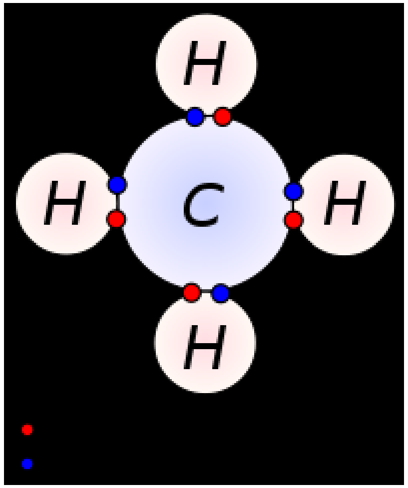
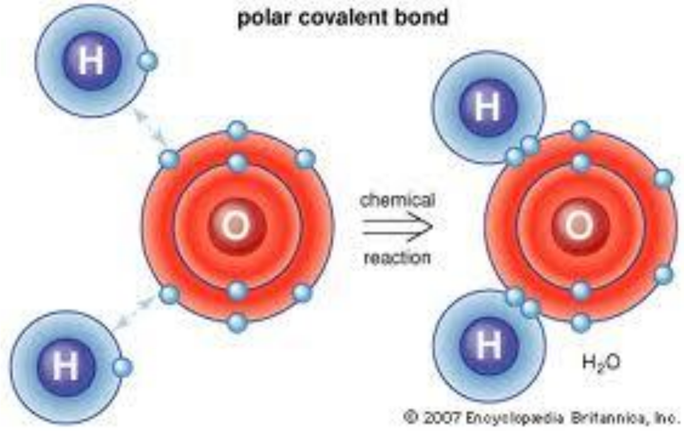
Ionic and Covalent Bonds



Section #2



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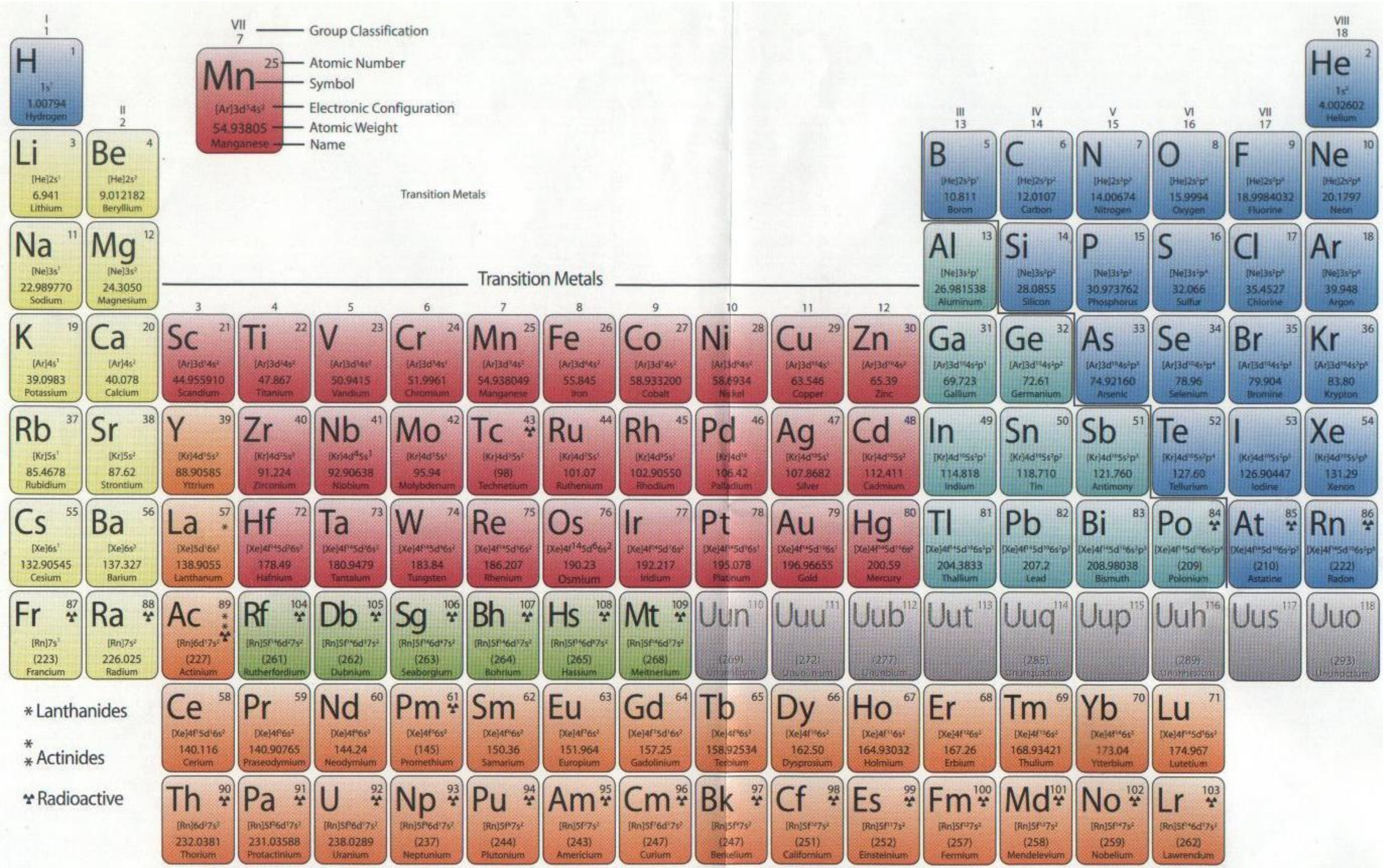


Atomic weights are based on $^{12}C = 12$ and conform to the 1995 IUPAC reported values. Number in ( ) indicates the isotope of longest half-life.

