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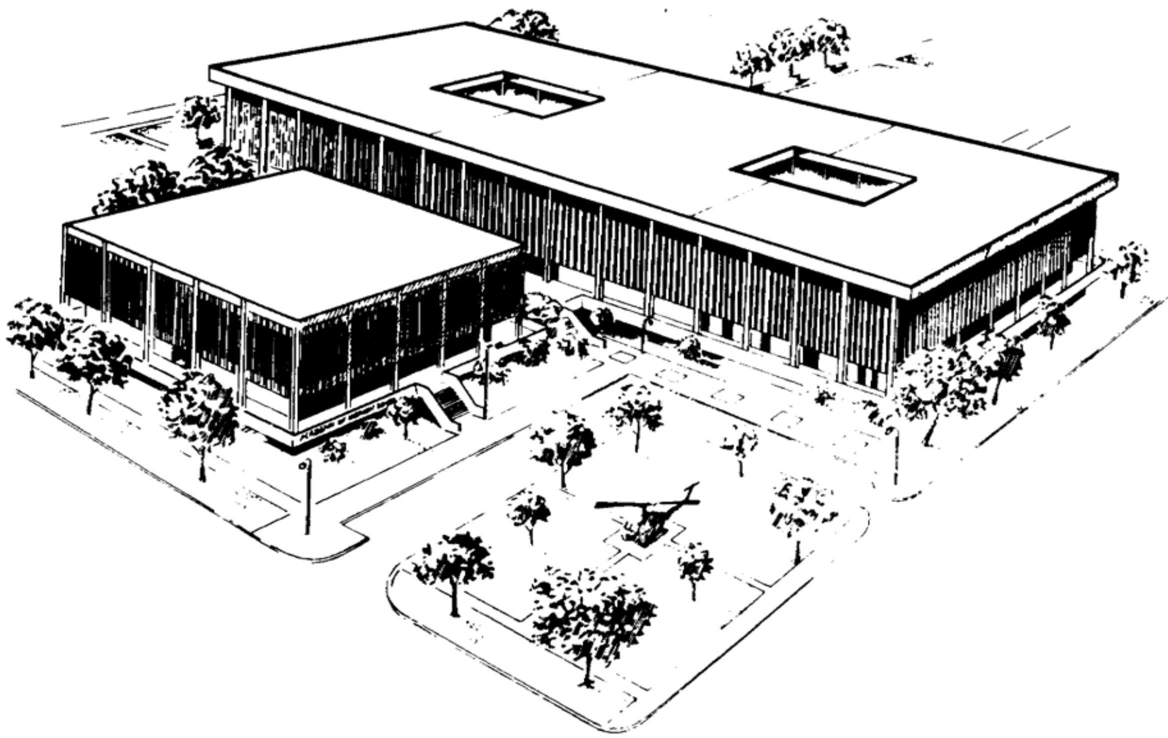

GENERAL CHEMISTRY

SUBCOURSE MD0803 EDITION 100

DEVELOPMENT

This subcourse is approved for resident and correspondence course instruction. It reflects the current thought of the Academy of Health Sciences and conforms to printed Department of the Army doctrine as closely as currently possible. Development and progress render such doctrine continuously subject to change.

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Be sure your social security number is on all correspondence sent to the Academy of Health Sciences.

CLARIFICATION OF TRAINING LITERATURE TERMINOLOGY

When used in this publication, words such as "he," "him," "his," and "men" are intended to include both the masculine and feminine genders, unless specifically stated otherwise or when obvious in context.

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SUBCOURSE MD0803

GENERAL CHEMISTRY

INTRODUCTION

In the process of achieving and maintaining proficiency in your military occupational specialty (MOS), you will be learning concepts and performing tasks that are based on important chemical principles. As you become more proficient with these principles, you may reach the point where you will not need to give them much conscious thought. Meanwhile, however, you should study this subcourse to gain a working knowledge of the fundamental principles of chemistry.

Subcourse Components:

This subcourse consists of 3 lessons. The lessons are:

Lesson 1, Elements of Chemical Structure and Inorganic Nomenclature.

Lesson 2, Elements of Chemical Change.

Lesson 3, Elements of Organic Chemistry.

Credit Awarded:

To receive credit hours, you must be officially enrolled and complete an examination furnished by the Nonresident Instruction Branch at Fort Sam Houston, Texas. Upon successful completion of the examination for this subcourse, you will be awarded 14 credit hours.

You can enroll by going to the web site http://atrrs.army.mil and enrolling under "Self Development" (School Code 555).

A listing of correspondence courses and subcourses available through the Nonresident Instruction Section is found in Chapter 4 of DA Pamphlet 350-59, Army Correspondence Course Program Catalog. The DA PAM is available at the following website: http://www.usapa.army.mil/pdffiles/p350-59.pdf.

LESSON ASSIGNMENT

LESSON 1 Elements of Chemical Structure and Inorganic Nomenclature.

LESSON ASSIGNMENT Paragraphs 1-1 through 1-18 and exercises.

LESSON OBJECTIVES After completing this lesson, you should be able to:

1-1. Define: atom, molecule, element, compound matter, energy, atomic number, atomic weight, electron configuration, isotope, valence octet rule, ion, cation, anion, radical.

1-2. List the three states of matter and the characteristics of each.

1-3. List the three basic particles in an atom and the charge and mass of each.

1-4. State the maximum number of electrons a given electron shell may contain.

1-5. Given a block for an element from the periodic table, write the name of each piece of information which may be obtained about the element.

1-6. Given the name of an element or radical commonly encountered in medicine, state the symbol or formula and common valence(s) for that element or radical.

1-7. List the three types of chemical bonds and state whether the electrons are shared or transferred.

1-8. Given the name of an inorganic compound commonly encountered in medicine, write the chemical formula for the compound.

1-9. Given a chemical formula of an inorganic compound commonly encountered in pharmacy, state the name for that compound.

SUGGESTION After completing the assignment, complete the exercises at the end of this lesson. These exercises will help you to achieve the lesson objectives.

LESSON 1

ELEMENTS OF CHEMICAL STRUCTURE AND INORGANIC NOMENCLATURE

Section I. ELEMENTS OF CHEMICAL STRUCTURE

1-1. INTRODUCTION

Chemistry is the science that studies the composition and changes in composition of the substances around us. Man's natural curiosity about the things and transformations that he observed was the original impetus for the development of this science, but its true beginning was in the work of the alchemists of the Middle Ages. These men searched for a way to change the base metals such as lead into gold. In the large span of time since then, chemistry has developed into a true science and we have amassed a tremendous volume of knowledge. To facilitate the study of chemistry, we can divide it into two divisions: Inorganic chemistry, which deals with the elements and mineral materials, and organic chemistry, which deals with compounds containing carbon. More divisions of chemistry exist, but we will be primarily concerned with these two.

1-2. IMPORTANCE OF CHEMISTRY

Why do we study chemistry? The answer to this question will be obvious when you consider the various classes of compounds we encounter in medicine and in our daily lives. For example, we are concerned with compounds such as drugs and the changes they undergo. Here are some things chemistry will tell us about drugs.

a. **Actions**. Chemistry may tell us about the actions of drugs on the body. Drug effects are determined by the chemical structure of a drug; changes in structure may alter the actions of the drug.

b. **Safety and Storage Procedures**. Special safety or storage precautions may be necessary for particular drugs. These can be identified by the chemical structure.

c. **Incompatibilities.** Sometimes, two or more drugs cannot be mixed because of undesirable consequences. There are three types of incompatibilities:

(1) Chemical. Alterations of chemical properties may occur when two or more drugs are mixed.

(2) Physical. Physical properties of ingredients may produce a mixture unacceptable in appearance or accuracy of dosage.

(3) Therapeutic. When two or more drugs are given to a patient, they may interact in some way to change the effects of one of the drugs.

1-3. MATTER

Matter is anything that occupies space and has weight. If you look around you, you will see matter. The table, books, walls, and your body are all composed of matter. Obviously, the matter around you is not all the same.

a. **Physical States of Matter**. In general, we can group all matter into three groups called states of matter.

(1) Solids. Solids have a definite shape and volume. Examples of solids are books, rocks, pieces of steel, and sand.

(2) Liquids. Liquids have a definite volume but indefinite shape. That is, they take the shape of their container. Water, mercury, alcohol, and oils are liquids.

(3) Gases. Gases have neither a definite shape nor a definite volume. They assume not only the shape of their container, but also the volume of their container. Gases may be expanded or compressed to fit the container in which they are being placed. Therefore, the air in an automobile tire would, if released, expand to fill a large weather balloon.

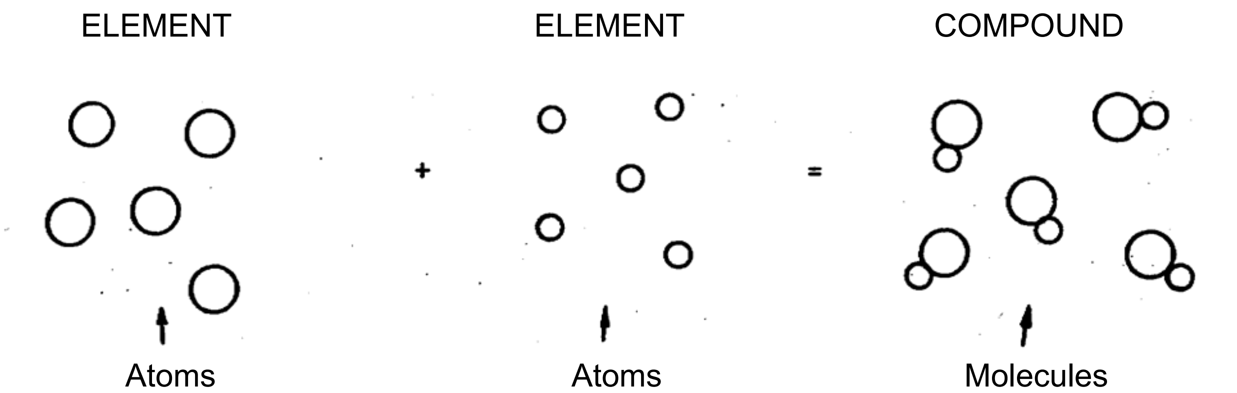
b. **Properties of Matter**. Matter possesses two types of properties, physical and chemical. Characteristics such as smell, color, shape, freezing point, boiling point, and solubility are said to be physical properties of matter. Energy content, reactions with other substances, and chemical reactions due to light, heat, and electricity are said to be chemical properties of matter. From the physical and chemical properties exhibited by a substance, it is possible to isolate, identify, and classify the particular substance.

c. **Classification of Pure Matter**. Matter that cannot be separated into two or more types of matter by physical means is called pure matter. Pure matter consists of two types, elements and compounds.

(1) Elements. Elements are substances that cannot be separated into two or more types of matter by physical or chemical methods. Another way to say this is that elements consist of only one type of atom. An atom is a chemical building block and can be defined as the smallest part of an element that remains unchanged during any chemical reaction and exhibits or displays the chemical properties of that element. Examples of common elements are oxygen, gold, iron, mercury, hydrogen, and carbon. Table 1-1 lists the elements with their symbols, atomic numbers, and atomic weights.

(2) Compounds. Compounds are composed of two or more elements chemically combined. Compounds are substances that have been purified by physical means, but not by chemical methods. They can be separated into two or more types of matter by chemical methods because their basic unit, the molecule, is a combination of

two or more types of atoms. A molecule is composed of two or more atoms and is the smallest part of a compound that can exist and still retain the chemical properties of that compound. Illustrated in Table 1-1 are the relationships of these building blocks and classifications of matter.



	ELEMENT	SYMBOL	ATOMIC NUMBER	ATOMIC WEIGHT
	Actinium	Ac	89	227
*	Aluminum	Al	13	26.9815
	Americium	Am	95	243
	Antimony	Sb	51	121.75
	Argon	Ar	18	39.948
*	Arsenic	As	33	74.9216
	Astatine	At	85	210
*	Barium	Ba	56	137.34
	Berkelium	Bk	97	247
	Beryllium	Be	4	9.0122
*	Bismuth	Bi	83	208.980
*	Boron	B	5	10.811
*	Bromine	Br	35`	79.909
	Cadmium	Cd	48	112.40
*	Calcium	Ca	20	40.08
	Californium	Cf	98	249

* Denotes elements most common to medicine.

Table 1-1. Elements, symbols, atomic numbers, and atomic Weights in alphabetical order (continued).

	ELEMENT	SYMBOL	ATOMIC NUMBER	ATOMIC WEIGHT
*	Carbon	C	6	12.01115
	Cerium	Ce	58	140.12
	Cesium	Cs	55	132.905
*	Chlorine	Cl	17	35.453
	Chromium	Cr	24	51.996
*	Cobalt	Co	27	58.9332
*	Copper	Cu	29	63.54
	Curium	Cm	96	247
	Dysprosium	Dy	66	162.50
	Einsteinium	Es	99	254
	Erbium	Er	68	167.26
	Europium	Eu	63	151.96
	Fermium	Fm	100	253
*	Fluorine	F	9	18.9984
	Francium	Fr	87	223
	Gadolinium	Gd	64	157.25
	Gallium	Ga	31	69.72
	Germanium	Ge	32	72.59
*	Gold	Au	79	196.967
	Hafnium	Hf	72	178.49
	Helium	He	2	4.006
	Holmium	Ho	67	164.930
*	Hydrogen	H	1	1.00797
	Indium	In	49	114.82
*	Iodine	I	53	126.9044
	Iridium	Ir	77	192.2
*	Iron	Fe	26	55.847
	Krypton	Kr	36	83.80
	Kurchatovium	Ku	104	257
	Lanthanum	La	57	138.91
	Lawrencium	Lw	103	257
*	Lead	Pb	82	207.19
*	Lithium	Li	3	6.939
	Lutetium	Lu	71	174.97

* Denotes elements most common to medicine.

Table 1-1. Elements, symbols, atomic numbers, and atomic Weights in alphabetical order (continued).

	ELEMENT	SYMBOL	ATOMIC NUMBER	ATOMIC WEIGHT
*	Magnesium	Mg	12	24.312
*	Manganese	Mn	25	54.9380
	Mendelevium	Md, Mv	101	256
*	Mercury	Hg	80	200.59
	Molybdenum	Mo	42	95.94
	Neodymium	Nd	60	144.24
	Neon	Ne	10	20.183
	Neptunium	Np	93	237
	Nickel	Ni	28	58.71
	Niobium	Nb, Cb	41	92.906
*	Nitrogen	N	7	14.0067
	Nobelium	No	102	254
	Osmium	Os	76	190.2
*	Oxygen	O	8	15.9994
	Palladium	Pd	46	106.4
*	Phosphorus	P	15	30.9738
	Platinum	Pt	78	195.09
	Plutonium	Pu	94	242
	Polonium	Po	84	210
*	Potassium	K	19	39.102
	Praseodymium	Pr	59	140.907
	Promethium	Pm	61	147
	Protactinium	Pa	91	231
*	Radium	Ra	88	226
	Radon	Rn	86	222
	Rhenium	Re	75	186.2
	Rhodium	Rh	45	102.905
	Rubidium	Rb	37	85.47
	Ruthenium	Ru	44	101.07
	Samarium	Sm	62	150.35
	Scandium	Sc	21	44.956
*	Selenium	Se	34	78.96
*	Silicon	Si	14	28.086
*	Silver	Ag	47	107.870

* Denotes elements most common to medicine.

Table 1-1. Elements, symbols, atomic numbers, and atomic Weights in alphabetical order (continued).

ELEMENT	SYMBOL	ATOMIC NUMBER	ATOMIC WEIGHT
* Sodium	Na	11	22.9898
* Strontium	Sr	38	87.62
* Sulfur	S	16	32.064
Tantalum	Ta	73	180.948
Technetium	Tc	43	99
Tellurium	Te	52	127.60
Terbium	Tb	65	158.924
Thallium	Tl	81	204.37
Thorium	Th	90	232.038
Thulium	Tm	69	168.934
Tin	Sn	50	118.69
Titanium	Ti	22	47.90
Tungsten	W	74	183.85
Uranium	U	92	238.03
Vanadium	V	23	50.942
Xenon	Xe	54	131.30
Ytterbium	Yb	70	173.04
Yttrium	Y	39	88.905
* Zinc	Zn	30	65.37
Zirconium	Zr	40	91.22

* Denotes elements most common to medicine.

Table 1-1. Elements, symbols, atomic numbers, and atomic Weights in alphabetical order (concluded).

d. **Classification of Mixed Matter**. Matter that can be separated by physical means is called mixed matter and may be homogeneous or heterogeneous.

(1) Homogeneous mixtures. Mixtures that are uniform throughout are called homogeneous. An example of a homogeneous mixture is a solution of sugar in water. Any small part of this solution would exhibit the same properties as any other small part; therefore, it would be uniform throughout the mixture.

(2) Heterogeneous mixtures. Mixtures that are not uniform are called heterogeneous. An example of a heterogeneous mixture is a mixture of water and oil. If a small sample is taken, it may not be the same as another small sample taken from elsewhere in the mixture. This is because oil and water do not mix well--they give a nonuniform mixture.

1-4. ENERGY

There are many things in our surroundings that we know exist, yet are not matter. They are forms of energy. Heat, light, and electricity are examples of energy. Energy may be simply defined as the ability to do work or overcome resistance.

1-5. ATOMIC STRUCTURE

Early scientists felt that all matter must be built from some basic unit, just as a wall may be constructed from a basic unit, the brick. In trying to find this basic unit, they separated matter by all the methods (chemical and physical) available to them until they could not separate it any further. They felt this separation must result in the building block of matter, which they called the atom (from the Greek word for indivisible). They also observed that the basic units or atoms for various elements differed in their properties, as iron was certainly different from carbon. This led them to try to find the structure of the atom. The difficulty of this problem can be seen when you consider that one cubic centimeter of gold contains as many as 59,000,000,000,000,000,000,000 atoms. The atom is so small that it defies conception. Through ingenious methods, particularly in the last 100 years, we have discovered many facts about this tiny particle, which enables us to understand many of the changes that occur around us.

a. **Atomic Model**. In order for us to picture what an atom looks like, we can use a description with which most people are familiar--the solar system model. In this model, the atom is thought of as a tiny solar system in which there is a central core (like the sun) with other particles traveling in circular paths or orbits (like the planets). While more complex and exact models have been developed, this is the best approximation for general use.

b. **The Nucleus**. The central core from the solar system model is called the nucleus (which is derived from the Latin word <u>nucis</u> meaning nut or kernel). The nucleus contains two types of particles, the proton and the neutron.

(1) <u>The proton</u>. The proton is a particle that has a mass (or weight) of one amu (atomic mass unit) and a positive one (+1) electrical charge. The symbol for the proton is <u>p</u>, <u>p+</u> or <u>H+</u>.

(2) <u>The neutron</u>. The neutron has a mass of one amu (atomic mass unit) but has no electrical charge; that is, it is a neutral particle. In an atom that has more than one proton, the positive charges tend to repel each other. The neutrons serve to bind the protons so that this electrical repulsion does not cause them to fly off into space. The symbol for the neutron is <u>n</u>.

(3) Atomic number and atomic weight. Two important figures commonly used when discussing an atom are its atomic number and its atomic weight.

(a) Atomic number. The atomic number of an atom is equal to the number of protons in the nucleus of the atom. For example, a carbon atom has six protons in its nucleus; therefore, the atomic number of carbon is six.

(b) Atomic weight. The atomic weight of an atom is equal to the number of protons in the nucleus of the atom (one amu each) plus the number of neutrons in the nucleus of the atom (one amu each). Therefore, a carbon atom with six protons and six neutrons has an atomic weight of 12.

c. **The Outer Structure**. The particles that orbit the nucleus (as the planets orbit the sun) are called electrons. These particles have an electrical charge of negative one (-1), but their mass is so small that it is considered to be zero. Actually, the mass of the electron is 1/1837 of the mass of a proton, but the mass, which contributes to the atom is so small that it is not important. The symbol for the electrons is e^- or $^-$.

(1) Electron configuration. Since we may have many electrons going around the nucleus, It might appear that there could be a collision of electrons. Collisions do not occur because the electrons are located in orbits, which are different distances from the nucleus and because of the repulsion between like charges. The number of electrons and their locations are called the electron configuration. This electron configuration is different for each element.

(2) Electron shell. The term electron shell (or energy level) describes where electrons are located (i.e., a specific region around the nucleus). Since electrons can be forced to leave their atoms, the term energy level indicated the amount of energy required to remove the electrons from the various levels or shells. A nucleus can have seven shells, but more chemicals of medicinal importance contain electrons in the first four, which are labeled the K, L, N, and N shells. The K shell is the closest to the nucleus and the N shell is the farthest from the nucleus (figure 1-1). These shells contain different numbers of electrons. The maximum number each shell can hold is equal to $2N^2$, where N is the number of the shell (K=1, L=2, M=3, and so forth.). Thus the maximum number of electrons that each of the first four shells can hold Is:

K = 2(1)2 = 2
L = 2(2)2 = 8
N = 2(3)2 = 18
N = 2(4)2 = 32

Since, for example, the M shell can contain as many as 18 electrons, the possibility for collision might still appear to exist. The reason collisions do not occur is that a shell is subdivided into smaller energy levels, called subshells and orbitals, which we will not need to consider.

(3) Number of electrons. What determines the number of electrons an atom will contain? For an atom to exist freely in nature, it must be electrically neutral (without a charge). There are two particles in an atom that have charges--the proton, which is positive, and the electron, which is negative. For electrical neutrality, the sum of the charges must be zero. In other words, the number of electrons (negative charges) must equal the number of protons (positive charges).

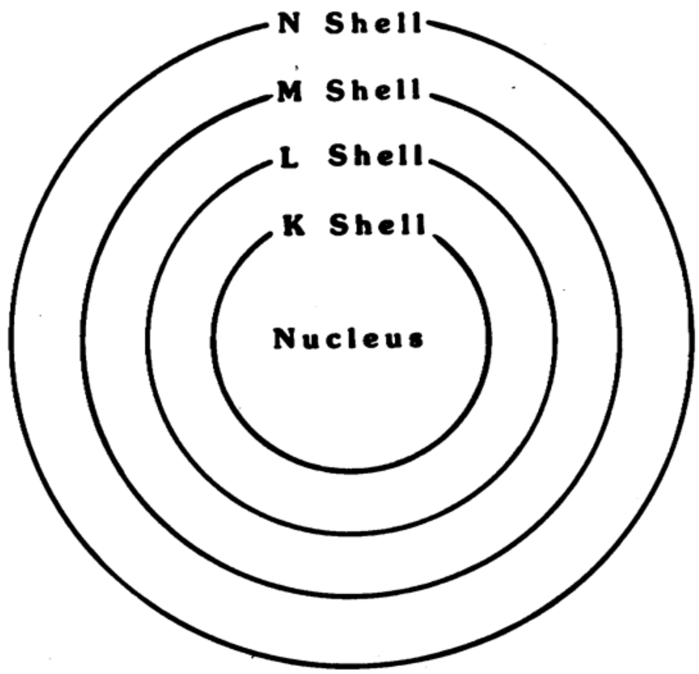


Figure 1-1. First four electron shells.

d. **Atomic Structure of Elements**. As previously stated, each element consists of a single type of atom. Since all atoms consist of the three basic particles we have just discussed (except hydrogen, which usually has no neutrons), the only ways in which elements can differ are atomic number (the number of protons) and atomic weight, (the number of protons and neutrons). There are over 106 different elements which scientists know to have a different atomic number and atomic weight. These elements have a large assortment of properties. Two elements are liquids at room temperature, eleven are gases, and all others are solids.

e. **Periodic Law**. While investigating the properties of the elements, scientists discovered an interesting fact that is now called the periodic law. This law states that the properties of the elements are periodic functions of the atomic number. As the atomic number increases, the properties of the elements repeat themselves at regular Intervals.

f. **Periodic Table**. The periodic law allowed the scientists to group together the elements that had similar properties and form a systematic table of the elements. This table is the periodic table (Table 1-2). The vertical columns are called groups, and the horizontal rows are called periods. This table contains a lot of information that we will not generally use; however, we are concerned with the basic information we can obtain about the elements. Figure 1-2 includes four blocks for elements from the periodic table showing the information, which can be obtained from it. You should note that the number of neutrons is not given in the periodic table. This can be determined by subtracting the atomic number from the atomic weight.

