DOCUMENT RESUME

ED 070 916 AC 014 058

TITLE	Atomic Structure and Valence: Level II, Unit 10, Lesson 1; Chemical Bonding: Lesson 2; The Table of Elements: Lesson 3; Electrolysis: Lesson 4. Advanced General Education Program. A High School Self-Study Program.
INSTITUTION	Manpower Administration (DOL), Washington, D. C. Job Corps.
REPORT NO	PM-431-73; PM-431-74; PM-431-75; PM-431-76
PUB DATE	Nov 69
NOTE	72p.
EDRS PRICE	MF-$0.65 HC-$3.29
DESCRIPTORS	*Academic Education; Achievement Tests; Atomic Structure; *Autoinstructional Aids; Chemical Bonding; Chemical Nomenclature; Chemical Reactions; Chemistry; *Course Content; Credit Courses; *General Education; *Independent Study; Secondary Grades

ABSTRACT

This self-study program for high-school level contains lessons on: Atomic Structure and Valence, Chemical Bonding, The Table of Elements, and Electrolysis. Each of the lessons concludes with a Mastery Test to be completed by the student. (DB)

PM 431 - 73

ED 070916

ADVANCED
GENERAL EDUCATION PROGRAM

A HIGH SCHOOL SELF-STUDY PROGRAM

ATOMIC STRUCTURE AND VALENCE

LEVEL: II
UNIT: 10
LESSON: 1

U.S. DEPARTMENT OF LABOR
MANPOWER ADMINISTRATION, JOB CORPS
NOVEMBER 1969

U.S. DEPARTMENT OF LABOR
MANPOWER ADMINISTRATION, JOB CORPS
NOVEMBER 1969

You have already learned some basic facts about the structure of different atoms. You have also seen that the reactivity of atoms depends upon their structures.

In the following lesson units you will study the relation between the structure and the degree of chemical reactivity of different atoms in greater detail and how this influences the physical properties of the compounds produced. You will gain a basic understanding of the chemical characteristics which are common to all elements in certain categories, such as the metallic elements. In addition, you will study some of the useful processes involved in chemical reactions between different substances.

The material in the following lesson units is presented in prose sections which are separated by questions. The prose and questions resemble the material which you can expect to find in high school equivalency examinations. However, in those examinations the questions are all found at the end of a complete prose selection.

READ each prose section carefully, then READ the questions which follow it. First try to answer the questions without looking through the preceding prose. If you do not know the answer to one or more questions READ through the prose again, keeping in mind the questions which you have to answer. You should always answer each question, even if you can only guess at the correct answer.

After you have answered all of the questions in a group, turn to the correct answers given on the next page. Check your answers. If your answers are wrong and you cannot understand the correct answers, look through the prose again to see what you have missed.

Never consult the correct answers before you give your own answer to each of the questions -- even if you must guess an answer.

ANSWER the following questions. Try to give the correct answers without referring to the passage above.

1. The following lesson units all deal with:

   a. ☐ only the chemical properties of different substances

   b. ☐ only the physical properties of different substances

   c. ☐ the physical and chemical properties of different substances

2. According to the instructions in the prose passage, you should:

a. ☐ read the entire lesson unit before reading each set of questions

b. ☐ read each prose section and try to answer the questions which follow it before proceeding to the next prose section

c. ☐ try to answer the questions at the end of each prose section before you read the prose

3. According to the instructions, if you have difficulty in answering the questions following the first reading of a prose passage, you should:

a. ☐ guess at the correct answers without reading the passage again

b. ☐ leave those questions blank and go on to the next passage

c. ☐ look at the correct answers given on the page following the questions

d. ☐ read the passage again to gain a better understanding and then try to answer the questions

ANSWERS:

1. c

2. b

3. d

As you have learned, atoms consist of positively charged protons, negatively charged electrons, and neutrons, which are neither positively nor negatively charged. The weight of an atom depends on the number of protons and neutrons in the nucleus. The chemical properties of an element depend on the number and arrangement of the atom's electrons.

In previous lessons, we said the electrons circle the nucleus in orbits. We refer to these orbits by using the letters of the alphabet from K to Q. The electrons are kept in their orbits by the force of attraction exerted by the positively charged nucleus. The further an orbit is from the nucleus, the more room it has for electrons, but the more loosely those electrons are held. For example, the first orbit can hold only 2 electrons, but these electrons are more strongly attracted to the nucleus than 2 electrons in the fourth or N orbit. Although the electrons in inner orbits are held so tightly that they can only be removed by exerting a great amount of energy, the electrons in the outermost shell are held so loosely that they are removed in chemical reactions. In fact, chemical reactions always involve the electrons in the outermost orbit; sometimes, the next inner orbit is also involved. The outermost orbit is called the valence orbit and the electrons in it are referred to as the valence electrons.

As you know, atoms with 8 electrons in their outermost orbits have a greater degree of stability. They are relatively nonreactive. Atoms that do not have 8 electrons in their outermost orbits are not as stable. The atoms of these elements react in order to gain the extra electrons they need to have 8 valence electrons, or they react in order to lose the valence electrons they already have. The number of electrons to be gained or lost is called the valence number. The valence number is positive if electrons are lost and negative if electrons are gained. For example, an atom of bromine (whose elemental symbol is Br) has 7 electrons in its valence orbit. A bromine atom needs to take 1 electron away from another atom. It therefore has a valence of -1. The valence number can be indicated in the upper right-hand corner of an element's symbol; thus, the fact that bromine, Br, has a valence of -1 may be shown as Br^{-1}. An oxygen atom (symbol 0) has 6 electrons in its outermost orbit; it needs 2 more electrons. Thus, its valence number is -2. The symbol 0^{-2} shows the valence number of an oxygen atom. Sodium atoms, on the other hand, have 1 electron in their outer orbits. The sodium atom (symbol Na) loses its extra electron in a chemical reaction; therefore, its valence is +1, written Na^{+1}. Atoms that have gained or lost one or more electrons have a charge and are called ions. Atoms that have gained electrons have a negative charge and are called anions; while atoms that have lost electrons have a positive charge and are called cations.

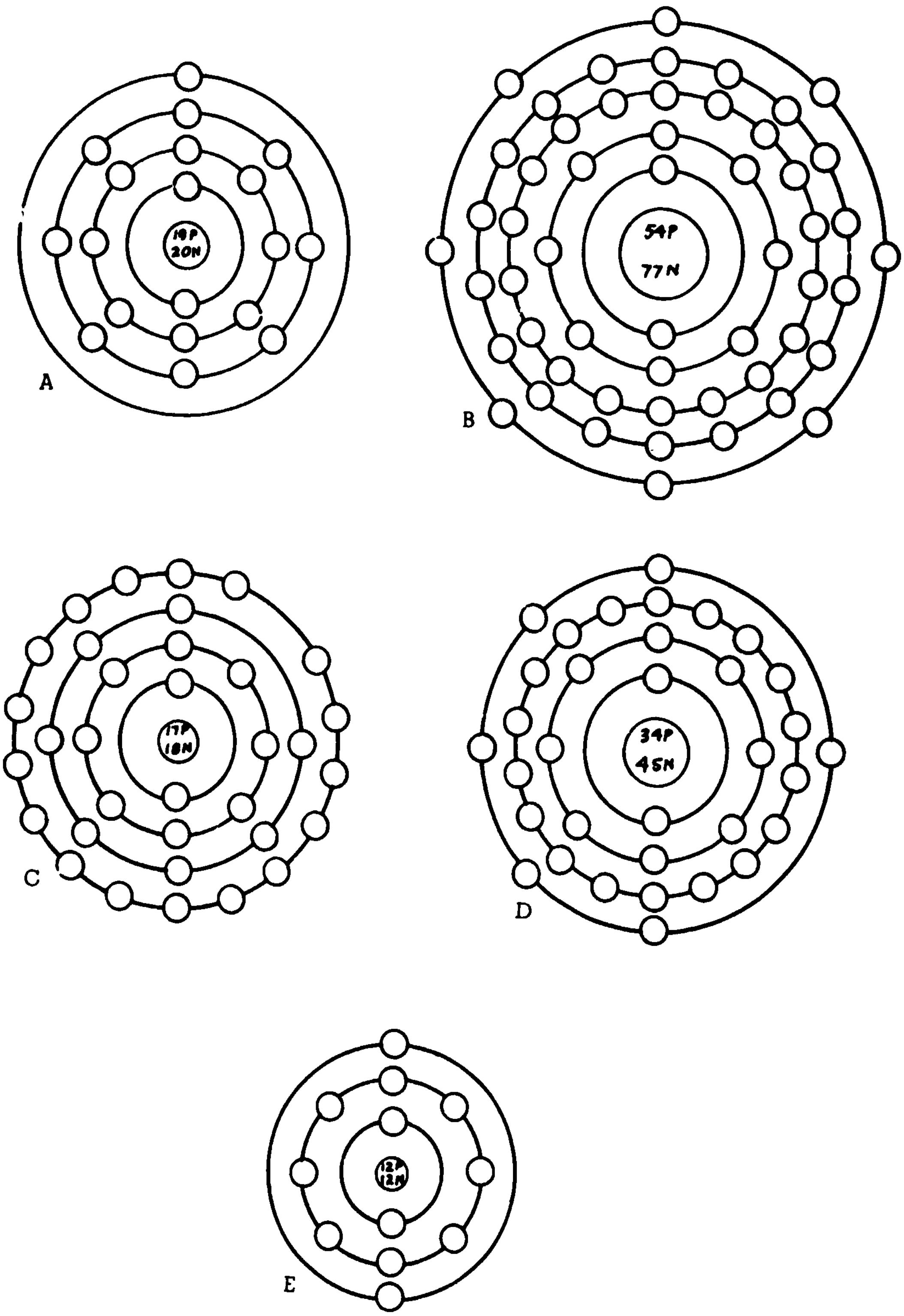
1. The valence orbit is:

   a. ☐ the innermost orbit

   b. ☐ next to the outermost orbit

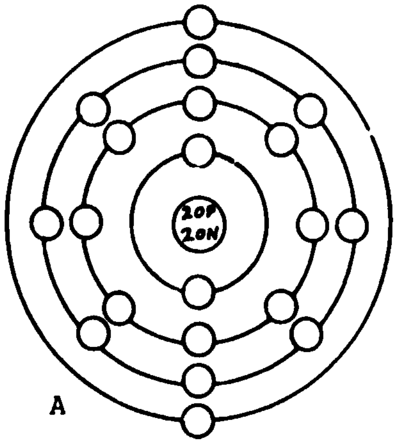
   c. ☐ the outermost orbit

2. REFER to the drawings below. How many valence electrons does each atom represented have?

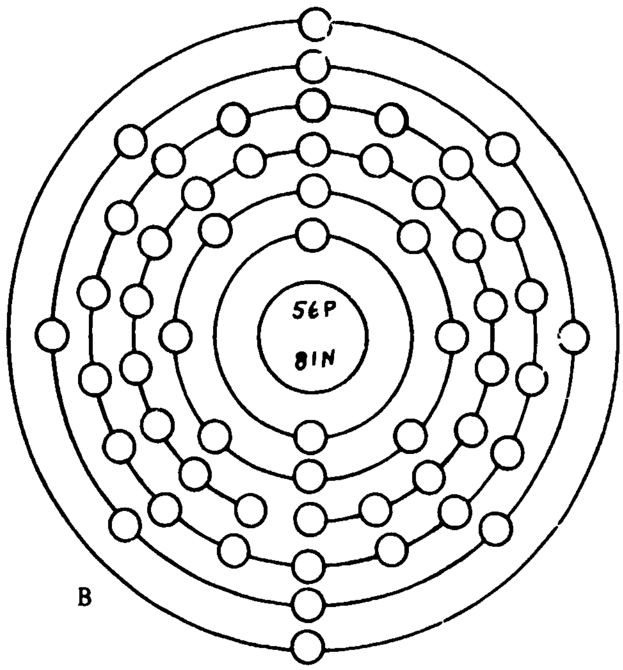


NOTE: THESE ARE DRAWINGS OF FICTIONAL ATOMS.

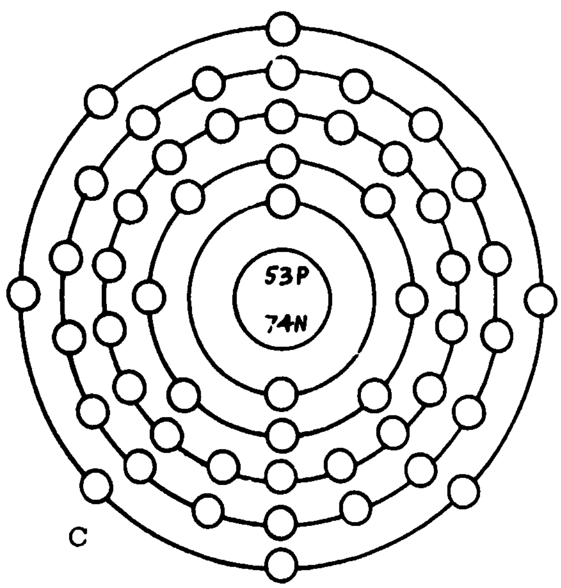
3. What is the valence number of the atoms represented below.



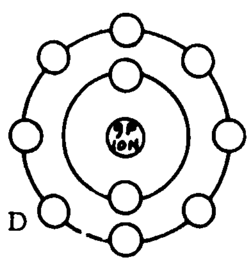
A



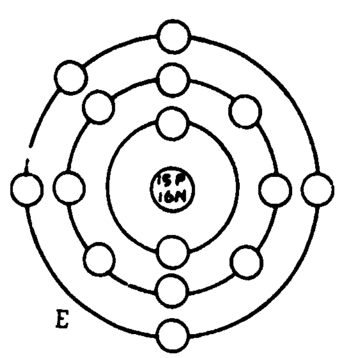
B



C



D



E

4. MATCH the following columns:

A. positively charged atom

B. negatively charged atom

C. atom without a charge

D. neutral atom with 2 electrons in its valence orbit

E. neutral atom with 7 electrons in its valence orbit

1. ____ anion

2. ____ cation

3. ____ ion

ANSWERS:

1. c

2. Atom A - 1
   Atom B - 8
   Atom C - 7
   Atom D - 6
   Atom E - 2

3. Atom A - +2
   Atom B - +2
   Atom C - -1
   Atom D - -1
   Atom E - -3

4. 1. B, 2. A, 3. A, B

When two atoms react so that one becomes an anion and the other a cation, the ions formed attract each other strongly because they have opposite charges. This electrical attraction, called electrovalence or an electrovalent bond, holds the ions together. Thus, compounds made up of ions are called electrovalent compounds or ionic compounds.

Atoms with 1 or 2 electrons to be gained or lost are very active and have a great tendency to form stable ionic compounds. Atoms with 3, 4, or 5 electrons are not as active and do not form ionic compounds. Although it is not hard to take a valence electron from an atom, it is difficult to remove another electron from a positive ion. If an atom has lost 2 electrons, it is twice as hard to take away another. So atoms with 3 or more electrons in the valence shell do not lose them all. The same principle holds true for adding electrons to atoms. It is easy to add 1 electron, harder to add 2, and harder still to add 3 or 4. Therefore, atoms with 3, 4, and 5 electrons do not gain or lose electrons and, thus, do not form ions. Instead, these atoms share electrons with other atoms. In other words, two atoms attracted to each other by their need for extra valence electrons, come together so that their outer orbits are in contact. The shared electrons (one or more from each atom) become associated with each other and, as a pair, they are attracted to the positively charged nuclei of both atoms at the same time. This is what we mean when we say that atoms are sharing electrons. Compounds formed by atoms sharing one or more pairs of electrons are called covalent compounds, and the attraction holding the atoms together is called covalence or covalent bonding. Covalent bonding is easier to understand if you think of the nuclei of each atom holding on to the same electrons like two men holding on to a chain. Unless the chain breaks, the two men will be bound together until one of them lets go. If neither lets go, they remain tied to each other.

For electrovalent compounds, there is a simple connection between the valences of the atoms and the number of atoms forming the compound. The valence number of each atom donating electrons multiplied by the number of atoms donating electrons gives the total positive charge. Similarly, the valence number of each atom gaining electrons multiplied by the number of atoms gaining electrons gives the total negative charge. Since the compound has to be electrically neutral, the number of the positive charge and the number of the negative charge must be the same. For example, when calcium (symbol Ca), which has a valence of +2, reacts with chlorine (symbol Cl), which has a valence of -1, the resulting compound, calcium chloride, must contain 1 atom of calcium and 2 atoms of chlorine because it takes 2 atoms of chlorine (each has a charge of -1) to neutralize the one atom of calcium (which has a charge of +2). The resulting formula is written $CaCl_2$ (the small number below the line in the formula, the subscript, indicates that there are 2 atoms of Cl in the compound). We can write the symbol for calcium showing its valence number; it is Ca^{+2}. The symbol for chlorine showing its valence number is Cl^{-1}. We can now construct a simple mathematical equation to show the relationship between the atoms and their valence numbers and the charge on the resulting compound.