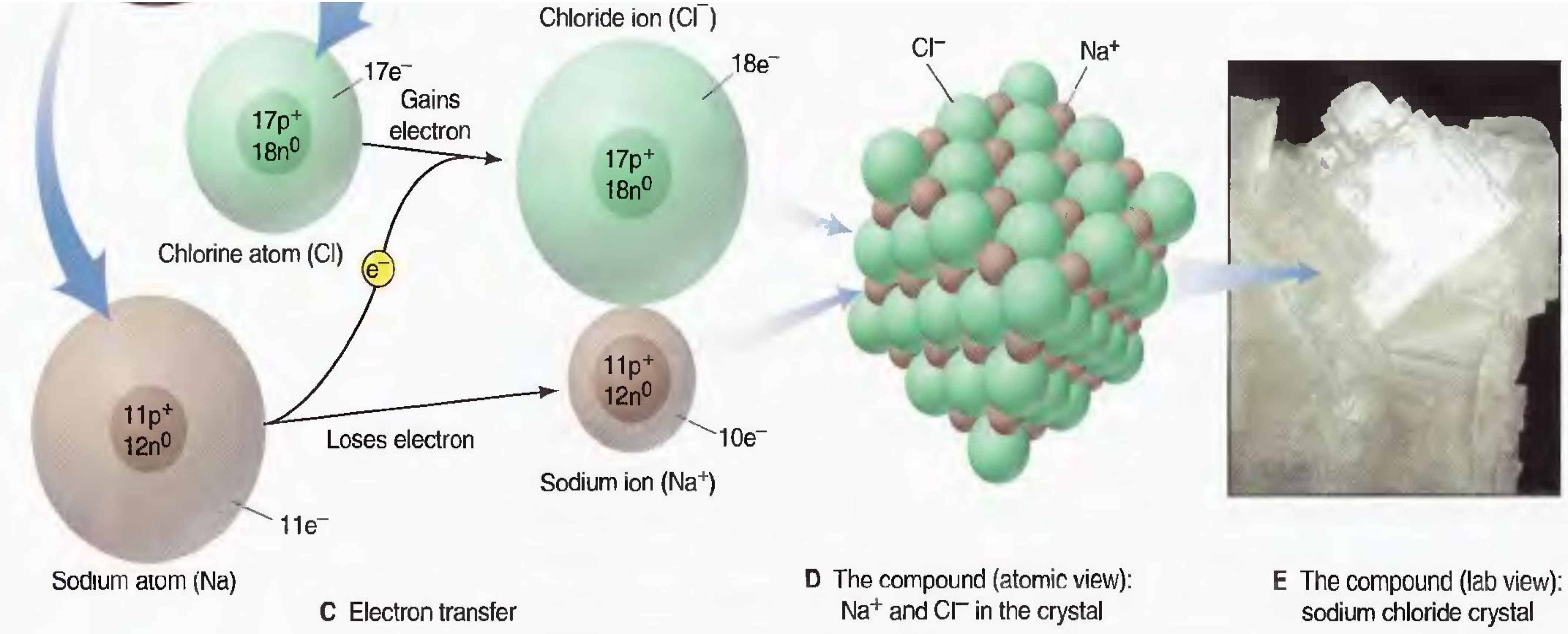
Compounds: Introduction to Bonding

- The noble gases - helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn) – occur in air as separate atoms
- Other molecules: O_2, N_2, H_2, S_8, C, Cu, Ag, Pt, Au
- Elements combine to form a compound in 2 general ways:
  - ***Transfering electrons*** from the atoms of one element to those of another to form ***ionic compounds***
  - ***Sharing electrons*** between atoms of different elements to form ***covalent compounds***.
- These processes generate **chemical bonds**, the forces that hold the atoms of elements together in a compound.

Formation of Ionic Compounds

- **Ionic compounds** are composed of ions – charged particles that form when an atom (or small group of atoms) gains or losses one or more electrons.
- **Binary ionic compound**: one composed of just two elements, typically forms when a metal reacts with a nonmetal.
- Each metal atom loses a certain number of its electrons and becomes a **CATION**, a positively charged ion.
- The non metal atoms gain the electron loss by the metal atoms and becomes **ANIONS**, negatively charged ion.
- The Cations and Anions attract each other through electrostatic forces and form the ionic compounds
- All binary ionic compounds are solid
- A cation or anion derived from a single atom is called a **MONATOMIC ION**

Example of the formation of ionic compound -- NaCl



IONIC BONDING ANIMATION

IONIC & COVALENT BONDING

- Sodium atom losses 1 electron and forms a sodium cation (Na^+).
- A Chlorine atom gains the electron and becomes a chloride anion (Cl^-).
- The oppositely charge ions (Na^+ and Cl^-) attract each other
- The resulting solid aggregation is a regular array of alternating Na^+ and Cl^- ions that extends in all three dimensions