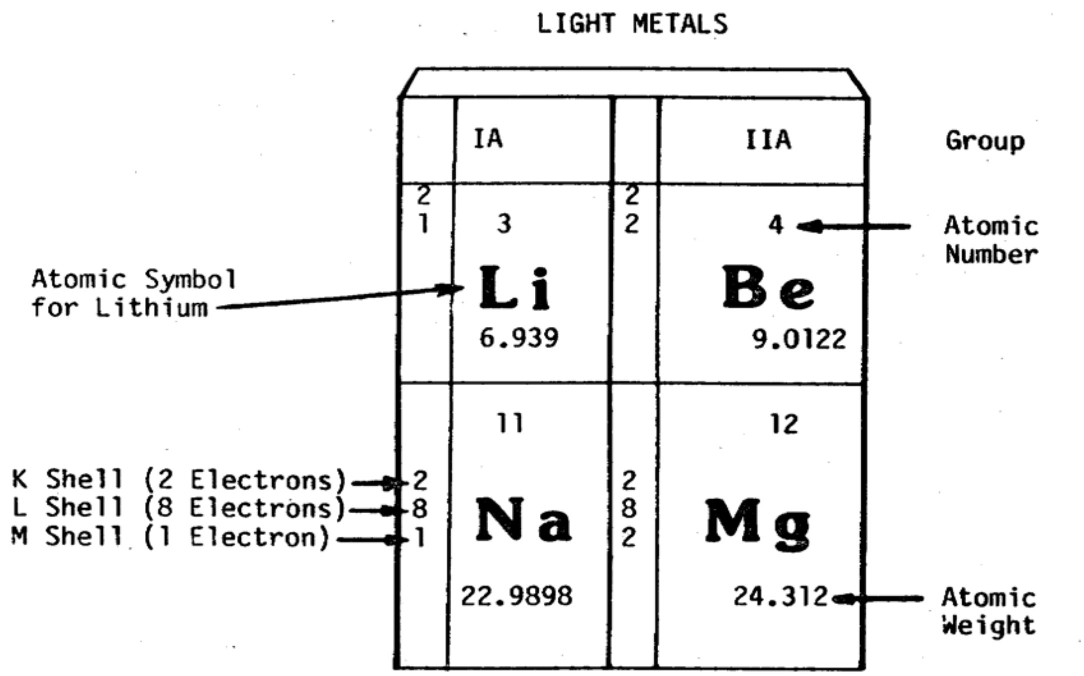


Figure 1-2. Identifying the components of the periodic table.

g. **Isotopes**. All the atoms of a particular element are not identical. Slight variations in the number of neutrons are found to occur naturally. Variations can also be produced in reactors. Atoms that have the same number of protons but a different number of neutrons (same atomic number, but different atomic weights) are called isotopes. Sometimes isotopes are referred to by their mass numbers, H^2, H^3, U^{239}, and so forth. All of the isotopes of a particular element have identical electronic configurations; and since electronic configurations determine chemical properties, isotopes of an element

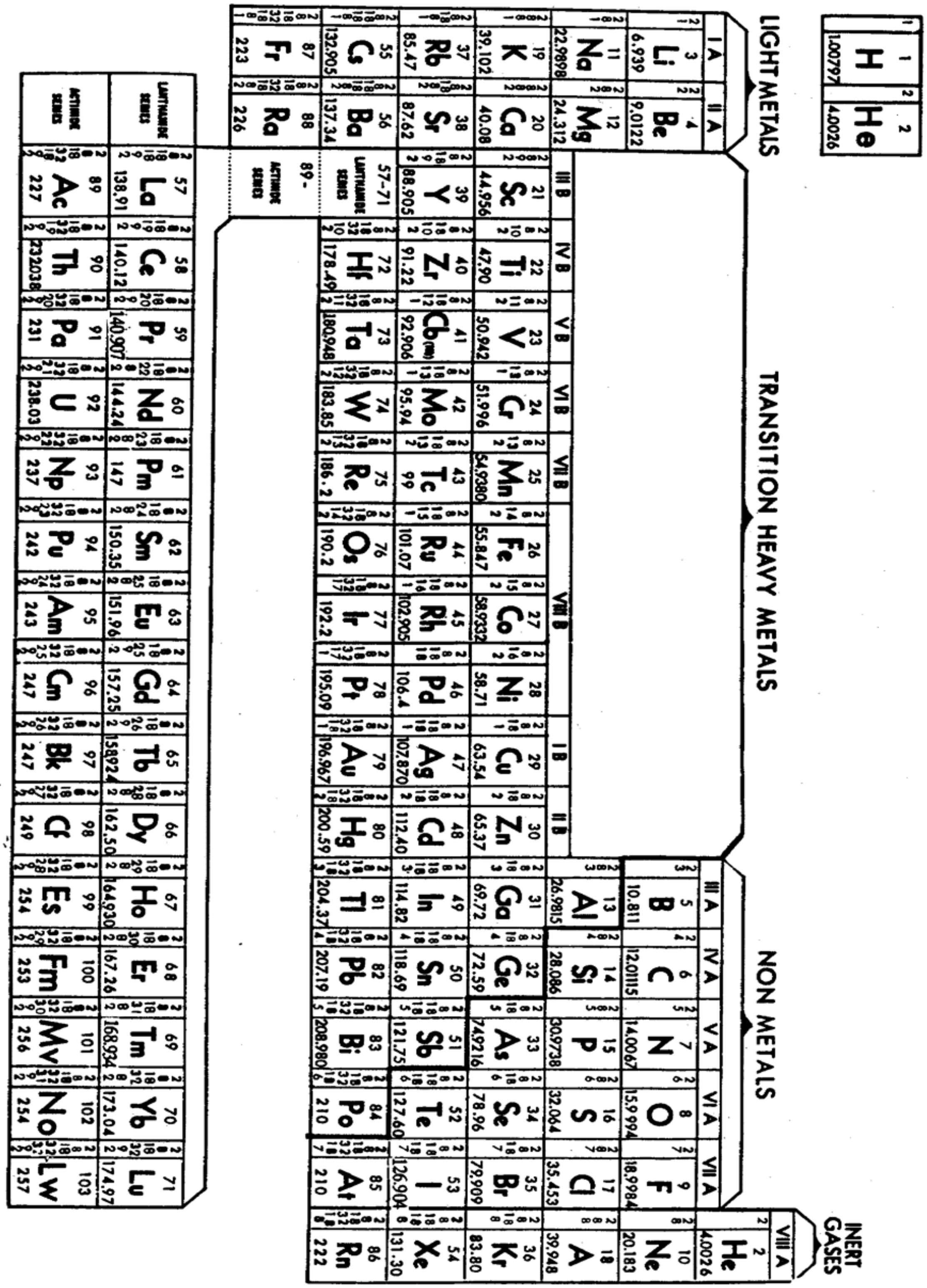


Table 1-2. Periodic table of the elements.

exhibit identical chemical behavior. Induced nuclear reactions can produce both stable and radioactive nuclei. If the nucleus of the atom is unbalanced during the bombardment reaction the atom is called a radioisotope. Radioisotopes, such as cobalt^{66} for treatment of cancer and iodine^{131} for diagnosing of thyroid tumors, are of vital importance in the medical field. The presence of isotopes helps to explain why many atomic weights in the periodic table are not whole numbers since all of the isotopes must be considered when computing the average atomic weight of the element.

1-6. VALENCE AND CHEMICAL BONDING

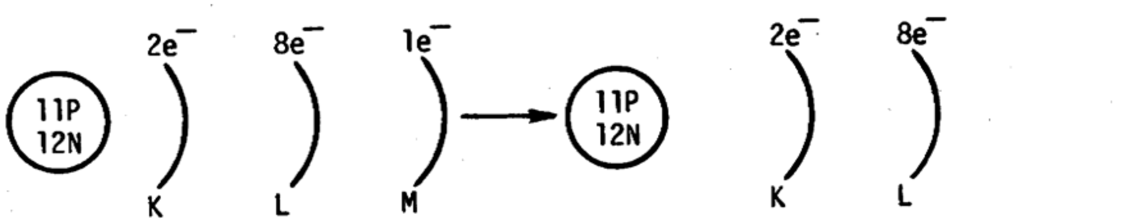
We have now developed the concept that matter was built from a basic unit called the atom, and we have discussed the nature of the atom. We know, however, that very little matter exists as free elements. Most of the things around us are combinations of elements. Logically, the next step is to consider how things combine.

a. **Valence**. The valence of an element can be defined as a measure of its combining power or the number of electrons an atom must gain, lose, or share to have a full or stable outer electron shell. The reason atoms combine is contained in this definition. There are certain electron configurations in nature that are unusually stable (unreactive). The elements that have these configurations are in Group VIII A of the periodic table. They are sometimes referred to as the inert or noble gasses because they are found in very few combinations in nature. Other elements, by gaining, losing, or sharing electrons, can try to make their outer electron shells resemble the shells of the noble gases and hence become very stable. We can see how this works by considering the two simplest elements, hydrogen, and helium. Hydrogen has one electron in the K shell since it has only one proton. Therefore, hydrogen is a very reactive element, occurring naturally in many compounds. Helium, a noble gas, has two electrons in the K shell since it has two protons. Helium is very unreactive. Note that helium, by having two electrons, has a completed outer shell, since the K shell can hold only two electrons. Hydrogen would like to be as stable as helium and could be if it could gain or share one more electron to give it a completed outer shell. Hydrogen seeks this electron in nature by combining with other elements.

b. **Octet Rule**. If you examine the noble (inert) gases (like helium), you will see that not all have a completed (full) electron shell. Except for helium, the noble gases have eight electrons in their outer shell, yet they are still very stable. Chemists have observed that other elements sometimes gain, lose, or share electrons in order to have eight electrons in their outer shell. This observation led to the development of the octet rule, which states that outer electron shells prefer to have eight electrons even though the shell may not be full. (Octet means a group of eight.) On the next page are some examples of the electron configurations for various elements which indicate to us how many electrons they can gain, lose, or share to fit the octet rule or have a completed outer shell.

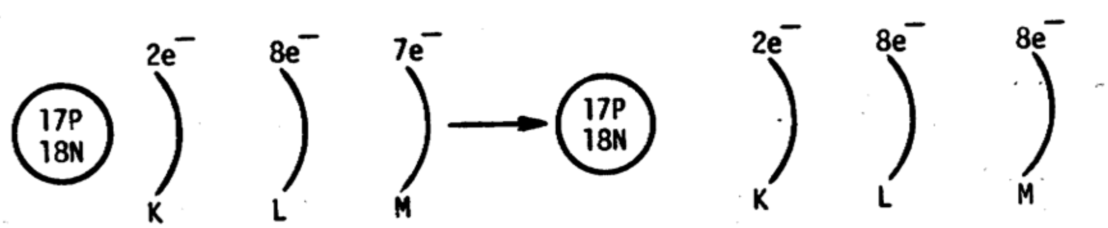
ELEMENT	ATOMIC NUMBER	SHELLS K(2)	L(8)	M(18)	N(32)
H	1	1			
He	2	2			
Li	3	2	1		
Be	4	2	2		
Na	11	2	8	1	
K	19	2	8	8	1

c. **Positive Valence.** An atom that must give up electrons to become stable will have more protons than electrons in its stable configuration and will not be electrically neutral. It will be positively charged since there are more positive charges than negative charges. This is indicated by a + sign. The number of electrons it gives up is the numerical value of its valence. Consider, for example, the element sodium, which has 11 protons and 11 electrons in its free state. It has one electron in the M shell, which it loses easily to become stable. After it loses the electron (that is, gives up a negative charge), it will have a positive one charge and its valence will be +1.



Sodium (free state) Sodium (+1 valence state)

d. **Negative Valence.** An atom that must gain electrons to become stable will have more electrons than protons in its stable configuration and will not be electrically neutral. It will be negatively charged since there are more negative than positive charges. This is indicated by a "-" sign. The number of electrons it gains is the numerical value of its valence. Consider, for example, the element chlorine, which has 17 protons and 17 electrons in its free state. It is one electron short of fitting the octet rule in the M shell as that shell contains 7 electrons. After it gains the electron, it will have a negative one charge and its valence will be -1.



Chlorine (free state) Chlorine (-1 valence state)

e. **Important Symbols and Valences.** Since it is very tedious to continually write complete names for elements, chemists developed the symbols for the elements which you observed on the periodic table. It will not be necessary for you to know all the symbols for your work but a number of them appear frequently enough that they should be memorized. Table 1-3 lists important elements with their symbols and valences. These should be committed to memory. (Please note that most, but not all, valences conform either to the completed shell or octet rules.)

f. **Ions.** Any atom that gains or loses electrons becomes charged (electrical charge) and is called an ion. An ion can be defined as any charged atom or group of atoms. If the ion is positively charged, it is called a cation. If it is negatively charged, it is called an anion. A group of atoms that has a charge and goes through a reaction unchanged is called a radical. Whenever we write the symbol for an element and wish to indicate it is an ion, we write the charge as a superscript to the symbol, for example, Cl^{-1} or Na^{+1}.

g. **Chemical Bonding.** When elements combine to form chemical compounds, the electrons in the outer shell may be transferred from one atom to another or there may be a mutual sharing of the electrons. In either case, a chemical bond is produced. This means the two atoms do not travel or react independently of one another but are held together by the exchange or sharing of the electrons. Both atoms involved in the reaction attain a completed outer orbit, and stability results. There are three types of chemical bonds--electrovalent, covalent, and coordinate covalent.

(1) Electrovalent (ionic) bonding. A transfer of one electron from one atom to another resulting in opposite charges on the two atoms that holds them together by electrostatic (opposite charges attract) attraction is called an electrovalent or ionic bond. A good example of this is the bond formed between a Na (sodium) and a Cl (chlorine) atom.

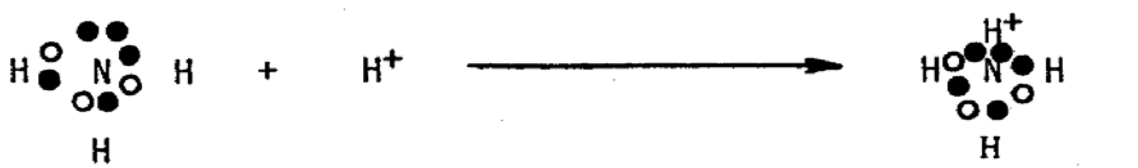
1 e^- in M shell	7 e^- in M Shell	Sodium has a complete outer shell and a charge of +1. Chlorine has met the octet rule in the M shell and has a charge of –1.

(2) Covalent bond. If two atoms each donate an electron that is shared with the other atom, the bond is a covalent bond. An example of this is the bond between two H (hydrogen) atoms. Double and triple covalent bonds are also possible.

Both atoms have 1 e^- in the K shell.

By sharing their electrons each hydrogen has 2 e^- in the K shell and both are stable because of the completed shell.

(3) Coordinate covalent bond. If one atom donates two electrons for sharing with another atom (which donates no electrons), it is called a coordinate covalent bond. An example of this type of bond between N (nitrogen) in ammonia and a hydrogen ion (proton).



Ammonia -N has a complete outer shell

H^+ has an empty K shell

By sharing the two e^-, H+ has a completed K shell and N still has a completed outer shell.

NAME	SYMBOL	VALENCE
Acetate	$C_2H_3O_2$	-1
Aluminum	Al	+3
Ammonium	NH_4	+1
Antimony	Sb	-3, $\underline{+3}$, +5
Arsenic	As	-3, $\underline{+3}$, +5
Barium	Ba	+2
Bicarbonate	HCO_3	-1
Bismuth	Bi	$\underline{+3}$, +5
Bromine	Br	$\underline{-1}$, +1, +3, +5, +7
Calcium	Ca	+2
Carbon	C	+2, $\underline{+4}$, $\underline{-4}$
Carbonate	CO_3	-2
Chlorine	Cl	$\underline{-1}$, +1, +3, +5, +7
Copper	Cu	$\underline{+1}$, $\underline{+2}$
Fluorine	F	-1
Gold	Au	+1, +3
Hydrogen	H	+1
Hydroxide (Hydroxyl)	OH	-1
Iodine	I	$\underline{-1}$, +1, +3, +5, +7
Iron	Fe	$\underline{+2,}$ $\underline{+3}$
Lead	Pb	$\underline{+2}$, $\underline{+4}$
Lithium	Li	+1
Magnesium	Mg	+2
Manganese	Mn	$\underline{+2}$, +3, +4, +6, +7
Mercury	Hg	$\underline{+1,}$ $\underline{+2}$
Nickel	Ni	$\underline{+2}$, +3
Nitrate	NO_3	-1
Nitrogen	N	+1, $\underline{-3}$, +3, +5
Oxygen	0	-2
Permanganate	MnO_4	-1
Phosphate	PO_4	-3
Phosphorus	P	-3, +3, +5
Potassium	K	+1
Silver	Ag	+1
Sodium	Na	+1
Strontium	Sr	+2
Sulfate	SO_4	-2
Sulfur	S	$\underline{-2}$, +2, +4, +6
Zinc	Zn	+2
NOTE: The most common valences are underlined where there may be more than one valence.		

Table 1-3. Valences.

1-7. FORMULA WRITING

a. **Formulas.** Formulas are combinations of symbols that represent a compound. A formula indicates which elements are involved and the number of atoms of each element contained in the compound. In writing formulas, we use subscripts, coefficients, and parentheses in addition to the symbols of the elements. Subscripts indicate the number of atoms of an element, as in H_2 where two is the subscript meaning two hydrogen atoms. If there is no subscript with a symbol, it is assumed there is only one atom of that element. Coefficients, numbers in front of the formula, indicate the number of molecules of compound, as in 4HCl where four is the coefficient indicating four molecules of HCl. Parentheses are used to separate a radical from the rest of the formula when it would be confusing not to do so. In HNO_3, it is not necessary to include parentheses for the NO_3 - radical since there is little chance for confusion. However, we use parentheses for the same radical if it appears NO_3 in a compound such as $Hg(NO_3)_2$ where the 2 indicates that we have two NO_3 - radicals.

b. **Steps in Formula Writing.** In writing formulas for compounds, there are four steps that should be followed.

(1) Determine the symbols for the elements in a compound.

(2) Determine the valence of each of the atoms or radicals.

(3) Write the positive element's symbol first, followed by that of the negative element.

(4) Make the compound electrically neutral by using subscripts.

c. **Example**. Write the formula for calcium chloride.

(1) Calcium = Ca, Chloride = Cl.

(2) Ca valence is +2, Cl valence is -1.

(3) $Ca^{+2}Cl^{-1}$. If we add the charges, we find that this compound is not neutral (+2 - 1 = +1). Therefore, we must proceed to step (4).

(4) To have two negative charges to balance the two positive charges, we must have two Cl^{-1} ions (-1 x 2 = -2). Thus, the formula would be $CaCl_2$.

d. **Rule of Crossing Valences.** A convenient rule for determining what subscripts are necessary in writing formulas is the rule of crossing valences. This rule states that one can take the valence of the element at the left and make it the subscript of the element at the right, and in like manner take the valence of the element at the right and make it the subscript of the element at the left. For example:

Fe^{+3} SO_4^{-2} becomes $Fe_2(SO_4)_3$

Section II. RULES OF INORGANIC NOMENCLATURE

1-8. INTRODUCTION

a. This section discusses how to name a compound from its formula. The interrelationship of names and formulas is very important to you. You will be required to recognize both, in interpreting, preparing, and using these chemicals.

b. This section is in the format of programmed instruction. Each frame presents some material, and then asks some questions in which you apply the material presented. The correct answers follow so that you can check your answers for accuracy. It is important that you use a piece of paper to cover the answers as you work the program. You should fill in the answers as you work each frame and then check your answers. If you answered any questions incorrectly, go back and review the frame so that you understand the correct answer.

1-9. GENERAL TERMS

There are several general terms we use that give us information about inorganic compounds. To describe the number of different elements in a compound we use the terms binary, ternary, and quaternary. A <u>binary</u> compound contains <u>two</u> different elements, such as NaCl. A <u>ternary</u> compound contains <u>three</u> different elements, such as H_2SO_4. A <u>quaternary</u> compound contains four different elements such as $NaHCO_3$.

a. **Questions**.

(1) CO_2 is a ______________________ compound because it contains ________ different elements.

(2) $Al(OH)_2Cl$ is a ______________________ compound because it contains ________ different elements.

(3) KNO_3 is a ______________________ compound because it contains ________ different elements.

b. **Answers**.

(1) Binary, two (C,O).

(2) Quaternary, four (Al,O,H,Cl).

(3) Ternary, three (K,N,O).

1-10. NUMBER PREFIXES

We often use prefixes to denote the number of atoms of an element in a compound. For example, CO contains one oxygen atom and is named carbon monoxide. Mon or mono indicates one atom. Here is a list of the commonly used number prefixes.

		Examples	
Mono, mon	= one	CO	Carbon monoxide
Di	= two	CO_2	Carbon dioxide
Tri	= three	SO_3	Sulfur trioxide
Tetra	= four		
Penta	= five		
Hexa	=six		
Hepta	= seven		
Octa	= eight		
Nona	= nine		
Deca	= ten		

a. **Questions**.

(1) NCl_3 is named nitrogen _chloride.

(2) SO_2 is named sulfur ______________________ oxide.

(3) CF_4 is names carbon ______________________ flouride.

b. **Answers**.

(1) Tri.

(2) Di.

(3) Tetra.