Calcium with a valence of +2 times the number of calcium atoms, 1; plus chlorine with a valence of -1 times the number of chlorine atoms, 2; yields a total charge of 0.

$$(+2) \times 1 + (-1) \times 2 = 0$$
$$+2 + (-2) = 0$$

Let's consider another example. When sodium (symbol Na), with a valence of +1, reacts with chlorine (symbol Cl), again, valence -1, the compound formed is NaCl. Thus, when 1 atom of sodium reacts with 1 atom of chlorine, the total positive charge, +1, is neutralized by the total negative charge, -1. If you know the valences of the atoms, you can predict the formula of the resulting compound; or, if you have the formula, you can tell the valences. In later lessons, you will learn to find the valences from a specially arranged chart of the elements called a periodic table.

5. Anions and cations are held together:

   a. ☐ only by an electrovalent bond

   b. ☐ only by a covalent bond

   c. ☐ by either a covalent or an electrovalent bond

6. If an atom has 4 valence electrons, it will react with other atoms by:

   a. ☐ gaining electrons

   b. ☐ losing electrons

   c. ☐ sharing electrons

7. In a covalent bond, atoms may:

   a. ☐ share only 1 pair of electrons

   b. ☐ share 1 or more pairs of electrons

   c. ☐ gain or lose electrons

8. The symbols of several elements with the valence number of the atom shown in the upper right-hand corner are given below. Keeping in mind that the positive charge of a compound must neutralize the negative charge, WRITE the chemical formula of the compound indicating the number of atoms of each element. (Note: it is not necessary to use a subscript when only one atom of an element is involved in the formula.) WRITE the correct chemical formula:

H^{+1} combined with Cl^{-1} __________

H^{+1} combined with S^{-2} __________

Fe^{+2} combined with Cl^{-1} __________

H^{+1} combined with O^{-2} __________

Al^{+3} combined with O^{-2} __________

ANSWERS:

5. a

6. c

7. b

8. HCl

   H_2S

   $FeCl_2$

   H_2O

   Al_2O_3

Atoms that tend to take up electrons are called oxidizing agents. The name comes from the fact that when oxygen attacks a metal, forming an oxide, it takes electrons from the metal. Atoms that give up electrons are called reducing agents. Reactions in which one atom is oxidized and the other is reduced a-e called oxidation-reduction reactions. Every oxidation must be accompanied by a reduction since there cannot be a winner of electrons without a loser. (Acids are often defined as substances which need electrons, and bases are defined as having excess electrons. An oxidizing agent is, therefore, also an acid and a reducing agent is also a base.) Metals such as sodium and magnesium have 1 or 2 electrons in their valence orbit so they tend to lose electrons. They are reducing agents. Nonmetals, such as chlorine and sulfur need 1 or 2 electrons to have 8 electrons in their outer orbit, so they tend to gain electrons. They are oxidizing agents. In the next lessons, you will learn how to predict when oxidation-reduction reactions will occur, and how to predict some of the properties of the compounds resulting from these chemical reactions.

9. MATCH the following columns:

A. gain electrons	1. _____ acids
B. lose electrons	2. _____ bases
	3. _____ metals
	4. _____ nonmetals
	5. _____ oxidizing agents
	6. _____ reducing agents

10. A. Sodium (symbol Na) has a valence of +1. Sodium is:

a. ☐ a metal

b. ☐ a nonmetal

B. Beryllium (symbol Be) has a valence of +2. Beryllium is:

a. ☐ an oxidizing agent

b. ☐ a reducing agent

10. C. Sulfur (symbol S) has a valence of -2. Sulfur is:

a. ☐ a metal

b. ☐ a nonmetal

D. Iodine (symbol I) has a valence of -1. Iodine is:

a. ☐ an oxidizing agent

b. ☐ a reducing agent

Time completed ______________

ANSWERS:

9. 1. A

   2. B

   3. B

   4. A

   5. A

   6. B

10. A. a

    B. b

    C. b

    D. a

WHEN YOU HAVE FINISHED THIS LESSON, WRITE DOWN THE TIME. THEN TAKE THE LESSON TO YOUR INSTRUCTOR OR HIS ASSISTANT FOR CHECKING. WAIT UNTIL THE LESSON IS APPROVED BEFORE GOING ON TO THE NEXT LESSON.

U. S GOVERNMENT PRINTING OFFICE 1970 O - 369-281

PM 431 - 74

ADVANCED GENERAL EDUCATION PROGRAM

A HIGH SCHOOL SELF-STUDY PROGRAM

CHEMICAL BONDING
LEVEL: II
UNIT: 10
LESSON: 2

U.S. DEPARTMENT OF LABOR
MANPOWER ADMINISTRATION, JOB CORPS
NOVEMBER 1969

U.S. DEPARTMENT OF LABOR
MANPOWER ADMINISTRATION, JOB CORPS
NOVEMBER 1969

Let us now consider the forces holding material together, and how these forces affect chemical properties.

Two atoms can share electrons to form a covalent bond, or one atom can take one or more electrons from another atom to form an electrovalent compound. For two atoms to share a pair of electrons, each nucleus must attract the pair of electrons with a force that is just about as strong as that of the other. The size of this force depends on two factors: the number of charges (protons) in the nucleus and the distance between the nucleus and the shared electrons. Remember, the further the nucleus is from the electrons, the weaker the attraction. So large atoms, those with a high atomic number, whose nuclei are far away from the outermost orbit, attract their valence electrons only weakly. On the other hand, the more protons (positive charges) in the nucleus, the greater the force of attraction for the electrons. In the next lesson, you will learn how to tell which atoms have large radii (the distance between the center of the nucleus and the outermost orbit) and which have highly charged nuclei. At present it is sufficient to know that the atomic radius and the charge of the nucleus vary so that atoms of different elements have different degrees of attraction for electrons. We call this attraction electronegativity. The reason for the name is that atoms attracting electrons away from other atoms become negatively charged.

If the electronegativities of two atoms are the same, or almost the same, the two atoms attract the electrons with equal force, and the bond is covalent. As you know, two or more atoms of the same element sometimes form molecules. Examples of these elemental molecules are oxygen (O_2), nitrogen (N_2), and hydrogen (H_2). Each molecule of these substances consists of two identical atoms; and thus the electronegativities are exactly the same, and the bond is covalent.

In some molecules in which the two atoms are not alike, such as carbon dioxide (CO_2), water (H_2O), and benzene (C_6H_6), the difference in electronegativities is so small that the electrons are still shared just about equally, so that the bond is still covalent.

On the other hand, if the difference in electronegativity is great, the electrons are not shared but are pulled completely away from one atom to the other. Thus, one atom completely loses its electron and becomes a positive ion. In other words, the atom is oxidized. The other atom gains an electron and becomes a negative ion. In other words, this atom is reduced. (You will remember that oxidation is the losing of one or more electrons and reduction is the gaining of one or more electrons.) For example, in the reaction:

$$Na + Cl = Na^+ + Cl^-$$

the sodium atom (Na) has been oxidized by the chlorine atom (Cl) and the chlorine has been reduced by the sodium.

