REVIEW: TOPIC 4

NAME: ______

DUE: ______

UNIT 4 – MAJOR UNDERSTANDINGS

- ☆ 3.1dd Compounds can be differentiated by their physical and chemical properties.
- ☆ 5.2g Two major categories of compounds are ionic and molecular (covalent) compounds.
- ☆ 5.2a Chemical bonds are formed when valence electrons are:
  - transferred from one atom to another (ionic)
  - shared between atoms (covalent)
  - mobile within a metal (metallic)
- ☆ 5.2e In a multiple covalent bond, more than one pair of electrons are shared between two atoms. Unsaturated organic compounds contain at least one double or triple bond.
- ☆ 5.2l Molecular polarity can be determined by the shape of the molecule and distribution of charge. Symmetrical (nonpolar) molecules include CO_2, CH_4, and diatomic elements. Asymmetrical (polar) molecules include HCl, NH_3, and H_2O.
- ☆ 5.2c When an atom gains one or more electrons, it becomes a negative ion and its radius increases. When an atom loses one or more electrons, it becomes a positive ion and its radius decreases.
- ☆ 5.2i When a bond is broken, energy is absorbed. When a bond is formed, energy is released.

UNIT 4 – MAJOR UNDERSTANDINGS (CONTINUED)

- ☆ 5.2b Atoms attain a stable valence electron configuration by bonding with other atoms. Noble gases have stable valence configurations and tend not to bond.
- ☆ 5.2n Physical properties of substances can be explained in terms of chemical bonds and intermolecular forces. These properties include conductivity, malleability, solubility, hardness, melting point, and boiling point.
- ☆ 5.2d Electron-dot diagrams (Lewis structures) can represent the valence electron arrangement in elements, compounds, and ions.
- ☆ 5.2j Electronegativity indicates how strongly an atom of an element attracts electrons in a chemical bond. Electronegativity values are assigned according to arbitrary scales.
- ☆ 5.2k The difference in electronegativity between two bonded atoms is used to assess the degree of polarity in the bond.
- ☆ 5.2h Metals tend to react with nonmetals to form ionic compounds. Nonmetals tend to react with other nonmetals to form molecular (covalent) compounds. Ionic compounds containing polyatomic ions have both ionic and covalent bonding.

BONDING

A **chemical bond** results from simultaneous attraction of electrons (either single or paired) to two or more nuclei. Two major categories of compounds are ionic or molecular (covalent) compounds.

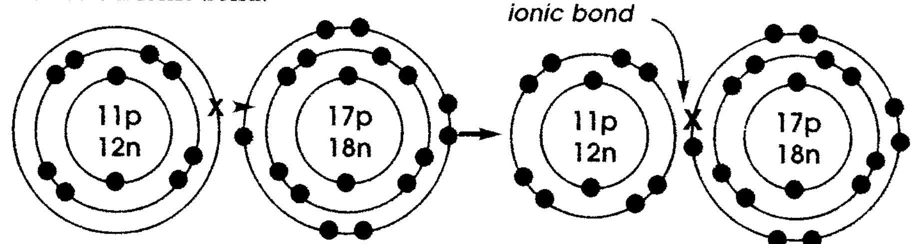
A – THE NATURE OF CHEMICAL BONDING

BONDS BETWEEN ATOMS

The electrons involved in bond formation may be transferred from one atom to another or may be shared equally or unequally between two atoms. When atoms of the elements enter into a chemical reaction, they do so in a manner that results in their becoming more like **inert** (noble gas) gas atoms. In this state, they contain their maximum complement of valance electrons, and they are in a condition of maximum stability.

IONIC BONDS

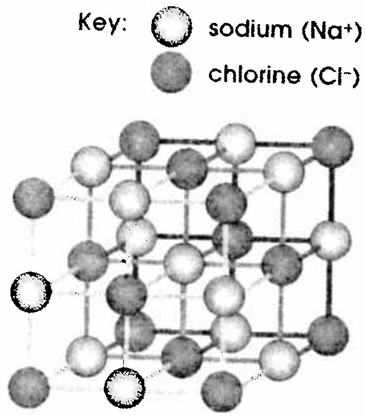
An **ionic bond** is formed by the transfer of one or more electrons from metals to nonmetals. This transfer of electrons results in the formation of ions. The attraction between a positive ion and a negative ion is called an ionic bond.