1-11. NAMING METALLIC CATIONS

Many metallic elements have only one possible valence. The names for the cations formed by these metals are given the name of the element. For example, Na^{+1} is called sodium ion; Ca^{+2} is called calcium ion. Other metallic elements, however, may have more than one valence. Since valence is a measure of combining power, these elements may form more than one compound with the same anion. Therefore, we must have some way to differentiate between the varying valences when we name them. There are two common methods for doing this.

a. The first method uses a root word from the name of the element (or the Latin name for the element) with a suffix to indicate the valence state. The suffix --ous indicates the <u>lower</u> valence; the suffix --<u>ic</u> indicates the <u>higher</u> valence. For example, Hg^{+1} is called mercurous ion, but Hg^{+2} is called mercuric ion.

(1) <u>Questions</u>. (You may wish to refer to table 1-3.)

(a) $Al+^3$ is called ______________________ ion.

(b) $Fe+^2$ is called ferr-- ______________________ ion.

(c) $Fe+^3$ is called ferr-- ______________________ ion.

(d) $K+^1$ is called ______________________ ion.

(e) $Cu+^1$ is called cupr- ______________________ ion.

(f) $Cu+^2$ is called cupr- ______________________ ion.

(g) $Ba+^2$ is called ______________________ ion.

(2) <u>Answers.</u>

(a) Aluminum.

(b) --ous.

(c) --ic.

(d) Potassium.

(e) --ous.

(f) --ic.

(g) Barium.

b. The second method for naming metallic cations uses the name of the element followed by a roman numeral in parentheses to indicate the valence. For example, Cu^{+1} is written as copper (I) and Cu^{+2} is written as copper (II). Remember, these methods for specifying valence need be used only when there is more than one valence possible.

(1) Questions.

(a) Fe^{+2} is written _____ion.

(b) Fe^{+3} is written _____ion.

(c) Mg^{+2} is wrjtten _____ion.

(d) Hg^{+1} is written _____ion.

(e) Ag^{+1} is written _____ion.

(f) Pb^{+4} is written _____ion.

(2) Answers.

(a) Iron (II) (ferrous).

(b) Iron (III) (ferric).

(c) Magnesium.

(d) Mercury (I) (mercurous).

(e) Silver.

(f) Lead (IV) (plumbic).

1-12. NAMING ANIONS

There are generally two types of anions. Many anions are elemental; that is they are made of only one atom of one element. Others are composed of groups of atoms of one or more elements that pass through a reaction unchanged in most cases. This latter group of anions is called radicals. We will concern ourselves first with the naming of elemental or monatomic anions.

a. The names of the elemental anions are made by adding the --ide suffix to the root of the element's name. Thus anions formed by chlorine (Cl^{-1}) are called chloride ion; anions formed by oxygen (O^{-2}) are called oxide ion.

(1) Questions.

(a) Br^{-1} is called ______________________ion.

(b) S^{-2} is called ______________________ ion.

(c) H^{-1} is called _ion.

(d) N^{-3} is called _ ion.

(2) Answers.

(a) Bromide.

(b) Sulfide.

(c) Hydride.

(d) Nitride.

b. The most common type of anionic radicals consists of a central atom covalently bonded to a number of atoms of oxygen. Monovalent anionic radicals (Valence = -1) normally contain three oxygen atoms; radicals with negative valences greater than one normally contain four oxygen atoms. The names for these normal types of radicals are formed from the root for the name of the central atom plus the suffix -ate. Thus, ClO_3^{-1} is named chlorate and SO_4^{-2} is named sulfate. It is important to note that these generalizations have exceptions. The best way to remember the names and formulas for the radicals is to memorize the common ones. Most of these are listed in this subcourse.

(1) Sometimes a central atom may be bonded to a different number of oxygen atoms than normal; in other words, a series of radicals may be formed with the same central atom. Different suffixes and prefixes are used to name these different radicals. When there is one less oxygen atom than normal, the suffix -ite is used. The name for ClO_2 -1 is chlorite; SO_3 -2 is called sulfite.

(2) Occasionally, there are other radicals in a series. This is especially true of the halides (fluoride, chloride, bromide, and iodide ions). If there are two less oxygen atoms than usual, the -ite suffix is used with the prefix hypo-. For example, ClO^{-1} is called hypochlorite. If there is one more oxygen atom than normal, the -ate suffix is used in combination with the prefix per-, so ClO_4^{-1} is named perchlorate.

(3) A chart summarizing the use of the prefixes and the series of radicals formed by chlorine as examples follows:

PREFIX	SUFFIX	NAME OF ION	RADICAL
hypo-	--ite	hypochlorite	ClO^{-1}
	--ite	chlorite	ClO_2^{-1}
	--ate	chlorate	ClO_3^{-1}
per-	--ate	perchlorate	ClO_4^{-1}

(a) Questions.

1 IO_3^{-1} is called ___________________ ion.

2 IO_2^{-1} is called ___________________ ion.

3 IO_4^{-1} is called ___________________ ion.

4 PO_4^{-3} is called __________________ ion.

5 PO_3^{-3} is called __________________ ion.

6 NO_3^{-1} is called __________________ ion.

7 CO_3^{-2} is called __________________ ion.

(b) Answers.

1 Iodate.

2 Iodite.

3 Periodate.

4 Phosphate.

5 Phosphite.

6. Nitrate.

7 Carbonate. (Be sure to learn the exceptions!)

c. There are several significant exceptions to the rules for the naming of anionic radicals. The most important is the previously mentioned carbonate radical (CO_3^{-2}). Several others bear mentioning because you are likely to see them in medicine.

(1) Certain radicals are derived when hydrogen is removed from an acid to form a charge group of atoms (radical). If one hydrogen is removed, the radical gets the prefix "bi." This indicates that <u>one</u> hydrogen is missing.

(a) Example 1: When carbonic acid (H_2CO_3) gives up one hydrogen ion, it loses a positively charged hydrogen atom. It becomes a radical HCO_3^{-1} and is assigned the name bicarbonate. "Bi" indicates one hydrogen was removed.

(b) Example 2: $H_2PO_4^{-1}$ is called the <u>bi</u>phosphate radical because it was derived from phosphoric acid (H_3PO_4) by removing <u>one</u> hydrogen atom.

(2) Several radicals do not follow any of the above rules. Their names and formulas must be memorized. Some of the most common are hydroxide (OH^{-1}), peroxide (O_2^{-2}), and thiosulfate ($S_2O_3^{-2}$).

(3) Occasionally, metals with valences higher than +1 will form salts that contain oxide or hydroxide ion. When these occur in the middle of the formula, they are referred to as either <u>oxy</u>- or <u>hydroxy</u>-, respectively. Number prefixes are used to denote the number of them.

1-13. NAMING SALTS

A salt is an ionic compound containing some cation other than hydrogen and some anion other than hydroxide and oxide. Since the compound must be electrically neutral, the total positive valence (from all of the cations) must equal the total negative valence (from all the anions). This gives us a method for determining the valence of any particular ion in the formula. The names for salts are made by writing the name of the cation followed by the name of the anion. For example, $CaCl_2$ has calcium as the cation and chloride as the anion, so the compound is called calcium chloride. $FeSO_4$ has sulfate as the anion, but we need to know whether the ion is ferrous ion or ferric ion. This is easy for us to do: since we know the total negative valence (from sulfate) is -2, the total positive valence (for iron) must be +2; therefore, it is ferrous ion. The compound is ferrous sulfate.

a. **Questions**.

(1) KBr is __________ .

(2) $Mg(NO_3)_2$ is ______ .

(3) $BaSO_4$ is _______ .

(4) BiOCl is ________ .

(5) $HgCl_2$ is ________ .

(6) $CuSO_4$ is ________ .

(7) $Al(OH)_2Cl$ is _____ .

(8) $NaHCO_3$ is ______ .

(9) $PbSO_4$ is ______ .

(10) $KBrO_3$ is ______ .

b. **Answers**.

(1) Potassium bromide.

(2) Magnesium nitrate.

(3) Barium sulfate.

(4) Bismuth oxychloride.

(5) Mercuric chloride (mercury (II) chloride).

(6) Cupric sulfate (copper (II) sulfate).

(7) Aluminum dihydroxychloride.

(8) Sodium bicarbonate (sodium hydrogen carbonate).

(9) Plumbous sulfate (lead (II) sulfate).

(10) Potassium bromate.

1-14. NAMING BINARY ACIDS

All acids have hydrogen as the only cation. Binary acids are those acids that are composed of only two elements; that is, they consist of hydrogen in combination with some elemental anion. Usually the anion is a halide (F, Cl, Br, I), but binary acids with other anions also occur.

a. The names for the binary acids are formed by using the prefix hydro-, the root name for the anion, and the suffix -ic, followed by the word “acjd.” For example, HCl is called hydrochloric acid.

b. An exception to this rule is hydrocyanic acid which has the formula HCN. Although this is a ternary acid, the cyanide radical (CN^{-1}) is usually treated like a halide ion when naming its compounds.

c. The binary acids are really covalent compounds which act as acids only when they are in solution, especially in water. When you know that one of the binary acids is by itself, you can properly name it in a similar manner to the salts; thus, HCl as a pure gas would be called hydrogen chloride.

(1) Questions.

(a) HBr is called ______________________________________ .

(b) HI is called ____ .

(c) H_2S is called ______________________________________ .

(d) HF gas is called ____________________________________ .

(2) Answers.

(a) Hydrobromic acid.

(b) Hydriodic acid.

(c) Hydrosulfuric acid.

(d) Hydrogen fluoride.

1-15. NAMING TERNARY ACIDS

a. The ternary acids generally are made of hydrogen ion combined with one of the radicals that contain oxygen. For this reason, they are often referred to as “oxyacids.”

b. When naming the ternary acids, the suffixes on the names of the radicals are changed and followed by the word "acid" to show the presence of the hydrogen. Radicals ending in -ate change their suffix to -ic; radicals ending in -ite change their suffix to -ous. The prefixes, if there are any, are not changed. Occasionally, an extra syllable is added in the middle of the name for pronunciation purposes--these do not follow any pattern and must be learned. Here are some examples of naming ternary acids from the radicals:

RADICAL	NAME OF RADICAL	ACID	NAME OF ACID
SO_4^{-2}	Sulfate	H_2SO_4	Sulfuric acid
SO_3^{-2}	Sulfite	H_2SO_3	Sulfurous acid
ClO^{-1}	Hypochlorite	$HClO$	Hypochlorous acid

(1) Questions.

(a) HNO_3 is called ____ .

(b) HNO_2 is called ____ .

(c) $HClO_4$ is called ____ .

(d) H_2CO_3 is called ____ .

(e) H_3PO_3 is called ____ .

(f) H_3PO_4 is called ____ .

(2) Answers.

(a) Nitric acid.

(b) Nitrous acid.

(c) Perchloric acid.

(d) Carbonic acid.

(e) Phosphorous acid.

(f) Phosphoric acid.

1-16. NAMING BASES

a. The most common bases are those included by the Classical Theory of Acids and Bases; that is, they are hydroxyl ion (OH^{-1}) donors. Thus most of the bases are composed of the hydroxyl radical combined with a metallic cation.

b. The names for these bases are made by writing the name of the cation followed by "hydroxide." It is not normally necessary to use number prefixes because the valence of the cation tells us the number of hydroxyl radicals in each molecule. You can see that this method of naming bases is very similar to the method used for naming salts, except that the anion is always hydroxide. For example, NaOH is called sodium hydroxide and $Ca(OH)_2$ is called calcium hydroxide.

(1) Questions.

(a) KOH is called ______________________________________ .

(b) $Mg(OH)_2$ is called ______________________________________ .

(c) $Fe(OH)_2$ is called ______________________________________ .

(d) $Al(OH)_3$ is called ______________________________________ .

(e) $Fe(OH)_3$ is called ______________________________________ .

(2) Answers.

(a) Potassium hydroxide.

(b) Magnesium hydroxide.

(c) Ferrous hydroxide.

(d) Aluminum hydroxide.

(e) Ferric hydroxide.

1-17. NAMING COVALENT INORGANIC COMPOUNDS

There are a number of inorganic compounds that are bonded into molecules by covalent bonds. Most of these are the oxides, sulfides, and halides of the nonmetallic elements.

a. Generally, these compounds are named by writing the name of the central atom (usually the first one in the formula) followed by the name of the anion formed by the other element. Number prefixes are used when necessary to avoid confusion between different compounds formed by the same elements. Here are some examples:

COMPOUND	NAME OF COMPOUND
H_2S gas (see para 1-14a(3))	Hydrogen sulfide
CO	Carbon monoxide
CO_2	Carbon dioxide

b. There are two very important exceptions to this which you have probably already seen. These are water (H_2O) and ammonia (NH_3). Both of these have common names, which are firmly established in the nomenclature, property of these two compounds which makes them different from almost all others is the ability to readily accept a proton (coordinate covalent bond with a hydrogen cation) to form cations. Thus water becomes hydronium ion (H_3O+^1); ammonia becomes ammonium ion (NH_4+^1) very easily in the right conditions.

(1) Questions.

(a) SO_2 is called ____ .

(b) SO_3 is called ____ .

(c) CCl_4 is called ____ .

(d) NI_3 is called ____ .

(e) CS_2 is called ____ .

(f) NH_3 is called ____ .

(g) NH_4^{+1} is called ____ .

(h) NH_4Cl is called ____ .

(2) Answers.

(a) Sulfur dioxide.

(b) Sulfur trioxide.

(c) Carbon tetrachloride.

(d) Nitrogen triiodide.

(e) Carbon disulfide.

(f) Ammonia.

(9) Ammonium.

(h) Ammonium chloride.

1-18. WATERS OF HYDRATION

Many times when a substance crystallizes into a solid, molecules of water are included in the crystal. These molecules of water combine with the substance in a fixed ratio, similar to the fixed ratios between the atoms in a molecule. Whenever weighing or doing calculations based on compounds that have waters of hydration, the amount of water in the crystals must be taken into consideration.

a. When writing formulas for these compounds, the waters of hydration are shown by placing a dot (or dash) after the formula for the compound, followed by the formula for water with a coefficient to indicate the number of waters of hydration. For example, cupric sulfate forms crystals that contain five molecules of water for each molecule of cupric sulfate--its formula is written $CuSO_4.5H_2O$.

b. Compounds that contain waters of hydration are called <u>hydrates</u>. (If all the water has been removed by drying, they are called <u>anhydrous.</u>) When writing the names for these compounds, the number of waters of hydration is indicated by using number prefixes. Thus, the name for $CuSO_4.5H_2O$ is cupric sulfate pentahydrate. Another number prefix seen occasionally in the names of hydrates is <u>hemi</u>-, which means one-half (1/2).

(1) <u>Questions</u>.

(a) $AlCl_3.6H_2O$ is called ____ .

(b) $Mg_3(PO_4)_2.5H_2O$ is called ____________________________________ .

(c) $Na_2HPO_4.7H_2O$ is called ________________________________ .

(d) $FeSO_4.7H_2O$ is called ____ .

(e) $Na_2CO_3.1OH_2O$ is called ___________________________________ .

(f) $CaSO_4.1/2H_2O$ is called ____________________________________ .

(2) <u>Answers</u>.

(a) Aluminum chloride hexahydrate.

(b) Magnesium phosphate pentahydrate.

(c) Disodium hydrogen phosphate heptahydrate.

(d) Ferrous sulfate heptahydrate.

(e) Sodium carbonate decahydrate.

(f) Calcium sulfate hemihydrate (two molecules of calcium sulfate for each molecule of water).

Continue with Exercises

EXERCISES, LESSON 1

INSTRUCTIONS. Write the word, words, symbols, or numbers that properly completes the statement in the space provided or mark the correct word/phrase from those given. After you complete the exercises, turn to Solutions to Exercises and check your answers. Reread the material referenced for each exercise answered incorrectly.

1. An atom is a c__________________ building block. An atom is the __________________est part of an (element) (compound) that remains unchanged during any __________________ reaction and exhibits the (chemical) (physical) properties of that (element) (compound).

2. A compound is a combination of two or more types of __________________ .

3. An element is a substance that (can)(cannot) be separated into two or more types of matter by __________________ or __________________ methods.

4. A compound is a substance that has been __________________ by physical methods, but not by __________________ methods.

5. Matter is anything which occupies ______________ and has ______________ .

6. Energy is the ability to do __________________ . Examples of energy are e__________________y, h ________________ , and l ________________ .

7. The three physical states of matter are s _______________ , l _____________ , and g _______________ . A s __________________ has a definite shape and a/an __________________ volume. A liquid has a/an __________________ shape and a/an __________________ volume. A gas has a/an __________________ shape and volume.

8. Examples of physical properties are s _______________ l, c _______________ , s _______________ e, f _______________ p _______________ , b _______________ p _______________ , and s _______________ y.

9. An example of a chemical property is ________________ content. To observe a chemical property, it is necessary that a c ________________ r ________________ occur. Such reactions may be due to l ________________ , h ________________, or e ________________ .

10. The three basic parts of an atom are the p ________________ , e ________________ , and n ________________ . A p ________________ has a mass of ________________ and a charge of (-1) (0) (+1). A n ________________ has a mass of ________________ and a charge of (-1) (0) (+1). An e ________________ has a mass, rounded to the nearest unit of ________________ and a charge of ________________ .

11. The atomic number of an element is the number of ________________ in each atom of the element. In an electrically neutral atom, it is also the number of ________________ in that atom.

12. The atomic weight of an element is the number of ________________ and ________________ in each atom of the element. An atom containing six protons and six neutrons has an atomic weight of ________________ .

13. The term electron configuration refers to the ________________ and ________________ of electrons in the atoms of an element.

14. Isotopes of an element have in their atoms the same number of ________________ but different numbers of ________________ .

15. The maximum number of electrons in the K shell is ________________ , in the L shell is ________________ , in the M shell is ________________ , and in the N shell is ________________ .

16. Below is a block from the periodic table.

	IV A
2 4	6 **C** 12.01115

The number 12.01115 is the atomic _________________ of the element. The numbers ${}^{2}{}_{4}$ represent the e _______________ c _______________ of the element. There are (two) (four) electrons in the K shell and (two) (four) electrons in the L shell. The number 6 in the block is the atomic ____________ of the element. Each carbon atom has (six)(twelve) protons. The letter C is the ____________ for the element carbon. IV A is the ____________ .

17. If an element's atomic number is 18 and its atomic weight is 40, the number of neutrons in each atom is _______________ .

18. The valence of an element is a measure of its c _______________ power. Valence is the number of e _______________ that an atom must g _______________ , l _________ , or sh _______________ to have a full or stable outer electron _______________ .

19. An ion is any _______________ atom or group of _______________ . It has g _______________ or l _______________ at least one electron.

20. A cation is an ion with a (positive) (negative) charge.

21. An anion is an ion with a (positive) (negative) charge.

22. A radical is a charged _______________ of atoms that goes through many reactions without being _______________ .

23. According to the octet rule, the outer electron shell of an atom “prefers” to have _______________ electrons.

24. What are the symbols the following?

a. Barium: ______ , + __

b. Iron: _____ e, + _____ , + ______

c. Sulfate: ___ ____4, ___ ___

d. Phosphorus:_____, +_____, _____5

e. Hydrogen: ______, +______

f. Potassium: ______, ______1

g. Oxygen: ______ ______

h. Copper: ______ u, +1, + ______

i. Bromine: _____ _____, _____ 1

j. Mercury: _____ g, +_____, _____2

k. Iodine: _____, _____ _____

l. Sulfur: _____, _____ _____

m. Silver: _____ _____, +_____

n. Calcium: _____ a, _____ 2

o. Nitrate: N _____ , - _____

p. Aluminum: A _____ , + _____

q. Chlorine: _____ _____, _____ _____

r. Zinc: _____ _____, _____ _____

25. The three types of chemical bonds are e__________ (i _____________), c_____________, and c__________ c_____________ .

26. In __________________ (I _____________) bonding, the electrons are not shared. They are donated by (one element) (both elements).

27. In c___________ bonding, the electrons (are) (are not) shared. They are donated by (one element) (both elements).

28. In c_________ c________ bonding, the electrons (are) (are not) shared. They are donated by (one element) (both elements).

29. Listed below are chemical symbols for parts of 12 different molecules. First, label each part with its valence. Then, write the formula of the molecule with the proper subscripts to make it electrically neutral.

PARTS FORMULA

a. H, Cl ______________________

b. H, SO4 ______________________

c. Na, Br ______________________

d. H, NO3 ______________________

e. Ca, Cl ______________________

f. Na, Cl ______________________

g. Mg, CO3 ______________________

h. Ca, NO3 ______________________

i. NH4, SO4 ______________________

j. K, PO4 ______________________

k. Al, SO4 ______________________

l. Zn, PO4 ______________________

30. Listed below are the names of 10 compounds. First, write down the symbols for each part of the compound. Then, label each part with its valence. Add subscripts to make each molecule electrically neutral. The result is the formula for the compound.

NAME	PARTS	FORMULA
a. Calcium bromide	_________	_________
b. Sodium carbonate	_________	_________
c. Aluminum hydroxide	_________	_________
d. Calcium hydroxide	_________	_________
e. Barium hydroxide	_________	_________
f. Potassium bromide	_________	_________
g. Silver chloride	_________	_________
h. Magnesium phosphate	_________	_________
i. Sulfuric acid	_________	_________
j. Sodium sulfide	_________	_________

31. The number of atoms of oxygen in H_2SO_4 is. ______. The number of atoms of chlorine in NaCl is ___________. The number of atoms of iron in $Fe_2(SO_4)_3$ is _______________. The number of atoms of sulfur in K_2S is ________. The number of atoms of oxygen in $Al_2(SO_4)_3$ is _____________. The number of atoms of hydrogen in $(NH_4)_2SO_4$ is ____________.

32. Listed below are the names of 10 compounds. First, write down the symbols for each part of the compound. Then, label each part with its valence. Add subscripts to make each molecule electrically neutral. The result is the formula for the compound.