1. All atoms have:

   a. ☐ equal numbers of protons and electrons

   b. ☐ the same atomic radius

   c. ☐ the same electronegativities

   d. ☐ the same number of protons

2. Electronegativity refers to:

   a. ☐ an atom's degree of attraction for electrons

   b. ☐ a positively charged atom

   c. ☐ the distance between the center of the nucleus and the valence electrons

   d. ☐ the number of protons in the nucleus

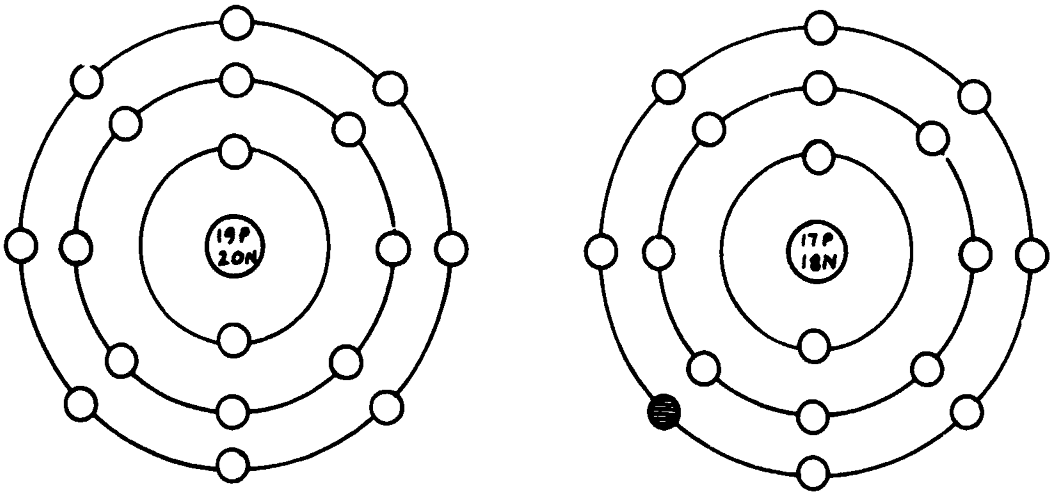
3. Two atoms of different elements held together by a covalent bond:

   a. ☐ have different electronegativities

   b. ☐ have similar electronegativities

   c. ☐ may have either similar or different electronegativities

4. LOOK AT the diagram below. The drawing represents the potassium atom after it has donated an electron to the chlorine atom to form the compound, potassium chloride.



Which statement can you correctly make concerning the electronegativities of potassium and chlorine?

a. ☐ The electronegativity of chlorine is greater than that of potassium.

b. ☐ The electronegativities of potassium and chlorine are the same.

c. ☐ The electronegativity of potassium is greater than that of chlorine.

d. ☐ Nothing can be determined about the electronegativity of either potassium or chlorine.

ANSWERS:

1. a

2. a

3. b

4. a

23

Sometimes the atoms of different elements are bound together by strong covalent bonds, so that they act as if they were a single ion with one valence number. An example of this is SO_4^{-2}; here, one sulfur atom (S^{-2}) shares electrons with 4 oxygen atoms (O^{-2}) to form a group of atoms with a valence of -2. A group of atoms held together by covalent bonds, behaving almost like a single ion, is called a radical. Other examples of radicals are the carbonate radical, CO_3^{-2}, the phosphate radical, PO_4^{-3}, and the permanganate radical, MnO_4^{-1}.

All reactions in which ions or radicals are formed are oxidation-reduction reactions. All oxidation-reduction reactions take place because there are large differences in the electronegativities of the reacting atoms. However, not all oxidation-reduction reactions produce ions or radicals. In some cases, ions or radicals are oxidized or reduced to form neutral atoms or molecules. For example, the chloride ion, Cl^{-1}, is oxidized by the permanganate radical, MnO_4^{-1}, in acid solution, forming chlorine, Cl_2, and manganese dioxide, MnO_2, both of which are neutral. The chloride ion is oxidized and the permanganate radical is reduced.

5. MnO_4^{-1} is:

   a. ☐ an ion that behaves like a radical

   b. ☐ an ion that can be reduced

   c. ☐ a radical that can be oxidized

   d. ☐ a radical that can be reduced

6. The carbonate radical, CO_3^{-2}, is held together:

   a. ☐ because the electronegativities of the elements involved are different

   b. ☐ by a covalent bond

   c. ☐ by an electrovalent bond

   d. ☐ by neither an electrovalent nor a covalent bond because the radical has a valance number

24

ANSWERS:

5. d

6. b

25

There is one other important type of bond between atoms, in addition to the covalent and electrovalent bonds. This is the metallic bond, which gives metals their unusual and important properties. In the metallic bond, electrons are neither shared by two atoms at the same time nor are they transferred completely from one atom to another. Instead, they wander freely back and forth from atom to atom. In a piece of metal, the valence electrons move around freely, so that at one moment an atom may be without a valence electron; at the next moment, it may be neutral or have two extra electrons and be negative. Because of this wandering of electrons, there are many positive ions in the metal at any given time and these positive ions attract the electrons moving around. This attraction constitutes the metallic bond which holds the metal together. The proof that electrons move freely in metals lies in the fact that metals conduct electricity. An electric current is a flow of charge. Since ions are relatively large and cannot move through solids, the charge moving through the metals must be due to the electrons. Another proof is the fact that when metal filaments are heated, they give off electrons. This phenomenon is the basis of the electronic tube.

7. An elemental molecule is held together by:

   a. ☐ a covalent bond

   b. ☐ an electrovalent bond

   c. ☐ a metallic bond

   d. ☐ none of the above

8. A substance held together by a metallic bond contains:

   a. ☐ free electrons and positive ions, but not negative ions

   b. ☐ only negative ions and free electrons

   c. ☐ only positive and negative ions

   d. ☐ positive ions

ANSWERS:

7. a

8. d

Another type of attraction we should consider is the attraction between molecules. Intermolecular attractions can be due to a variety of forces, all of which are weaker than covalent bonding, ionic bonding, and metallic bonding. As a result, the forces between molecules are weak while the forces between atoms and between ions are strong. For example, it does not take much energy to separate O_2 from other molecules. At very low temperatures, oxygen boils, going from a liquid, which consists of molecules packed tightly together, to a gas, which consists of molecules more widely separated. On the other hand, to break the O_2 molecule down into 2 single oxygen atoms takes a lot of energy. Oxygen does not break down until it is heated hundreds of degrees above normal temperatures.

The physical properties of a substance depend largely on the type of bond that holds it together. When a substance is made up of molecules held together by a covalent bond, such as carbon dioxide molecules or benzene molecules, the forces holding it together are weak. These substances melt at a low temperature, boil easily, dissolve readily in solvents, and, in the solid state, are soft and break easily. When a substance consists only of atoms held together by a metallic bond, as iron does, the substance is harder and stronger, has a high melting point, does not boil until very high temperatures are reached, and does not dissolve in most solvents. When a material is held together by electrovalent bonds, as sodium chloride is, it too is hard and strong, does not melt easily, and does not boil at all; but it does dissolve in certain solvents, such as water. These will be discussed in a later lesson.

One very good way to distinguish electrovalent compounds, covalent compounds, and metals from one another is to test their ability to conduct electricity, both in the solid and in the liquid states. Covalent compounds do not form ions and do not conduct electricity at all, either in the solid or in the liquid states (unless they are combined with water or another liquid in which they form ions). Metals have free electrons which move about in both liquids and solids, so metals can conduct electricity in both the liquid and the solid states. Electrovalent compounds conduct electricity when they are melted, but not when they are in the solid state. This is because electrovalent compounds consist of ions which cannot move in the solid state, but which can move when given sufficient heat energy, that is, when melted.

9. Molecules can be separated at a low boiling point because:

a. ☐ all molecules held together by covalent bonds have a low boiling point

b. ☐ the intermolecular force of attraction is stronger than a covalent bond

c. ☐ the molecule does not break apart at a low boiling point

d. ☐ they are held together by covalent bonds

10. MATCH the columns below:

A. covalent compounds

B. electrovalent compounds

C. metals

1. _____ conduct electricity in the liquid state

2. _____ conduct electricity in the solid state

3. _____ do not conduct electricity in the liquid state

4. _____ do not conduct electricity in the solid state

Time completed ______________

ANSWERS:

9. a

10. 1. B, C

    2. C

    3. A

    4. A, B

WHEN YOU HAVE FINISHED THIS LESSON, WRITE DOWN THE TIME. THEN TAKE THE LESSON TO YOUR INSTRUCTOR OR HIS ASSISTANT FOR CHECKING. WAIT UNTIL THE LESSON IS APPROVED BEFORE GOING ON TO THE NEXT LESSON.

☆ U. S. GOVERNMENT PRINTING OFFICE 1970 O - 369-882

U.S. DEPARTMENT OF LABOR
MANPOWER ADMINISTRATION, JOB CORPS
NOVEMBER 1969

When a flower is placed in water, it appears to be refreshed, but if the same flower were placed in a strong solution of sugar, it would quickly become wilted in appearance. This fact is due to a process called <u>osmosis</u>. Osmosis plays an important role in the life processes of both plants and animals.

Osmosis is caused by the phenomenon called <u>diffusion</u>, which in turn is caused by the movement of molecules. At all temperatures, molecules move.* In solids, molecules vibrate back and forth; in other words, they stay in the same place but are in constant motion in that place. In gases and liquids, however, the molecules move around from place to place at very high speeds, constantly bumping into each other and bouncing off in all directions. For example, in air, the average rate of movement of the molecules is 700 miles an hour. On the average, one molecule collides with other molecules at a rate of 100,000 times a second. Every time there is a collision, the molecules bound off in different directions. In regions where there are fewer molecules, there are fewer collisions; so the molecules can travel faster and farther apart where there are fewer of them.

