General overview of Categories of Chemical Bonds

distinction based on bond locality → **Intrastructural** vs. **Interstructural** (Intermolecular)

Mr. Stone

Intrastructural:
Covalent
Ionic
Metallic

bonds that link atoms to build molecules or independent structures such as metal or salt crystals

found within (between individual atoms)

Interstructural:
(Van der Waals forces
Dipole-Dipole interactions (includes Hydrogen Bonds)
London Dispersion forces

- bonds that link molecules together

(found between molecules)

CHEMICAL BONDING

IONIC and COVALENT BONDS

Objectives:

1. Describe the nature of the chemical bond and its relationship to valence electrons.
2. Compare ionic and covalent bonding.
3. Use dot diagrams to represent ionic and covalent compounds.
4. Use comparisons of electronegativities of atoms in order to predict the nature of bonding.
5. Compare and contrast the properties of ionic and covalent compounds

Main Principles of Bonding:

1) Bonding is a byproduct of atoms making changes in their valency to achieve maximum electron ~~stabl~~ stability
2) all chemical bonds are electrostatic forces of attraction (attractive force between opposing electrical charges (+, −))

Valence and Bonding

distinctions or categorization based upon: 1) locality of bond (intra vs. inter) 2) origination or mode of establishment of electrostatic force)

Atoms combine to form compounds in two general ways. In both cases, the position of one or more electrons in the valence shell of each combining atom is altered.

chemical bond-- a force of attraction that holds two atoms together and which develops between them when shifts in the location of their valence electrons occur

The formation and breaking of chemical bonds occur during chemical reactions.

During chemical reactions, atoms will lose, gain, or share valence electrons so that each type of atom involved in the reaction aquires the stable electron configuration of the noble gas nearest the element in the periodic table. ← this is the driving force for chemical bonding

Except for helium, the electron configuration of the valence shell of all the noble gases is s^2p^6. This arrangement of valence electrons is referred to as the **stable octet** of electrons (meaning the atom has the maximum number of eight valence electrons).

octet rule (rule of eight)– the tendency of valence electrons to rearrange themselves during chemical reactions so that each atom involved in the reaction ends up with a stable octet

Ionic Bonding

In one of the ways by which atoms combine to form compounds, electrons from the valence shell of an atom of one element are transferred to the valence shell of an atom of another element.

The atom that gives away its valence electrons will become an ion with a net positive charge (cation) equal to the number of electrons lost by the atom.

The atom that accepts electrons into its valence shell will become an ion with a net negative charge (anion) equal to the number of electrons gained by the atom.

The opposite electrical charges on the cation and anion will cause them to be attracted to each other and to form an *ionic compound*. The chemical bond that holds the cation and anion together is called an *ionic bond*.

The simplest kind of chemical reaction leading to the formation of ionic bonds occurs between a metal on the far left of the periodic table and a nonmetal on the far right. On one hand, the atoms of metals lose electrons readily because metals have low ionization energies (energy required to remove electrons from a neutral atom). On the other hand, the atoms of nonmetals gain electrons because most nonmetals have a strong attraction for electrons.

During a reaction between a metal and a nonmetal, electrons in orbitals in the valence shell of the metal are transferred to orbitals in the valence shell of the nonmetal so that each atom is able to form a stable octet.

Metals with atomic numbers lower than 5 **do not** follow the octet rule. They will form ionic bonds by losing enough valence electrons so that their electron configurations match that of the noble gas helium.