NAME	PARTS	FORMULA
a. Sulfur dioxide	______________	______________
b. Mercurous oxide	______________	______________
c. Hydrobromic acid	______________	______________
d. Mercury (1) chloride	______________	______________
e. Potassium bicarbonate	_____________	______________
f. Ammonium iodide	______________	______________
g. Aluminum oxychloride	______________	______________
h. Nitrous acid	______________	______________
i. Potassium permanganate	___________	______________
j. Magnesium nitrate hexahydrate	____________	______________

33. Listed below are the formulas for eight different compounds. Write the name for each one.

FORMULA	NAME
a. $Fe(HCO_3)_3$	fer ____ ___ carbon ___
b. $MgCl_2 . 6H_2O$	________ ium chlor ___ _______ hydrate
c. HI	___dri _____acid
d. KOH	___ t ____ ium ___ dro ______
e. K_2HPO_4	___ tassium _____ogen phos______
f. $FeCO_3$	fer _____ _____ bon _____
g. $Ca(OH)NO_3$	cal ____ ____ dro _____ nit _____
h. NaOH	sod _____ hydro ______

34. Listed below are the formulas for 12 different compounds. Write the name for each one.

FORMULA	NAME
a. HNO_3	n ___________ a ___________
b. FePO4	f ____________ p ___________
c. Al(OH)2Cl	a ____________ d ___________
d. (NH4)2SO3	a ____________ s ___________
e. Hg3PO4	_____________ p ___________
f. NaHCO3	s _____________ b ___________
g. NCl	h _____________ a ___________
h. MgO	m _____________ o ___________
i. Ba(OH)2	b _____________ h ___________
j. CaHPO3	c ________ h ________ p ________
k. CaCO3	c ___________ c ___________
l. KCl	p ___________ c ___________

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 1

1. chemical,
   small(est), element, chemical, chemical, element (para 1-3c(1))

2. elements (para 1-3c(2))

3. cannot, physical, chemical (para 1-3c(1))

4. purified, chemical (para 1-3c(2))

5. space, weight (para 1-3)

6. work, electricity, heat, light (para 1-4)

7. solid, liquid, gas
   solid, definite
   indefinite; definite
   indefinite, (para 1-3a)

8. smell, color; shape; freezing point; boiling point; solubility (para 1-3b)

9. energy
   chemical reaction
   light; heat; electricity (para 1-3b)

10. proton; electron, neutron
    proton; one; +1
    neutron; one; 0
    electron; zero; -1 (para 1-5b,c)

11. protons, (para 1-5b(3)(a))
    electrons (para 1-5c(3))

12. protons, neutrons
    12 (para 1-5b(3)(b))

13. Number, locations (para 1-5c(1))

14. Protons; neutrons (para 1-5g)

15. 2, 8, 18, 32 (para 1-5c(2))

16. weight
electron configuration
two, four,
number
six
symbol
group (para 1-5; fig 1-2)

17. 22 (para 1-5f)

18. combining
electrons; gain, loose, share, shell (para 1-6a)

19. charged atoms
gained, lost (para 1-6f)

20. positive (para 1-6f)

21. negative (para 1-6f)

22. group, changed (para 1-6f)

23. eight (para 1-6b)

24. a Ba: +2
b Fe; +2, +3
c SO_4: -2
d P +3 +5
e H +1
f K +1
g O; -2
h Cu +1 +2
i Br -1
j Hg +1 +2
k I; -1
l S: -2
m Ag; +1
n Ca; +2
o NO_3; -1
p Al; +3
q Cl; -1
r Zn; +2 (para 1-6e; Table 1-3)

25. electrovalent, ionic; covalent, coordinate covalent (para 1-6g)

26. electrovalent (ionic); one element (para 1-6g(1))

27. covalent; are, both elements (para 1-6g(2))

28. coordinate covalent, are; one element (para 1-6g(3))

29. a H^{+} , Cl^{-} ; HCl
    b H^{+} , SO_4^{-2}; H_2SO_4
    c Na^{+}, Br^{-} ; NaBr
    d H^{+}, NO_3^{-} ; HNO_3
    e Ca^{+2} , Cl^{-} ; $CaCl_2$
    f Na^{+} , Cl^{-} ; NaCl
    g Mg^{+2} , CO_3^{-2} ; $MgCO_3$
    h Ca^{+2} , NO_3^{-1} ; $Ca(NO_3)_2$
    i NH_4 + , SO_4^{-2}; $(NH_4)_2 SO_4$
    j K^{+} , PO_4^{-3} , K_3PO_4
    k Al^{+3} , SO_4^{-2} ; $Al_2(SO_4)3$
    l Zn^{+2} , PO_4 -3; $ZN_3(PO_4)_2$ (para 1-7; Table 1-3)

30 Formula
    a $CaBr_2$
    b Na_2CO_3
    c $Al (OH)_3$
    d $Ca(OH)_2$
    e $Ba(OH)_2$
    f KBr
    g AgCl
    h Mg3(PO4)2
    i H2SO4
    j Na2S (paras 1-7, 1-12, 1-13, 1-15, Table 1-3)

31. 4
    1
    2
    1
    12
    8 (para 1-7a)

32. Formula
   a SO2 (Table 1-3, paras 1-10, 1-17)
   b Hg2O (Table 1-3; para 1-11a)
   c HBr (Table 1-3; para 1-14)
   d HgCl (Table 1-3; para 1-11b)
   e KHCO3 (Table 1-3; para 1-12c(1))
   f NH4I (Table 1-3; Para 1-17b)
   g AlOCl (Table 1-3; para 1-12c(3))
   h HNO2 (Table 1-3; para 1-15)
   i KMnO4 (Table 1-3; para 1-12b)
   j Mg(NO3)2.6H2O (Table 1-3; paras 1-12b, 1-18)

33. Name
   a ferric bicarbonate (Table 1-3; paras 1-11, 1-13)
   b magnesium chloride, hexahydrate (Table 1-3; paras 1-10, 1-18)
   c hydriodic acid (para 1-14a(2))
   d potassium hydroxide (para 1-16a(1))
   e potassium monohydrogen phosphate or dipotassium phosphate (paras 1-12b(1), 1-13)
   f ferrous carbonate (para 1-13)
   g calcium hydroxynitrate (paras 1-12c(3), 1-13)
   h sodium hydroxide (para 1-16)

34. Name
   a nitric acid (para 1-15b(1)(a)
   b ferric phosphate (paras 1-12, 1-13)
   c aluminum dihydroxy-chloride (para 1-12c(3))
   d ammonium sulfite (paras 1-12b, 1-17)
   e mercurous phosphate (paras 1-11, 1-12, 1-13)
   f sodium bicarbonate (para 1-13a(8))
   g hydrochloric acid (para 1-14)
   h magnesium oxide (para 1-17)
   i barium hydroxide (para 1-16)
   j calcium monohydrogen phosphite (paras 1-12, 1-13)
   k calcium carbonate (para 1-13)
   l potassium chloride (para 1-13)

End of Lesson 1

LESSON ASSIGNMENT

LESSON 2 Elements of Chemical Change.

LESSON ASSIGNMENT Paragraphs 2-1 through 2-13 and exercises.

LESSON OBJECTIVES After completing this lesson, you should be able to:

2-1. Given a chemical equation, describe the chemical events occurring in the reaction, to include names of reactants and products and any special conditions indicated.

2-2. Given a description of a chemical reaction, write and balance the equation for the reaction.

2-3. Define equilibrium exothermic, endothermic, milligram molecular weight, and milliequivalent weight.

2-4. Given the name of an inorganic compound commonly encountered in medicine, calculate the milligram formula weight and milliequivalent weight of that compound.

2-5. Define oxidation and reduction.

2-6. Given an inorganic chemical reaction, identify the oxidizing and reducing agents.

2-7 Define an acid and a base according to the classical theory and according to the Bronsted-Lowry theory.

2-8. Given a chemical formula, indicate whether it is an acid or a base.

2-9. List three properties of bases and five properties of acids.

2-10. State the antidotes for external or internal contact with a strong acid or a strong base.

2-11. Correctly define the term salt.

2-12. Given a chemical formula, indicate whether it is a salt, acid, or base.

2-13. Given a pH value, indicate whether it is acidic, basic, or neutral.

2-14. List three methods for measuring pH.

2-15. State the function and two general components of a buffer system.

2-16. From a list of pairs of compounds, select which represents a buffer system.

2-17. State four important properties of water and two major methods of water purification.

2-18. Define solute, solvent, solubility, dissociation, and electrolyte.

SUGGESTION

After completing the assignment, complete the exercises at the end of this lesson. These exercises will help you to achieve the lesson objectives.

LESSON 2

ELEMENTS OF CHEMICAL CHANGE

2-1. CHEMICAL REACTIONS

As a provider of health care, you will not be required in most cases, to write and balance chemical equations. You will, however, be using and/or seeing the effects of chemical reactions on a daily basis. Chemical reactions are frequently used to explain various concepts of pharmacology and physiology. Consider drugs. All drugs are chemicals and any pharmacological reference you consult will refer to the chemical changes drugs undergo in the body. Consequently, it is essential that you have a basic knowledge of what a chemical reaction involves and how that chemical reaction can be expressed as a chemical equation.

a. **Definite Composition**. When atoms combine, they do so in definite ratios of intact atoms to produce compounds with definite composition. Note that this combination is by <u>number</u> of atoms, not by weights of atoms. What the individual atoms happen to weigh is not important. Atoms do not know what they weigh. When they do interact and combine, it is always as whole particles, and the particle-to-particle or atom-to-atom ratio can always be expressed in simple, whole numbers. Chemical changes do not split atoms into fractional pieces. This is the reason we are able to write a formula such as HCl for the compound hydrochloric acid. Hydrochloric acid is always formed from one atom of hydrogen and one atom of chlorine. Since a chemical reaction is merely a change in matter, and matter consists of atoms or molecules, we can discuss chemical reactions by talking about interactions of individual molecules or atoms.

b. **Chemical Equations**. In discussing a chemical reaction, it would be very cumbersome to write it out in the same manner as we state it verbally. To get around this problem, chemists have developed chemical equations. Chemical equations are abbreviated ways of writing chemical reactions. They save much writing and effort and give at least as much information as a verbally stated reaction. Chemical equations show:

(1) The kinds of atoms or molecules reacting.

(2) The products formed.

(3) The number of atoms entering the reaction.

(4) The number of molecules formed in the product.

(5) The proportion in which the substances react to give definite products.

c. **Chemical Symbols.** In writing chemical equations, we use a number of symbols. The most common symbols are shown below with their meanings.

<u>SYMBOL</u>	<u>MEANING</u>
Δ	Heat (a form of energy)
$\rightarrow$	"yields," indicates direction of reaction
$\uparrow$	given off as a gas
$\downarrow$	given off as a precipitate

As we illustrate several types of reactions, the uses of these symbols will become apparent.

d. **Types of Reactions.** There are four types of chemical reactions, which are possible: Combination reactions, decomposition reactions, single replacement reactions, and double replacement reactions.

(1) <u>Combination reactions</u>. A combination reaction can be represented by the chemical equation A + B --> AB (one atom of A plus one atom of B yield one molecule of AB). A specific example of this type of reaction is the combination of a metal with oxygen to yield a metallic oxide.

$$2\ Mg + O_2 \text{ --> } 2\ MgO$$

This equation tells us that two atoms of magnesium and one molecule of oxygen react to form two molecules of magnesium oxide.

(2) <u>Decomposition reactions</u>. The general equation representing decomposition reactions is $AB \rightarrow A + B$. Here is a good example:

$$CaCO_3 \underset{\Delta}{\rightarrow} CaO + CO_2 \uparrow$$

This equation tells us that calcium carbonate will yield calcium oxide and carbon dioxide. The Δ also tells us that this reaction occurs when heat is applied to calcium carbonate. The $\uparrow$ indicates that the carbon dioxide is given off as a gas.

(3) Single replacement reactions. The general equation for a single replacement reaction is $A + BC \rightarrow AC + B$. An example is:

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

This equation tells us that one atom of zinc and one molecule of cupric sulfate yield one molecule of zinc sulfate and one atom of copper.

(4) Double replacement reactions. The most commonly occurring reaction is the double replacement reaction. The general equation for this reaction is $AB + CO \rightarrow AD + CB$. Double replacement reactions can be further subdivided into several classes. The most common of these classes are the precipitation reaction, the acid-base reaction, and the oxidation-reduction reaction. An example of the precipitation reaction is:

$$BaCl_2 + Na_2SO_4 \rightarrow 2\ NaCl + BaSO_4$$

This equation tells us that one molecule of barium chloride reacts with one molecule of sodium sulfate to yield two molecules of sodium chloride and one molecule of barium sulfate as a precipitate. Acid-base and oxidation-reduction reactions will be covered later.

2-2. WRITING CHEMICAL EQUATIONS

At this point, you have seen several examples of chemical equations and should be familiar with the symbols used in an equation. We will now examine the process of writing an equation when we are given a verbal description of the reaction. One general rule that must be kept in mind is that there will always be the same number and kinds of atoms in the products of a reaction as in the reactants. This is because matter can neither be created nor destroyed in a chemical reaction and atoms always combine in certain proportions. When given a written verbal description of a chemical reaction, the following steps are used to write the equation for the reaction.

a. Write the symbols for all elements involved.

b. Write the correct formulas for any compounds and check for diatomic molecules. (Some elements never exist as single atoms but only as diatomic molecules. These elements can be identified from their names, which end in -gen or -ine. The common diatomic molecules are hydrogen (H_2), nitrogen (N_2), oxygen (O_2), chlorine (Cl_2), fluorine (F_2), and bromine (Br_2).)

c. Balance the equation by placing coefficients where appropriate. Remember that there must be equal numbers of atoms of each kind on both sides of the equation. In this step, the subscripts that were used in writing the correct formulas cannot be changed.

2-3. EXAMPLE

For application of these steps, consider this description of a reaction. Calcium metal and water react to yield calcium hydroxide and hydrogen gas.

a. Write the symbols for all elements involved.

Ca, O, H

b. Write the correct formulas for any compounds and check for diatomic molecules.

$$Ca + H_2O \rightarrow Ca(OH)_2 + H_2 \uparrow$$

c. Balance the equation by placing coefficients where appropriate. Look at the number of atoms of each element in the products and reactants.

REACTANTS	PRODUCTS
1 Ca	1 Ca
1 O	2 O
2 H	4 H

It is apparent here that there are twice as many oxygen and hydrogen atoms in the products as reactants. How can this equation be balanced to give equal numbers of atoms on both sides? Fill in the coefficients of the molecules in the equation below.

$$\_\_\_\_ \ Ca \ + \ \_\_\_\_\_ \ H_2O \rightarrow \_\_\_\_\_\_\_ \ Ca(OH)_2 \ + \ \_\_\_\_\_\_\_\_ \ H_2 \uparrow$$

Since there are twice as many hydrogen and oxygen atoms on the right as on the left, if we could double the numbers of these atoms on the left, we would have a balanced equation. This can be done by placing a two in front of H_2O. All the other coefficients would be one (if there is no coefficient, we assume it is one, so there is no need to write it in front of each molecule).

2-4. EQUILIBRIUM REACTIONS

We have implied that all reactions only go in the direction of the products, but this is not always the case. Sometimes as products are formed, they react with one another or decompose to form the reactants. Thus, the reaction is going in both directions at the same time, and if allowed to continue indefinitely, would result in a constant amount of products and reactants. Reactions that go in both directions are called equilibrium reactions, and when the rate of formation of product is the same as the rate of formation of reactant, they are said to be in equilibrium. In writing an equation, we indicate equilibrium by drawing arrows pointing in opposite directions

$$\left(\overrightarrow{\quad} \right) \left(\overleftarrow{\quad} \right).$$

As an example of an equilibrium reaction, consider the dissociation of a compound into ions:

$$Na_2CO_3 \rightleftharpoons 2Na^+ + CO_3^{-2}$$

Sodium carbonate in solution dissociates into sodium ions and carbonate ions. Some of the ions come back together to form sodium carbonate. Thus, an equilibrium is established.

2-5. EXTERNAL CONDITIONS AFFECTING CHEMICAL REACTIONS

External conditions that affect reactions are usually types of energy that are put into a reaction, such as heat or light. Chemical reactions are always accompanied by an energy change. Either energy is released or it is acquired. When the amount of energy is changed, so is the amount of matter. This is called the Law of Conservation of Matter and Energy. However, ordinary chemical reactions involve such small matter changes that they go undetected and may be ignored.

a. **Heat**. Generally, heat is the form of energy we are most concerns us most. It may affect a reaction in one of two ways.

(1) Exothermic reactions. If a reaction gives off heat, it is called an exothermic reaction. External heat, if supplied to this type of reaction, will slow down the rate of reaction.

(2) Endothermic reactions. If a reaction takes in heat, it is an endothermic reaction. If heat is added to an endothermic reaction, the rate of reaction will increase. This may be of value in the preparation of medicinal products.

b. **Light**. Light is a form of energy that may cause many chemicals to decompose. For this reason, it is necessary to protect some drugs from contact with light by placing them in dark-colored or opaque containers. These containers prevent most or all of the outside light from coming into contact with the drug.

2-6. REACTING QUANTITIES

It has already been emphasized that all reactions occur on an atom-to-atom level. This presents a small problem to us, since we cannot hold an atom in our hand, or count out a specific number of atoms to put into a reaction. How then do we measure amounts of material that will react together? Chemists long ago solved this problem by learning how to count particles indirectly. They did this by measuring samples of the chemicals in particular ratios by their weights. To understand the means of doing this, we need to expand our concept of atomic weight to compounds in the form of the formula (or molecular) weight.

a. **Milligram Formula (Milligram Molecular) Weight**. When atoms combine to form compounds, the atomic nuclei are not affected. There is no net loss of weight. Regardless of whether the particle formed is a molecule or an ion group, it will have a formula and a formula weight. The formula weight of a compound is the sum of the atomic weights of all the atoms that appear in its chemical formula. Consider, for example, carbon dioxide:

Atoms: $C + O + O = CO_2$ (molecule)

Atomic weights: 12 + 16 + 16 = 44 (formula weight)

While we have arrived at a formula weight which is in terms of atomic mass units, it is much more useful to express it in terms of milligrams. This is known as the milligram formula weight. For the example above, CO_2, the milligram formula weight is 44 mg. This is a quantity that we can measure and see, and thus can easily work with. It also represents a reacting unit of the compound.

b. **Molarity.** A molar solution, or a one molar (1M) solution, consists of one-gram molecular weight (GMW) of solute dissolved in enough water to make 1 liter of finished solution. Molarity, then, is the number of GMWs dissolved in enough water to make a finished solution of 1000 ml. Molar solutions may have as a solute a solid, a liquid, or a gas. Later in this subcourse, we will use the concept of molarity to explain the measurement of acidity, called the pH.

(1) Calculating the gram molecular weight. One-gram molecular weight of a substance is its molecular weight expressed in grams. Thus, a GMW of NaOH would be 40 grams, where the atomic weights are as follows: Na = 23, O = 16, and H = 1. Thus, .5 GMW of NaOH would be 20 grams, and so forth. A mole is one-gram molecular weight of a substance. Thus, a mole of NaOH is 40 grams of NaOH; a half-mole (.5 mole) is 20 grams; two moles of NaOH are 80 grams, and so on.

(2) Calculating the molarity of a solution. To find the molarity of a solution, we divide the number of gram molecular weights of solute by the number of liters of total solution. The formula may be written:

$$\text{Molarity} = \frac{\text{no. of GMWs of solute}}{\text{no. of liters of solution}}$$

Since many problems are stated in terms of the weight of solute and require you to determine the number of gram molecular weights (moles), the following formula will be of benefit:

$$\text{No. of GMWs} = \frac{\text{weight of solute}}{\text{GMW}}$$

(3) Example. What is the molarity of a solution containing 29.25 grams of sodium chloride in 500 ml. of total solution?

Step 1. Find the number of GMWs.

GMW of NaCl = 58.4 grams

$$\text{No. of GMWs} = \frac{\text{weight of solute}}{\text{GMW}}$$

$$\text{No. of GMWs} = \frac{29.25}{58.4} = 0.5$$

Step 2. Find the molarity.

$$\text{Molarity} = \frac{\text{no. of GMWs of solute}}{\text{no. of liters of solution}}$$

500 ml = 0.5 liter

$$\text{Molarity} = \frac{0.5}{0.5} = \text{1 molar or 1M}$$

c. **Milligram Equivalent Weight (Milliequivalent Weight)**. Sometimes we are interested in more than just the weight ratios of reacting compounds. Since the valence of an element is a measure of that element's combining power, the valences in a compound should be indicative of their reactivity. Therefore, chemists have modified the milligram formula weight to include the positive or negative valence of a compound. This value is called the milligram equivalent weight and is defined as the milligram molecular weight divided by the total positive or negative valence. Consider, for example, sodium hydroxide:

Milligram molecular weight = 40 mg
Total positive valence = 1
Milligram equivalent weight = $\frac{40 \text{ mg}}{1}$ = 40 mg

Another example is potassium phosphate (K_3PO_4):

Milligram molecular weight = 212 mg
Total positive valence = 3
Milligram equivalent weight = $\frac{212}{3}$ mg = 70.7 mg

In a reaction, one milliequivalent (mEq) weight of one compound will react with one milliequivalent weight of another. If we are reacting two compounds, then, we can determine how much of each compound should be used to obtain a desired amount of product.

2-7. OXIDATION-REDUCTION REACTIONS

Previously, we have examined the processes involved in writing, balancing, and interpreting reactions and looked at examples of several types of reactions. One type of reaction we did not examine closely was the oxidation-reduction reaction (sometimes called redox reaction). Even though this type of reaction is very important in the chemistry of drug molecules, it is beyond the scope of our instruction to study them in detail. However, a basic understanding of this process will be valuable to you in understanding many of the incompatibilities, storage problems, and some disease states that you will encounter later.

a. **Review of Valence**. Before these reactions are studied, valence should be reviewed briefly. The following two valence concepts are especially important in oxidation-reduction reactions:

(1) All elements in their free and uncombined state are considered to have a valence of zero. This holds even for those elements that are diatomic molecules in their free state.

(2) All atoms can exist in a number of valence states. The common valences which you learned previously are the preferred and most stable valences under normal conditions, but other valences can and do occur.

(3) These two concepts are important because oxidation-reduction reactions always involve a change in the valence numbers of some of the elements involved in the reaction.

b. **Oxidation**. Oxidation, in inorganic chemistry, is defined as the loss of electrons or an increase in the valence of an element. Consider, for example, the oxidation of elemental iron:

$$Fe^{O} - 2e^{-} \longrightarrow Fe^{+2}$$

Iron in its free state has a valence of zero and is very reactive since its common valence state is +2 or +3. It loses two electrons to become the ferrous ion. The valence has gone from 0 to +2, thus iron has been oxidized. It can undergo further oxidation to the +3 valence state:

$$Fe^{+2} \text{ -le}^{-} \text{ ---> } Fe^{+3}$$

Here the ferrous ion has lost another electron to become a ferric ion.

c. **Reduction**. In inorganic chemistry, reduction is defined as the gain of electrons or a decrease in the valence of an element. Consider the reduction of elemental oxygen:

$$O_2 + 4e^{-} \text{ -----> } 2\ O^{-2}$$

Observe that oxygen is a diatomic molecule in its free elemental form and has a valence of zero. Since the most common valence state of oxygen is -2, oxygen accepts electrons readily to become the oxygen anion. The valence of each oxygen atom has gone from 0 to -2, thus oxygen had been reduced. If the valence is made smaller (reduced), reduction has occurred.

d. **Oxidizing and Reducing Agents**. For all practical purposes, it is impossible to simply add or subtract electrons from an element except in an electrolytic cell. In fact, the oxidation of one element and the reduction of another always occur simultaneously. One element loses the electrons; the other element gains the electrons that are lost by the first. Consider these two reactions when they are combined:

$$2Fe - 4e^{-} \text{ -----> } 2Fe^{+2}$$

$$O_2 + 4e^{-} \text{ ------> } 2O^{-2}$$

$$2Fe + O_2 \text{ ------> } 2FeO$$

This is an oxidation-reduction reaction that is very common in our industrialized society. The oxidation of iron by atmospheric oxygen gives us iron oxide, commonly known as rust. In this reaction, oxygen was reduced, going from a zero to a -2 state by receiving electrons from iron. Because it accepted the electrons from iron and allowed the iron to oxidize, oxygen is called an oxidizing agent. Iron, which gave up electrons, is called the reducing agent. General characteristics of reducing and oxidizing are shown in the following table.