If some gas is put into one end of a long tube and there is a vacuum at the other end, the gas molecules, since they go faster in the direction where there are fewer of them, will move towards the vacuum, filling it. Very rapidly, the concentration of gas molecules will be the same all over the tube. A similar process can take place in a liquid, although the molecules of a liquid move more slowly. If a dye is put in a tube of liquid, the dye molecules, by bumping around, slowly distribute themselves uniformly throughout the liquid, just as if the tube were stirred or shaken. The process whereby the molecules in a liquid or a gas move around until they are distributed uniformly is called diffusion. Molecules always diffuse from a region of high concentration toward a region of low concentration.

1. Molecules of solids:

   a. ☐ are in motion only above the boiling point

   b. ☐ are never in motion

   c. ☐ are sometimes in motion

   d. ☐ vibrate back and forth but do not diffuse through the solid

*Theoretically, if the temperature of a substance could be lowered to -273° C., the molecules of the substance would no longer be moving. For this reason, a temperature of -273° C. is referred to as absolute zero. The theory that all molecular motion would cease at -273° C. was proposed by Lord Kelvin.

ANSWERS:

1. d

2. d

3. b

ANSWERS:

4. a

5. c

7. In the process of <u>osmosis</u>:

a. ☐ solute molecules flow through an impermeable membrane to an area of low concentration

b. ☐ solute molecules flow through a permeable membrane

c. ☐ solvent molecules flow across a permeable membrane

d. ☐ solvent molecules flow through a membrane which prevents the flow of solute molecules

Consider our first example of osmosis: when a flower is placed in water, water will pass through the cell membranes (cell membranes are semipermeable membranes) of the petals, and, thus, the petals will lose some of the withered look they may have had and appear fresher. If the flower, however, is placed in a strong sugar solution, the water in the cells of the petals will pass through the cell membranes into the more highly concentrated sugar solution and the flower will appear withered.

There are many other examples of osmosis. Consider a swimmer immersed in salt water: the membrane is the skin and water flows out of the dilute salt solution in the body to the more concentrated solution in the ocean. With the loss of water, the fingers appear shrunken and wrinkled.

Another example of osmosis: Jams and jellies are made with large quantities of sugar. In other words, they are very concentrated solutions of sugar. Any bacteria or yeasts that come in contact with the jams or jellies are killed. Water flows through the semipermeable membrane of the bacterial cells into the concentrated sugar solution. (Remember: a concentrated sugar solution has fewer water molecules than a dilute sugar solution.) When the bacterial cells lose their water, they decompose and die. This is why jams and jellies are called "preserves"; they are preserved from bacterial attack.

8. The fluids in a plant and the water in the soil surrounding the roots both contain dissolved salts. Water flows into the roots of the plant across the cell membranes in the roots. You would guess that:

   a. ☐ the concentration of dissolved salts is greater in the fluids of the plant than in the solutions in the soil

   b. ☐ the concentration of dissolved salts in the plant is the same as the concentration of the water solution in the soil

   c. ☐ the concentration of solutes is greater in the water solution in the soil than in the fluids of the plant

9. Certain types of fish, such as herring, are preserved in a strong salt solution. When bacterial cells come in contact with these salted fish, they are destroyed because:

   a. ☐ salt crosses the semipermeable cell membranes of the bacteria and kills the bacteria

   b. ☐ water osmoses from the salt solution in which the fish is preserved through the semipermeable membrane of the bacterial cells

   c. ☐ water osmoses out of the bacterial cells into the salt solution in which the fish is preserved

PM 431 - 75

ADVANCED GENERAL EDUCATION PROGRAM

A HIGH SCHOOL SELF-STUDY PROGRAM

THE TABLE OF ELEMENTS

LEVEL: II

UNIT: 10

LESSON: 3

U.S. DEPARTMENT OF LABOR
MANPOWER ADMINISTRATION, JOB CORPS
NOVEMBER 1969

U.S. DEPARTMENT OF LABOR
MANPOWER ADMINISTRATION, JOB CORPS
NOVEMBER 1969

GROUPS

PANEL 1

Periodic Table of Elements

PERIODS	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIII A
1	1 H 1.00797							2 He 4.0026
2	3 Li 6.939	4 Be 9.0122	5 B 10.811	6 C 12.01115	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.183
3	11 Na 22.9898	12 Mg 24.312	13 Al 26.9815	14 Si 28.086	15 P 30.9738	16 S 32.064	17 Cl 35.453	18 Ar 39.948
4	19 K 39.102	20 Ca 40.08	31 Ga 69.72	32 Ge 72.59	33 As 74.9216	34 Se 78.96	35 Br 79.909	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.9044	54 Xe 131.30
6	55 Cs 132.905	56 Ba 137.34	81 Tl 204.37	82 Pb 207.19	83 Bi 208.980	84 Po (210)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)						

	III B	IVB	VB	VI B	VII B	VIII B	VIII B	VIII B	IB	IIB
4	21 Sc 44.956	22 Ti 47.90	23 V 50.942	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.71	29 Cu 63.54	30 Zn 65.37
5	39 Y 88.905	40 Zr 91.22	41 Nb 92.906	42 Mo 95.94	43 Tc (99)	44 Ru 101.07	45 Rh 102.905	46 Pd 106.4	47 Ag 107.870	48 Cd 112.40
6	57 La 138.91	72 Hf 178.49	73 Ta 180.948	74 W 183.85	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.09	79 Au 196.967	80 Hg 200.59
7	89 Ac (227)									

No groups designated

	The Rare Earth Elements, or the Lanthanides													
6	58 Ce 140.12	59 Pr 140.907	60 Nd 144.24	61 Pm (145)	62 Sm 150.35	63 Eu 151.96	64 Gd 157.25	65 Tb 158.924	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.97
	The Heavy Rare Earth Elements, or the Actinides													
7	90 Th 232.038	91 Pa (231)	92 U 238.03	93 Np (237)	94 Pu (242)	95 Am (243)	96 Cm (247)	97 Bk (249)	98 Cf (251)	99 Es (254)	100 Fm (253)	101 Md (256)	102 No (253)	103 Lw (257)

1

The Periodic Table and Its Uses

There are innumerable chemical compounds, each with different properties and different uses. It is obviously impossible to remember them all, and so it becomes necessary to organize the information we have. By far, the most useful and important tool for understanding and predicting the properties of elements, and the compounds they form, is the periodic table. The periodic table is a special chart of the elements. On it, the elements are arranged in a way which makes it easy to see which elements will react with each other, if the compound they form will be covalent or electrovalent, and what the formula will be.

The periodic table is shown on Panel 1. Each of the 103 elements known to man is represented, in a separate box, by the symbol of the element. Each box also contains two numbers: the number in the upper left-hand corner is the atomic number of the element. As you know, the atomic number is the number of protons, or positive charges, in the nucleus (and, since an atom has an equal number of negative charges, the number of electrons circling the nucleus). The number at the bottom of the box is the atomic weight of the element. You already know that the atomic weight of an atom is based on the number of protons and neutrons in the nucleus. You may now be wondering why many of the atomic weights are decimals.

All the atoms of an element have the same atomic number, but not all the atoms of an element have the same atomic weight. For example, all the atoms of the first element, hydrogen (symbol H), have at least 1 proton and 1 electron. There are some atoms of the first element, however, that also have 1 neutron in their nucleus; and others still, that have 2 neutrons in their nucleus. The atomic weight of the atoms of the first element that have 1 proton and 1 electron but no neutrons is 1. We call these atoms hydrogen. The atomic weight of the atoms of the first element that have 1 proton, 1 electron, and 1 neutron is 2. We call these atoms deuterium. All the atoms with an atomic number of 1 and an atomic weight of 3 (atoms with 1 proton, 1 electron, and 2 neutrons) are called tritium. Atoms that have the same atomic number but different atomic weights are called isotopes. Usually, the isotopes of an element have the same name; in fact, we can refer to deuterium as an isotope of hydrogen. Most of the elements have isotopes. Actually, most of the elements exist in nature as a mixture of several different isotopes.

The atomic weight of an element given on the periodic table is an average weight of this mixture. For example, the element chlorine (symbol Cl) exists as two isotopes. One of these, with a mass number (atomic weight) of 35, makes up 75.4 percent of all natural chlorine. The other, with a mass number of 37, makes up 24.6 percent of the element. The average atomic weight of chlorine is 35.453. (It is a remarkable fact that all natural chlorine contains these two isotopes in exactly this proportion.) This is why many of the atomic weights are decimals.