STRENGTH OF IONIC BONDING

- depends on net strength of attractions and repulsions and is described by Coulomb's Law
- *The energy of attraction (or repulsion) between two particles is directly to the product of the charges and inversely proportional to the distance between them*
- *Ions with higher charges attract (or repel) each other more strongly than ions with lower charges*

$$Energy \propto \frac{\text{Charge1} * \text{Charge 2}}{\text{Distance}}$$

- Smaller ion attract (or repel) each other more strongly than larger ions, because their charges are closer together.
- Ionic compounds are neutral, that is, they posses no net charge
- Figure 2.13

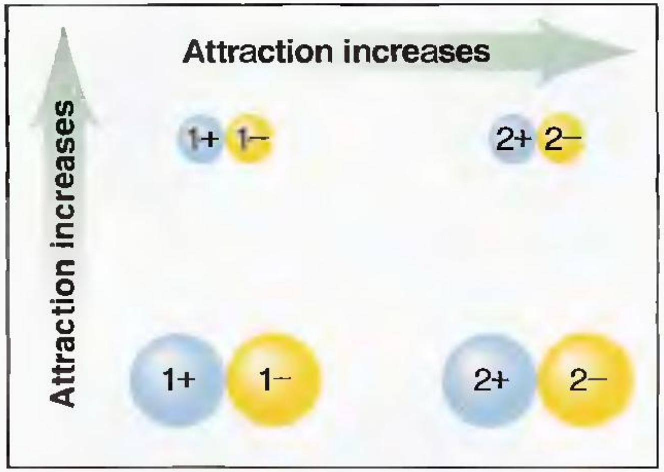
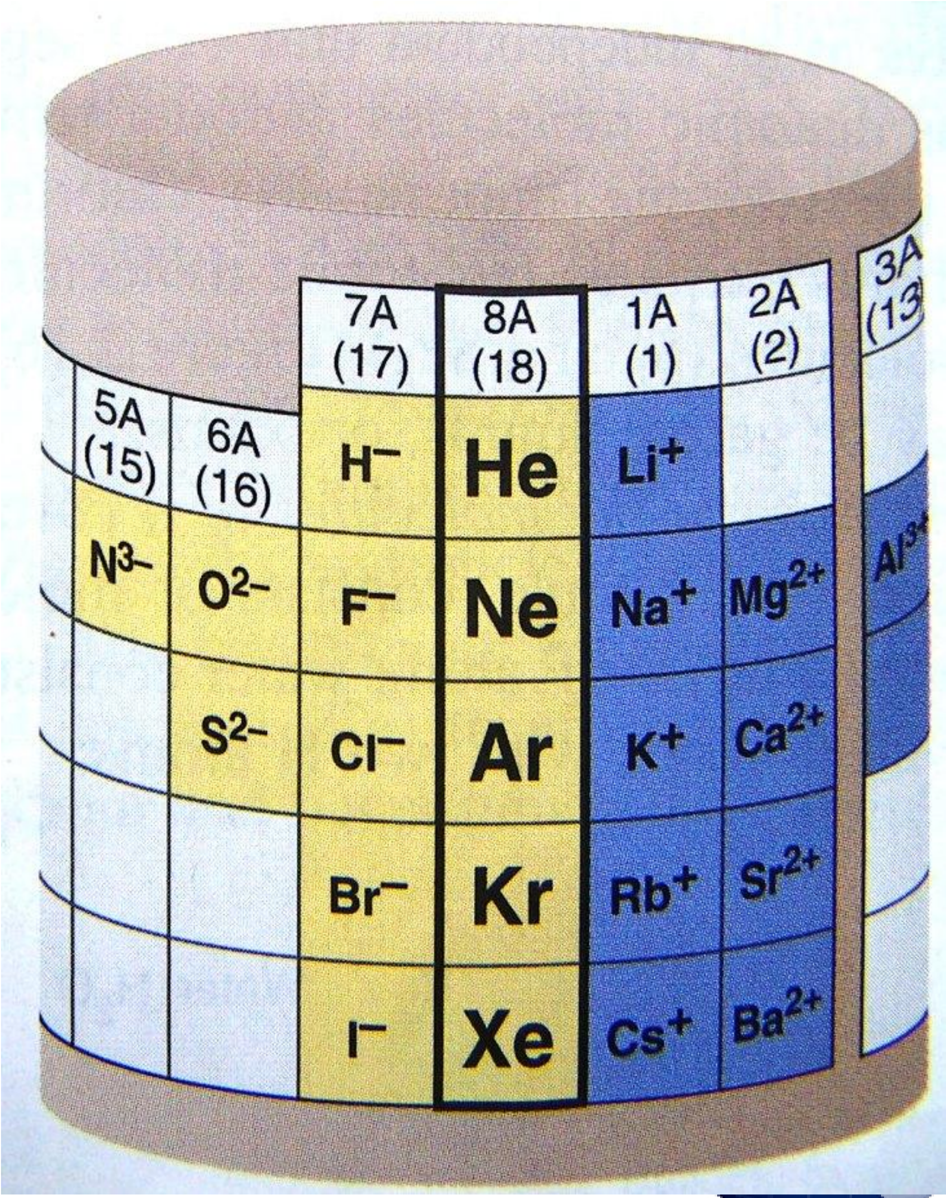


Figure 2.13 Factors that influence the strength of ionic bonding. For ions of a given size, strength of attraction *(arrows)* increases with higher ionic charge *(left to right)*. For ions of a given charge, strength of attraction increases with smaller ionic size *(bottom to top)*.