REDUCING AGENT	OXIDIZING AGENT
(1) Gives up electrons	(1) Gains electrons
(2) Oxidized during reaction	(2) Reduced during reaction
(3) Unusually low valence state compared to most common state	(3) Unusually high valence state compared to most common state

2-8. ACIDS AND BASES

The two most important classifications of compounds in inorganic chemistry are acids and bases. The following discussion forms the groundwork for understanding some of the most important chemical changes which you will encounter.

a. **Classical Acid-Base Theory**. Svante Arrhenius, in 1887, published the first satisfactory explanation of the acid-base phenomena that had been observed by chemists.

(1) Acids. Arrhenius defined an acid as a compound that donates protons (H^+) in solution. Examples would be any of the compounds you learned to name as acids earlier in this subcourse.

$$HCl \xrightarrow{HOH} H^+ + Cl^-$$

$$H_2SO_4 \xrightarrow{HOH} H^+ + HSO_4^-$$

NOTE: The HOH (H_2O), which indicates that water is the solvent in these reactions. Both HCl and H_2SO_4 contribute protons in solution.

(2) Bases. Arrhenius defined a base as any compound that donates hydroxyl (OH^-) ions in solution. Again, you should be familiar with several examples from your nomenclature studies.

$$NaOH \xrightarrow{HOH} Na^+ + OH^-$$

$$KOH \xrightarrow{HOH} K^+ + OH^-$$

(3) Discussion. These classical definitions are based on the dissociation of the compounds into ions in solution. This implies that all acids and bases must contain exchangeable hydrogen and hydroxyl ions, respectively, in their

formulas. This theory did explain the majority of the compounds known at the time, but there were some exceptions. Chemists knew, for example, that metal oxides (MgO, CaO, etc.) dissolved in water exhibited base-like properties. Also, ammonia (NH_3) in solution exhibited the properties of a base. The attempts to explain these exceptions led to new definitions of acids and bases.

b. **Modern Acid-Base Theory**. In 1923, Bronsted and Lowry, two chemists in different countries, independently derived new definitions of acids and bases to explain the exceptions to the classical theory. The new theory they developed was named, appropriately, the Bronsted-Lowry theory. This theory differs from the classical theory in that the dissociation of water is considered as well as the dissociation of the compound.

(1) Dissociation of water. Even though we often think of water as merely being an inert solvent, it does dissociate into ions.

$$H_2O \rightleftharpoons H^+ + OH^-$$

This is an equilibrium type reaction as indicated by the double arrow. Actually, very few ions exist at any time since they rapidly recombine to form molecular water. If we put numbers in this reaction, there are 500 million molecules of water for each hydrogen or hydroxyl ion.

(2) Bronsted-Lowry acid. By the Bronsted-Lowry theory, an acid is any compound (charged or uncharged) capable of donating a proton. This is essentially the same as the classical definition.

(3) Bronsted-Lowry base. The real value of the Bronsted-Lowry theory is in the definition of a base. A base is defined as a charged or uncharged substance capable of accepting a proton. Generally, the proton a base accepts comes from the dissociation of water.

(a) Consider, for example, ammonia dissolved in water:

$$NH_3 + H_2O \rightleftharpoons NH_3 + H^+ + OH^- \rightleftharpoons NH_4^+ + OH^-$$

By accepting a proton from water, ammonia has effectively increased the concentration of hydroxyl ions in the solution. This would account for the properties like those of a classical base.

(b) A second example would be magnesium oxide dissolved in water.

$$MgO + H_2O \rightleftharpoons MgOH^+ + OH^-$$

By accepting a proton from water, magnesium oxide has likewise increased the concentration of hydroxyl ions in the solution.

NOTE: The two theories explain all the properties of acids and bases that will be utilized in medicine. It deserves mention that there are other theories of acids and bases that explain more complex phenomena. If these are of interest to you, a college chemistry text should have a discussion of some of them.

c. **Properties of Acids**. We have defined all acids based on one common property, the ability to donate hydrogen ions in solution. Therefore, you should expect them all to exhibit a set of common properties, which they do. The properties we are concerned with are as follows:

(1) Acids change blue litmus paper to red. Litmus paper, which contains dyes sensitive to hydrogen ion concentration, turns red when there is a high concentration, blue when there is a low concentration.

(2) Acids have a sour taste. This property is familiar to you if you have ever tasted a lemon. Lemons contain citric acid, which gives them their sour taste.

(3) Acids react with metals to release hydrogen gas. For example:

$$Zn + 2H^+ \longrightarrow Zn^{++} + H_2$$

You will notice that this reaction is an oxidation-reduction reaction. For practice, pick out the oxidizing and reducing agents.

(4) Acids react with carbonates and bicarbonates to form carbon dioxide. For example:

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2\uparrow$$

(5) Acids react with bases to form salts and water (neutralization reaction). For example:

$$HCl + NaOH \longrightarrow \underset{\text{(salt)}}{NaCl} + H_2O$$

d. **Properties of Bases**. In the same manner that all acids had certain properties in common, all bases have related properties. The ones that are important to the medical personnel are as follows:

(1) Bases change red litmus paper to blue. This is just the opposite of the change which acids cause in litmus paper.

(2) Bases possess a bitter taste and feel soapy when in contact with the skin.

(3) Bases react with acids to form salts and water (neutralization reaction). This is the same type reaction as previously discussed under acids.

e. **Classification of Acids and Bases**. Even though all acids possess certain properties in common, as do bases, not all possess them to the same degree. Some acids, for example, will completely neutralize sodium hydroxide with equal concentrations while others will only partially neutralize this base. As you might suspect, the differences in the strengths of acids results from differing abilities to donate hydrogen ions and the differences in bases from differing abilities to donate hydroxyl ions or accept hydrogen ions.

(1) Some acids and bases dissociate more readily than others when placed in solution. Those that dissociate at a rate greater than 50 percent are considered to be strong acids or bases. Weak acids and bases dissociate at a rate that is less than 50 percent. Examples:

(a) When hydrochloric acid (HCl) is placed in solution, <u>most</u> of the molecules will dissociate to form free H+ ions and Cl^- ions. Hydrochloric acid is therefore considered a strong acid.

(b) When carbonic acid (H_2CO_3) is placed in solution, less than 50% will ionize into free H^+ ions and HCO_3^- ions. Most of the molecules will remain in molecular form.

$$H_2CO_3 \xrightarrow{HOH} H^+ + HCO_3 \longrightarrow H_2CO_3$$

(2) This means one mole (gram molecular weight) of HCl will produce more hydrogen ion in solution than will one mole of H_2CO_3 and will consequently exhibit acidic properties to a greater degree than will carbonic acid. A simpler way to say this is that HCL is a stronger acid than H_2CO_3.

(3) The same rationale holds for bases as well as acids. Therefore, we can divide or classify acids or bases into groups based on their dissociation--strong acids or bases (those that dissociate completely) and weak acids or bases (those that dissociate to a small degree).

f. **Acids and Bases of Medicinal Importance**. One may come in contact with a number of important acids and bases. You must be able to identify them as acids or bases and know their relative strengths. Table 2-1 shows these acids and bases. There is not an easy way to differentiate between strong and weak acids, but strong and weak bases can be differentiated based on valence. Strong bases have a positive valence of one; weak bases have a positive valence greater than one.

g. **Safety and Antidotes**. Acids and bases should be handled with care to avoid spilling on skin. They should not be taken internally unless intended for that purpose. If the skin is exposed to these compounds or is ingested, the following antidotes are recommended for first aid treatment.

(1) Acids.

(a) External. Use large amounts of water to wash acids off the skin. Exception: If phenol (an organic acid) is spilled on the skin, wash off with alcohol.

(b) Internal. Give an antacid, other than a carbonate or bicarbonate, such as milk of magnesia or magnesium oxide. DO NOT give an emetic or induce vomiting.

(2) Bases.

(a) External. Wash the area with large amounts of water.

(b) Internal. Give a weak acid such as vinegar or fruit juice. Weak acids (or weak bases) are only effective if administered within 10-15 minutes of ingestion of strong base (or strong acid). DO NOT give an emetic or induce vomiting.

2-9. SALTS

Previously, it has been stated that one of the properties associated with acids and bases is the neutralization reaction. This reaction involves the production of a salt and water from the reaction of an acid and a base. We will now examine various types of salts produced in neutralization reactions. Salts are the third major classification of inorganic compounds (acids and bases being the first two). They are important in the physiology of the body and are often used as therapeutic agents.

a. **Definition**. We have already given one definition of a salt in our discussion, that is, the product of a reaction between an acid and a base. A more specific definition, however, would be an ionic compound formed by the replacement of part or all of the acid hydrogen of an acid by a metal or a radical acting like a metal. It is an ionic compound that contains a positive ion other than hydrogen and a negative ion other than hydroxyl (OH^-) or "O^{-2}," as in MgO.

b. **Types of Salts**. There are four types of reactions possible between acids and bases as we classified them (strong or weak) earlier. These are as follows:

Relative strength of common acids and bases

STRONG ACIDS		WEAK ACIDS	
HCl	Hydrochloric acid	$HC_2H_3O_2$ (HAC)	Acetic acid
H_2SO_4	Sulfuric acid	H_2CO_3	Carbonic acid
H_3PO_4	Phosphoric acid	H_3BO_3	Boric acid
STRONG BASES		**WEAK BASES**	
KOH	Potassium hydroxide	$Fe(OH)_2$	Ferrous hydroxide
NaOH	Sodium hydroxide	$Al(OH)_3$	Aluminum hydroxide
		NH_3	Ammonia

* $Ca(OH)_2$	Calcium hydroxide
$Mg(OH)_2$	Magnesium hydroxide
MgO	Magnesium oxide

* Notice that $Ca(OH)_2$, $Mg(OH)_2$, and MgO which forms $Mg(OH)_2$ in water are chemically classified as strong bases because of their high degree of dissociation. Because they are only slightly soluble in water, they produce low concentrations of the hydroxide (OH^-) ion in solution. Since calcium hydroxide and magnesium hydroxide do not produce tissue damage, they can be safely used as therapeutic agents (e.g., antacids).

Table 2-1. Relative strength of common acids and bases.

(1) Strong acid and strong base.

$$HCl + NaOH \dashrightarrow NaCl + H_2O$$

(2) Weak acid and weak base.

$$2H_2CO_3 + Fe(OH)_2 \dashrightarrow Fe(HCO_3)_2 + 2H_2O$$

(3) Strong acid and weak base.

$$2HCl + Fe(OH)_2 \dashrightarrow FeCl_2 + 2H_2O$$

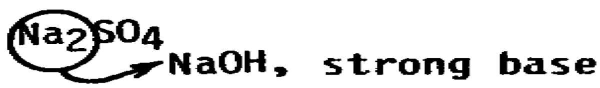
(4) Weak acid and strong base.

$$H_2CO_3 + NaOH \text{ -----> } NaHCO_3 + H_2O$$

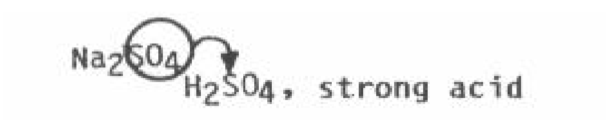
NOTE: These four reactions result in three types of salts. Reactions (1) and (2) result in neutral salts (that is, in terms of pH), which means a solution of the salt in water will be a neutral solution. Reactions such as (3) result in acidic salts, which produce acidic solutions. Reaction (4) results in basic salts, which produce basic solutions.

c. **Determination of Salt Type**. To determine the type of salt from a chemical formula, we employ the following steps:

(1) The first element comes from a base. Determine which base and whether it is weak or strong.



(2) The remainder of the formula comes from the acid. Determine which acid and whether it is weak or strong.



(3) By knowing the strengths of the acid and base that formed the salt, the salt type can be assigned. Table 2-2 is a summary of salt types resulting from various acid-base combinations.

d. **Example**. $Al_2(SO_4)_3$.

(1) The first element, aluminum, comes from the base $Al(OH)_3$. Since it has a valence of +3, it is a weak base.

(2) The sulfate radical comes from H_2SO_4, sulfuric acid, which is a strong acid.

(3) This compound is an acidic salt since it is the product of a reaction between a strong acid and a weak base.

e. **Example**. $FeBO_3$.

(1) The first element, iron, comes from the base $Fe(OH)_3$, and since its valence is +3, it is a weak base.

(2) The borate radical comes from boric acid, which is a weak acid.

(3) Thus, this is a neutral salt, since it is the product of a reaction between a weak acid and a weak base.

f. **Importance of Type of Salt**. The type of salt is very important when a salt is used medicinally, since the body maintains a specific acidity in the tissues and fluids. The type of salt is also important in the prediction and understanding of incompatibilities. It is important for you to identify the type of salt from its formula. The importance and use of the type will become clear to you as you progress through the course.

REACTANTS---> ↓	WEAK ACID	STRONG ACID
WEAK BASE	Neutral Salt	Acidic Salt
STRONG BASE	Basic Salt	Neutral Salt

Table 2-2. Salt types resulting from various acid-base combinations.

2-10. pH AND ACIDITY

In discussing acids, bases, and salts, we often refer to a solution or compound being acidic, neutral, or basic in a qualitative manner. This concept is useful to us in a general sense, but would be of much greater value if we could speak in quantitative terms. It would be valuable if we could answer the question of how acidic one solution is in relation to another solution.

a. **pH.** The solution to this problem is not as difficult as it may seem. Acids donate protons (hydrogen ions, H +) in solution. Thus, the acidity of a solution must be related to this property.

(1) In fact, the acidity of a solution is the concentration of hydrogen ions in that solution. Since we can calculate the hydrogen ion concentration, as you will learn later, we can now determine a numerical value of the acidity of a solution. The concentrations of hydrogen ions in both acidic and basic solutions are generally very small. A strong solution of HCl, for example, may contain only 0.01 mole of hydrogen ions per liter of solution. A solution of NaOH may have as little as 0.00000000001 mole of hydrogen ion per liter of solution.

(2) To simplify the expression of such terms, chemists have transformed the concentration values into numbers, called pH numbers, which are easier to utilize. This is done according to the following equation:

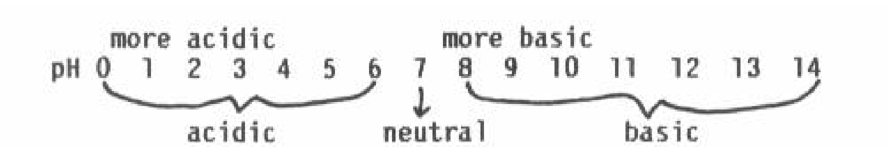
$$pH = -\log[H^+]$$

The abbreviation log stands for logarithm. (For example, log 1 = 0, log 0.1 = -1, log 0.01 = -2, log 0.001 = -3, log 0.0001 = -4.) The expression $[H^+]$ here is the concentration of hydrogen ions in moles per liter. If we consider the two previous examples, you can see how this transformation aids us. The pH of the HCl solution would be -(-2.0) = 2.0; the pH of the NaOH solution would be -(-11.0) = 11.0. These numbers, 2 and 11, are certainly easier to work with than 0.01 and 0.00000000001.

b. **pH Scale**. This transformation results in a range of pH numbers from 0 to 14, which is called the pH scale.

(1) The limits of the scale are related to the dissociation; how they are arrived at is beyond our scope. Further information on this relationship can be found in an inorganic chemistry textbook.

(2) While you will not need to calculate a pH value, you will need to interpret what a pH value means at times. To learn this function, examine the following pH scale:



(3) A pH value less than 7.0 means the solution is acidic; the lower the number, the more acidic. A solution with a pH of 2.0 is more acidic than one with a pH of 4.0. Any pH value greater than 7.0 means the solution is basic with larger numbers indicating solutions that are more basic. The only value on the scale that indicates a neutral solution is 7.0. The pH values for some common pharmaceutical products are given below.

PRODUCT	pH
Cherry Syrup	3.5 - 4
Benylin® Expectorant	5.0 - 5.5
Glycyrrhiza® Syrup	6.0 - 6.5
Iso-Alcoholic Elixir	5.0
Orange Syrup	2.5 - 3.0
Terpin Hydrate Elixir with Codeine	8.0

c. **Measurement of pH.** There are three common methods for measuring pH, which you may encounter in medicine.

(1) Litmus paper. Litmus paper is a paper coated with a dye, which is red in an acid pH or blue in a basic pH. It will only indicate whether a solution is acidic or basic; it will not give an actual pH value.

(2) pH paper. pH paper works on the same principle as litmus paper but uses several different dyes. By comparing paper color with a chart, the pH of a solution can be determined within one pH unit. If a closer measurement is needed, special narrow-range papers can be used to determine the pH within 0.1-pH unit.

(3) pH meter. The most accurate tool for pH measurement is the pH meter. This makes use of an electrode dipped into solution and is accurate to about 0.01-pH unit, depending on the particular machine.

2-11. BUFFERS

Many drugs are stable in solution only at certain pHs or in narrow pH ranges. If a solution of one these drugs is desired, the manufacturer must find a way to maintain this certain pH over a period of time. This is accomplished by the use of buffer systems. A buffer is a solution of a weak acid and the salt of that weak acid (weak bases could also be used, but usually are not practical). The function of the buffer is to resist changes in pH by reacting with any hydrogen or hydroxyl ions that are added to the solution. Two of the most common buffer systems are:

a. **Acetic Acid/Sodium Acetate**. This is a common buffer found in many drug solutions.

b. **Carbonic Acid/Sodium Bicarbonate**. This is the buffer system that is most common in the fluids and tissues of the body and is used to keep the pH of the blood and body fluids constant.

2-12. WATER

Water is the most important liquid to all living organisms. It comprises about 57 percent of your body weight. It is the solvent for or is contained in most of the nutrients your body requires for growth or maintenance. It is also the primary vehicle for almost all liquid pharmaceutical preparations. Because of the inherent importance of water in the practice of medicine, it is essential to acquire some knowledge of the properties of water and aqueous (water-based) solutions.

a. **Properties of Water**. All of us are familiar with some properties of water. We know that generally water is a bland-tasting, colorless liquid. Other specific properties of water are of importance in medicine.

(1) Its boiling point is $100^{o}C$ ($212^{o}F$).

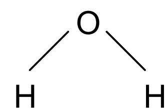
(2) Its freezing point is $0^{o}C$ ($32^{0}F$).

(3) It is a polar solvent (dissolves ionic compounds).

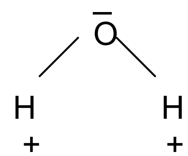
(4) Generally, it is chemically inert (unreactive) in biological or drug systems.

b. **Importance of Properties**. The properties above are the specific reasons that water is so valuable to living systems and to pharmaceutical preparations. The wide difference between the freezing point (water as ice) and the boiling point (water as steam or vapor) indicates that water will be a liquid at most of the temperatures encountered under normal conditions. An example should help emphasize the importance of these properties. If we wanted to prepare a liquid drug solution for a patient who could not swallow capsules, we used a liquid vehicle with a freezing point of 25^{o} C (77^{o} F) and a boiling point of 30^{o} C (86^{o} F), we would be giving the patient a worthless product. As the patient left home, the drug solution would boil if it were a normal summer day (temperature = 86^{o} F), and when the patient entered his air-conditioned home, the remaining solution might become a solid which could not be poured from the bottle. We also want our vehicle to be as unreactive as possible so that only the drug is exerting a pharmacological effect.

c. **Structure of Water**. The properties of water may best be explained by examining the structure of the water molecule. The water molecule consists of two hydrogen atoms bonded covalently to one oxygen atom. The three atoms are bound together as shown below.

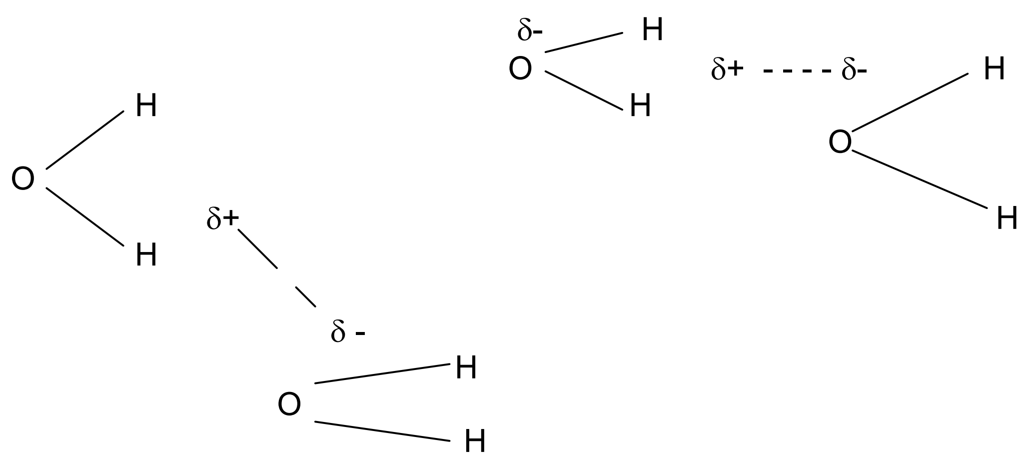


This arrangement leads to an electron-rich atom, oxygen, on one end and two electron-poor atoms, hydrogen, on the other end. This results in a molecule that resembles a bar magnet in that it has a negative pole and a positive pole, as shown below.



Actually, there are not distinct electrical charges on the molecule, only partial charges, referred to as δ+ and δ- (the Greek letter delta, δ, meaning partial). While these charges are only partial, they are still strong enough for water to be referred to as a polar molecule, meaning that it has a positive and negative end.

d. **Hydrogen Bond**. The polarity of the water molecule gives rise to an unused type of bond between water molecules, the hydrogen bond. This bond is the electrical attraction between the partially negative oxygen atom of one molecule and the partially positive hydrogen atom of another molecule.



The hydrogen bond is a very weak attraction about 1/10 to 1/20 the strength of the hydrogen-oxygen covalent bond. The hydrogen bond explains why water has such a high boiling point in relation to other compounds of similar molecular weight. For example, methane (CH_4, molecular weight = 16) boils at a temperature below $0^{\circ}C$, while water (molecular weight = 18) boils at $100^{\circ}C$. Methane does not exhibit hydrogen bonding.

e. **Water Purification**. We are all familiar with some of the ecological problems facing the world today. Water is subject to mineral and biological contamination. Since we will often be using water in the preparation of our products, we must be concerned with its purity and the methods utilized for its purification. There are two common methods of water purification used at Army medical treatment facilities--distillation and ion exchange.

(1) Distillation. Distillation is the process of boiling water, collecting the vapor, and then condensing the vapor back into water. Minerals and some of the bacterial contamination will remain in the boiling vessel as a residue. Very pure water may be prepared by repeating the distillation process several times. If sterile water is desired, the water must be sterilized, because the process of distillation does not necessarily sterilize water.