Isotopes of an element have different atomic weights but their atomic number (number of protons, and consequently, electrons) is the same. The chemical properties of an element, which depend on the number and arrangement of electrons, are the same for all the isotopes of an element.

1. An atom of deuterium:

   a. ☐ has the same mic weight as an atom of hydrogen

   b. ☐ has the same atomic weight as an isotope of hydrogen, but a different atomic number than hydrogen

   c. ☐ has the same atomic number as an atom of tritium

   d. ☐ has the same atomic number as hydrogen, but not the same atomic number as tritium

2. The isotopes of an element have:

   a. ☐ the same chemical properties because they have the same atomic weight

   b. ☐ different chemical properties because they have the same number of electrons

   c. ☐ different chemical properties because they have a different number of neutrons

   d. ☐ the same chemical properties because they have the same number of electrons

REFER to Panel 1. Notice that the elements are arranged from left to right, row after row, according to their increasing atomic numbers. Thus, going from left to right on each row of the table and continuing at the left-hand side of the next row, the atoms of each element have 1 more proton and 1 more electron than the atoms of the element they follow. Since the weight of an atom increases with the number of protons, an increase in atomic number is also an increase in atomic weight; thus, elements in the periodic table are also arranged in order of increasing atomic weight. However, there are a few exceptions to this, because of the existence of isotopes. In some cases, an element with a smaller atomic number has an average atomic weight that is greater than the average atomic weight of an element that follows it on the

ANSWERS:

1. c

2. d

periodic table.

The columns of the periodic table are numbered from IA to VIIIA, starting at the left, as shown on the Panel. The number at the top of each column is the same as the number of valence electrons in the atoms of the elements in the column. All the elements whose atoms have 1 valence electron are in column IA, all those with 5 valence electrons are in column VA, etc. Thus, from their location on the periodic table, one can tell how many electrons the atoms of a pair of elements have, and work out the formula of the compound they may react to form. For example, aluminum, symbol Al (in column IIIA) has a valence of +3; and sulfur, symbol S (in column VIA) has a valence of -2. Since the positive and negative charges of the compound must be equal, it takes 2 aluminum atoms to react with 3 sulfur atoms. That is, $2 \times (+3) + 3 \times (-2) = 0$. The compound formed has the formula Al_2S_3.

Since all the atoms of the elements in one column have the same number of valence electrons, their chemical properties are similar and they constitute a <u>group</u> or <u>family</u>. By using the periodic table, it is not necessary to learn the chemical properties of each of the 103 elements, separately. We can study the chemical properties of one element in each family. For example, sodium (Na) is an active metal with a valence of +1, it forms ionic compounds, it is a base, a good reducing agent, it is soft and has a low melting point, it reacts violently with oxygen and with water, and it is too active to be found naturally in an uncombined state. Since the atoms of the other elements found in column 1A also have 1 electron in their valence orbit, they have similar chemical properties. Thus, the elements of the first group are highly reactive metals, form ionic compounds, are good reducing agents, and so forth. Let's consider another example -- group VIIIA. All of the atoms in group VIIIA (except helium) have 8 electrons in their valence orbit, and so they are relatively nonreactive gases and are not found in a combined state. Helium is included in this group because even though it has only 2 electrons, its valence orbit is filled and so it is not reactive.

Each row or line in the table is called a period. Since the element at the extreme right of each line has a filled valence orbit, that is 8 electrons, the element with the next atomic number begins a new period, and has an additional orbit. The number of periods is therefore equal to the number of orbits. Going from left to right across the periodic table, each additional electron is usually added to the outermost orbit, but this is not always the case. In the 3rd, 4th, and 5th periods, additional electrons can go into inner orbits as well. You will note that in the table there is a gap between group II and group III, from the third line down. The table is drawn this way because after there are 2 electrons in the valence orbit, the next 10 electrons get put into the orbit underneath the valence orbit. After these 10 electrons, additional electrons go into the valence orbit until it is built up to 8; and then a new period is started. This is how inner orbits are built up to 18 electrons. The elements that fit into the gap in the periodic table between groups II and III are called the <u>transition elements</u>. The properties of the transition elements are slightly different from those of the atoms in the rest of the table. Notice that these elements are also grouped together from I to VIII followed by the suffix B. The elements in each group followed by the suffix B resemble each other more closely than they resemble the elements in the main group. Thus, the metals in group IB (copper, silver, and gold) resemble one another more closely than they resemble the metals in group IA. This is due to the fact that the electrons in the next inner orbit are distributed in a similar manner. The groups followed by the suffix B are referred to as subgroups.

There are two other transition series: the rare earth elements, called the lanthanides, and the heavy rare earth elements called the actinides. As the periodic table shows, there are no groups for these two series.

3. Which of the following is a true statement?

   a. ☐ Aluminum, (Al, atomic number 13) belongs to the same group as silver, (Ag, atomic number 47).

   b. ☐ Potassium, (K, atomic number 19) has the same number of valence electrons as copper, (Cu, atomic number 29).

   c. ☐ Sodium, (Na, atomic number 3) and magnesium, (Mg, atomic number 12) belong to the same group.

   d. ☐ Helium, (He, atomic number 2) and neon, (Ne, atomic number 10) do not have the same number of electrons in their valence orbit.

4. According to the periodic table, tellurium, (Te, atomic number 52) has a greater atomic weight than iodine, (I, atomic number 53). Which one of the following statements is the reason for this?

   a. ☐ There is an error in the table.

   b. ☐ Tellurium has more isotopes than iodine.

   c. ☐ The average atomic weight of the isotopes of iodine is greater than the average atomic weight of the isotopes of tellurium.

   d. ☐ The more abundant isotopes of tellurium are heavier than the more abundant isotopes of iodine.

5. By studying their positions on the periodic table, CHOOSE the correct formula for a compound of lithium, (Li, atomic number 3) and oxygen, (O, atomic number 8).

a. ☐ LiO

b. ☐ 2LiO

c. ☐ Li_2O

d. ☐ LiO_2

6. The elements of which collection have similar chemical properties?

a. ☐ groups IA to VIIIA

b. ☐ subgroups IB to VIIIB

c. ☐ group A periods 1 to 7

d. ☐ period 4 - groups IA to VIIIA

In each group, going from top to bottom, the size of the atom increases because of the increase in the number of orbits. For example, in column IA, lithium (Li) is smaller than cesium (Cs); and in column VIIA, fluorine (F) is smaller than iodine (I). The additional orbits are further away from the nucleus, and so the electrons are at a higher energy level and are not held as tightly. Therefore, the atoms of the elements at the bottom of column IA have smaller electronegativities than the atoms of the elements at the top of the column, and the atoms of the elements at the top of column VIIA have greater electronegativities than the atoms of the elements at the bottom of the column.

To sum up, toward the upper right-hand corner are nonmetals that are more electronegative and better oxidizing agents. Toward the lower left-hand corner are the metals that are less electronegative and good reducing agents. Fluorine (F), in row 2 column VIIA, is the most active nonmetal; while cesium (Cs), row 6 column IA, is the most active metal.

Since the periodic table indicates relative electronegativities (in some charts, numerical values for the electronegativities are given), one can predict the nature of the bond of two reactive elements. If the two elements are in the middle of the chart, the bond will be covalent and the compound will have the properties expected of a covalent molecule. If the elements are at opposite sides of the table, and the compound formed will probably be ionic. Elements

ANSWERS:

3. d

4. d

5. c

6. c

8

such as chlorine (Cl) and sodium (Na) which are far apart form an ionic compound; while elements such as sulfur (S) and carbon (C) which are close together form a covalent compound.

7. CHECK the information that cannot be obtained from a periodic table.

   a. ☐ the number of protons an atom has

   b. ☐ the number of isotopes an atom has

   c. ☐ the usual electron configuration of an atom

   d. ☐ the type of bond an element would have a tendency to form when reacting with another element

ANSWER:

7. b

TO SCORE THIS LESSON AS IF IT WERE A MASTERY TEST, SEE PAGE 86 OF THE SAK.

WHEN YOU HAVE FINISHED THIS LESSON, WRITE DOWN THE TIME. THEN TAKE THE LESSON TO YOUR INSTRUCTOR OR HIS ASSISTANT FOR CHECKING. WAIT UNTIL THE LESSON IS APPROVED BEFORE GOING ON TO THE NEXT LESSON.