In ionic bonding, the number of electrons transferred is such that the atoms involved achieve an "inert" gas configuration, except for some transition elements. Since the ion has a different electron configuration than the atom, the properties of the ion differ from those of the atom. Also, ionic bonds may form between ions that were formed in a previous reaction as a result of a transfer of ions. For example:

$$\overset{+\ -}{AgNO_3(aq)} + \overset{+\ -}{NaCl(aq)} \rightarrow \overset{+\ -}{AgCl(s)} + \overset{+\ -}{NaNO_3(aq)}$$

Characteristics of Ionic Solids – Ionic solids have high melting points and do not conduct electricity. In the geometric structure of the solid ionic crystal, ions form the crystal lattice and are held in relatively fixed positions by electrostatic attraction. When melted or dissolved in water, the crystal lattice is destroyed and the ions move freely. This free movement of ions permits electrical conductivity. Examples of ionic solids are sodium chloride and magnesium oxide.

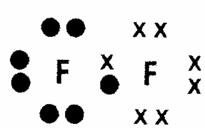


NaCl Crystal Lattice

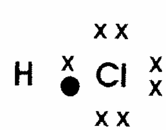
Covalent Bonds

A **covalent bond** is a simultaneous attraction of two nuclei for the same electrons resulting in the sharing of those electrons. A covalent bond is formed when two atoms share electrons, instead of transferring them. In order to form this type of bond, the electronegativity difference between the two atoms forming the bond must be less than 1.7. Covalent bonds are classified as two types. Their structures can be demonstrated using Lewis dot structures.

Nonpolar Covalent Bond – When electrons are shared between atoms of the same element, they are shared equally, and the resulting bond is a **nonpolar bond**. An example of a nonpolar covalent bond is found in the fluorine molecule. Since the electronegativity of both fluorine atoms in the molecule is the same, the difference is zero, and the electron density of the molecule is symmetrical.

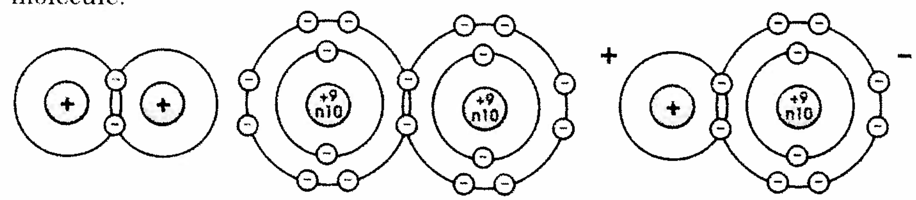


Polar Covalent Bonds – When electrons are shared between atoms of different elements, they are usually shared unequally. The resulting bond is polar. An example of a **polar covalent bond** is found in the hydrogen chloride molecule.



Chlorine, having an electronegativity value of 3.2 will attract the bonding electrons to a greater extent than hydrogen, which has an electronegativity value of 2.2. The difference of 1.0 denotes a covalent bond. Since the chlorine end of the molecule will show a greater electron density probability than the hydrogen end, the molecule will be asymmetrical and therefore polar.

In the illustration below, the hydrogen and fluoride molecules are both considered nonpolar, because both hydrogen atoms in the hydrogen molecule have the same electronegativity; therefore, the molecule is symmetrical. The same nonpolar symmetrical situation occurs in the fluorine molecule.

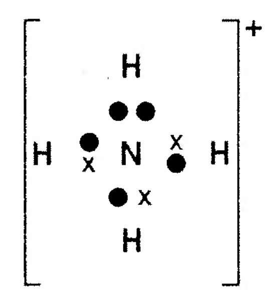


Hydrogen Gas H_2 — Fluorine Gas F_2 — Hydrogen Fluoride HF

In the HF molecule, hydrogen has an electronegavity rating of 2.1 and fluorine 4.0; therefore, the fluorine atom attracts the shared electrons to a greater extent than hydrogen. The fluorine portion of the molecule becomes more negative, but unlike the hydrogen, it becomes positive. This arrangement makes for an unsymmetrical molecule called a **dipole**.

