Energy and Chemical Bond: Chemical bonds are forces that holds atoms together in a compound. Energy is required to overcome these attractive forces and separate atoms in a compound. Breaking of a chemical bond, endothermic process. If energy is required to break a bond, then the opposite process of forming a bond must release energy. Formation of a bond is an exothermic process

- Chemical bond is formed, resulting compound has less potential energy than the substances from which it was formed. Energy is always released when a bond is formed. Greater the energy released during formation of a bond, greater its stability.

Lewis Electron Dot Structures

TABLE 5.5 Lewis structures for the elements of the first three periods and Group II

Period	Group I	II	III	IV	V	VI	VII	VIII
1	H·							He:
2	Li·	Be:	B:	·C:	·N:	·O:	:F:	:Ne:
3	Na·	Mg:	Al:	·Si:	·P:	·S:	:Cl:	:Ar:
4		Ca:						
5		Sr:						
6		Ba:						
7		Ra:						

Lewis dot diagram/Electron dot diagram: Consists of a chemical symbol surrounded by 1-8 dots representing valence electrons. Kernel of every atom is positively charged. Valence electrons surround kernel and are usually represented by small dots, x's or o's. Dots go in a clockwise formation after the first two dots are placed on the top.

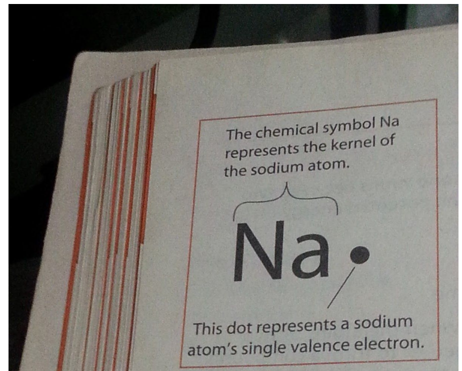


Figure 6-2. The Lewis dot diagram of a sodium atom: The kernel, represented by the symbol Na, contains 11 protons and 10 electrons; the kernel has a 1+ overall charge. The 1+ charge of the kernel is balanced by the 1− charge of the single valence electron. Overall, the sodium atom is neutral.

- Atoms gain/lose electrons, become charged particles, ions. Atoms that lose electrons don't have any dot in electron dot diagram of ion. Square bracket around kernel and ionic charge written as a superscript outside brackets indicate diagram is that of an ion.
- Nonmetals that gain electrons to form positive ions, the valence electrons gained are shown in its electron dot diagram. It is common to show the added valence electrons with a different symbol than that used to represent the atom's originally present valence electrons. These positive ions have a complete outer level of electrons, resulting in 8 dots in the electron dot structure.
- EX: Na listed as 2-8-1. This notation indicates that there is only 1 valence electron. The other 10 belong to the kernel of the atom. The corresponding electron dot diagram for the sodium atoms would show this 1 valence electron as a single dot, outside the kernel

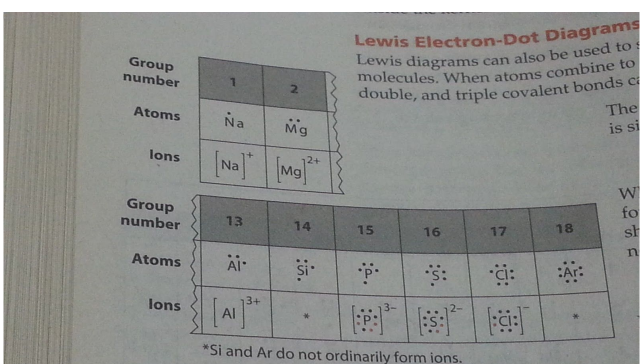


Figure 6-3. Electron dot diagrams of some Period 3 elements and ions

Lewis Electron- Dot Diagrams of Compound: Can be used to show how atoms combine to form molecules. When atoms combine to form diatomic molecules, single, double and triple covalent bonds can be shown.