* U. S. GOVERNMENT PRINTING OFFICE : 1970 O - 369-883

PM 431-76

ADVANCED GENERAL EDUCATION PROGRAM

A HIGH SCHOOL SELF-STUDY PROGRAM

ELECTROLYSIS
LEVEL: II
UNIT: 10
LESSON: 4

49

U.S. DEPARTMENT OF LABOR
MANPOWER ADMINISTRATION, JOB CORPS
NOVEMBER 1969

U.S. DEPARTMENT OF LABOR
MANPOWER ADMINISTRATION, JOB CORPS
NOVEMBER 1969

In oxidation-reduction reactions, the substance that loses electrons is oxidized and the substance that gains electrons is reduced. For example, in the reaction:

$$Cu + I_2 \longrightarrow Cu^{++} + 2I^-$$

a copper atom loses 2 electrons and becomes a positive cupric ion. Each of the atoms of the iodine molecule gains an electron and becomes a negative iodide ion. In this reaction the copper is oxidized to cupric ions and the iodine is reduced to iodide ions.

The transfer of electrons between the copper and iodine is quite simple. A piece of copper, as you may remember from our discussion of metallic bonding in a previous lesson, is made up of copper atoms, copper ions, and free electrons. Thus, when iodine molecules come in contact with a piece of copper, each of the atoms of the iodine molecules gains a free electron and becomes an iodide ion. As the free electrons are taken up by the iodine, the copper atoms lose more free electrons, thus forming more copper ions. Also, some of the iodine atoms react directly with the copper atoms by pulling electrons away from the copper atoms when they collide. The reaction is complete when all of the copper atoms have become copper ions and all of the atoms of the iodine molecules have become iodide ions.

All metals have free electrons present, and so the electron gained by the iodine does not even have to come from the copper. The electron could come from another piece of metal, which is itself in contact with the copper. In this case, the iodine gains an electron from the second metal which in turn gains an electron from the copper. The copper can even be in a separate vessel. The metal which dips into the iodine is called an <u>electrode</u>. An electrode that supplies electrons to a substance has a negative charge and is called the <u>cathode</u>. An electrode that takes electrons from a substance has a positive charge and is called the <u>anode</u>. A little thought will convince you that oxidation takes place at the anode and reduction takes place at the cathode. Also, since oppositely charged particles attract one another, the positively charged cation moves to the negative cathode, and the negatively charged anion goes to the anode.

1. In oxidation-reduction reactions, the substance that loses electrons:

   a. ☐ is not affected by the reaction; the substance that gains electrons is oxidized and reduced

   b. ☐ is oxidized; the substance that gains electrons is reduced

   c. ☐ is reduced; the substance that gains electrons is oxidized

2. Which statement can be correctly made about the reaction between copper (Cu) and chlorine (Cl_2) as shown in the equation:

$$Cu + Cl_2 \longrightarrow Cu^{++} + 2\ Cl^{-}$$

a. ☐ Each copper atom gains 2 electrons and becomes a positive ion, while each atom of chlorine loses an electron and becomes a negative ion.

b. ☐ Each copper atom loses 2 electrons and each chlorine atom gains one electron.

c. ☐ The copper is reduced and the chlorine is oxidized.

d. ☐ The copper is unchanged by the reaction while the chlorine gains electrons.

3. In the oxidation-reduction reaction between copper and iodine, the iodine is reduced as a result of gaining electrons from:

a. ☐ copper atoms and positive cupric ions

b. ☐ copper atoms only

c. ☐ positive copper ions

d. ☐ the free electrons in the copper and from copper atoms

4. MATCH the following:

A. anode
B. cathode

1. _____ positively charged electrode
2. _____ negatively charged electrode
3. _____ supplies electrons
4. _____ takes electrons
5. _____ where oxidation takes place
6. _____ where reduction takes place

5. anion
cation

CHOOSE one of the terms listed above and use it to COMPLETE the sentences below:

An ion formed by the gain of electrons is called a (n) ______________.

The ion with a positive charge is attracted to the negative electrode. The ion is called a (n) ______________.

The ion formed by the loss of electrons is attracted to the cathode. This ion is known as a (n) ______________.

ANSWERS:

1. b

2. b

3. d

4. 1. A, 2. B, 3. B, 4. A, 5. A, 6. B

5. anion

   cation

   cation

PANEL 2

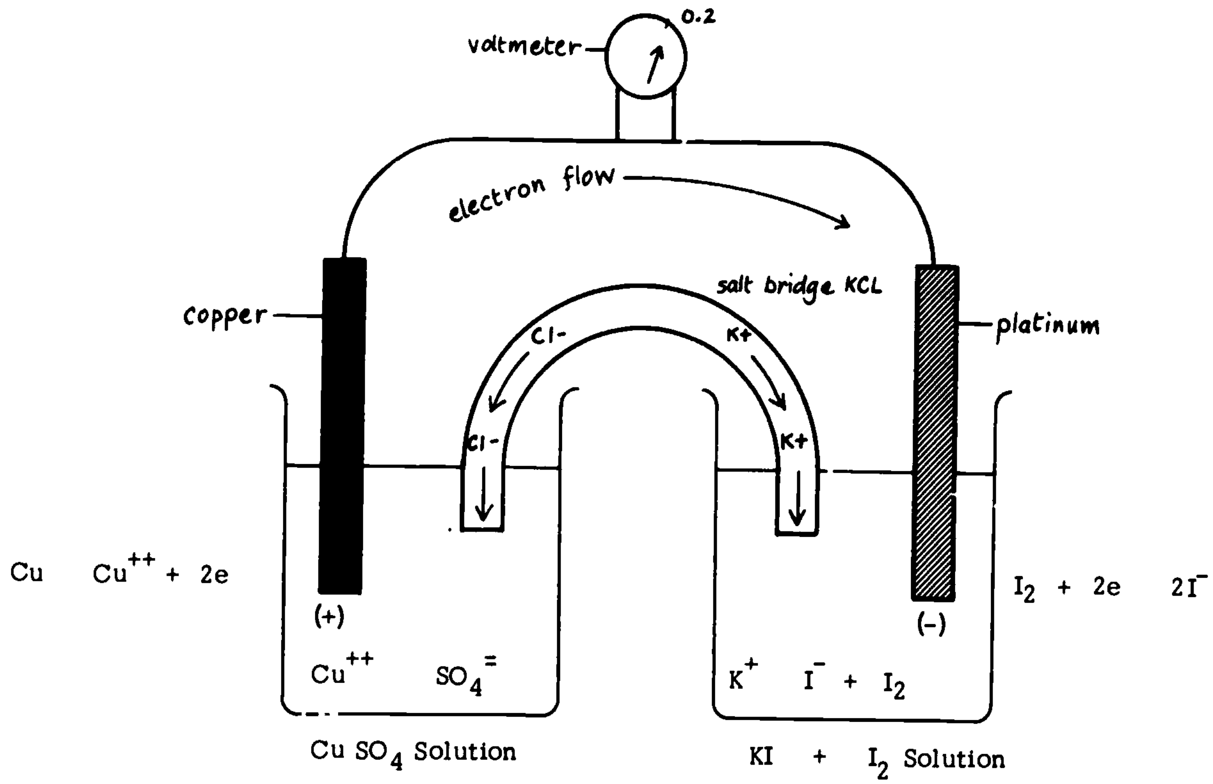
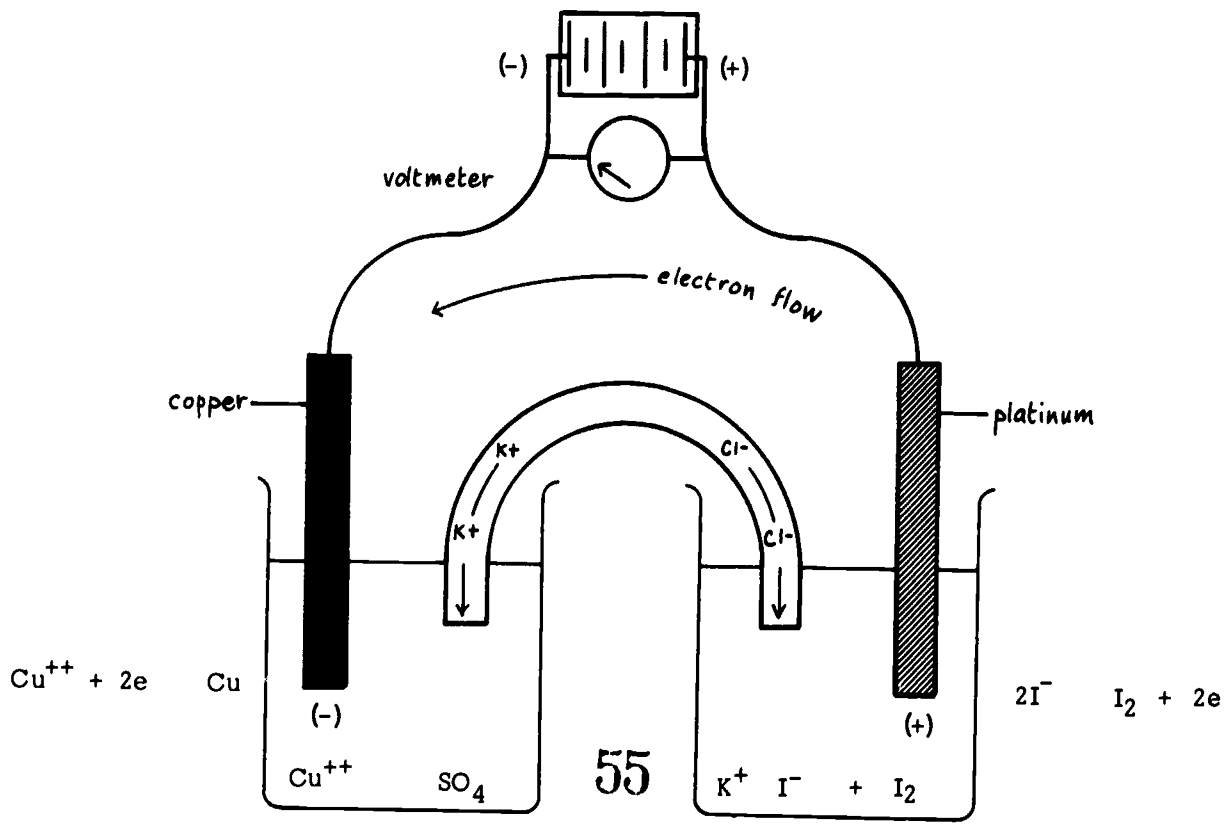


