ^{238}U?

c) What is the mass and the name of the element that contains 47 protons, 61 neutrons, and 47 electrons?

7) <u>Sec 3.7 – Isotopes</u> – Atoms that have the same number of protons but different numbers of neutrons. (Atoms of the same element with different masses. Remember: It is the number of protons which determines which element we have.)

<u>Example</u> – How many protons, neutrons, and electrons are present in one atom of each of the following isotopes of chlorine? ^{35}Cl and ^{37}Cl

protons:
electrons
neutrons

8) <u>Sec 3.8 – Introduction to the Periodic Table</u>
In all chemistry classrooms and most laboratories you will see a chart called the Periodic Table. It is our map for almost everything we do in Chemistry. As the year progresses we will learn more and more about the Table and we will also learn why some things happen the way they do. For right now, the details are not important. All we need to know are some simple truths about the Table and its elements.

A) Groups and Periods – The Table is organized in columns and rows. The columns are called groups and these contain elements with similar chemical properties. They are often referred to by the number at the top of the column although there are at least two (2) different ways that chemists number the columns.

Four of the groups have special names:
a) Group 1 elements are called the **alkali metals**.
b) Group 2 elements are called the **alkaline earth metals**.
c) Group 17 elements are called the **halogens**.
d) The elements in group 18 are the **noble gases**.

The rows are called periods and we will learn that here is a progression as one moves across a row in the Periodic Table. We can notice at this point that the atomic numbers increase as we move across the period.

B) Representative Elements – The elements in groups 1, 2, 13, 14, 15, 16, 17 and 18 are called the representative elements. These elements behave as we would expect them to if we were to be able to completely predict behavior. These groups are easy to identify because the groups are usually shown to stick up above the others.

C) Transition Elements – The other elements on the main part of the table (groups 3 through 12) are called the transition elements. These behave less predictably than the representative elements (although we will learn more details about their behaviors as time goes on). These are easy to identify because they are "in the valley" as you look at the Periodic Table.

D) Inner-Transition Elements – The elements at the bottom of the Period Table are most often called the inner-transition elements. They behave like the transition elements but their columns do not have numbers when we talk about groups.

E) Alternative Group Numberings – The other most common numbering system for the groups differentiates between the representative and transition elements. The representative elements are given an "A" designation and the representative groups are numbered from 1 - 8 (there are 8 representative groups). The corresponding numbers then become: 1A(1), 2A(2), 3A (13), 4A (14), 5A (15), 6A (16), 7A (17) and 8A (18).

The transition elements are given a "B" designation. The corresponding numbers are: 3B (3), 4B (4), 5B (5), 6B (6), 7B (7) 8B (8, 9, 10), 1B (11) and 2B (12).

F) Metals and Non-Metals – Periodic tables usually have a bold or colored line with a "stair-step" going down groups 13-16. This line separates the metals and the non-metals. The metals are all the elements to the left of this line. The non-metals are all the elements to the right of the line. All elements are either metals or non-metals.

Metals:
* conduct heat
* conduct electricity
* are malleable (can be beaten into thin foils)
* are ductile (can be drawn into wires)

* have a luster (can be polished)

Non-Metals do not do any of the above. They also are much more variable in their properties than the metals.

G) Metalloids – The elements that lie close to the "stair-step" line often show a mixture of metal and non-metal properties. These are often referred to as metalloids or semi-metals. Although they may be confused about their properties and so we have this special name for them, they are ultimately either metals or non-metals.

9) Sec 3.9 – Natural States of the Elements
The majority of the elements on the Periodic Table are solids under standard conditions which most often is considered to be room conditions. On the Periodic Table in our room the symbols for these are colored **BLACK**. Other elements are gases at room conditions and these are RED. There are two elements, Mercury and Bromine, that are liquid at room conditions and these are BLUE.

There are seven (7) elements that appear in nature as two atoms when they are elements. These are referred to as the diatomic elements or diatomic molecules. They are: Hydrogen - H_2, Nitrogen - N_2, Oxygen - O_2, Fluorine - F_2, Chlorine - Cl_2, Bromine - Br_2, and Iodine - I_2. I often refer to these as the "magnificent 7" – in part because, except for hydrogen, they form a 7 as you look at how they are located on the Periodic Table. (This does not mean that they have to be in pairs in compounds but we will get to that.)

10) Section 3.10 – Ions
Atoms have certain numbers of protons and electrons. We have seen how atoms are neutral (protons = electrons). If an atom either loses or gains one or more electrons we call the new thing an ion.

* If an atom loses 1 electron, we designate that as a "+1" ion. Remember: We look at the relationship between the total numbers of protons and electrons, and in this case we now have 1 more of protons than electrons. This leads to a net charge of +1. The same holds true for however many electrons an atom loses.

* If an electron gains 1 electron, we designate that as a "–1" ion. Similar argument.

A positive ion is called a cation and a negative ion is an anion.

We should remember that isolated atoms do not form ions on their own. For an atom to lose one or more electrons there has to be another atom around capable of gaining one or more electrons.

11) Naming Ions
Metals **ONLY** form cations. When we name these ions, we simply use the name of the metal and add the word "ion." Thus, sodium goes to sodium ion.

Non-Metals **ONLY** form anions. When we name these ions from the simple representative atom, we take the root of the name and add an "ide" ending. In this case, chlorine becomes chloride and oxygen becomes oxide.

12) Section 3.11 – Ion Charges and the Periodic Table

Some things are VERY predictable based on the Periodic Table. You do not have to know all the details now (eventually you will) but you can still recognize which ions can be formed for the representative metals and representative non-metals.

Group 1 (1A) metals form **+1** ions.
Group 2 (2A) metals form **+2** ions.
Group 13 (3A) metals form **+3** ions.
Group 15 (5A) non-metals form **–3** ions.
Group 16 (6A) non-metals form **–2** ions.
Group 17 (7A) non-metals form **–1** ions.

Group 18 (8A) noble gases do not form ions.
Group 14 (4A) non-metals (carbon and silicon) do not form ions.
Group 14 (4A) metals usually form **+4** ions.

I have a litany or mantra which describes this fact. You will hear it over and over and over and ... It usually starts out: "Always, absolutely, forever, (there should be) no doubt in your mind, pass this course, ..." If you know that a metal element is in groups 1,2 or 13 you KNOW what ion it forms (+1, +2, or +3). If you have a non-metal in group 15, 16 or 17 you KNOW what ion it forms (-3, -2, or -1). Those are the only ions that those elements form. Position of a representative element on the Periodic Table tells you everything.