- When 2 hydrogen atoms combine to form the hydrogen molecule (H2), they share the two electrons between them in a nonpolar covalent bond: H:
- Instead of using dots to show the pair of electrons, a single dash may be used to show the covalent bond. H-H

Members of group 17, halogens, form single bonds between them in their diatomic molecules. In addition to the single covalent bond, each atoms has 6 additional valence electrons for a complete octet of 8

- Oxygen has 6 valence electrons with 2 atoms combining to form O2, in this case, the 2 oxygen atoms share 2 pairs of electrons

Nitrogen, 5 valence electrons and forms diatomic N2 molecule by forming a triple covalent bond
When drawing Lewis diagrams for compounds, pair of electrons forming a covalent bond usually represented by a dashed line, any unpaired electrons are represented by a pair of dots

Steps useful for determining the dot diagrams of compounds

1. Determine the total number of valence electrons of the atoms in the compound

1 carbon with 4 valence electrons = 4
3 hydrogens with 1 electron each = 3
1 chlorine with 7 valence electrons = 7
Total Valence electrons = 14

2. Arrange atoms to show bonds between them. Central atoms often have smallest electronegativity value, and generally appear once in the formula. Remember that hydrogen cannot be the central atom as it can only form one single covalent bond. If more than one atom of an element is present, they generally surround the central atom. Use a dash to show a covalent bond between the atoms. **Each dash represents 2 electrons in a covalent bond.**

- These 4 bonds account for 8 of 14 electrons needed. Carbon atoms has a complete octet of electrons. Distribute remaining 6 electrons around chlorine. Check to see that each atom has an octet of electrons except for hydrogen which needs only 2. In this ex. the 6 remaining electrons should be distributed around chlorine to provide it with an octet of electrons
- If each atom doesn't have octet of electrons, probably some "leftover" electrons. These can be placed as double bonds/triple bonds between atoms that don't have complete octets.

Consider the compound C2H2
2 carbon atoms with
4 electrons = 8 electrons
2 hydrogen atoms with
1 electron = 2 electron
Total = 10 electron
Join the carbons and hydrogen with bonds
H-C-C-H

- 3 bonds account for 6 of 10 electrons. Distributing 4 remaining electrons around carbons will not provide an octet of electrons, but forming triple bonds between the 2 carbons will

Metallic Bonds: Bonds holding metallic atoms together in solid and liquid phases are strong, as metals have high melting and boiling points. Metallic atom considered to have a central kernel, made up of its nucleus and its non-valence electrons. Atom's valence electrons surround kernel. Kernels of metallic atoms making up a metallic solid are arranged in fixed positions of a crystalline lattice. Valence electrons move freely throughout crystal and do not belong to any given atom. Freely moving valence electrons give metals properties of good electrical and thermal conductivity.
Metallic Bond: Results from the force of attraction of the mobile valence electrons for an atom's positively charged kernel.

- Hammering a metal forces the kernels to move to new locations, they are still surrounded by the valence electrons. The freedom of the valence electrons often leads to the use of the term “sea of mobile electrons” to describe metallic bonding.

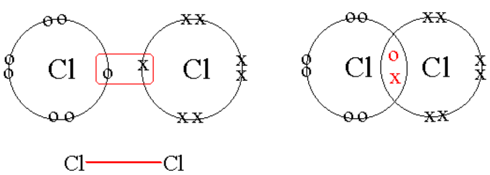
Octet Rule: Configuration of 8 valence electrons is known as an octet. An octet represents the maximum number of valence electrons that an atom can have

- Ordinary chemical reactions result in changes to valence electron configuration of atoms involved. Complete octet of 8 v.e results in exceptionally stable electron configuration.