Why Why Why Why ??????

- Why does each sodium atom give up only 1 of its 11 electrons?
- Why doesn't each chlorine atom gain two electrons, instead of just one?
- ==> ***PERIODIC TABLE will answer***
- In general, metals lose electrons and non metals gain electrons to form ions with the same number of electrons as in the nearest noble gas [GROUP 8A(18)]

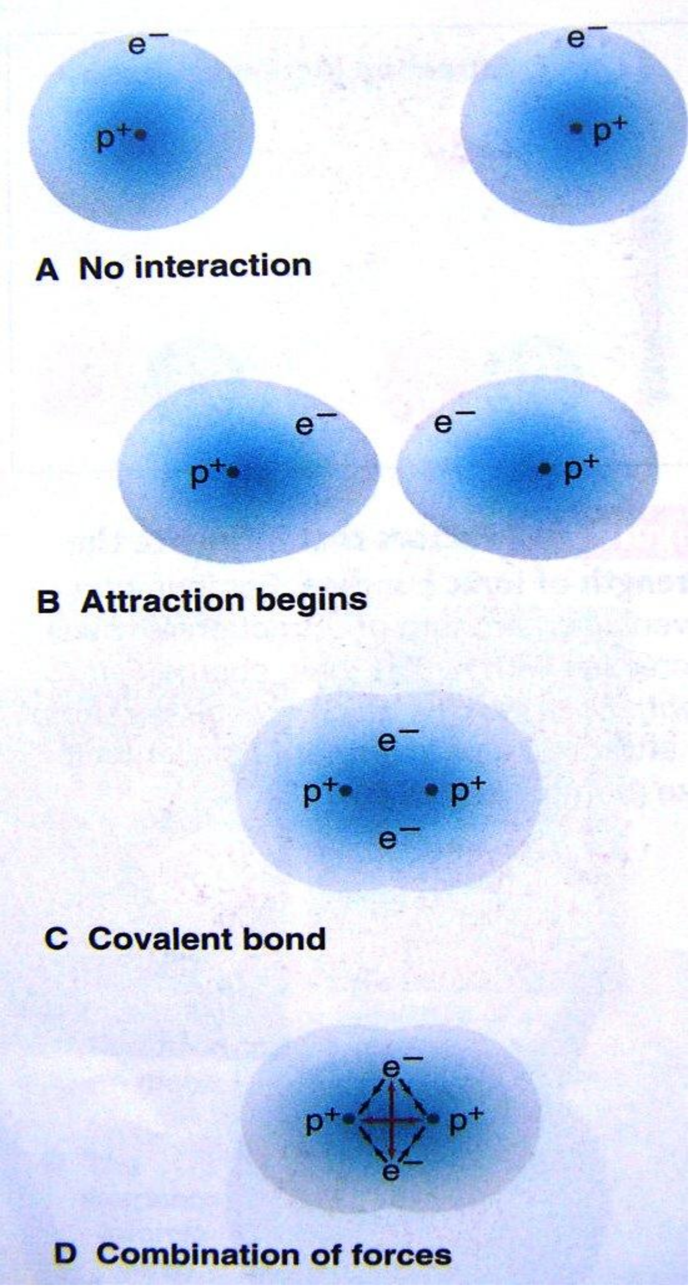
- Noble gases have a stability (low reactivity) that is related to their number of electrons
- Sodium ion (11e-) can attain the stability of neon (10e-) by losing one electron
- By gaining one electron, a chlorine atom (17e-) attains the stability of argon (18e-).
- Thus, when an element located near a noble gas forms a monatomic ion, it gains or loses enough electrons to attain the same number as that noble gas



In Periodic Table

- The elements in Group 1A(1) lose one electron, those in Group 2A(2) lose two, and aluminum in Group 3A(13) loses three
- The elements in Group 7A(17) gain one electron, oxygen and sulfur in Group 6A(16) gain two, and nitrogen in Group 5A(15) gains three
- Fluorine (F, Z=9) has one electron fewer and sodium (Na, Z=10) has one electron more than the noble gas neon (Ne, Z=10), they form the F- and Na+ ions
- Similarly, oxygen (O, Z=8) gains two electrons and magnesium (Mg, Z=12) loses two to form the O^{2-} and Mg^{2+} ions and attain the same number of electrons as neon

Formation of Covalent Compounds

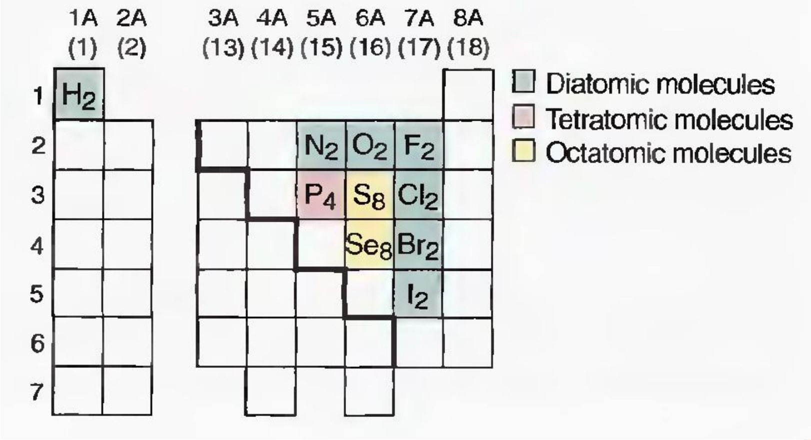


- Covalent compound form when elements share electrons, which usually occurs between nonmetals
- Example: Two Hydrogen atoms (H, Z=1):
- The nucleus of each atom attracts the electron of the other atom more and more strongly
- The separated atoms begin to interpenetrate each other
- At some optimum distance between the nuclei, the atoms form a **COVALENT BOND, where a pair of electrons mutually attracted by the two nuclei**

