Chapter 2 – Chemical Foundations: Elements, Atoms and Ions

Section 2-1: 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.

1) Sec 2-1.1 – Free Elements in Nature]
Sec 2-1.2 – The Names of Elements] – Read these, ask questions.
Sec 2-1.3 – The Distribution of the Elements]

2) Sec 2-1.4 – The Symbols of 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 earlier or alternate Latin names (iron = Fe, from Latin name of ferrum; gold = Au from aurum) even though the modern name is an English name. The majority of the symbols are from the English names (except for Tungsten whose symbol is "W" for the German Wolfram).

Section 2-2: The Composition of Elements: Atomic Theory

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. This is often also described as: matter is neither created nor destriyed in 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. So $CaCO_3$ is always 40.0% Ca, 12.0% C and 48.0% O.

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 2-2.1 – [Dalton's] Atomic Theory

The smallest particle of an element that retains the properties of the element. Greek philosophers (principally Democritus) had proposed that matter is composed of tiny particles that were called *atoms*. Although this idea was not commonly held throughout the centuries, people who followed this way of thinking were often referred to as *atomists*. In 1803, John Dalton proposed a theory of matter based on the ideas of Democritus and scientific discoveries made up to his time.

This is often known as **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.)

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

This also explains why the alchemists could not make gold from lead.

5) 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) CO represents the formula for a compound that contains 1 carbon atom and 1 oxygen atom while CO_2 represents the formula for a compound that contains 1 carbon atom and 2 oxygen atoms. Different formulas, different compounds.

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

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

5) Sec 2-2.2 – The Size of the Atom

Atoms are extremely small. The diameter of a typical atom is about 1×10^{-8} cm. We cannot see these or, perhaps, even comprehend them.

Section 2-3: Composition of the Atom (Electrons, Protons, and Neutrons)

6) Sec 2-3.1 – The Electron and Electrostatic Forces
   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).

7) Sec 2-3.3 – The Particles in the Nucleus

   a) The Proton – Shortly after electrons were discovered, scientists began to think about the particles left over when a hydrogen atom loses an 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.

   b) 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.

8) Sec 2-3.2 – The Nuclear Model of the Atom

   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. The prevailing hypothesis was based on a "plum pudding" model - that is, electrons and protons were randomly distributed within an atom. 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 radiation, found that some substances spontaneously emit three kinds of rays (alpha, beta, and gamma). Some substances (such as ZnS, zinc sulfide) 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

9) The Particles in the Nucleus (continued) [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

Section 2-4: Atomic Number, Mass Number, and Atomic Mass

10) Sec 2-4.1 – Atomic Number, Mass Number, and Isotopes

11) Atomic Number

In 1913 H. G. J. 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

12) 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) 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 of 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

Examples:

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?

12) Isotopes – 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.)

Example – 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

13) Sec 2-4.2 – Isotopic Mass and Atomic Mass

Isotopc Mass is determined by comparison to a standard, ^{12}C, which is defined as having a mass of *exactly* 12 atomic mass units. Therefore, **one atomic mass unit** (amu) is a mass of exactly 1/12 of the mass of ^{12}C.

The **atomic mass** of an element is obtained from the weighted average of the atomic masses of all the isotopes of that element present in nature.

14) Introduction to the Periodic Table

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 there is a progression as one moves across a row in the Periodic Table. We can also 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. Their placement in the modern periodic table was defined in the 1940's as more of these elements were discovered. You do, however, need to recognize where they are situated in a complete or wide-form view of the table.

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).

There generally is no special number description given to the Inner Transition elements.

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. On our classroom periodic table this line is **red**.

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) <u>Metalloids</u> – The elements that lie close to either side of 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 in their chemical behaviors.

15) <u>Natural States of the Elements</u>

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**.

<u>Section 2.5: Molecular Compounds</u>

16) <u>Sec 2-5.1 – Recognizing the Names of Compounds</u>

The names of the simplest compounds are usually based on the elements from which the are composed. Some compounds have common names, such as water, ammonia, etc. In Chapter 4 we will learn about naming compounds.

17) <u>Sec 2-5.2 – Molecules, Molecular Compounds, and Covalent Bonds</u>

a. A **molecule** is formed by the chemical combination of two or more atoms.

b. Molecules composed of different atoms are **molecular compounds**.

c. The atoms in a molecule are joined and held together by a force called a **covalent bond**. In this bond the electrons are shared between the atoms.