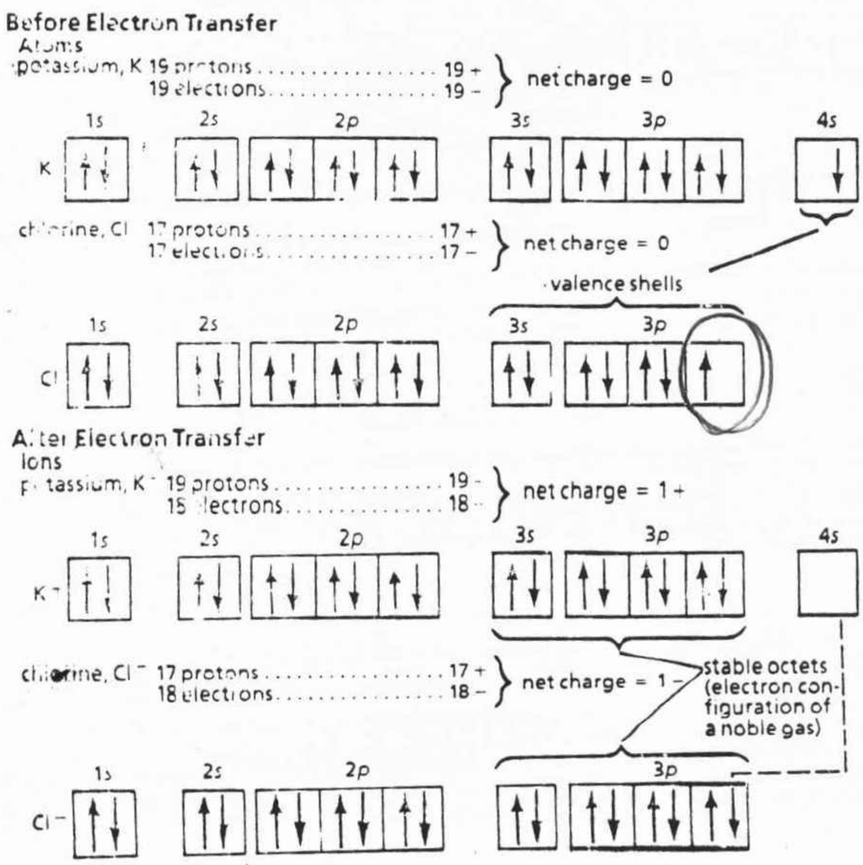
Ionic bonds involve the loss or gain of valence electrons. Valence Electron Transfer ! 6

Primarily:

Ionic bonding occurs between metallic and nonmetallic elements.

Except for polyatomic ions
charged molecules

A good example of ionic bonding can be seen in the reaction between potassium(K), which is a metal, and chlorine(Cl), which is a nonmetal as is seen in the figure below.



Note that the potassium atom has one valence electron. It needs to lose this electron to have the electron configuration of the nearest noble gas, argon. The chlorine atom has seven valence electrons. It needs to gain one electron to have the electron configuration of the nearest noble gas, argon. The transfer of the 4s electron from the potassium atom to the half-filled 3p orbital of the chlorine atom produces the stable octet of electrons in the valence shells of both atoms.

The loss of an electron by the potassium atom produces the K^+ cation.
The gain of an electron by the chlorine atom produces the Cl^- anion.

$K^+ + Cl^-$ -------------------> KCl (potassium chloride)

It is the **electrostatic force of attraction** between oppositely charged particles in an ionic compound that bonds cations to anions. This force is the **ionic bond.** Ionic bonds are also called **electrovalent bonds.**

Covalent Bonding

Another way in which atoms combine to form compounds involves one or more pairs of electrons that move close enough to the nuclei of two atoms so that they are attracted by the nuclei of both atoms simultaneously.

Usually--but not always--one of the electrons in such a pair comes from the valence shell of an atom of one element, and the other electron comes from the valence shell of an atom of the other element. Electrons that are positioned between two atoms in this way are said to be **shared electrons**.

When electrons are shared between two atoms, they are in a lower energy state than when they are attracted to the nucleus of only one atom. This sharing of electrons, which is called a **covalent bond**, produces a force of attraction that holds the atoms together. When such a bond is formed, the energy of each of the covalently bonded atoms decreases.