Octet Rule: States that atoms generally react by gaining, losing or sharing electrons in order to achieve a complete octet of 8 valence electrons- configuration of a noble gas

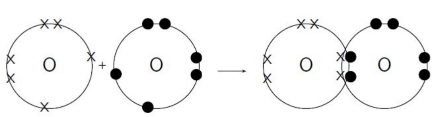
Covalent Bonds: When 2 atoms approach 1 another, electrons repel each other, tending to push atoms apart. Their positive nuclei also repel each other. Attractive force between atom comes from attraction of positively charged nucleus of 1 atom for negatively charged electrons of other atom. Chemical bonds occur when attractive forces between atoms are greater than repulsions forces

Covalent Bond: Formed when 2 nuclei share electrons in order to achieve a stable arrangement of electrons. Covalent bonds often between 2 nonmetal atoms of same element. The diatomic chlorine molecule (Cl2) is an example of a covalent bond. Atoms of different elements may combine to form covalent bonds. Sulfur and oxygen combine covalently to form sulfur dioxide (SO2)



Nonpolar Covalent Bonds: Equal sharing of electrons. Nonpolar covalent bonds are formed between atoms having equal or close electronegativity values. Similar nonpolar bonds exist between each of the other Group 17 atoms in their diatomic molecules, namely F2, Br2, I2.

Diatomic chlorine molecule



Multiple Covalent Bonds: Atoms may share more than 1 pair of electrons resulting in formation of a multiple covalent bond. Oxygen atoms combine with other oxygen atoms to form O2, achieve stable octet configuration. Oxygen nuclei must share 2 **Diatomic Oxygen Molecule** pairs of electrons. The sharing of 2 pairs of valence electrons results in the type of multiple covalent bonds called double covalent bonds.

- Diatomic nitrogen (N2) forms when 2 nitrogen atoms share 3 pairs of valence electrons. The sharing of 3 pairs of valence electrons results in a triple covalent bond

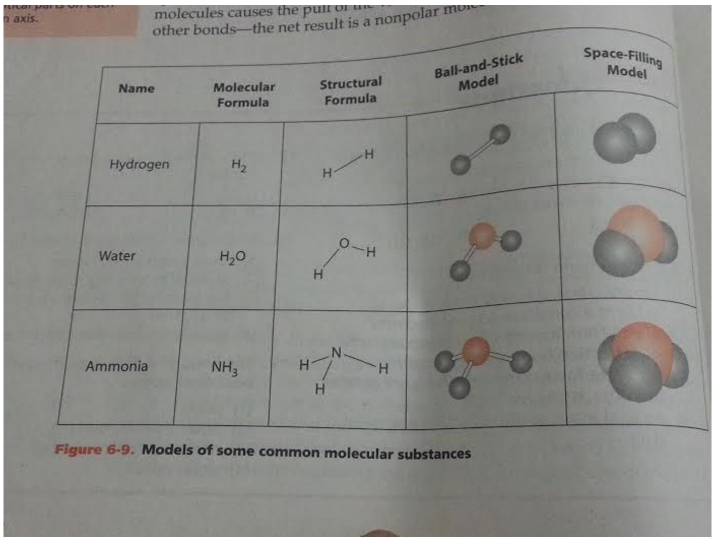
Polar Covalent Bonds: Different atoms involved in a bond, likely unequal electronegativity values. Sharing of electrons in bond is unequal. Higher electronegativity elements attracts shared electrons. Hydrogen Chloride (HCl) and Hydrogen Iodide (HI), examples of polar covalent compounds.