Coordinate Covalent Bonds – When the two shared electrons forming a covalent bond are both donated by one of the atoms, this bond is called a coordinate covalent bond. A coordinate covalent bond, once formed, is not different from an ordinary covalent bond. The difference lies in the source of the electrons involved in the bond. This type of bond is frequently involved in the bonding within polyatomic ions and is very important in modern acid-base theories.

A classic example of a coordinate covalent bond is the ammonium ion. In the ammonium ion, the nitrogen atom has five valence electrons. Three electrons are unpaired but are shared with the electrons from three hydrogen atoms. The other two form a full pair and are not shared.



ammonium ion

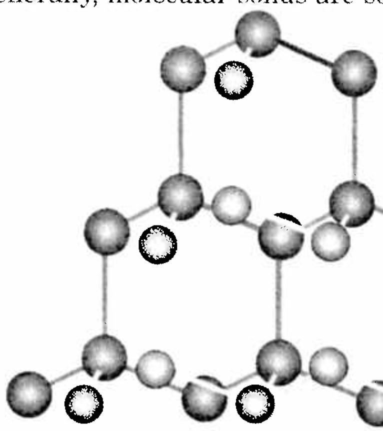
It is at this unshared pair of electrons that the electron density is so great that the molecule may attract a hydrogen ion (proton). When that occurs, the ion is formed and takes on a charge of positive one (+1). For example, water (H_2O) can attract a proton (hydrogen nucleus) to become a hydronium ion (H_3O^+)

Molecular Substances – A molecule may be defined as a discrete particle formed by covalently bonded atoms. A molecule has also been defined as the smallest particle of an element or compound capable of independent existence. When a stable molecule is formed, a covalent bond is established. The atoms that form the bond usually assume electronic structures of inert gases by sharing electrons. Examples of molecules include:

H_2 NH_3 H_2O HCl CCl_4 S_8 $C_6H_{12}O_6$

Characteristics of Molecular solids – Molecular substances may exist as gases, liquids, or solids, depending on the attraction that exists between the molecules. Generally, molecular solids are soft, good electrical insulators, poor heat conductors, and have low melting points.

Network Solids – Certain solids consist of covalently bonded atoms linked in a network that extends throughout the sample with an absence of simple discrete particles. Such a substance is said to be a network solid.



Carbon Atom Crystal lattice

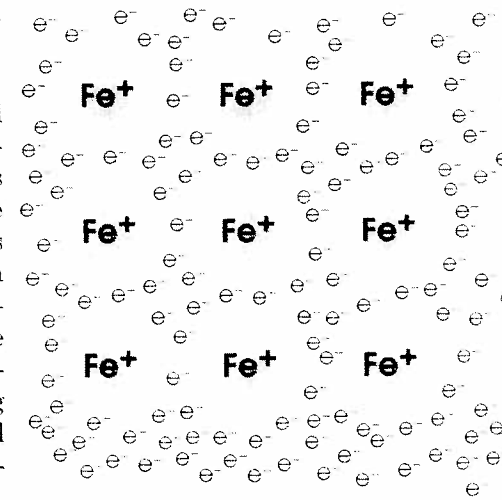
Generally, network solids are hard, are poor conductors of heat and electricity, and have high melting points. Examples of network solids are diamond (C), silicon carbide (SiC), and silicon dioxide (SiO_2).

Note 3: Metals tend to react with nonmetals to form ionic compounds (i.e. NaCl, MgBr, KI). Nonmetals tend to react with nonmetals to form covalent (molecular) compounds; i.e. CO, NF_3, SCl_2. Compounds containing polyatomic ions have both ionic and covalent bonding. By definition these compounds are charged ionic substances; but, within the molecule they have covalent bonds; i.e. SO_4^{2-}, PO_4^{3-}, CN^-.