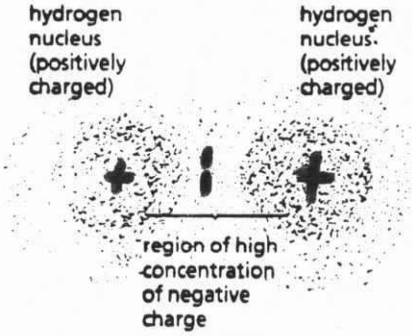
* Covalent Bonding involves the sharing of valence electrons in pairs between the nuclei of bonding atoms

* - shared pairs of valence electrons are established by the overlapping of valence shell orbitals between bonding atoms

Covalent bonding primarily occurs between nonmetals (nonmetal to nonmetal bonding)

When atoms are involved in covalent bonding, a valence orbital containing an electron from one atom overlaps a valence orbital containing a valence electron from another atom. Because the nuclei of both atoms have a positive charge, they are attracted to the shared concentration of negative charge that exists between them. **This attraction of the nuclei of both atoms for the concentration of negative charge is the covalent bond that binds the atoms together.**

The simplest example of covalent bonding can be seen in the hydrogen molecule(H_2) :



When a covalent bond is formed between two hydrogen atoms, there is an overlapping of two *s* electron clouds. These two electrons are attracted to both hydrogen nuclei at the same time and are therefore shared between the atoms. In this arrangement, the two electrons and the two nuclei are in a lower energy state than when arranged as separate atoms. Because of this lowering of energy, a stable covalent bond forms between them. The energy that appears to be lost by both atoms is actually stored in the covalent bond.

The result of this covalent bond, a hydrogen molecule, can be represented in any of the following ways:

H_2	H:H	H—H
Molecular formula. It shows that there are two atoms of H in the molecule.	**Dot diagram.** Each dot represents one valence electron. Each hydrogen atom contributes one valence electron to the bond.	**Structural formula.** The dash represents a covalent bond, or one pair of shared electrons.

As is true of ionic bonding, atoms usually require a stable octet of electrons in their valence shells during covalent bonding. However, the stable octet contains the shared electrons as well as those that are unshared. **When counting electrons in a stable octet, a shared pair is counted twice, once in the octet of the atom on one side of the covalent bond and once in the octet of the atom on the other side of the bond.**

Other examples of covalent bonding:

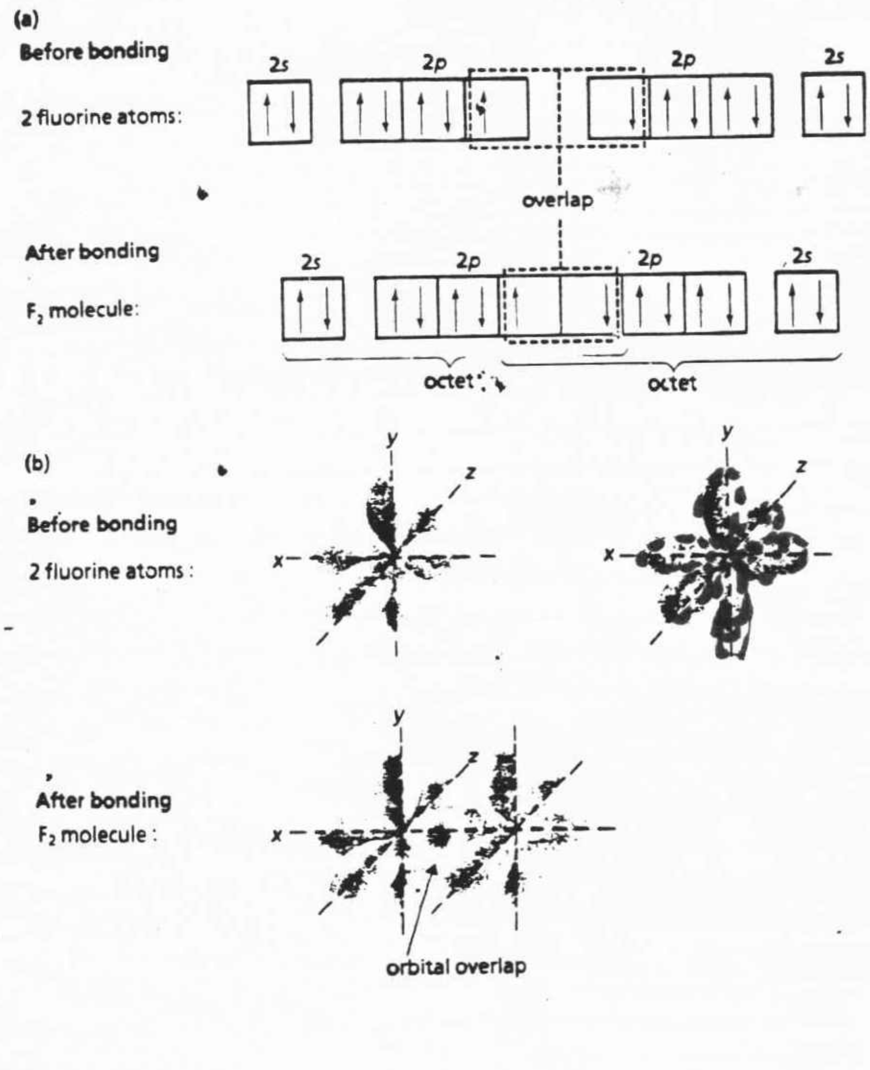


Figure 15-8
Bonding in the fluorine, F_2, molecule. Overlap of the half-filled *2p* orbitals in each atom accounts for the formation of a covalent bond in F_2. **(a)** Orbital diagram representation of the formation of the covalent bond in F_2. Note that by sharing a valence electron, each atom acquires a stable octet in its valence shell. **(b)** Electron-cloud representation.

Substances with atoms held together by covalent bonds are called **molecular substances**. Molecular substances are composed of molecules, not ions. There exist many more molecular substances than ionic substances.

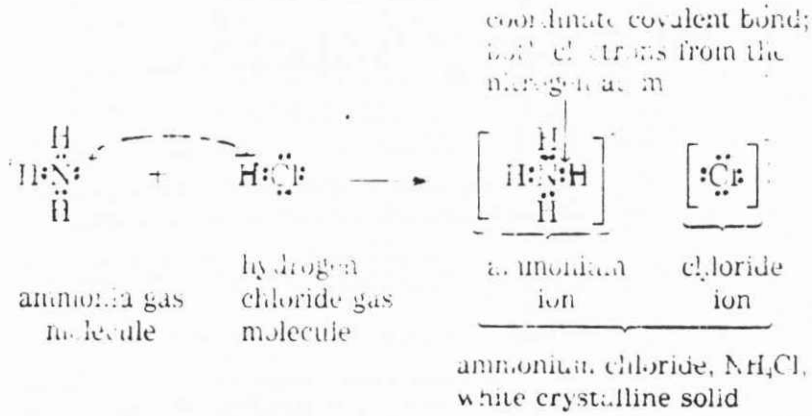
<u>Polyatomic ions:</u>

An ion that is made up of more than one atom is called a **polyatomic ion**. Like other ions, they can have a net positive or net negative charge. A polyatomic ion with a negative charge has aquired electrons from an outside source (from atoms not contained in the polyatomic ion). A polyatomic ion with a positive charge has lost electrons to other atoms not contained in the polyatomic ion.

The atoms in polyatomic ions are covalently bonded to each other. In reactions, polyatomic ions generally behave as if they were single atoms. Examples of polyatomic ions are the hydroxide ion, OH^-, nitrate ion, NO_3^-, and the carbonate ion, CO_3^{2-}.