(2) Ion exchange (deionization). Less common than distillation because it is less efficient, ion exchange involves passing water through a column containing a charged resin. Ions in the water are held by electrical attraction and are thus removed from the water.

2-13. SOLUTIONS

We are seldom concerned with just water in the hospital. We are generally more concerned with substances dissolved in water. These are solutions. When we speak of a solution, there are several terms, which are important to understand.

a. **Solute**. A solute is the substance, which is dissolved in a solution.

b. **Solvent**. The solvent is, the substance, which dissolves the solute. It is usually water in pharmaceutical solutions, but not always.

c. **Solubility**. The maximum amount of a compound, which will dissolve in a given amount of solvent at a given temperature is the solubility of that compound.

d. **Dissociation**. (Ionization). In general, two things can happen to a solute in a solution. It can dissolve and exist in solution as molecules or it can dissociate and exist entirely or partially as ions. The process of splitting a molecule into ions is known as dissociation.

e. **Electrolyte**. When a substance dissociates to a fair extent in water, it will produce enough ions to support an electric current. We can use this property to differentiate between substances that are molecular and substances that are ionic in solution. An electrolyte is a substance, which dissociates sufficiently in solution to carry an electric current. It is therefore ionic in nature (figure 2-1).

f. **Non-Electrolyte**. A substance that does not dissociate or carry an electric current in solution is called a non-electrolyte.

g. **Hydrolysis**. Some compounds form ions in solution by reacting with water. This reaction with water is called hydrolysis. Hydrolysis is the dissociation of a compound through the splitting and incorporation of water. Hydrolysis occurs when acidic, basic, or neutral (weak acid/weak base) salts are dissolved in water. Consider, for example, the basic salt, sodium bicarbonate.

$$NaHCO_3 + H_2O \text{ ---> } Na^+ + HCO_3- + H_2O \text{ ---> } Na^+ + OH^- + H_2CO_3$$

The hydroxyl ion from water is associated with the sodium ion from the salt. The hydrogen ion from water is associated with the bicarbonate radical, and these two exist primarily as undissociated carbonic acid. The net result of this hydrolysis reaction is a basic solution containing sodium and hydroxyl ions and undissociated carbonic acid.

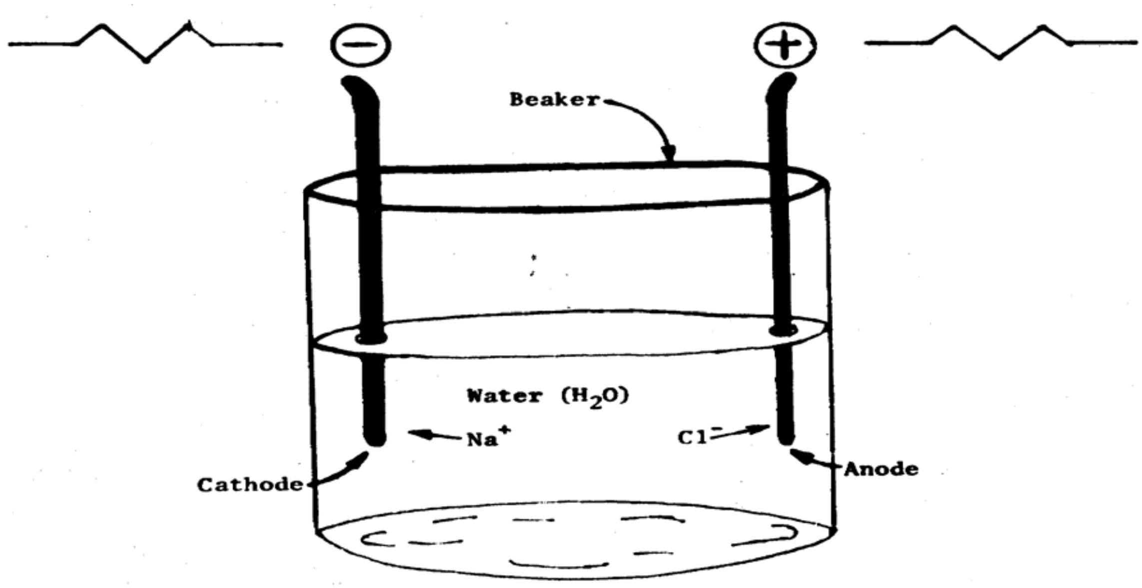


Figure 2-1. Flow of electric current through electrolyte solution.

h. **Electrolyte Strength**. It should be apparent, recalling the discussion of acids and bases that not all electrolytes will dissociate to the same extent in solution.

(1) Those that dissociate and exist entirely as ions in solution are called strong electrolytes. Strong electrolytes include strong acids, strong bases, and their neutral salts.

(2) Compounds that dissociate to a small extent and exist only partially as ions in solution are called weak electrolytes. Weak electrolytes include weak acids, weak bases, and salts of weak acids and/or weak bases.

(3) To identify whether a compound is a strong or weak electrolyte, it is first necessary to identify what type of compound it is. For example, consider NaCl. This compound is a salt formed from a strong base (NaOH) and a strong acid (HCl) and is, therefore, a neutral salt. Since it is a neutral salt of a strong acid and a strong base, it is a strong electrolyte as defined above. Other salts can be determined in a like manner.

Continue with Exercises

EXERCISES, LESSON 2

INSTRUCTIONS. Write the word, words, symbols, or numbers that properly completes the statement in the space provided or mark the correct word/phrase from those given. After you complete the exercises, turn to Solutions to Exercises and check your answers. Reread the material referenced for each exercise answered incorrectly.

1. We are given the following chemical equation:

$$NaOH + HCl \dashrightarrow NaCl + H_2O$$

It means that the base s ________ h_________ reacts with the acid h___________ a__________ to yield the salt s ____________ c____________ and the compound ___________ . To form one molecule of sodium chloride and one molecule of water, we need (one) (two) molecule(s) of sodium hydroxide and (one) (two) molecule(s) of hydrochloric acid.

2. We are given the following chemical reaction:

$$AgNO_3 + KCl \dashrightarrow KNO_3 + AgCl \downarrow$$

This means that s___________ n__________ and p____________ c__________ react to yield p____________ n__________ and s____________ c____________ . The arrow next to AgCl means that AgCl p_______________ s.

3. Calcium hydroxide and nitric acid react to yield calcium nitrate and water. The formula for calcium hydroxide is ______. The formula for nitric acid is ______. The formula for calcium nitrate is ________. The formula for water is ________. Before balancing, the equation for this reaction is _________________________, In the columns below are listed atoms and radicals of the reactants and products. Indicate the number in the formula of each compound. Consider water to be HOH.

REACTANTS	PRODUCTS
_____ Ca	_____ Ca
_____ OH	_____ OH
_____ H	_____ H
_____ NO_3	_____ NO_3

Thus, on the left side of the equation, we need twice as much _________________. On the right side of the equation, we need twice as much _____________. In order to satisfy this requirement, the equation becomes __________________.

On each side of the equation, we now have (one) (two) calcium atoms(s), (one) (two) hydroxide radical(s), (one) (two) hydrogen atoms, and (one) (two) (three) nitrate radicals. Since we now have an equal number of each type of atom and radical on both sides of the equation, we can say that the equation is __________________.

4. Sulfuric acid and ferric hydroxide react to produce ferric sulfate and water. The equation without balancing is:

_________+________ ------> _________+________

Now, list the reactants and the products as in the previous exercise.

REACTANTS	PRODUCTS
_______ H	_______ H
_______ SO_4	_______ SO_4
_______ Fe	_______ Fe
_______ OH	_______ OH

Now, balance the equation. (Before looking at the answer, be sure you have the same number of each type of atom or radical on both sides of the equation.)

__H_2SO_4 + __ Fe (OH)3 ----> __$Fe_2(SO_4)$3 + __ HOH

5. Iron metal and sulfur react to yield ferrous sulfide. The balanced equation for this reaction is ______________________________.

6. Sodium carbonate and hydrochloric acid react to yield sodium chloride, water, and carbon dioxide gas. Use a piece of scratch paper to write the balanced equation. The balanced equation is ________________________.

Did you forget the arrow after CO_2? The arrow means that CO_2 is given off as a _______________.

7. An exothermic reaction (gives off) (takes in) heat.

8. An endothermic reaction ______________heat.

9. A reaction is in equilibrium if the reactants are being formed from the ___________ at the same rate as the products are being formed from the _______________. The symbol used to indicate that a reaction may go in both directions is _________________________.

10. The formula weight of a compound is also known as its ________________ weight. If the molecular weight is expressed in milligrams, it is known as the _____________ molecular weight. If it is expressed in grams, it is known as the _______________ molecular weight. The formula weight of carbon dioxide is 12 + 16 + 16 = 44. The milligram formula weight of carbon dioxide is ______. The gram formula weight of carbon dioxide is __________.

11. The milligram molecular weight of a compound is the sum of the a________ w_______s of all the atoms that appear in its chemical formula, with the weights expressed in ________________.

12. Using Table 1-1 in this subcourse and rounding the atomic weights to the nearest whole number, what is the milligram molecular weight of sodium bicarbonate, $NaHCO_3$?

    Atoms: Na + H + C + O + O + O

    Atomic weights: __+ __+ __+ __+ + __

    Adding the atomic weights, we get a formula weight of __________.

    Thus, the milligram molecular weight is _____________.
    The gram milecular weight is ______________________.

13. A mole of $NaHCO_3$ weighs _______ grams. A liter of a 1M solution of $NaHCO_3$ contains ________grams of $NaHCO_3$.

14. The milligram equivalent weight of a compound is its milligram molecular weight divided by the total ____________ or _______________v ___________.

15. The bicarbonate radical has a total negative valence of______. Sodium bicarbonate has a total positive valence of ________. Since the milligram molecular weight of $NaHCO_3$ is 84 mg, its milligram equivalent weight is ______ ______.

16. Sulfuric acid, H_2SO_4, has a milligram molecular weight of _________ ________ . Its total positive valence is ___________. Its milligram equivalent weight is ______ _______.

17. If an element is oxidized, it (gains) (loses) electrons and its valence (increases) (decreases). If an element is reduced, it (gains) (loses) electrons and its valence (increases) (decreases). If an element is oxidized, there is an increase in its _______________. If it is reduced, there is a decrease in its ____________. When elemental iron reacts with diatomic oxygen, the valence of Fe goes from______________ to _____________. The valence of O goes from ___________ to _____________. Iron (gains) (loses) electrons. Oxygen (gains) (loses) electrons. Therefore, iron is (oxidized) (reduced) and oxygen is (oxidized) (reduced). The element that gains electrons and loses valence is said to be ___________________. The element that loses electrons and gains valence is said to be _________________________.

18. Oxidation may be defined as a ________________ of electrons or a ______________of valence. Reduction may be defined as a _____________ of electrons or a _____________ of valence.

19. In an oxidation-reduction reaction, the oxidizing agent is (oxidized) (reduced) and the reducing agent is ____________________.

20. When elemental magnesium Mg reacts with diatomic iodine, we have the following reaction:

$$Mg + I_2 \text{ ---> } MgI_2$$

In this reaction, the valence of magnesium goes from _______ to ______. The valence of iodine (I) goes from _______ to ____________. Since the valence of magnesium increases, magnesium is said to be ______________. Since the valence of iodine decreases, iodine is said to be __________. The reducing agent in this reaction is _______. The oxidizing agent is _____________

21. According to the classical theory, an acid is a compound that donates ____________________ and a base is a compound which donates___________ _____________. The symbol for a proton is _________. The symbol for a hydroxyl (hydroxide) ion is _____. According to the classical theory, an acid donates ____________________ ions and a base donates ____________________ ions.

22. According to the Bronsted-Lowry theory, an acid is a compound which donates________________ and a base is a compound which ___________________s protons.

23. According to the Bronsted-Lowry theory, which substance acts as an acid in the following reaction?

$$H_2CO_3 + OH^- \text{--->} HCO_3^- + H_2O$$

Which substance acts as a base in the following reaction?

$$H^+ + HCO_3^- \text{--->} H_2CO_3$$

24. Below is a list of nine compounds. For each, indicate whether it is best described as an acid, a base, or salt.

a. HCl _____________

b. NaOH _____________

c. HNO_3 ____________

d. $Ca(OH)_2$

e. KOH _____________

f. H_3PO_4 _____________

g. $Fe(OH)_3$ ______________

h. MgO ____________

i. NaCl ____________

25. Below are listed five properties of acids with some words missing. Complete the list.

    a. Acids change the color of litmus paper from __________ to ___________.

    b. Acids taste _______________.

    c. Acids react with metals to release ________gas.

    d. Acids react with carbonates and bicarbonates to form ____________ ___________ gas.

    e. Acids react with bases to form __________ and water.

26. Below are listed four properties of bases some words missing. Complete the list.

    a. Bases change the color of litmus paper from ___________ to ___________.

    b. Bases taste _________________________.

    c. Bases feel ____________________.

    d. Bases react with acids to form ______________ and ______________.

27. If a person spills a strong acid (except phenol) or a strong base on his skin, he should wash the acid or base off with large amounts of ___________________ .

28. If a person swallows a strong acid, you should give him an antacid such as ____________ of ____________ or _____________________ __________________. The antacid should NOT be a __________________ or ______________________. Vomiting (should) (should not) be induced.

29. If a person swallows a large amount of a strong base, such as sodium hydroxide, you should give him a _________ such as v__________ or f ______________ j__________. Vomiting (should) (should note) be induced.

30. A salt is a compound formed by the replacement of ______________ions in an___________ with a __________ or a radical acting like a metal. For example, the salt NaCl is formed by the replacement of the hydrogen in______ by the metal ________. The same salt NaCl results when NaOH reacts with _______. In general, we can say that an ionic compound is a salt if it contains a positive ion other than _________________ and a negative ion other than _____________________ or ___________________ .

31. For each of the eight compounds listed below; indicate whether it is an acid, a base, or a salt.

a. $NaNO_3$ _______________

b. $Al(HCO_3)_3$ _________________

c. K_2CO_3 _______________________

d. $FePO_4$ ______________________

e. $Ca(OH)_2$ ______________________

f. $Na(C_2H_3O_2)$ ____________________

g. NH_4Cl _______________________

h. H_2SO_4 ___________________

32. If the pH of a solution is 6.9, it is slightly ____________. If the pH is 7.5, it is slightly _____________. If the pH is 8.8, the solution is __________ .

33. According to the subcourse, the devices used to measure pH are ____________________ paper, and a ___________ ____________

34. A mixture used to keep the pH of a solution nearly constant is called a ___________________ system. Such a system usually contains a weak __________ and a ___________ of that weak _______________. The function of a buffer system is to keep the __________ at a nearly _____________ level.

Which of the following represents a buffer system?

$HCl/NaHCO_3$
HBr/KBr
$NaOH/NaCl$
$H_2CO_3/NaHCO_3$

35. The boiling point of water is ____________. The freezing point of water is___________. Because water dissolves ______________ compounds, it is called a __________ solvent. In biological and drug systems, water is generally (inert) (reactive).

36. Two major methods of water purification for medical use are di _______________________ and I ___________ e _____________ (de _______________).

37. A solute is a substance that is _______________ in a solution. A solvent is the substance in which the ____________ is dissolved. The solubility of a compound is the maximum amount that will ____________ in a given amount of _________________ at a given ____________. The substance in which the solute is dissolved is called the ________________. The amount of a substance that will dissolve under specified conditions is called its _________________.

38. The dissociation of a molecule means that it is ______________ into ions. When a compound is dissolved and its molecules split into ions, we call this process _______________. When an electrolyte is dissolved, many of its molecules ________________ and form _________. If a compound forms enough ions in solution to make the solution capable of carrying an electric current, then we say that compound is an _____________________ .

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 2

1. sodium hydroxide; hydrochloric acid; sodium chloride; water
   one, one (para 2-1)

2. silver nitrate; potassium chloride; potassium nitrate; silver chloride
   precipitates (para 2-1)

3. $Ca(OH)_2$
   HNO_3
   $Ca(NO_3)_2$; H_2O
   $Ca(OH)_2 + HNO_3$ ---->
   $Ca(NO_3)_2 + H_2O$

   1 Ca; 1 Ca
   2 OH; 1 OH
   1 H; 1 H
   1 NO_3; 2 NO_3

   NO_3
   OH
   $Ca(OH)_2 + 2HNO_3$ ---->$Ca(NO_3)_2$ + 2 HOH

   one, two; two, two
   balanced (paras 2-2, 2-3)

4. H_2SO_4 + $Fe(OH)_3$ --->$Fe_2(SO_4)_3$ + HOH

   2 H; 1 H
   1 SO_4; 3 SO_4
   1 Fe; 2 Fe
   3 OH; 1 OH

   $3H_2SO_4$ + $2Fe(OH)_3$ ---->Fe2$(SO_4)_3$ + 6HOH (paras 2-2, 2-3)

5. Fe + S ----> FeS (paras 2-2, 2-3)

6. Na_2CO_3 + HCl ----> 2 NaCl + H_2O + $CO_2 \uparrow$
   gas (paras 2-2, 2-3)

7. gives off (para 2-5a(1))

8. takes in (para 2-5a(2))

9. products, reactants,
 ------>
 <-------) (para 2-4)

10. molecular
milligram
gram
44 mg
44 grams (para 2-6)

11. atomic weights, milligrams (para 2-6a)

12. 23 + 1 + 12 + 16 + 16 + 16
84
84 mg
84 grams (para 2-6, table 1-1)

13. 84
84 (para 2-6b)

14. positive, negative valence (para 2-6b)

15. -1
+1
84 mg (para 2-6c)

16. 98 mg
+2
49 mg (para 2-6, Table 1-1)

17. loses, increases
gains; decreases
valence
valence
0; +2
0; -2
loses
gains
oxidized, reduced
reduced
oxidized (para 2-7)

18. loss, gain
gain, loss (para 2-7b, c)

19. reduced; oxidized (para 2-7d)

20. 0; +2
0; -1
oxidized
reduced
Mg
I_2 (para 2-7)

21. protons
hydroxyl ions
H^+
OH^-
H^+ (hydrogen), OH^- (hydroxyl) (para 2-8a)

22. protons accepts (para 2-8b)

23. H_2CO_3
HCO_3 - (para 2-8)

24. a acid
b base
c acid
d base
e base
f acid
g base
h base (MgO + H_2O ----> MgOH+ + (OH -)
i salt (para 2-8a, b)

25. a blue to red
b sour
c hydrogen
d carbon dioxide
e salts (para 2-8c)

26. a red to blue
b bitter
c soapy
d salts, water (para 2-8d)

27. water (para 2-8d)

28. milk of magnesia; magnesium oxide
carbonate, bicarbonate;
should not (para 2-8g(1)(b))

29. weak acid, vinegar; fruit juice
should not (para 2-8g(2))

30. hydrogen; acid; metal
HCl, Na
HCl
hydrogen (H^+); hydroxyl (OH^-); oxide (O^{-2}) (para 2-9a)

31. a salt
b salt
c salt
d salt
e base
f salt
g salt
h acid (para 2-9a)

32. acidic
basic
basic (para 2-10b)

33. litmus; pH; pH meter (para 2-10c)

34. buffer
acid, salt; acid
pH; constant
$H_2CO_3/NaHCO_3$ (HCl and HBr are strong acids; NaOH is a strong base) (para 2-11)

35. 100°C (212 °F)
0° C (32°F)
ionic, polar
inert (para 2-12e)

36. distillation; ion exchange (deionization) (para 2-12a)

37. dissolved
solute
dissolve; solvent temperature
solvent
solubility (para 2-13c-c)

38. split
dissociation
dissociate; ions
electrolyte (para 2-13d, e)

End of Lesssson 2

LESSON ASSIGNMENT

LESSON 3 Elements of Organic Chemistry.

LESSON ASSIGNMENT Paragraphs 3-1 through 3-18 and exercises.

LESSON OBJECTIVES After completing this lesson, you should be able to:

3-1. State the type of bond most prevalent in inorganic chemistry and the type of bond most prevalent in organic chemistry.

3-2. State the reason for the importance of structural formulas for organic compounds.

3-3. List the three types of carbon-carbon bonds and state the class of organic compounds representative of each type.

3-4. Define the terms hydrocarbon, aliphatic, saturated, unsaturated, and aromatic.

3-5. Given a list of structural formulas for organic compounds, match each to the name of the class of compounds to which it belongs.

3-6. Given a list of structural formulas for organic compounds, select those whose groups are acidic or basic.

3-7. Given a list of classes of organic compounds, select those that have the ability to form hydrogen bonds between themselves.

3-8. Given a list of classes of organic compounds, select the special reactions which each will undergo, to include oxidation, reduction, salt formation, esterification, amide formation, and hydrolysis.

3-9. Given a list of definitions for chemical terms, select the definitions for oxidation and reduction as they apply to organic chemistry.

SUGGESTION After completing the assignment, complete the exercises at the end of this lesson. These exercises will help you to achieve the lesson objectives.

LESSON 3

ELEMENTS OF ORGANIC CHEMISTRY

3-1. INTRODUCTION

Carbon is one of the most abundant elements in our world. It is part of the molecular structure of all living organisms. It is the basis for our fuel and energy production and it plays a large role in the chemistry of many of the synthetic fabrics and plastics that have become so important to our lifestyle. Carbon compounds also account for a vast majority of today's drugs. It is very important that you, as a health care provider, have a basic understanding of the chemistry of carbon compounds and organic chemistry due to the roles that carbon plays.

3-2. CONTRAST WITH INORGANIC CHEMISTRY

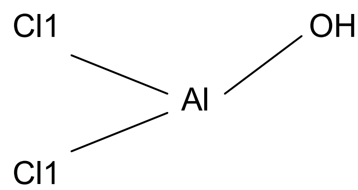
There are several general differences between the chemistries of carbon compounds and inorganic compounds, which will help give you an overall view of organic chemistry (Table 3-1).

	NORGANIC CHEMISTRY	ORGANIC CHEMISTRY
TYPE OF BONDING	Ionic	Covalent
MOLECULAR SIZE	Small	Large
WATER SOLUBILITY	Soluble	Insoluble
SOLUBILITY IN ORGANIC SOLVENTS	Insoluble	Soluble
CLASSES OF COMPOUNDS	Acid, base, or salt groups)	Many (functional
STRUCTURAL FORMULAS	Unimportant	Very important

Table 3-1. Comparison of organic and inorganic chemistry.

3-3. STRUCTURAL FORMULAS

A structural formula is a chemical formula that shows how atoms are bonded to each other. For example, we might write $AlOHCl_2$ as



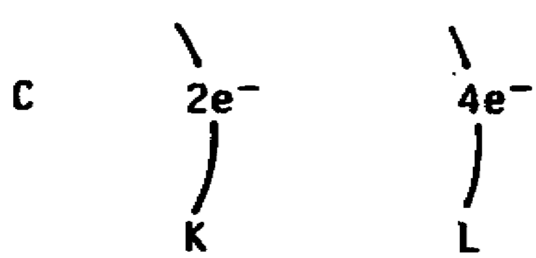
to show the bonds. However, in inorganic chemistry, the compounds are such that there is generally only one possible way to combine the atoms. This is not the case in organic chemistry, where very often there are many possible combinations for the atoms in the compound. Consider, for example, the formula C_4H_{10}. This formula could represent either of the following compounds.

$$\begin{array}{l} CH_3\text{-}CH\text{-}CH_3 \\ \quad\;\;| \\ \quad\; CH_3 \end{array} \qquad \text{or} \qquad CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_3$$

These compounds have slightly different properties. As the formulas become more complex, the differences are even greater. For this reason, it is often better to use a structural formula in organic chemistry rather than the simple chemical formula.