METALLIC BONDING

Metallic bonding occurs between atoms of metals that have a small number of valence electrons leaving them with many vacant valence orbitals and low ionization energies.

A **metallic bond** consists of an arrangement of positive ions that are located at the crystal lattice sites and are immersed in a "sea" of mobile electrons. These mobile electrons can be considered as belonging to the whole crystal rather than to individual atoms.



Metallic Bonding

Note: This mobility of electrons (previous page) distinguishes the metallic bond from an ionic or covalent bond and gives the metal the following characteristics:

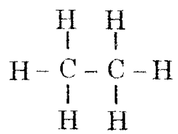
- good conductors of electricity and heat
- great strength
- malleability and ductility
- luster

Ionic Properties	Molecular Properties	Metallic Properties
• poor conductor of electricity in solid state • high melting point • when melted or dissolved in water, become good conductors	• exists in gas, liquid, or solid state • soft • good insulators • poor heat conductor • low melting points	• good conductor of heat and electricity • great strength • good malleability and ductility • luster

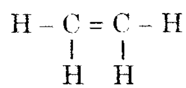
Saturated & Unsaturated Compounds

A bond formed between carbon atoms by the sharing of one pair of electrons is referred to as a **single bond**. Organic compounds, where carbon atoms are bonded by the sharing of a single pair of electrons, are said to be **saturated compounds**.

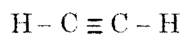
Organic compounds, containing two adjacent carbon atoms bonded by the sharing of more than one pair of electrons, are said to be **unsaturated compounds**. A bond between carbon atoms by the sharing of two pairs of electrons is referred to as a **double bond**. A bond formed between carbon atoms by the sharing of three pairs of electrons is referred to as a **triple bond**. Examples include



Single Bonds (Saturated)



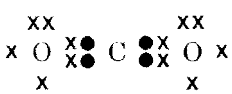
Double Bonds



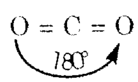
Triple Bonds

Non Polar Molecules

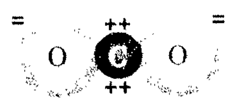
Because of the directional nature of the covalent bonds which form them, some molecules composed of more than two atoms may be nonpolar – even though the individual bonds are polar if the shape of the molecule is such that symmetric distribution of charge results (i.e., carbon dioxide – CO_2 is a nonpolar molecule). $\mathbf{CO_2}$ molecule can be shown in many ways. Examples follow:



electron dot notation



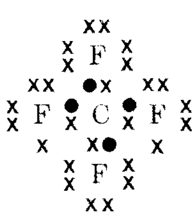
structural formula



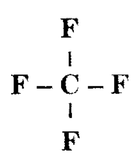
Electron Density Formula of Carbon Dioxide - CO_2

Because of the greater electronegativity of oxygen, the electron density is greater around the oxygen atoms. The bond with an electronegativity difference of 0.9 is polar, but because the bond angle is 180° and the molecule is symmetrical, it is considered a nonpolar molecule.

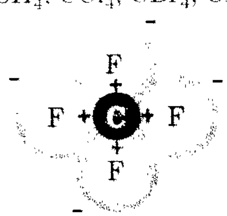
Carbon tetrafluoride molecule can be shown to be a nonpolar molecule with polar bonds (other examples include: CH_4, CCl_4, CBr_4, CI_4):



electron dot notation



structural formula



Electron Density Formula of Carbon Tetrafluoride - CF_4

Fluorine has an electronegativity of 4.0, and Carbon has an electronegativity of 2.6. The difference, 1.4, indicates a polar covalent bond. However, because the tetrahedral molecule is symmetrical, the molecule is nonpolar.

When two hydrogen atoms bond to form a hydrogen molecule (H_2) they form a nonpolar bond with an electronegativity difference of 0. The molecule they form is a nonpolar, symmetrical molecule. This is true of other diatomic molecules such as oxygen, chlorine, and fluorine.