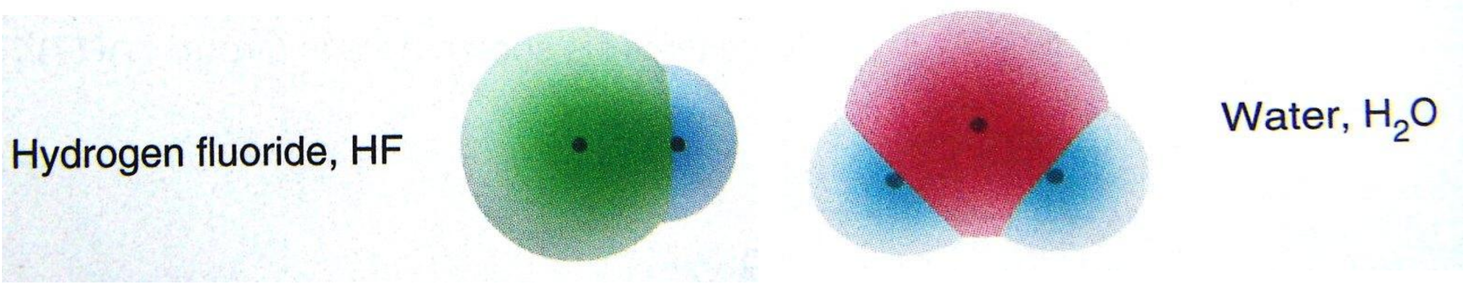
COVALENT BONDING

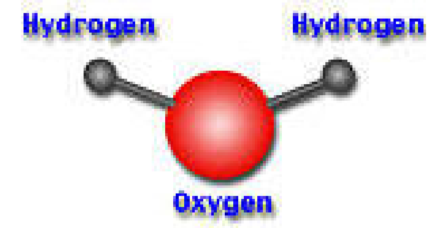
IONIC & COVALENT BONDING

- The result is a hydrogen molecule, in which each electron no longer “belongs” to a particular H atom; the two electrons are shared by the two nuclei
- Repulsions between the nuclei and between the electrons also occur, but the net attraction is greater than the net repulsion.
- Other nonmetals that exist as diatomic molecules at room temperature are nitrogen (N_2), oxygen (O_2), and the halogens (fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2)).
- Phosphorus exists as tetratomic molecules (P4), and sulfur and selenium as octatomic molecules (s_8 and Se_8).
- At room temperature, covalent compounds may be gases, liquids, or solids



- Atoms of different elements share electrons to form the molecules of a covalent bond
- Example: Hidrogen Fluoride (HF), Water (H_2O)