3-4. CARBON

Before we examine carbon compounds, we first need to examine the structure and mention some properties of the carbon atom. Carbon has an atomic number of six, meaning it has six protons, and consequently has six electrons. These electrons are distributed with two in the K shell and four in the L shell. In forming compounds, carbon would appear to gain or lose the four electrons in its outer shell. Thus, we have the +4, -4 valences you learned for carbon earlier in this subcourse.



In fact, carbon does not usually exchange electrons with other elements but prefers to share four electrons to complete its L shell. This is the reason that covalent bonding is predominant in organic chemistry.

3-5. CARBON-CARBON BONDING

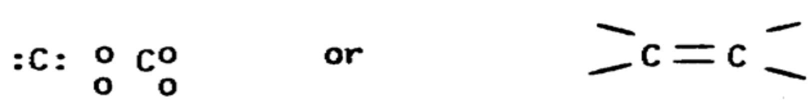
Carbon atoms have the unique ability to bond to other carbon atoms and form chains which may also have branches.

```
C-C-C-C-C-C-C          or          C-C-C-C-C-C-C-C-C-C-C
                                         I           I
                                         C           C
                                         I           I
                                          C-C         C
```

This is the reason that the molecular size is so great in organic chemistry. Molecular weights in the thousands are not uncommon. Three types of bonds are formed between carbon atoms.

a. **Single Bonds.** A single bond is a covalent bond formed by two carbon atoms sharing two electrons. Compounds that contain only single bonds between carbon atoms are called alkanes.

b. **Double Bonds.** A double bond consists of two covalent bonds formed by two carbon atoms sharing four electrons as shown below.



Compounds that contain at least one carbon-carbon double bond are referred to as alkenes.

c. **Triple Bonds**. A triple bond consists of three covalent bonds formed by two carbon atoms sharing six electrons as shown below.

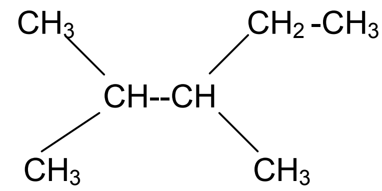
Compounds that contain at least one triple bond between carbon atoms are called alkynes.

3-6. HYDROCARBONS

The simplest organic compounds are the hydrocarbons, which are composed solely of carbon and hydrogen. Since there are only two elements involved, one might expect there would be only a few different compounds. However, carbon does bond to itself and form long chains. So there are many, many different hydrocarbons. They can be classed in two general groups, aliphatic and aromatic. These compounds are the starting point for all organic compounds.

a. **Aliphatic Hydrocarbons.** Aliphatic hydrocarbons consist of straight or branched chains of carbon atoms with the other valence electrons involved in bonds with hydrogen. Examples are:

$CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}CH_3$ $\qquad$ $CH_2 = CH_2$



We can subdivide aliphatic hydrocarbons into two groups based on the types of carbon-carbon bonds the compounds contain.

(1) <u>Saturated aliphatic hydrocarbons</u>. Saturated aliphatic hydrocarbons are hydrocarbons in which all of the carbon-carbon bonds are single bonds. These compounds are also referred to as alkanes, as mentioned for single bonds earlier. We often refer to the alkanes as the methane series. Methane is the simplest hydrocarbon with the formula CH_4. All other alkanes are formed by adding CH_2s to the formula (table 3-2). In this series, the names from C_5 to C_{10} all begin with the Greek prefix for the number (e.g., penta- for five) and end in -ane from "alkane." The two low-molecular-weight alkanes are gases. Alkanes are not very reactive chemically and are insoluble in water. About the most important reaction they undergo is that they burn to form carbon dioxide and water (combustion reaction). Some typical saturated compounds you might encounter are:

NAME	FORMULA	BOILING POINT (^oC)
Methane	CH_4	-161.5
Ethane	C_2H_6	- 88.3
Propane	C_3H_8	- 44.5
Butane	C_4HIO	- .5
Pentane	C_5H_{12}	+ 36.2
Hexane	C_6H_{14}	
Heptane	C_7H_{16}	
Octane	C_8H_{18}	+125.8
Nonane	C_9H_{20}	
Decane	$C_{10}OH_{22}$	+174.0

Table 3-2. Common alkanes.

(a) Liquid petrolatum (mineral oil). This liquid is used as a solvent and as a laxative.

(b) Petrolatum (petroleum jelly). This semisolid is used as an ointment base.

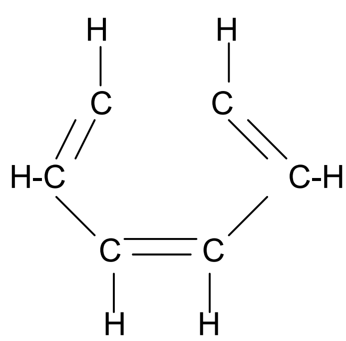
(c) Paraffin (wax). This solid is used in pharmacy as a stiffening agent.

(2) Unsaturated aliphatic hydrocarbons. The second type of aliphatic hydrocarbon is unsaturated hydrocarbons. These are hydrocarbons, which contain at least one double or triple bond (that is, they are alkenes or alkynes). An example of an alkene is ethene, the simplest alkene, which consists of two double-bonded carbon atoms and four hydrogen atoms.

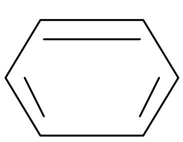
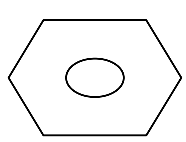
$$CH_2 = CH_2$$

Note that the name is similar to the saturated compound ethane. The -ene ending comes from the word alkene and denotes that it contains a double bond. Similarly, if there were a triple bond between the two carbon atoms, the name would be ethyne with the -yne ending denoting the triple bond (from alkyne). The physical properties of alkenes and alkynes are similar to the properties of alkanes of similar molecular weights. Chemically, the word unsaturated implies that these compounds can form additional bonds. This is the case, for alkenes and alkynes are much more reactive and undergo many reactions not possible with alkanes.

b. **Aromatic Hydrocarbons.** The second major group of the hydrocarbons is the aromatic hydrocarbons, which are hydrocarbons that contain a benzene ring as part of their structure. Benzene has the formula C_6H_6 and consists of six carbon atoms in a ring with three alternating double bonds.



The benzene ring is also represented with the following symbols:

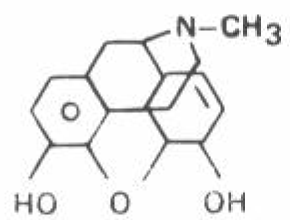
 or  or Ø

(1) Benzene is completely insoluble in water, it is a volatile liquid at room temperature, and it is fairly unreactive. The properties of other aromatics are reflective of benzene but vary according to the substituents added to the ring in place of one of the hydrogen atoms.

(2) The term “aromatic” has its origin in the fact that certain aromatic substances (for example: oil of bitter almonds, vanilla, and oil of wintergreen) contain the benzene ring. The possession of an odor is not characteristic, however, of all aromatic substances.

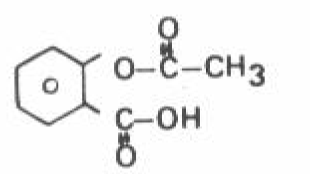
(3) Aromatic hydrocarbons are the starting point for many medicinally important compounds, as the following examples indicate.

Morphine



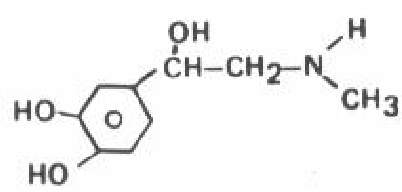
Narcotic Analgesic–For relief of severe pain.

Aspirin



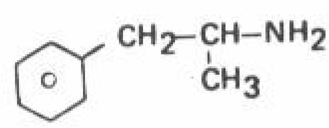
Non-Narcotic Analgesic–For relief of moderate pain.

Epinephrine



Adrenergic Drug–Used to treat shock.

Dextroamphetamine



Stimulant–For diet control, narcolepsy, and hyperkinesis.

(4) You will notice in the compounds above that they are not pure aromatic hydrocarbons because they contain elements other than carbon and hydrogen. These additional elements are the basis for the classification of substituted organic compounds and are called functional groups. The important functional groups will be considered in the following paragraphs.

3-7. INTRODUCTION TO FUNCTIONAL GROUPS

There are millions of organic compounds known to exist, and most are more complex than the simple hydrocarbons we have discussed. To facilitate the study of their reactions and properties, they are conveniently classed according to the functional groups they contain. A functional group is a group of atoms or a single atom that is substituted for a hydrogen or a hydrocarbon. These groups generally determine the types of reactions and properties of these more complex compounds. (A summary of the properties and reactions of the compounds contained in paragraphs 3-8 through 3-16 is tabulated in Table 3-3, pages 3-14 and 3-15.)

3-8. ALCOHOLS

Alcohols are hydroxyl (-OH) derivatives of hydrocarbons formed by replacing a hydrogen with the hydroxyl radical and are of the general form R-OH where R represents the hydrocarbon. There are three classes of alcohols: primary, secondary, and tertiary. A primary alcohol is one in which the hydroxyl group is attached to a carbon atom which, in turn, is attached to not more than one other carbon atom. A secondary alcohol is one in which the hydroxyl group is attached to a carbon atom which in turn, is connected to two carbon atoms. A tertiary alcohol is one in which the hydroxyl group is attached to a carbon atom which in turn, is attached to three other carbon atoms.

CH_3-CH_2-OH

Primary Alcohol
Ethanol
(ethyl alcohol)

$$\begin{array}{c} CH_3\text{-}CH\text{-}CH_3 \\ | \\ OH \end{array}$$

Secondary Alcohol
2-propanol
(isopropyl alcohol)

$$\begin{array}{c} CH_3 \\ | \\ CH_3\text{—}C\text{—}CH_3 \\ | \\ OH \end{array}$$

Tertiary Alcohol

Alcohols that contain two or more hydroxyl groups are referred to as polyhydroxy alcohols. An example you will encounter frequently in this course and on the job is glycerin, which is:

$$\begin{array}{ccccc} CH_2 & \text{—} & CH & \text{—} & CH_2 \\ | & & | & & | \\ OH & & OH & & OH \end{array}$$

a. **Properties of Alcohols.** The low-molecular-eight alcohols are volatile liquids, and the high-molecular-weight alcohols (more than 13 carbons) are solids. The first three alcohols (C_1 to C_3) are completely miscible (mix in any proportion) with water. The water solubility decreases as the number of carbons increases, and the large-molecular-eight alcohols are insoluble in water. Alcohols have higher boiling points and melting points than alkanes with the same or similar molecular weights (MW). For example:

	MW	MELTING POINT	BOILING POINT
CH_3-CH_2-CH_2-CH_2-OH	74	-9O^{0}C	118^{0}C
CH_3-CH_2-CH_2-CH_2-CH_3	72	-130^{0}C	36^{0}C

The water solubility and the high melting and boiling points of alcohols result from their ability to form hydrogen bonds with water and to form hydrogen bonds intermolecularly (between themselves).

b. **Reactions of Alcohols.** Chemically, the alcohols can be considered to be neutral (in terms of acids and bases) even though they can act as very weak acids or bases as water does. They undergo several kinds of chemical reactions, the most important of which is oxidation. <u>Oxidation</u> in organic chemistry is defined as the elimination of hydrogen from or the addition of oxygen to a compound.

(1) The oxidation of a primary alcohol can be expressed by the following example:

$$CH_3\text{-}OH \xrightarrow{KMnO_4} \overset{\overset{\displaystyle O}{\|}}{HCH} \xrightarrow[\Delta]{O_2} H\text{—}\overset{\displaystyle O}{C}\text{—}OH$$

NOTE: H╲C=O and O╲-C-OH are two more functional groups, indicating aldehydes and carboxylic acids, respectively.

The first step in this oxidation is the removal of two hydrogen atoms from the alcohol to form an aldehyde, and the second step is the addition of one oxygen atom to the aldehyde to form a carboxylic acid.

(2) Secondary alcohols undergo only the first step. For example, a three-carbon alcohol is oxidized to form $CH_3\ \text{–}\underset{\underset{\displaystyle O}{\|}}{C}\text{–}CH_3$, which is an example of a new class of compounds called ketones.

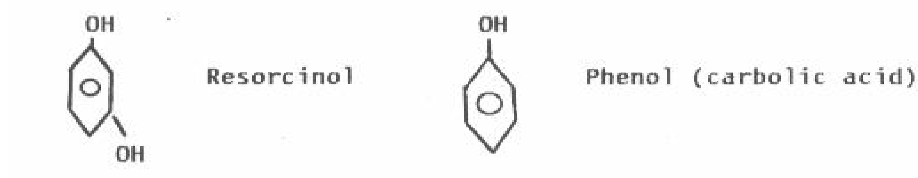
(3) Tertiary alcohols are not oxidized.

(4) One reason the oxidation reaction is important is that it is the means the body uses to eliminate the popular liquid, ethyl alcohol, or ethanol (CH_3-CH_2-OH).

c. **Uses of Alcohols.** Alcohols are most commonly used as solvents in the pharmacy. They are also used as disinfectants and antiseptics.

3-9. PHENOLS

Phenols are hydroxyl derivatives of hydrocarbons formed by replacing a hydrogen on the benzene ring of an aromatic hydrocarbon with the hydroxyl radical. Phenols have the general formula Ar-OH, where Ar represents a substituted or nonsubstituted aromatic hydrocarbon. Thus, phenols are really just a special class of alcohols. However, they have enough unique properties that they deserve to be considered as a separate class of compounds. Below are some examples of typical phenols.



a. **Properties of Phenols.** All phenols are white solids with moderately high melting points and are soluble in water. They also have the property of being able to form eutectics with camphor, menthol, or thymol, which are solid alcohols. (A eutectic is a uniform mixture formed from two compounds that melt at a temperature lower than the melting point of either of the two compounds.) Thus, phenol (a solid) and camphor (a solid) form a liquid mixture at room temperature which is called a eutectic.

b. **Reactions of Phenols.** Chemically, phenols are weakly acidic compounds. The hydrogen dissociates to a small degree from the hydroxyl radical to act as an acid as shown below.

Since phenols are weak acids, they will form salts with inorganic bases. Phenols with two hydroxyl groups also undergo oxidation reactions.

c. **Uses of Phenols.** Medicinally, phenolic compounds have three uses: as keratolytics (compounds that remove hornified or scaling outer layers of skin), antipruritics (relieve itching), and disinfectants. These uses arise from the fact that phenols are very caustic to animal tissues. Precautions must therefore be taken when you are using phenols in preparations. These properties, possessed to different degrees by various phenols, depend on which other functional groups are present and the number of hydroxyl groups.

3-10. ETHERS

An ether can be thought of as a hydrocarbon derivative of water where the two hydrogens of water are replaced by hydrocarbon groups. Thus, ethers have the general structural formula R-O-R' where R and R' represent any two hydrocarbons, which may be alike or different. Some examples of ethers are:

CH_3-CH_2-O-CH_2-CH_3 Ø-O-CH_3 CH_2=CH-O-CH_3

a. **Properties of Ethers.** Ether molecules are slightly polar, but cannot form hydrogen bonds with each other since they do not have a hydrogen atom attached directly to an oxygen atom. Therefore, they have about the same boiling points and melting points as alkanes of similar molecular weights.

	M.W.	Boiling Point
CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3	100	98^0C
CH_3-O-CH_2-CH_2-CH_2-CH_2-CH_3	102	100^0C

b. **Reactions of Ethers.** Since ether molecules are slightly polar and have an oxygen atom in their structure, they can form hydrogen bonds with water. This property accounts for the fact that ethers are slightly soluble in water. Chemically, ethers are inert except for the oxidation reaction. Ethers are oxidized in the presence of oxygen to form peroxides, which are explosive when concentrated.

c. **Uses of Ethers**. Medicinally, ethers are used as general anesthetics. They are also used as solvents. Many of you are involved with ordering and storing ethers.

3-11. AMINES

Amines result from the replacement of one or more of the hydrogen atoms of ammonia with hydrocarbons and have the general formula R-NH_2. There are four classifications of amines: primary, secondary, tertiary, and quaternary. Primary amines result from replacing one of the hydrogens of ammonia by a hydrocarbon, as in CH_3-NH_2; secondary amines result from the replacement of two hydrogens of ammonia by two hydrocarbons, as in CH_3-NH-CH_3; and tertiary amines result from the replacement of all three hydrogens of ammonia by hydrocarbon groups. The fourth classification of amines is sometimes encountered in drug structures. This classification is the quaternary amine that is formed by replacing the four hydrogens of the ammonium ion (NH_4 +) by hydrocarbon groups. Whenever one of the hydrocarbon groups connected to the nitrogen atom contains a benzene ring, the compound is referred to as an aromatic hydrocarbon.

a**. Properties of Amines.** The low-molecular-weight amines are all volatile liquids, and those having up to five carbons are soluble in water. The element nitrogen is in the same period of the periodic table as oxygen and has some similar properties--

the most significant being the ability to form hydrogen bonds. The formation of hydrogen bonds between amines, and between amines and water, accounts for their higher boiling points (than alkanes) and their water solubility.

b. **Reactions of Amines.** Since amines are derivatives of ammonia, they are bases as defined by the Bronsted-Lowry theory. The nitrogen of the amine can accept a proton to form a substituted ammonium ion.

$$CH_3-CH_2-NH_2 + H^+ \dashrightarrow CH_3-CH_2-NH_3^{\ +}$$

Amines will thus react with inorganic acids to form salts. (Amines react with organic acids to form amides, a class of organic compounds discussed later in this subcourse.)

$$CH_3-NH_2 + HCl \dashrightarrow CH_3-NH_3^{\ +}Cl^-$$

The reaction in the example above results in a hydrochloride salt of the amine and is a very important reaction in pharmacy. Many drugs contain an amine functional group, and if they contain many carbon atoms, they are not very soluble in water. The salts formed from amines, however, are very soluble in water. Therefore, if we wish to use a water solution of an amine drug that is insoluble, we can make it soluble by forming the salt of the amine.

c. **Use of Amines.** As already stated, the amine functional group is contained in many different drugs that have quite different actions in the body. Generally, these drugs are very complex and you would never be expected to draw or know the structure for these drugs. You should, however, recognize the $-NH_2$ group of an amine and be cognizant of its basic properties.

3-12. CARBOXYLIC ACIDS

Carboxylic acids are formed by the two-step oxidation of alcohols as stated previously and have the general structural formula $R-\overset{\overset{\displaystyle O}{\|}}{C}-OH$ or $R-COOH$. Some examples of carboxylic acids are:

$$CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-OH$$

Ethanoic Acid
(Acetic Acid)

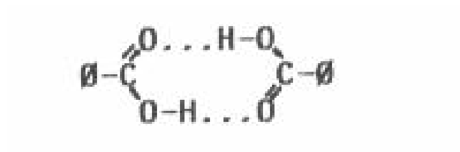
$$O-\overset{\overset{\displaystyle O}{//}}{C}-OH$$

Benozoic Acid

$$CH_3-\underset{\underset{\displaystyle CH_3}{|}}{CH}-\overset{\overset{\displaystyle O}{\|}}{C}-OH$$

2-Methylpropionic Acid

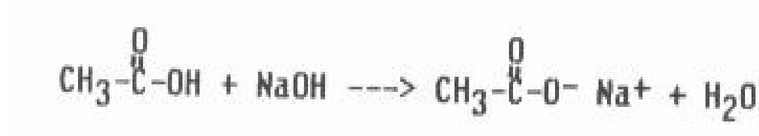
a. **Properties of Carboxylic Acids.** Carboxylic acids are very polar compounds due to the two oxygen atoms and can form two hydrogen bonds between themselves as shown below.



They have the highest melting points of any of the classes of compounds in table 3-3; a carboxylic acid has a higher melting point than a different type of organic compound with a similar molecular weight. Consequently, they are all solids under normal conditions. The compounds with four carbons or less are miscible with water; those with five carbons are slightly soluble, and those with more than five carbons are generally insoluble in water.

b. **Reactions of Carboxylic Acids.** As their name implies, carboxylic acids are the most acidic of all organic compounds but are still weak acids when compared to inorganic acids.

Carboxylic acids will form salts with inorganic bases, and as with the basic amines, this property is often used to make insoluble organic acids soluble in water as their salt. This pair, ethanoic acid (acetic acid) and its salt sodium ethanoate (sodium acetate), is used as a buffer system.



Carboxylic acids undergo three other important chemical reactions: reduction, ester formation, and amide formation.

(1) Reduction in organic chemistry is the opposite of oxidation and is the addition of hydrogen to or the elimination of oxygen from a compound. In the case of carboxylic acids, the removal of oxygen first results in an aldehyde, which may be reduced further by the addition of hydrogen to form an alcohol.

(2) Ester formation, as illustrated by the reaction below, is the reaction of a carboxylic acid with an alcohol to yield a new class of compound called an ester.

$$\underset{\text{Acid}}{CH_3-COOH} + \underset{\text{Alcohol}}{CH_3-CH_2-OH} \text{ ------> } \underset{\text{Ester}}{CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-O-CH_2-CH_3} + H_2O$$

CLASS OF COMPOUNDS	ALCOHOL	PHENOL	ETHER	AMINE	CARBOXYLIC ACID	ALDEHYDE	KETONE	ESTER	AMIDE
GENERAL STRUCTURE	R-OH	Ar-OH	R-O-R	$R-NH_2$	O ‖ R-C-OH	O ‖ R-C-H	O ‖ R-C-R	O R-C-O-R	O $R-C-NH_2$
	ROH	ArOH	ROR	RNH_2	RCOOH	RCOH	RCOR	RCOOR	$RCONH_2$
PRODUCT OR COMPOUND	CH_3CH_2OH	Ø-OH	$C_2H_5OC_2H_5$	$CH_3CH_2NH_2$	O ‖ CH_3COH	O ‖ CH_3CH	O ‖ CH_3CCH_3	O ‖ CH_3COCH_3	O ‖ CH_3CNH_2
NAME IN COMMON SYSTEM	Ethanol (Ethyl Alcohol)	Phenol	Ethyl Ether	Ethylamine	Ethanic Acid (Acetic Acid)	Ethanol (Acetal-dehyde)	2-Propanone (Methyl Ketone)	Methyl EthanoatE (Methyl Acetate)	Ethanamide (Acetamide)
OTHER COMMON NAMES	Grain Alcohol	Carbolic Acid	Diethyl Ether	-	-	-	Acetone Dimethyl Detone	Dimethyl Ester	-
pH	Neutral	Slightly Acidic	Neutral	Basic	Acidic	Neutral	Neutral	Neutral	Neutral
HYDROGEN BONDING BETWEEN THEMSELVES	Yes	Yes	No	Yes	Yes 2 – H Bonds	No	No	No	Yes
COMPARISON OF BOILING POINT TO CORRES-PONDING ALKANE	Higher	Higher	Same	Highest	Same	Same	Same	Same	High
OXIDIZED TO	1^o Aldehyde And/Or Acid 2^o Ketone	-	-	-	CO_2 + H_20	Acid	Acid (Very Difficult)	-	-
REDUCED TO		-	-	-	Alcohol	Alcohol	2^O Alcohol	Alcohol + Alcohol	-
HYDROLYSIS	-	-	-	-	-	-		Alcohol + Acid	Acid + Amine

Table 3-3. Summary of properties for functional groups

(3) Amide formation, as illustrated by the reaction below, is the reaction of a carboxylic acid with an amine to yield a new class called an amide.

$$\underset{\text{Acid}}{CH_3-COOH} + \underset{\text{Amine}}{CH_3-CH_2-NH_2} \longrightarrow \underset{\text{Amide}}{CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-NH-CH_2-CH_3} + H_2O$$

c. **Uses of Carboxylic Acids.** Many acids, such as acetic, salicylic, and lactic, are used topically to treat local conditions. Others are used systemically. Still others, like citric acid, which is found naturally in lemons, are used to flavor syrups for administration of other drugs. They are also used in many analytical procedures in the clinical laboratory.