Chemical Energy

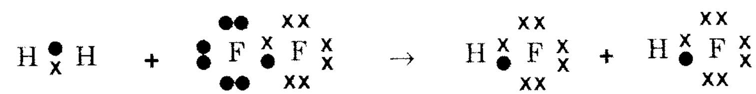
Potential energy is stored in molecules, and the transfer or release of some of this energy manifests itself in the form of chemical energy. Substances possess energy because of their composition and structure. Factors such as mass, types of bonding, and types of motion influence the absorption and storage of energy by molecules.

Energy Changes in Bonding

When two atoms are held together by a chemical bond, they are generally at a lower energy condition than when they are separated. Therefore, when a chemical bond is formed, energy is released, and when a chemical bond is broken, energy is absorbed.

Bonding & Stability

Because there is a release of energy when bonds are formed, systems at lower energy levels are more stable than systems at higher energy levels. So, it follows that bonding will more often occur among atoms if the changes lead to a lower energy condition and, therefore, a more stable structure. The more energy given off when a bond is formed, the stronger and more stable the bond will be. Also, the less energy given off in the formation of a bond, the weaker and less stable it will be. This is readily illustrated when hydrogen and fluorine combine to form hydrogen fluoride:



In order for the bonds of the reactants (ex.: hydrogen and fluoride molecules) to break and form the products, energy must be added:

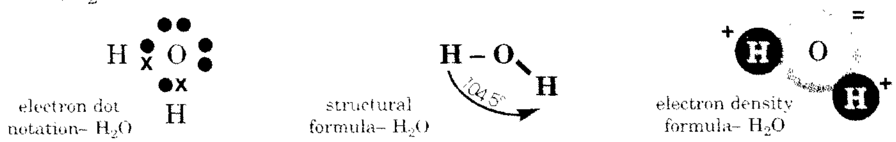
436.8 kJ/mole + 155.4 kJ/mole → 567 kJ/mole + 567 kJ/mole

total energy in = 592.3 kJ *total energy given off* = 1,134 kJ

Note 1: As bonds are formed, energy is given off. It is easily seen that more energy is given off in the formation of the products, than is required to break the bonds in the reactants. Therefore, the products are more stable than the reactants.

Polar Molecules

Generally, the geometric structure of covalent substances, which result from the directional nature of the covalent bond, helps to explain properties of the resulting molecule. The polarity of a water molecule is explained by the asymmetrical shape of the molecule. The water molecule (H_2O) is shown as follows:



Oxygen's electronegativity is 3.4, and hydrogen's electronegativity is 2.1. The difference, 1.3, indicates a polar covalent bond. The bond angle is such that there exists an unsymmetrical distribution of electron density. Therefore, a polar molecule results. Other examples of polar molecules include the following (see *Reference Table S*):

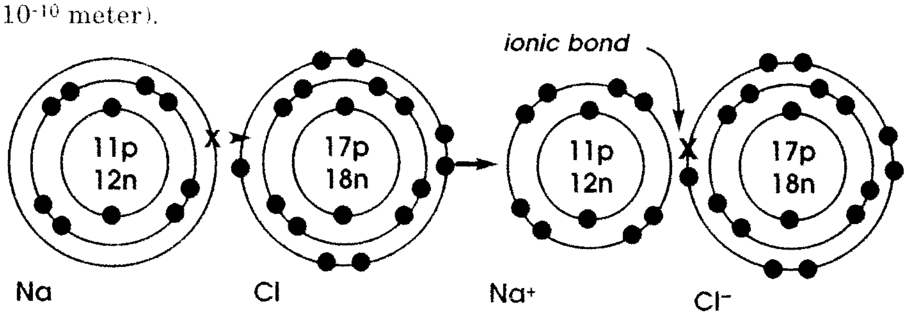
Hydrogen Chloride – electronegativity = hydrogen 2.1, chlorine = 3.2, difference = 1.1

Ammonia – electronegativity = hydrogen 2.1, nitrogen = 3.0, difference = 0.9

The electrons involved in bond formation may be transferred from one atom to another or may be shared equally or unequally between two atoms. When the atoms of these elements enter into a chemical reaction, they do so in a manner that results in their becoming more like "inert" gas atoms. In this state, they contain their maximum complement of valance electrons, and they are in a condition of maximum stability.