Molecular Substances

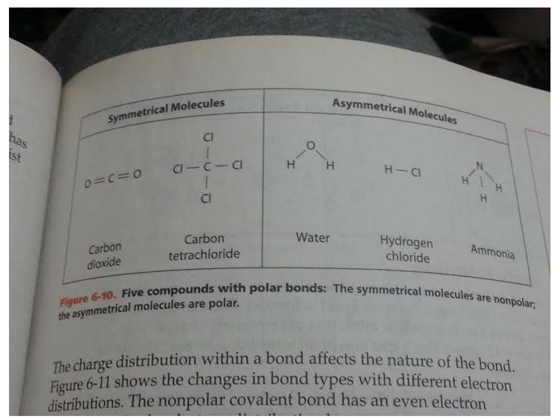
- **Molecule**: Smallest discrete particle of an element or compound formed by covalently bonded atoms. Each atom in a molecule usually has the electron configuration of a noble gas. Molecular substances may exist as solids, liquids and gases, depending on the strength of the forces of attraction between the molecules.
- Molecules generally have properties associated with covalent bonding. Molecules are generally soft, are poor conductors, and have low melting and boiling points.

- H_2O and CO_2 and ammonia (NH_3) are examples of molecular compounds. The diatomic gases, such as O_2 and N_2, are also molecular substances.

Polar Molecules: Molecule may contain polar bonds, not necessarily that molecule itself is polar. Molecules such as hydrogen chloride and water have polar bonds and are also polar molecules. Carbon dioxide and carbon tetrachloride,, contain polar bonds but are not polar molecules.

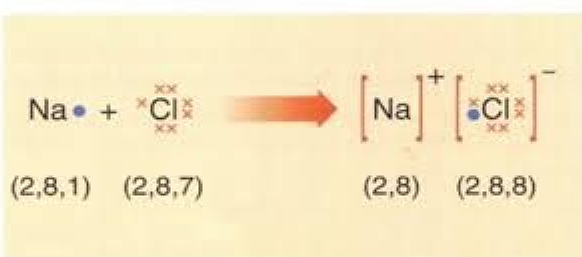


- Water, hydrogen chloride and ammonia molecules are asymmetric. The carbon dioxide, and carbon tetrachloride molecules are symmetrical. The symmetrical shape of the carbon dioxide and carbon tetrachloride molecules causes the pull of the various polar bonds to be offset by other bonds- the net results is a nonpolar molecule

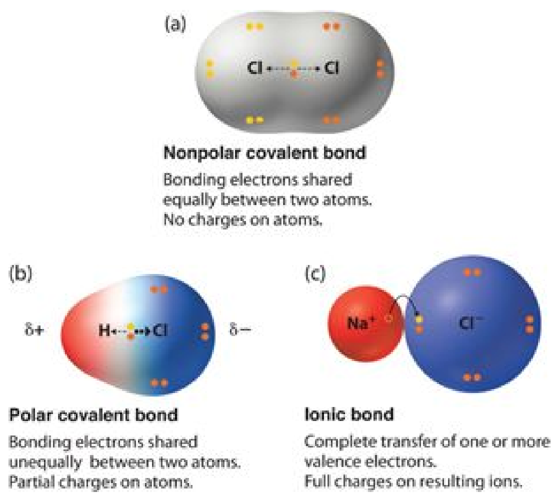


- Charge distribution within a bond affects nature of the bond. This figure shows changes in bond types with different electron distributions, Nonpolar covalent bond has an even electron distribution, As the electron distribution becomes unequal, a polar covalent bond is formed. When the electron is actually transferred , an ionic bond is formed.

Ionic Bonding



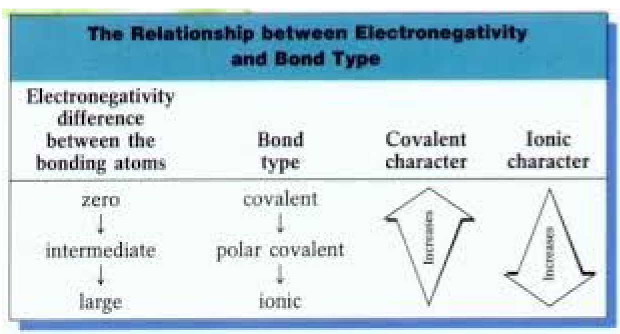
Electron transfer during the formation of sodium chloride



- Atoms that gain or lose electrons become charged particles,ions. Ionic bond is formed when ions bond together because of the electrostatic attraction of oppositely charged ions.