Diagram A



An arrangement in which the oxidizing agent and the reducing agent are separated from each other is shown in Diagram A on Panel 2. Here, copper is the reducing agent, in addition to acting as an electrode. When the two electrodes are connected by a wire, electrons removed at the copper electrode flow to the platinum electrode. At the copper electrode, the anode, the reaction:

$$Cu \longrightarrow Cu^{++} + 2 \text{ electrons}$$

goes from left to right, supplying more electrons as fast as they flow out through the wire. At the same time, the cupric ions which are formed go into solution around the electrode. At the platinum cathode, the electrons flowing in are picked up by the iodine in solution and the reaction:

$$I_2 + 2 \text{ electrons} \longrightarrow 2\,I^-$$

goes from left to right as fast as electrons flow in. The flow of electrical charge through the wire is called the <u>current</u>; and the energy of the reaction is called the <u>potential</u> or the <u>voltage</u>. The apparatus as a whole is called an <u>electrochemical cell</u>.

A further examination of Diagram A shows two features we have not yet discussed, the "salt bridge" and the solutions surrounding the two electrodes. The purpose of the salt bridge is to keep the two vessels electrically neutral. As electrons flow out of the copper through the platinum and pass on to the iodine, the left-hand vessel tends to become positively charged and the right-hand vessel, negatively charged. Since electrons flow toward a positive charge and away from a negative charge, unless the two vessels were neutralized somehow, the current would stop almost as soon as it started. However, the solution in the salt bridge keeps the current flowing. It contains positive and negative ions in equal quantity. The positive ions flow toward the cathode and the negative ions flow toward the anode, thus neutralizing the charge built up in the two solutions by the transfer of electrons from Cu to I_2. This neutralization of the charges in the solutions allows the current to continue. If the salt bridge were removed, the current would stop. The salt bridge is therefore said to complete the circuit.

The electrodes must have a solution around them in order for the circuit to be completed. These solutions must be electrolytic, that is, they must contain ions which are free to move. In Diagrams A and B, copper sulfate, $CuSO_4$, is in solution in the left-hand compartment and iodine and potassium iodide, $KI + I_2$, is in solution in the right-hand compartment. Although there are many other electrolytic solutions, these electrolytes have been selected to illustrate that this electrochemical process is a reversible reaction.

As you have already learned, ionic compounds do not conduct electricity in the solid state, but they do when they are melted or dissolved in certain solvents. For electricity to be conducted, the charges must be able to move in the direction of the current. In a solid, the ions attract each other so strongly that they cannot move. Melting increases the energy of the ions so that they can move away from each other. Water, and solvents like water, do not increase the energy of the ions at all, but they do decrease the forces of attraction holding the ions in place. Such solvents, called polar solvents, enable ions of the electrolytes dissolved in them to move and thus conduct electricity. Even with the salt bridge, if there were no ions present in the solution surrounding the electrodes, the current would flow very slowly if at all.

6. In an electrochemical cell, the flow of electrons:

   a. ☐ goes from the anode to the cathode

   b. ☐ goes from the cathode to the anode

   c. ☐ varies in direction unpredictably

7. The positive ions in the salt bridge migrate to the negative electrode of an electrochemical cell and the negative ions migrate to the positive electrode. These ions:

   a. ☐ increase the positive charge built up on the anode only

   b. ☐ increase the negative charge built up on the cathode only

   c. ☐ increase the negative charge on the cathode and the positive charge on the anode

   d. ☐ neutralize the negative charge on the cathode and the positive charge on the anode

8. Ionic compounds conduct electricity when they are:

   a. ☐ dissolved in polar solvents only

   b. ☐ either dissolved in polar solvents or melted

   c. ☐ in the melted state only

   d. ☐ in the solid state only

9. When ionic compounds are dissolved in polar solvents:

a. ☐ the energy of the ions is decreased

b. ☐ the energy of the ions is increased

c. ☐ the forces of attraction holding the ions together are decreased

d. ☐ the forces of attraction holding the ions together are increased

10. Electrolytic solutions:

a. ☐ contain ions which are closely bound and are not free to move

b. ☐ contain ions which are free to move

c. ☐ do not conduct electricity and so are used as insulation

ANSWERS:

6. a

7. d

8. b

9. c

10. b

If the two electrodes in Diagram A are connected by means of a wire, current flows and Cu^{++} is formed at the anode and I^- at the cathode. The reaction is spontaneous, producing about 0.2 volts. The cell is now a source of energy. It is a battery.

If the connecting wire is removed and another course of potential, whose energy is greater than 0.2 volts, is connected in opposition to the cell shown in the Diagram, the entire electrochemical process is reversed. See Diagram B. Electrons no longer flow from the copper, out of the platinum, to the iodine in solution. In the copper compartment, the copper ions now pick up electrons at the negative electrode, or cathode, and are being deposited; that is, as the copper ions gain electrons and become copper atoms they are deposited at the cathode. At the positive platinum electrode, the anode, iodide ions give up electrons and form iodine molecules. The electrochemical reaction now occurring is electrolysis. This special case in which a metal is deposited as a thin layer on top of another metal at the cathode is called electroplating.

To sum up, in a battery, an electrochemical reaction is used to produce energy; while, in electrolysis, energy from an outside source is used to produce an electrochemical reaction.

In an electrolysis cell, the solution must contain the particular ions we wish to deposit at the cathode; but in a battery cell there is a much wider choice of electrolytes. In any case, the solutions must contain ions in order to conduct electricity.

11.

battery cell
electrolysis cell
negative electrode
positive electrode

CHOOSE a term from the list above to COMPLETE each of the sentences below:

An electrochemical cell which is used to produce energy is a (n) ______________.

Energy from an outside source is used to produce a reaction in an electrochemical cell. The cell is a (n) ______________.

In electroplating, metallic ions in solution gain electrons and are deposited as atoms of metal at the ______________.

Time completed ______________

ANSWERS:

11. battery cell

    electrolysis cell

    negative electrode

WHEN YOU HAVE FINISHED THIS LESSON, WRITE DOWN THE TIME. THEN TAKE THE LESSON TO YOUR INSTRUCTOR OR HIS ASSISTANT FOR CHECKING. WAIT UNTIL THE LESSON IS APPROVED BEFORE GOING ON TO THE NEXT LESSON.

U. S GOVERNMENT PRINTING OFFICE 1970 O - 369-884