Ionic Radius

A loss or gain of electrons by an atom causes a corresponding change in size. Metal atoms lose one or more electrons when they form ions. Ionic radii of metals are smaller than the corresponding atomic radii. Nonmetal atoms gain one or more electrons when they form ions. Ionic radii of nonmetals are larger than the corresponding atomic radii. Atomic and ionic radii are usually measured in Angstrom (Å) units (1Å = 1 x 10^{-10} meter).



C – Intermolecular Forces

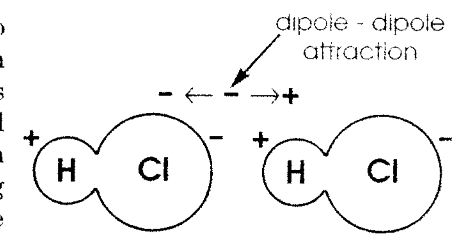
Physical properties of substances such as conductivity, malleability, solubility, hardness, melting point, and boiling point can be explained in terms of chemical bonds and intermolecular forces. The strength of attraction that an atom has for another atom is a measure of its electronegativity rating. It is an arbitrary scale proposed by American Linus Pawling (1901-1994). The electronegativity difference is used to assess the degree of polarity of a bond formed by two atoms. This polarity of intramolecular bonds results in polar molecules which are attracted to other polar molecules.

Dipoles

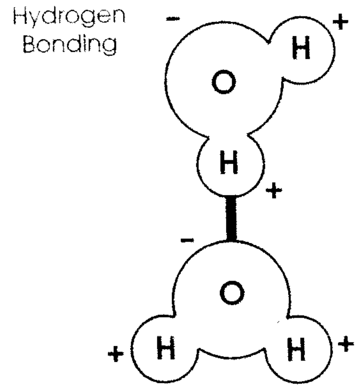
The asymmetric distribution of an electrical charge in a molecule causes a molecule that is polar in nature and is referred to as a **dipole**. That is, the uneven electron cloud density will cause one end of a molecule to be more negative than the other end. A molecule composed of only two atoms will be a dipole if the bond between the atoms is polar. For example, the hydrogen chloride molecule is a dipole because

- the chlorine atom is larger than the hydrogen atom, and
- the difference in electronegativity of hydrogen (2.1) and chlorine (3.2).

These factors allow chlorine to share electrons closer to itself than to hydrogen. When the bond is formed, the electron density around the chlorine atom is greater than around the hydrogen atom, leading to a polar molecule (a dipole). The bond between two hydrogen chloride molecules is a result of dipole–dipole attraction (at right)



Hydrogen Bonding

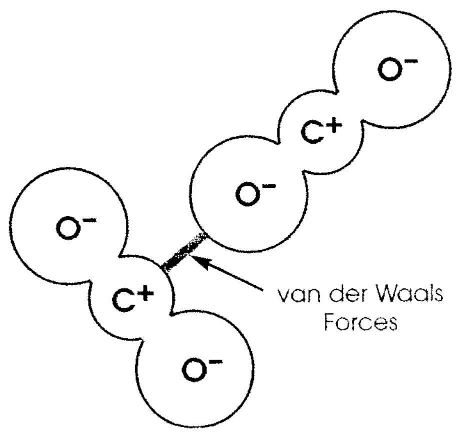


Hydrogen bonds are formed between molecules when hydrogen is covalently bonded to an element of small atomic radius and high electronegativity. When a hydrogen atom is bonded to a highly electronegative atom, the hydrogen has such a small share of the electron pair that it acts like a bare proton.

Hydrogen bonding (shown in a quantity of water) is important in compounds of hydrogen with fluorine, oxygen, or nitrogen. These compounds represent special cases of dipole to dipole attraction.

These forces also account for the meniscus when water or any other liquid is poured into a measuring instrument.