Molecules are distinct entities. They are often gases and liquids. If they are solids they have low melting points.

18) <u>Sec 2-5.3 – The Formulas of Molecular Compounds</u>

A compound can be represented by its name or by the symbols of the elements of which it is composed. This latter is called the **formula** of the compound. The formula for water is H_2O. This means that the water molecule contains two atoms of hydrogen and one of oxygen.

Sometimes the same formula can represent two or more distinct compounds that are different by how the atoms are put together. The formula C_2H_6O can describe two different compounds - ethanol and dimethyl ether. Formulas that show the order and arrangement of specific atoms are known as **structural formulas.** We will get much more into this in later chapters.

19) Sec 2-5.4 – Molecular Elements

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 or molecular elements*. 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.)

You will need to recognize these diatomic elements when we talk about them and use them in reactions.

Section 2-6: Ionic Compounds

20) Sec 2-6.1 – Cation and Anions

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*.

Sodium chloride is an ionic compound where the sodium is present as a cation and the chlorine as an anion.

We should remember that isolated atoms do not form ions on their own even though we may sometimes simply talk about them that way. For an atom to lose one or more electrons there has to be another atom or atoms around capable of gaining one or more electrons.

21) Sec 2-6.2 – The Origin of the Charge on Ions

Remember that ions come from atoms that have either lost (cations) or gained (anions) one or more electrons. The charge comes from a comparison between the total number of protons present and the total number of electrons after they have been lost or gained.

22) 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. The position of a representative element on the Periodic Table tells you everything.

23) 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. This easily word-wise distinguishes between the element and the ion.

Na	------->	Na^{+1}	+	1 e
sodium		sodium ion		1 electron
Ca	-------->	Ca^{+2}	+	2 e
calcium		calcium ion		2 electrons

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 chlor*ide* and oxygen becomes ox*ide*.

Cl	+	1 e	------->	Cl^{-1}
chlorine		1 electron		chloride
O	+	2 e	-------->	O^{-2}
oxygen		2 electrons		oxide

24) Sec 2-6.3 – The Formulas of Ionic Compounds

Since cations and anions have opposite charges the compounds are held together by these forces of attraction. Bonding in these compounds is much different than with molecular compounds in which the atoms come together to make distinct molecules.

Compounds consisting of ions are known as **ionic compounds**. The forces that hold the ions together are known as **ionic bonds**. The results are compounds that are almost all hard and rigid, and solid. They will have high melting points.

Because ionic compounds are made of ions, they are just groups of charges attracting each other. An ionic compound can consist of numerous cations and anions, but chemists simplify that be describing the ionic compound as the simplest ratio of the cations and anions present. This simplest, smallest, whole-number ratio of ions in an ionic compound is referred to as a **formula unit**.

This ratio is determined by balancing the charges of the cations and anions. Thus, sodium chloride has sodium ion with a +1 and a chloride ion with a –1 and the formula unit is NaCl. *Note that we do not write the charges when we write the formula unit.* Likewise, calcium chloride has a calcium ion with a +2 charge and chloride with still a –1, so its formula unit becomes $CaCl_2$ (with the 2 for the chlorides being represented as a *subscript*) because the charges have to balance, to cancel. However, calcium oxide will have a formula unit of CaO because calcium ion is still a +2 but the oxide is a –2.

25) <u>Sec 2-6.4 – The Formulas of Compounds Containing Polyatomic Ions</u>

Groups of atoms that are covalently bonded to each other may as a whole have a charge and be cations or anions. These groups are called **polyatomic ions**.
If a formula requires more than one polyatomic ion to balance the charges in a formula unit, *parentheses and subscripts* are used. Thus, sodium nitrate would be $NaNO_2$ while calcium nitrate is $Ca(NO_2)_2$.

We will have to be able to distinguish between molecular compounds and polyatomic ions that have the same general formulas. The difference to remember is that the molecular compound are neutral while the polyatomic ions have a charge.

The charge on the polyatomic ion results from an overall imbalance of electrons. Later we will discuss how this happens and even how to predict charges, etc.

26) <u>Determining Ionic Formulas from Ions</u>

For the representative elements or the polyatomic ions you can know the charges of the ions – either from the periodic table or from learning them. Determining the formulas is then simply a matter of balancing the total pluses with the total minuses in order to have the final formula be neutral. – If it is not neutral, then it is still an ion and you are not finished!

We will learn more about this in Chapter 4.