3-13. ALDEHYDES

Aldehydes result from the first oxidation of alcohols and have the general structural formula $R-\underset{\underset{\displaystyle O}{\|}}{C}-H$. Since aldehydes cannot form hydrogen bonds between themselves, they have lower boiling points than corresponding alcohols or acids. Again, as with the other classes or organic compounds in table 3-3, the lower-molecular-weight aldehydes (up to five carbons) are soluble in water. Aldehydes are neutral in pH and undergo both oxidation and reduction reactions. They are easily oxidized to acids and reduced to alcohols. Some aldehydes, such as vanillin and benzaldehyde, are frequently used in the pharmacy as flavoring agents.

$$\underset{\text{Benza ldehyde}}{Ø-\overset{\overset{\displaystyle O}{\|}}{C}-H} \qquad\qquad \underset{\text{Methanal (Formaldehyde)}}{H\overset{\overset{\displaystyle O}{\|}}{C}H}$$

Others, such as formaldehyde, are often used as disinfectants.

3-14. KETONES

Ketones result from the oxidation of a secondary alcohol and have the general structural formula $R-\overset{\overset{\displaystyle O}{\|}}{C}-R'$ where R and R' can be the same or different hydrocarbon groups.

a. Ketones are similar to aldehydes in their boiling points, which are lower than those of corresponding alcohols and carboxylic acids.

b. Ketones are neutral compounds, being neither acids nor bases. They undergo the process of reduction, by which they are converted to secondary alcohols.

c. The ketone functional group appears in the structure of many complex drugs, such as steroid compounds and vitamins. Simple ketones, with the exception of acetone, are seldom used. Acetone is used as a solvent and cleaning fluid.

O
‖
CH_3–C–CH_3
2-propanone
(Acetone)

3-15. ESTERS

Esters, as previously mentioned, are formed from the reaction of a carboxylic acid with an alcohol and have the general structural formula RCOOR' or O where R and R' can be the same or different hydrocarbon groups.

O
II
R-COR'

a. **Properties of Esters.** The simplest esters are liquids and have fragrant odors. An example is ethyl acetate, CH_3-CH_2-OOC-CH_3, which has the odor of pineapple. Esters cannot form hydrogen bonds between themselves; consequently, they have boiling points similar to alkanes of similar molecular weight. They can form hydrogen bonds with water. Therefore, esters that contain less than five carbon atoms are soluble.

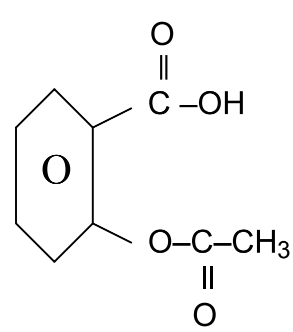
b. **Reactions of Esters.** Esters are neutral in pH and undergo two important chemical reactions, hydrolysis, and reduction. Hydrolysis is the splitting of an ester with the incorporation of water to form a carboxylic acid and an alcohol.

```
  O               O
  II      H2O     II
R-C-O-R   --->  R-C-OH + R'-OH
```

After hydrolysis, the acid product can undergo reduction to form a second alcohol as described previously.

c. **Uses of Esters.** The ester functional group is found in many complex molecules which you will be studying in the pharmacology subcourses if you take them.

Acetylsalicylic Acid (Aspirin)--an analgesic



Nitroglycerin--a cardiac drug

$$\begin{array}{l} CH_2\text{-}O\text{-}NO_2 \\ \vert \\ CH_2\text{-}O\text{-}NO_2 \\ \vert \\ CH_2\text{-}O\text{-}NO_2 \end{array}$$

3-16. AMIDES

Amides are formed from the reaction of a carboxylic acid with an amine or ammonia and have the general formula $R\text{-}\overset{\overset{\displaystyle O}{/\!/}}{C}\text{-}NH\text{-}R'$ where R and R' can be the same or different hydrocarbon groups.

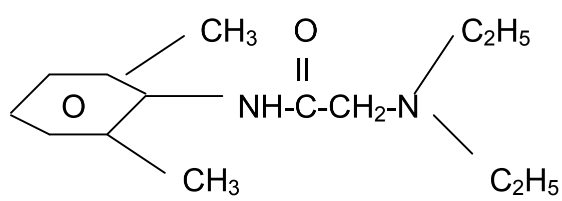
a. **Properties of Amides.** Amides, because of the hydrogen attached to the nitrogen atom, can form hydrogen bonds between themselves. They have higher boiling and melting points than corresponding alkanes. Since they can also form hydrogen bonds with water, amides containing up to five carbon atoms are soluble in water.

b. **Reactions of Amides.** Amides are neutral in pH and undergo the hydrolysis reaction. For amides, hydrolysis is the splitting of the compound with the incorporation of water to form a carboxylic acid and an amine.

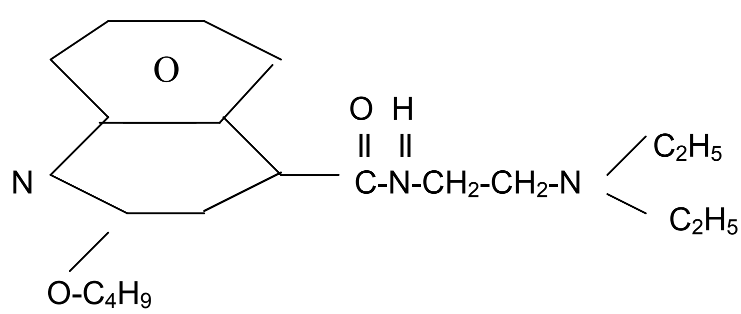
$$R\text{-}\overset{\overset{\displaystyle O}{\Vert}}{C}\text{-}O\text{-}NHR' \xrightarrow[H_2O]{} R\text{-}\overset{\overset{\displaystyle O}{\Vert}}{C}\text{-}OH + R'\text{-}NH_2$$

c. **Uses of Amides.** Some examples of drug molecules containing the amide functional group are shown below[1]

Lidocaine (Xylocaine®) = local anesthetic



Dibucaine (Nupercaine®) = local anesthetic



3-17. HALOGENATED HYDROCARBONS

Halogenated hydrocarbons are compounds with the general formula R-X where R is any hydrocarbon group and X is a halogen (Cl, Br, F or I). The most significant property of halogenated hydrocarbons is that as you increase the number of halogens on the compound, the flammability of the compound decreases. This property has been used to produce ethers that are nonflammable to be used as general anesthetics such as:

Methoxyflurane (Penthrane®)

```
   F  Cl    H
   |   |    |
 H–C–C–O–C–H
   |   |    |
   F  Cl    H
```

3-18. SUMMARY

Functional groups, when attached to various hydrocarbons, increase the reactivity and water solubility of the hydrocarbon. Carboxylic acids and phenols are the only organic acids; they are weak acids. Amines are the only significant organic bases. All functional groups that contain a hydrogen connected to a nitrogen or oxygen atom have the ability to form hydrogen bonds between themselves. All functional groups that contain a nitrogen or oxygen atom can form hydrogen bonds with water, which increase their solubility. In general, organic compounds of low molecular weight (less than five carbons) which contain functional groups, are soluble in water. Table 3-3 summarizes the properties and some reactions of the organic compounds we have studied.

Continue with Exercises

EXERCISES, LESSON 3

INSTRUCTIONS. Write the word, words, symbols, or numbers that properly completes the statement in the space provided or mark the correct word/phrase from those given. After you complete the exercises, turn to Solutions to Exercises and check your answers. Reread the material referenced for each exercise answered incorrectly.

1. The most common type of bonding in inorganic chemistry is __________ bonding. The most common type of bonding in organic chemistry is ______________ bonding. This means that in organic compounds, the electrons involved in bonds are usually ____________ by the atoms.

2. In inorganic chemistry, a compound is often completely identified by a simple formula such as H_3PO_4, NaCl, or Fe(OH)3. However, in organic chemistry, we frequently need formulas that indicate the molecule's ____________. In organic chemistry, several compounds may have the same makeup in terms of e__________ and number of each _________, but their chemical and physical properties may be quite _____________.

3. In water, organic compounds are generally (soluble) (insoluble) and inorganic compounds tend to be _____________. In organic solvents, organic compounds are generally _________ and inorganic compounds are generally __________ .

4. A hydrocarbon is a compound that contains only two elements, _________ and _________. If a hydrocarbon has only open chains (straight or branched), it is called an _______________ hydrocarbon.

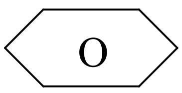
5. In an organic compound, two carbon atoms may share only one pair of electrons. This is called a ______ bond. If two carbon atoms share two pairs of electrons, it is called a _________ bond. If two carbon atoms share three pairs of electrons, it is called a __________ bond.

6. An aliphatic hydrocarbon, which has only single bonds between carbon atoms is called an ____________. An aliphatic hydrocarbon with at least one double bond between carbon atoms is called an ________. If an aliphatic hydrocarbon has at least one triple bond between two carbon atoms, the compound is called an __________ .

7. In an aliphatic hydrocarbon, the chains of carbon atoms are open, that is, either ________________ or _____________. If an aliphatic hydrocarbon has only single carbon-to-carbon bonds, the compound is said to be s______________. A saturated compound has only ____________ bonds. Saturated aliphatic hydrocarbons are also called _____________.

8. An unsaturated aliphatic hydrocarbon contains at least one bond, which is ___________ or __________. If it contains at least one double bond, it is called an ______. If it contains as least one triple bond, it is called an __.

9. As a part of their structure, aromatic compounds include a ___________ ________. The number of carbon atoms in a benzene ring is _________



is a comon symbol for the _____ _____.

Compounds including such a structure are called _________________ compounds.

10. Below is a list of organic compounds.

Amide   Carboxylic Acid
Ether   Phenol
Halogenated Hydrocarbon   Amine
Alcohol   Aldehyde
Ester   Ketone

Using this list, decide which class is most appropriate for each of the chemical formulas below. Write the name of the class beside the chemical formula.

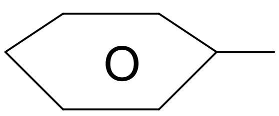
a. $CH_3\text{-}CH_2\text{-}NH_2$ ______________________________

b. $CH_3\text{-}CH_2\text{-}OH$ ______________________________

c. $CH_3\text{-}CH_2\text{-}O\text{-}CH_2\text{-}CH_3$ ______________________________

d. $CH_3\text{-}\overset{\overset{\displaystyle O}{\|}}{C}\text{-}O\text{-}CH_3$ ______________________________

e. $CH_3\text{-}\overset{\overset{\displaystyle O}{\|}}{C}\text{-}H$ ______________________________

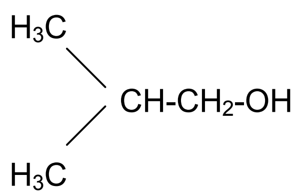
f.  OH ______________________________

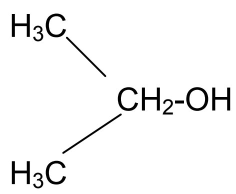
g. $CH_3\text{-}\overset{\overset{\displaystyle O}{\|}}{C}\text{-}CH_3$ ______________________________

h. $CH_3\text{-}CH_2\text{-}CH_2\text{-}Cl$ ______________________________

i. $CH_3\text{-}\overset{\overset{\displaystyle O}{\|}}{C}\text{-}NH\text{-}CH_3$ ______________________________

11. When a hydrogen atom in a hydrocarbon is replaced with a hydroxyl group (OH^-), the resulting compound is an __________. The general form of an alcohol is represented by the symbol __________. CH_3OH is an _________.


is a (primary) (secondary) alcohol.


is a (primary) (secondary) alcohol.

12. If we compare an alcohol with an alkane of similar molecular weight, we find that the boiling point and melting point of the alcohol are (lower) (higher). The solubility of the alcohol is (lower) (higher). These properties are due to the fact that alcohols can form ___________ bonds, both with w_________ molecules and with other _______________ molecules.

13. If an organic compound has a hydrogen atom bonded to an oxygen or nitrogen atom, then that compound can form ___________ bonds between its own molecules.

14. If an organic compound includes a nitrogen or oxygen atom, then the compound can form hydrogen bonds with __________. This makes the compound more __________ in water if its molecular weight is (low) (high). By low molecular weight, we mean that the number of carbons is generally less than ________. Since all of the classes of organic compounds in Table 3-3 (alcohols, phenols, ethers, amines, carboxylic acids, aldehydes, ketones, ester, and amides) have at least one oxygen or nitrogen atom, lower molecular-weight in these classes tend to be _____________ in water.

15. Write yes or no by each of the following classes or compounds to indicate whether compounds of that class generally form hydrogen bonds between their own molecules.

    a. Alcohol (ROH) ____________________

    b. Aldehyde (RCOH) ____________________

    c. Amide •($RCONH_2$) ____________________

    d. Amine (RNH_2) ____________________

    e. Carboxylic acid (RCOOH) ____________________

    f. Ester (RCOOR) ____________________

    g. Ether (ROR) ____________________

    h. Ketone (RCOR) ____________________

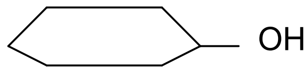
    i. Phenol (ArOH) ____________________

16. If compounds of a single class can form hydrogen bonds between their own molecules, then the boiling point of that compound is (lower) (higher) than the alkane of similar molecular weight. Compounds with one particular functional group are able to form two hydrogen bonds between a pair of its molecules; this functional group is ________________. Thus, carboxylic acids tend to have a very (high) (low) boiling point, compared to alkanes of similar molecular weight. As a result, carboxylic acids are usually (solid) (liquid).

17. The boiling points of ethers, aldehydes, ketones, and esters are (about the same as) (more than) the boiling points of alkanes of similar molecular weight.

18. The boiling points of alcohols, phenols, amines, amides, and carboxylic acids are (lower) (higher) than the boiling points of alkanes of similar molecular weight.

19. Label each of the following compounds as to whether it is acidic, basic, or neutral.

a.  ________________

b. $CH_3\text{-}CH_2\text{-}OH$ ________________

c. $CH_3\text{-}O\text{-}CH_3$ ________________

d. $CH_3\text{-}CH_2\text{-}NH_2$ ________________

e. $CH_3\text{-}\overset{\overset{\displaystyle O}{\|}}{C}\text{-}OH$ ________________

20. Of the compounds in the exercise above, salts would be mostly easily formed with the ones lettered ___, ___, and ___. The two classes of organic compounds which will form salts when reacted with inorganic bases are _________ and ___________ __________. The class of organic compounds which will form salts when reacted with inorganic acids is _________. This property of carboxylic acids and amines is most useful in pharmacy because it helps to make organic compounds that are very ________________ in water. Such salts tend to be quite soluble even when the molecular weight is high. They can thus be more easily utilized by the body.

21. In organic chemistry, a molecule is oxidized if it (gains) (loses) oxygen atoms or if it (gains) (loses) hydrogen atoms. Reduction is the opposite of oxidation. In organic chemistry, a molecule is reduced if it (gains) (loses) oxygen atoms or (gains) (loses) hydrogen atoms.

22. $$\underset{\text{ethyl alcohol}}{CH_3\text{-}CH_2\text{-}OH} \xrightarrow{[O]} \underset{\text{acetaldehyde}}{CH_3\text{-}\overset{\overset{\displaystyle O}{\|}}{C}\text{-}H} \xrightarrow{[O]} \underset{\text{acetic acid}}{CH_3\text{-}\overset{\overset{\displaystyle O}{\|}}{C}\text{-}OH}$$

As illustrated above, the first step in the oxidation of a primary alcohol is the formation of an ___________________________. The second step results in the formation of an _________________________________ ___________.

```
        OH                      O
         ||       [O]           ||
23.  CH3-CH-CH3  -------->  CH3-C-CH3
     isopropyl               acetone (dimethyl)
     alcohol                 ketone)
```

As illustrated above, the oxidation of a secondary alcohol results in the formation of a ____________ .

24. The first step in reduction of a carboxylic acid results in the formation of an ____________ . The second step results in the formation of an _____________ .

25. The reduction of ketones result in the formation of a _____________ ___________ .

```
26.     O                         H+        O
        ||                                  ||
   CH3-C-O-CH2-CH3  +  H2O  ----->  CH3-C-OH
   ethyl acetate                    acetic acid

                    +  CH3CH2OH

                      ethyl alcohol
```

The above reaction is an example of (oxidation) (reduction) (hydrolysis). The general form of an ester is _____________.

The ester in the above reaction is ___________ __________. Hydrolysis of an ester results in the formation of an ____________ and a ___________ ___________. We know that a carboxylic acid can be reduced to form another ___________. Therefore, the combined processes of hydrolysis and reduction of an ester results in the formation of two ____________.

Hydrolysis is the splitting of a compound with the incorporation of ___________ to form (two) (three) new compounds.

27.

$$CH_3\text{-}\overset{\overset{\large O}{||}}{C}\text{-}OH \; + \; CH_3CH_2OH$$

acetic acid ethyl alcohol

$$\xrightarrow{H+} CH_3\text{-}\overset{\overset{\large O}{||}}{C}\text{-}O\text{-}CH_2CH_3 \; + \; H_2O$$

ethyl acetate

Ethyl acetate is an example of the class of compounds called ____________. Therefore, the reaction above can be called ester ________________ or ____________________. It is almost exactly the opposite of a reaction between an ester and water called _______________. Esterification is an important property of both alcohols and carboxylic acids.

28.

$$CH3\text{-}\overset{\overset{\large O}{||}}{C}\text{-}NH2 \; + \; H2O \; \text{------->} \; CH3\text{-}\overset{\overset{\large O}{||}}{C}\text{-}OH$$

acetamide acetic acid

+ H-NH2
ammonia

The above reaction is an example of _____________. Acetamide is an example of the class of compounds called __________. Ammonia may be considered the simplest example of the class of compounds called ___________. Hydrolysis of an amide results in the formation of a ____________ ________ and an ____________.

29.

$$CH_3\text{-}\overset{\overset{\large O}{||}}{C}\text{-}OH \; + \; CH_3\text{-}NH_2 \; \text{----->} \; CH_3\text{-}\overset{\overset{\large O}{||}}{C}\text{-}NH\text{-}CH_3 \; + \; H_2O$$

acetic acid methylamine methylacetamide

Methylacetamide is an example of the class of compounds called ____________. Therefore, the reaction above can be called _____________ ______________. Whereas a carboxylic acid and an alcohol may react to form an _______________, a carboxylic acid and an amine may react to form an ________. This reaction is the opposite of a reaction between an amide and water called ______________. Amide formation is an important property of both ____________ _______ and ___________.

30. As hydrogen atoms in a hydrocarbon molecule are replaced by halogen atoms, the flammability of the molecule (decreases) (increases).

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 3

1. ionic
   covalent
   shared (para 3-4, Table 3-1)

2. structure
   elements, atom, different (para 3-3)

3. insoluble; soluble
   soluble, insoluble (Table 3-1)

4. hydrogen; carbon
   aliphatic (para 3-6)

5. single
   double
   triple (para 3-5)

6. alkane
   alkene
   alkyne (para 3-5)

7. straight, branched
   saturated;
   single
   alkanes (para 3-6a)

8. double, triple
   alkene
   alkyne (para 3-6a(2))

9. benzene ring
   six
   benzene ring
   aromatic (para 3-6b)

10. a Amine (para 3-11)
    b Alcohol (para 3-8)
    c Ether (para 3-10)
    d Ester (para 3-15)
    e Aldehyde (para 3-13)
    f. Phenol (para 3-9)
    g Ketone (para 3-14)
    h Halogenated hydrocarbon (para 3-17)
    i Amide (para 3-16)

11. alcohol
ROH
alcohol
primary
secondary (para 3-8)

12. higher;
higher
hydrogen, water, alcohol (para 3-8a)

13. hydrogen (para 3-18)

14. water
soluble low
five
soluble (para 3-18)

15 a yes (para 3-8a; Table 3-3)
b no (para 3-13; Table 3-3)
c yes (para 3-16a; Table 3-3)
d yes (para~3-11a; Table 3-3)
e yes (para 3-12a; Table 3-3)
f no (para 3-isa; Table 3-3)
g no (para 3-l0a; Table 3-3)
h no (Table 3-3)
i yes (Table 3-3)

16. higher
$$-\overset{\overset{\displaystyle O}{\|}}{C}-OH$$
(carboxyl)
high;
solid (paras 3-8a, 3-12a;Table 3-3)

17. about the same as (Table 3-3)

18. higher (Table 3-3)

19 a acidic
b neutral
c neutral
d basic
e acidic (Table 3-3)

20. a, d. ephenols; carboxylic acids,
amines
soluble (paras 3-9b, 3-11b, 3-12)

21. gains, loses
loses; gains (paras 3-8b; 3-12b(1))

22. aldehyde
carboxylic acid (para 3-8b(1))

23. ketone (para 3-8b(2))

24. aldehyde
alcohol (para 3-12b(1))

25. secondary alcohol (para 3-14b)

26. hydrolysis

$$\begin{array}{c} \quad O \\ \quad \| \\ R\text{-}C\text{-}O\text{-}R \end{array}$$

ehyl acetate
alcohol; carboxylic acid
alcohol
alcohols
water; two (para 3-15)

27. esters
formation; esterification
hydrolysis (paras 3-12b(2); 3-15)

28. hydrolysis
amides
amines
carboxylic acid, amine (para 3-16)

29. amides
amide formation
ester, amide
hydrolysis
carboxylic acids, amines (para 3-12b(3); 3-16)

30. decreases (para 3-17)

End of Lesson 3