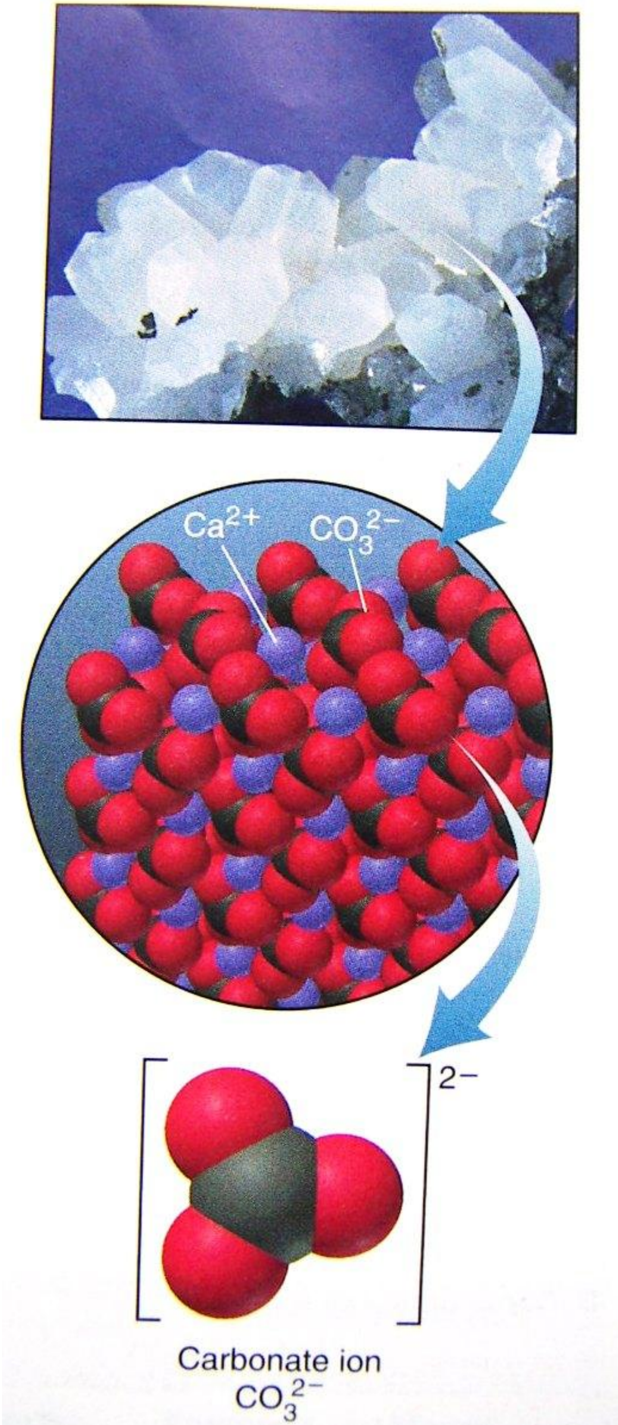


- ***Most covalent substances consists of molecules*** : a cup of water ==> consists of individual water molecules
- ***No molecules exist in a sample of ionic compound***: ==> NaCl is a continuous array of oppositely charged sodium and chloride ions, not a collection of individual NaCl molecules

Another Key Difference

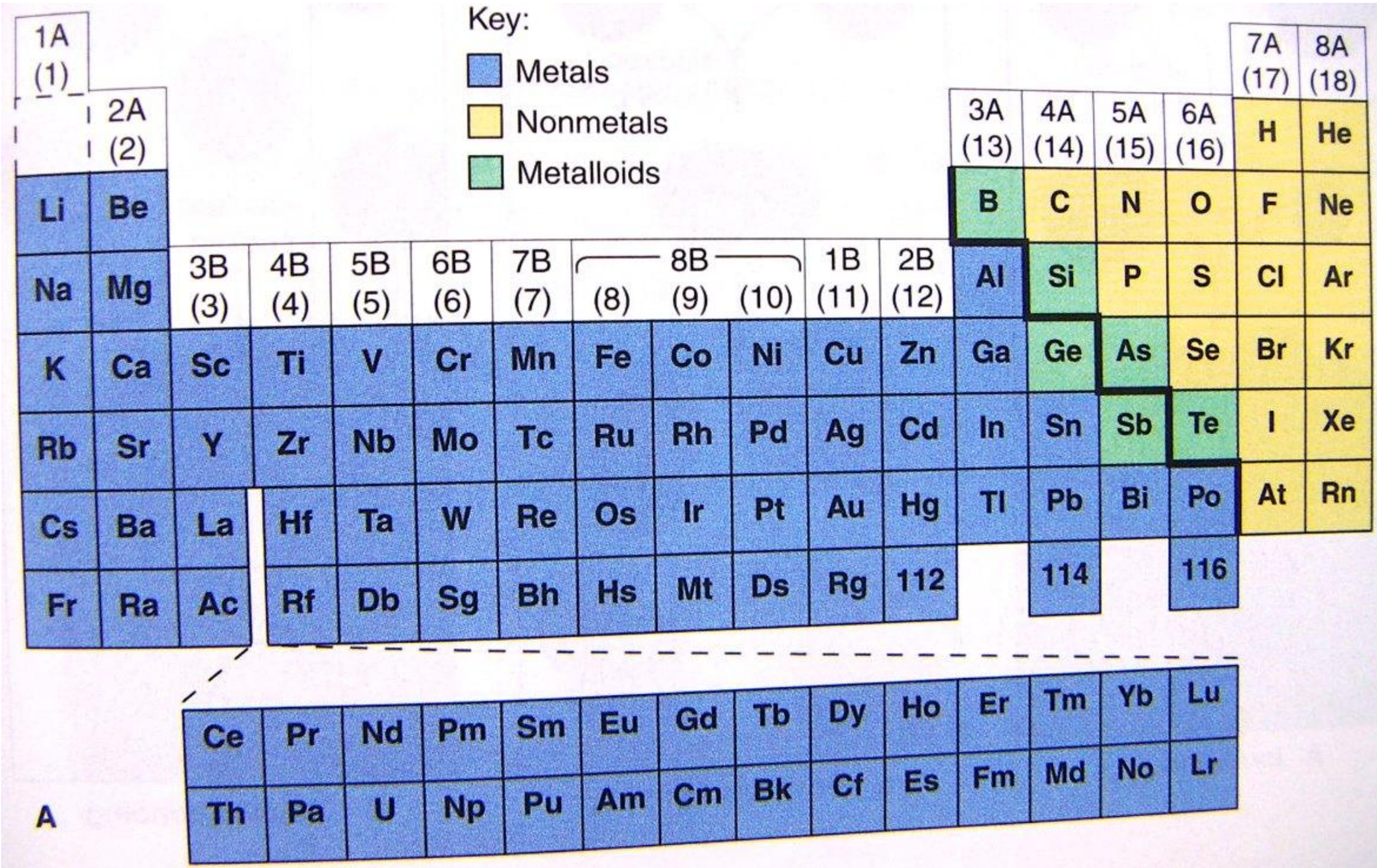
- ***Covalent bonding***: involves the mutual attraction between two (positivey charged) nuclei and the two (negatively charged) electrons that reside between them.
- ***Ionic bonding***: involves the mutual attraction between positive and negative ions