Van der Waals Forces



In the absence of dipole attraction and hydrogen bonding, as in nonpolar molecules, weak attractive forces exist between molecules. These forces are called **van der Waals forces**.

Van der Waals forces make it possible for species of small nonpolar molecules, such as hydrogen, helium, oxygen, etc., to exist in the liquid and solid phases under conditions of low temperature and high pressure.

Van der Waals forces appear to be due to chance distribution of electrons resulting in momentary dipole attractions. Therefore, these forces are momentary electrostatic forces that increase as the distance between molecules decreases. Also, as the size of the molecules increases, the greater the Van der Waals forces.

The effect of molecular size on the magnitude of the van der Waals forces accounts for the increasing boiling points of a series of similar compounds (such as the alkane series of hydrocarbons).

Molecule Ion Attraction

Polar solvents, when interacting with ionic compounds, attract ions from these compounds and form a solution. Ionic compounds are generally soluble in polar solvents such as water, alcohol, and liquid ammonia. The negative ion of the substance being dissolved is attracted to the positive end of the adjacent polar molecules, while the positive ion is attracted to the negative end of the polar molecules. Water is the polar substance most commonly used to dissolve these ionic compounds. When an ionic compound is dissolved in water, its crystal lattice is destroyed, and water molecules surround each ion, forming hydrated ions. It is because water is a dipole that this attraction between the water molecules and the positive or negative ion exists. The orienting of water molecules around ions is called the hydration of the ions. This process is important in aqueous chemistry.

Polyatomic Ions

A single atom with a charge is called a **monatomic ion**. A compound of two or more covalently bonded atoms with a charge is called a polyatomic ion.

A polyatomic ion is very stable and behaves like a monatomic particle, because it contains strong covalent bonds. These bonds are stronger than the bonds that hold it to the rest of the atoms in the compound. Therefore, during reactions, the polyatomic ion usually remains intact as it passes from the reactants to the products.

Some polyatomic ions include: NH_4^+ (ammonium ion), PO_4^{-3} (phosphate ion), and NO_3^- (nitrate ion). Other examples can be found in the *Reference Table E*.

Although the bonds which keep the atoms in a polyatomic ion are covalent bonds, the polyatomic ions possess a charge. When they attach themselves to a metal ion or another polyatomic ion, they do so by forming an ionic bond. The final compound contains both ionic and covalent bonds. Some examples include: $Na^+NO_3^-$ and $NH_4^+OH^-$. In general,

- metals tend to react with nonmetals to form ionic compounds;
- nonmetals tend to react with other nonmetals to form covalent (molecular compounds; and,
- ionic compounds containing polyatomic ions have both ionic and covalent bonding.

UNIT FOUR: Chemical Bonding

KEY CONCEPTS

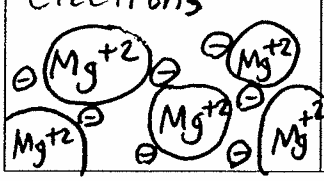
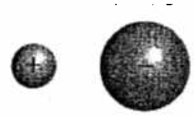
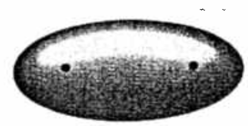
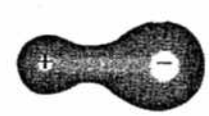
Valence electrons	Lewis dot diagrams
Ionic bond	physical properties
Metallic bond	polar covalent
Covalent bond	nonpolar covalent
electronegativity	polarity
Ionic substance	symmetrical molecules
Molecular substance	asymmetrical molecules
Metallic substances	Polyatomic ions

Atoms come together (bond) to form compounds.

Atoms get a stable valence electron configuration by bonding with other atoms. Noble gases (group 18) have a stable valence electron configuration *"8 is great"* and tend not to bond with other elements.