<u>Coordinate covalent bonds</u>

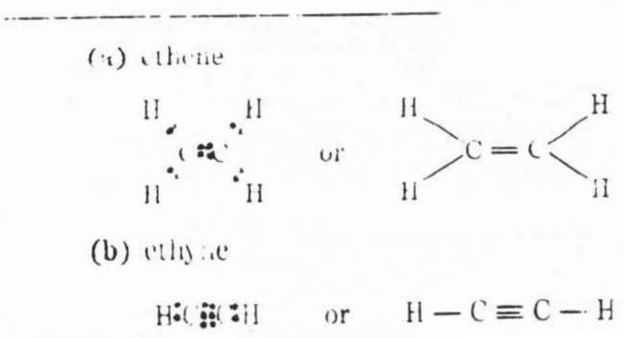
Coordinate covalent bonds are formed when the two shared electrons forming a covalent bond are both donated by the same atom. Once formed, these bonds are no different than ordinary covalent bonds except for the source of the electrons forming the bond



<u>Multiple bonds:</u>

Not all bonds are based on the formation of **single bonds** (the sharing of **one pair of electrons**). Some substances are bonded by sharing two or three pairs of electrons.

double bond- when two pairs of electrons are shared between atoms
triple bond- when three pairs of electrons are shared between atoms



Electronegativity and Bonding:

Electronegativity is a measure of the ability of an atom of an element that is chemically bonded with another element to attract electrons to itself.

Below is a table that lists the electronegativities of the elements:

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
2.1 **H**																
1.0 **Li**	1.5 **Be**											2.0 **B**	2.5 **C**	3.0 **N**	3.5 **O**	4.0 **F**
0.9 **Na**	1.2 **Mg**											1.5 **Al**	1.8 **Si**	2.1 **P**	2.5 **S**	3.0 **Cl**
0.8 **K**	1.0 **Ca**	1.3 **Sc**	1.5 **Ti**	1.6 **V**	1.6 **Cr**	1.5 **Mn**	1.8 **Fe**	1.8 **Co**	1.8 **Ni**	1.9 **Cu**	1.6 **Zn**	1.6 **Ga**	1.8 **Ge**	2.0 **As**	2.4 **Se**	2.8 **Br**
0.8 **Rb**	1.0 **Sr**	1.2 **Y**	1.4 **Zr**	1.6 **Nb**	1.8 **Mo**	1.9 **Tc**	2.2 **Ru**	2.2 **Rh**	2.2 **Pd**	1.9 **Ag**	1.7 **Cd**	1.7 **In**	1.8 **Sn**	1.9 **Sb**	2.1 **Te**	2.5 **I**
0.7 **Cs**	0.9 **Ba**	1.1–1.2 **La–Lu**	1.3 **Hf**	1.5 **Ta**	1.7 **W**	1.9 **Re**	2.2 **Os**	2.2 **Ir**	2.2 **Pt**	2.4 **Au**	1.9 **Hg**	1.8 **Tl**	1.8 **Pb**	1.9 **Bi**	2.0 **Po**	2.2 **At**
0.7 **Fr**	0.9 **Ra**	1.1–1.7 **Ac–No**														

Certain trends in this table should be obvious. Notice that electronegativities increase as you move from left to right across the periodic table and that they decrease as you move down a family or group.

The difference between the electronegativities of two atoms in a bond can be used as a guide to determine the degree of electron sharing in the bond. In other words, this difference can be used to predict the type of bond that will most likely form between two atoms.

As the difference increases, the degree of sharing decreases.

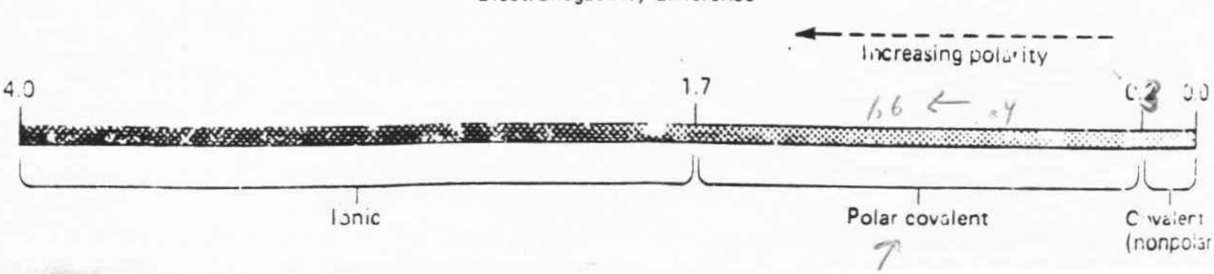
If the difference in electronegativities between the two bonding atoms is 0.2 or less, the pair of bonding electrons is shared equally and thus the type of bond they will form is a covalent bond.