Polyatomic Ions



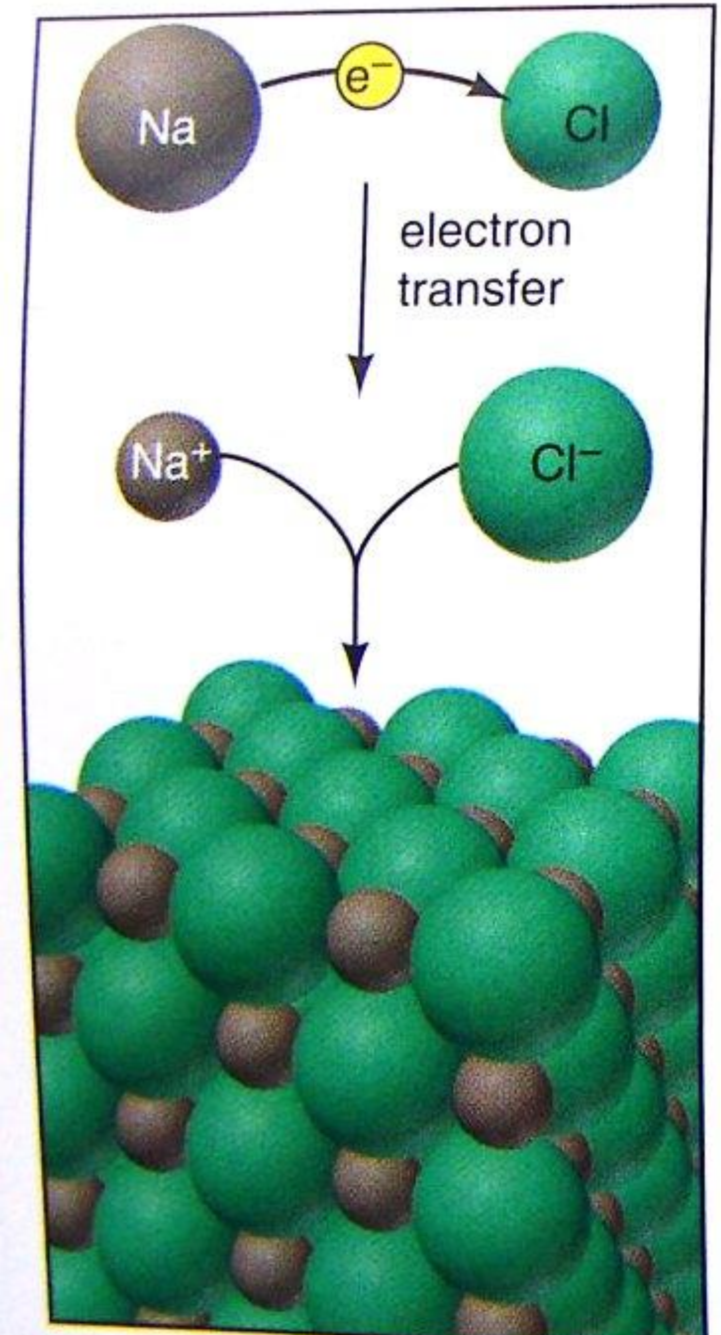
- Many ionic compounds contain polyatomic ions, which consist of two or more atoms ***bonded Covalently*** and have a net positive or negative charge
- ***Calcium carbonate*** is a 3D array of monatomic calcium cations and polyatomic carbonate anions.
- As the bottom structure shows, each ***carbonate*** ion consists of four covalently bonded atoms

Three Types of Chemical Bonding



Why atoms bond at all ?

- ***Bonding lowers the potential energy between positive and negative particles.***
- Just as the electron configuration and the strength of the nucleus-electron attraction determine the properties of an atom,
- the type and strength of chemical bonds determine the properties of a substance



A Ionic bonding

- ***(1). Metal with nonmetal***: electron transfer and ionic bonding
- Typically as IONIC, BONDING
- Such difference occur between reactive metals (1A(1) and 2A(2)) and nonmetals (7A(17)) and the top of Group 6A (16).
- The metal atom loses its one or two valence electrons, whereas the nonmetal atom gains the electron(s).
- Electron transfer from metal to nonmetal occurs, and each atom forms an ion with a noble gas electron configuration