Ion Formation in Metals: Metal losing a valence electron, atom acquires octet arrangement of a noble gas. Loss of the valence electron from atom’s outer energy levels results in the ions having a smaller radius than the atom from which it was formed
Ions formation in Nonmetals: A nonmetal will gain electrons when forming an ion with a 1- charge. The addition of the valence electron to the atom’s outer energy level results in the ions having a larger radius than the atom from which it was formed
Ion Formation and the Octet Rule:

When metals react they:	When nonmetals react they:
1. Lose electrons 2. Become positively charged 3. Have smaller radii 4. Acquire the electron configuration of a noble gas	1. Gain electrons 2. Become negatively charged 3. Have larger radii 4. Acquire the electron configuration of a noble gas



Electronegativity: As electronegativity between 2 atoms in a bond increases, bond becomes more polar in nature. As electronegativity difference increases, bond becomes ionic in character. At some point bond can no longer be considered as sharing of electrons, rather a bond in which 1 or more electrons have been transferred from 1 atom to another. Transfer of electrons from 1 atom to another results in an ionic bond.

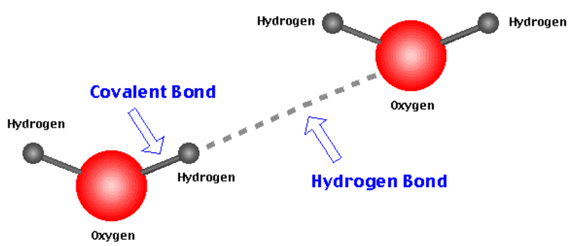
- If the electronegativity difference between bonding atoms is 1.7 or greater, bond is generally ionic. Ionic bonds are generally formed between metals and nonmetals.

Polyatomic Ions: Any compound containing a polyatomic ion must contain an ionic bond. All compounds with polyatomic ions contain both ionic and covalent bonds.
Distinguishing Bond Types: Metallic, covalent and ionic bonds different properties that can distinguish among them. Mercury, liquid standard conditions and an exceptions along with metals of Group 1. Ionic substances, conductors when they are melted (fused) or dissolved in aqueous solutions. Molecular substances held by covalent bonds. No charged particles to conduct an electric current. Molecular substances poor conductors, whether they are solid state, liquid state or aqueous solution
Properties of metallic, ionic and covalent bonds

Bond Type	Melting and Boiling Points	Hardness	Conductivity: Solid	Conductivity: Liquid	Conductivity: Aqueous
Metallic	High	Hard	Yes	Yes	Yes
Covalent	Low	Soft	No	No	No
Ionic	High	Hard	No	Yes	Yes

Intermolecular Forces: Just as atoms are held together by chemical bonds, molecules have attractive forces acting on them in solid and liquid states. As polar molecules have positive and negative ends, polar molecules are also dipoles. Positive area of 1 dipole molecule is attracted to the negative portion of an adjacent dipole molecule. These attractive forces are dipole-dipole forces.

Hydrogen Bonds: Is an intermolecular bond between a hydrogen atom in one molecule and a nitrogen, oxygen or fluorine atom in another molecule. Perhaps the most important example is water. Water is a polar molecule, so you can expect it to be held by other water molecule by dipole-dipole attractions.



- However, if only dipole-dipole forces acted to hold water molecules together, all of the water on Earth would have been boiled away. The water has not boiled away because there are relatively strong hydrogen bonding forces are acting on water molecules. Hydrogen bonding forces are much stronger than dipole-dipole attractions. Hydrogen bonding is responsible for the relatively high boiling point of water.
- Hydrogen bonding only occurs between hydrogen and nitrogen, oxygen and fluorine. It is worth noting that N,O,F are highly electronegative and have small radii and only 2 energy levels.