When a bond is formed, energy is released. (exothermic)
When a bond is broken, energy is absorbed. (endothermic)

Four types of bonding:

Metallic Bonds	Ionic Bonds	Nonpolar Covalent Bonds	Polar Covalent Bonds
Metals	Metal and Nonmetal	Two Nonmetals with *Equal* Electronegativities	Two Nonmetals with *Unequal* Electronegativities
Valence Electrons are mobile within a metal	Valence Electrons are Transferred from the Metal to the Nonmetal	*Equal* Sharing of Valence Electrons	*Unequal* Sharing of Valence Electrons
Mg, Au	NaCl, AlF_3	H_2, Cl_2	H_2O, NH_3
Metal Ions in a "sea" of valence electrons 	Electron Distribution  Ionic bond (transfer)	Electron Distribution  Nonpolar covalent bond (equal sharing)	Electron Distribution  Polar covalent bond (unequal sharing)

It is important to note that, covalent bonds can share more than one pairs of electrons. When 2 pairs of electrons are shared it is a double covalent bond. When 3 pairs of electrons are shared it is a triple covalent bond.

Electronegativity indicates how strongly an atom attracts electrons in a chemical bond. Electronegativity values for elements can be found on Table S of the Reference Tables. The difference in electronegativity between two elements is proportional to the degree of polarity in the bond.

Electron-dot diagrams (Lewis dot structures) can represent the valence electrons in atoms, ions and compounds.

Atoms	Ions	Compounds
Na• Ca: :N· :Cl·	Na^{+1} Ca^{+2} :N^{-3}:	Na^{+1} :Cl^{-1}: :Cl–Cl:

Ionic substances, molecular substances and metallic substances can be distinguished by their chemical and physical properties. **Physical properties** include: conductivity, solubility, malleability, hardness, melting pt. and boiling pt.

Molecules can be polar or nonpolar. The **polarity** of a compound can be determined by the shape of the molecule and the distribution of charge. **Symmetrical molecules** are nonpolar because the charges are evenly distributed around the molecule. **Asymmetrical molecules** are polar because the charge is not evenly distributed around the molecule.

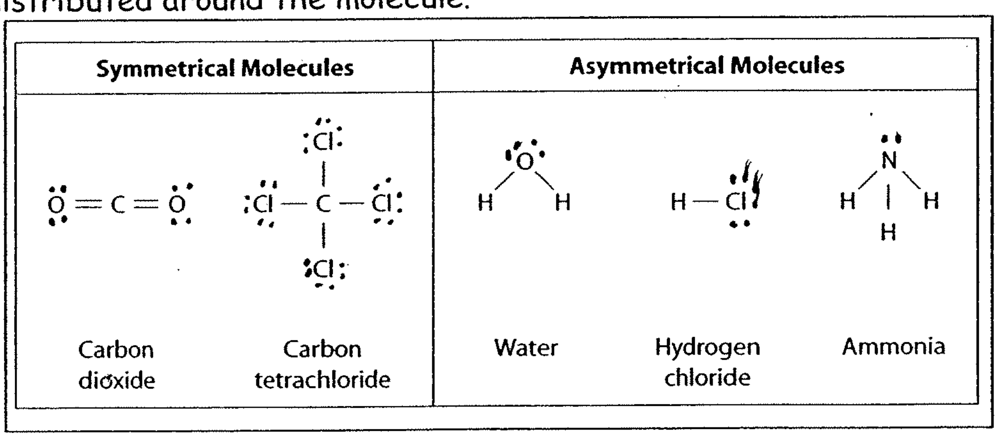


Figure 6-10. Five compounds with polar bonds: The symmetrical molecules are nonpolar; the asymmetrical molecules are polar.

Polyatomic ions have both ionic and covalent bonding when bonded in compounds.

	POLAR MOLECULES	NONPOLAR MOLECULES
Intermolecular Force	**STRONG**: The asymmetrical distribution of charge holds the molecules to each other better than nonpolar molecules	**WEAK**: The equal charge distribution throughout the molecule makes it hard for each one to hold onto one another.
Boiling & Melting point	Higher	Lower
Conductivity as a Solid	NO	NO
Conductivity when aqueous	Low	No
Solubility	Only in Polar Solvents	Only in Nonpolar Solvents
Examples