On the other hand, if the difference between electronegativities is 1.7 or greater, the electron is transferred from one atom to the next and thus the type of bond they will form is an ionic bond.

[the element with greater electronegativity will become the anion and the element with lesser electronegativity will become the cation]

Bonds with electronegativity differences between 0.3 and 1.6 are considered to be covalent but with an unequal sharing of electrons and thus they are called polar covalent bonds.

Electronegativity difference



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Polar covalent bonds:

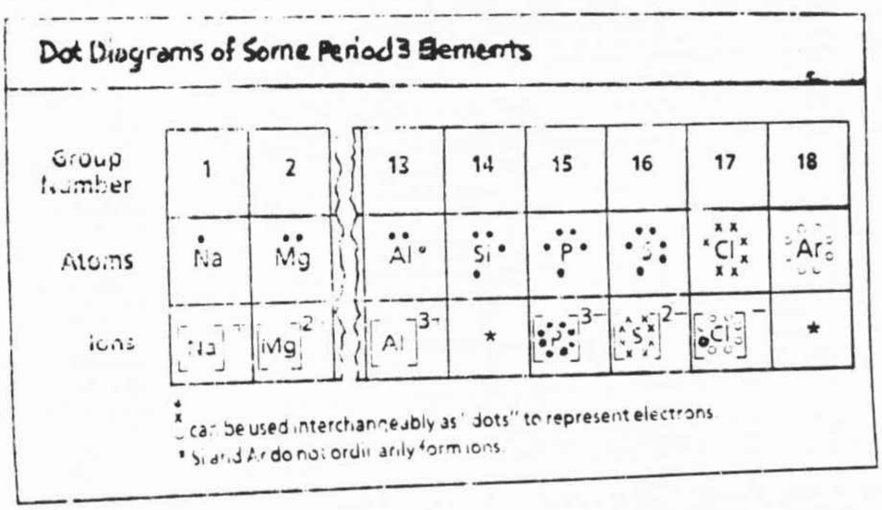
When electrons are shared unequally, the atom with the greater electronegativity takes on a partial negative charge and the atom with the lesser electronegativity takes on a partial positive charge. Thus a bond of this nature has somewhat negative and somewhat positive poles that together are said to make up a **dipole**. A bond with a dipole is called a **polar covalent bond** or **polar bond**.

Electron Dot Diagrams

Dot diagrams are used to show valence electrons. They were developed in the 1920s by G.N. Lewis so they are also called **Lewis structures** or **Lewis dot diagrams**.

A dot diagram for an atom consists of the chemical symbol for the element surrounded by one or more dots. The chemical symbol represents the **kernel** of the atom (the nucleus of the atoms and its inner electrons). Each dot represents one valence electron in the atom. It is customary to write the dots in pairs on the four sides of the symbol as a reminder that electrons are paired in orbitals.

Monatomic ions can also be represented using dot diagrams. As with the dot diagrams for atoms, the chemical symbol represents the kernel of the ion. Dots are used to represent valence electrons. Square brackets are used to show that an ion, rather than a neutral atom, is being represented with the symbol for charge located outside of the brackets.



Properties of Ionic Compounds:

Ionic compounds consist of different types of ions bonded to each other by ionic bonds

properties:

-- high melting point
-- conducts electricity when melted or in water solution
-- highly soluble in water
-- insoluble in covalent liquids

Properties of Covalent Compounds:

Covalent compounds , also known as **molecular compounds**, consist of molecules containing covalently bonded atoms.

properties:

-- low melting point
-- will not conduct electricity (insulator)
-- variable solubility in water
-- soluble in covalent liquids