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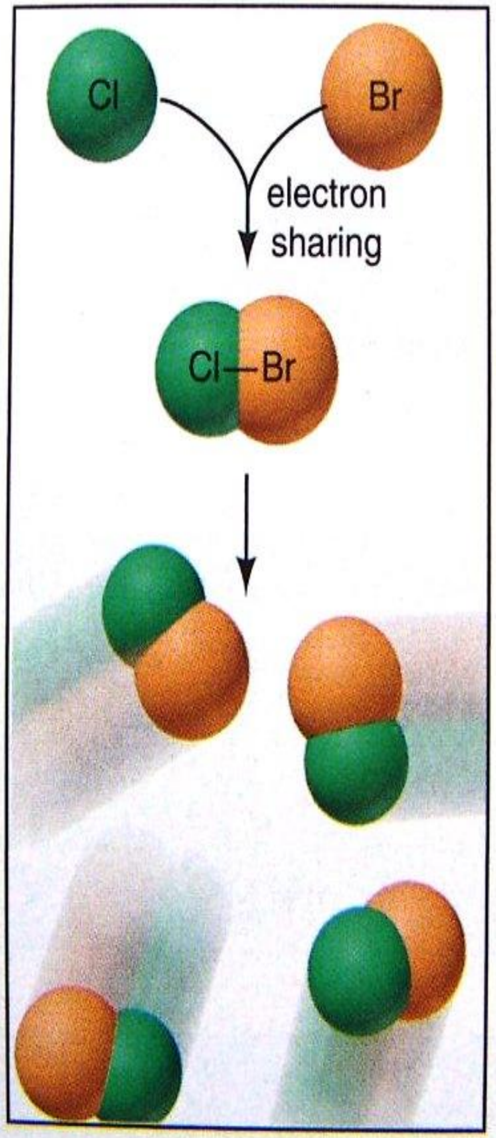
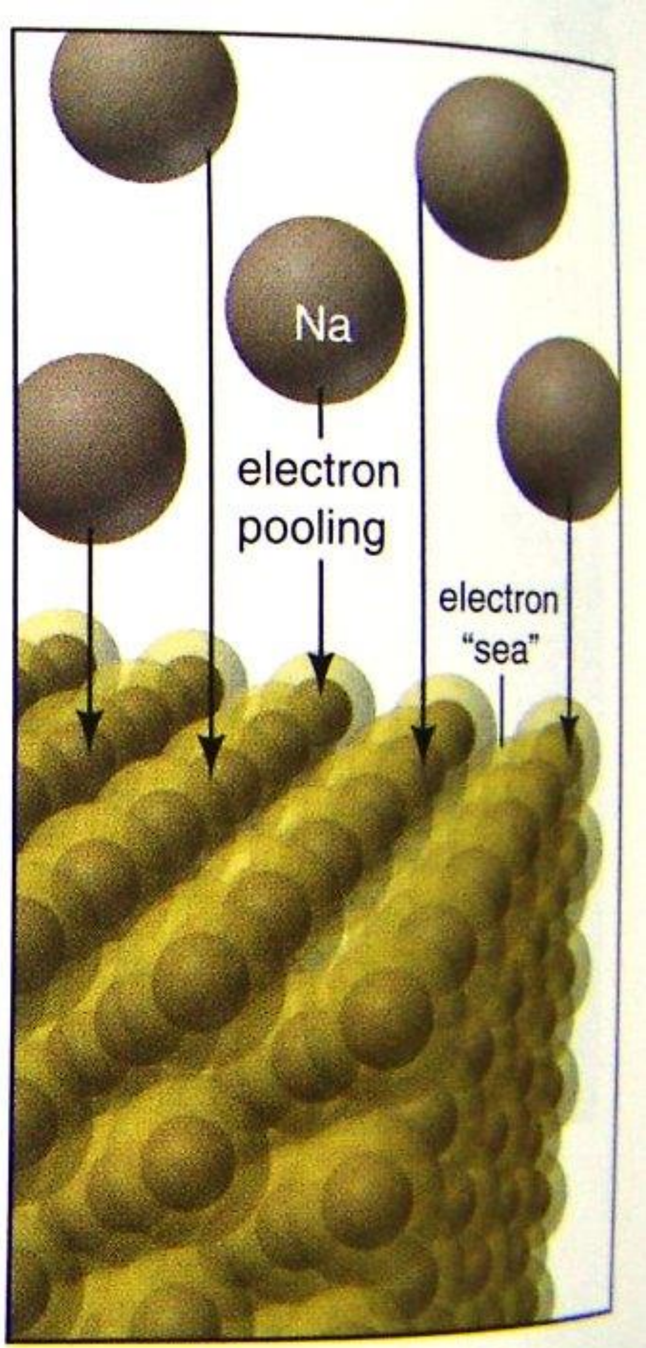
- The electrostatic attraction between these positive and negative ions draws them into the 3D array of an ionic solid
- whose chemical formula represents the cation-to-anion ratio (empirical formula)



B Covalent bonding

- ***Nonmetal with nonmetal***: electron sharing and covalent bonding.
- when two atoms have a small difference in their tendencies to lose or gain electrons, we observe *electron sharing* and *covalent bonding*
- This type of bonding most commonly occurs between nonmetal atom, Although a pair of metal atoms can sometimes form a covalent bond also
- The attraction of each nucleus for the valence electrons of the other draws the atoms together.
- A shared electron pair is considered to be “localized” between the atoms because it spends most of its time there, linking them in a covalent bond of a particular length and strength.



C Metallic bonding

- (3). ***Metal with metal***: *electron pooling* and *metallic bonding*
- In general, metal atoms are relatively large, and their view outer electrons are well shielded by filled inner levels
- Thus, they lose outer electrons comparatively easily but do not gain them very readily.
- These properties lead large numbers of metal atoms to share their valence electrons, but in a way that differs from covalent bonding

- All the metal atoms in a sample ***pool*** their valence electrons into an evenly distributed ***"sea" of electrons*** that "flows" between and around the metal-ion cores (nucleus plus inner electrons) and attracts them, thereby holding them together
- Unlike the localized electrons in covalent bonding, electrons in metallic bonding are "*delocalized*", *moving freely throughout the piece of metal*.
- In other word: Metallic bonding occurs when many metal atoms pool their valence electrons in a delocalized electron sea that holds all the atoms together.

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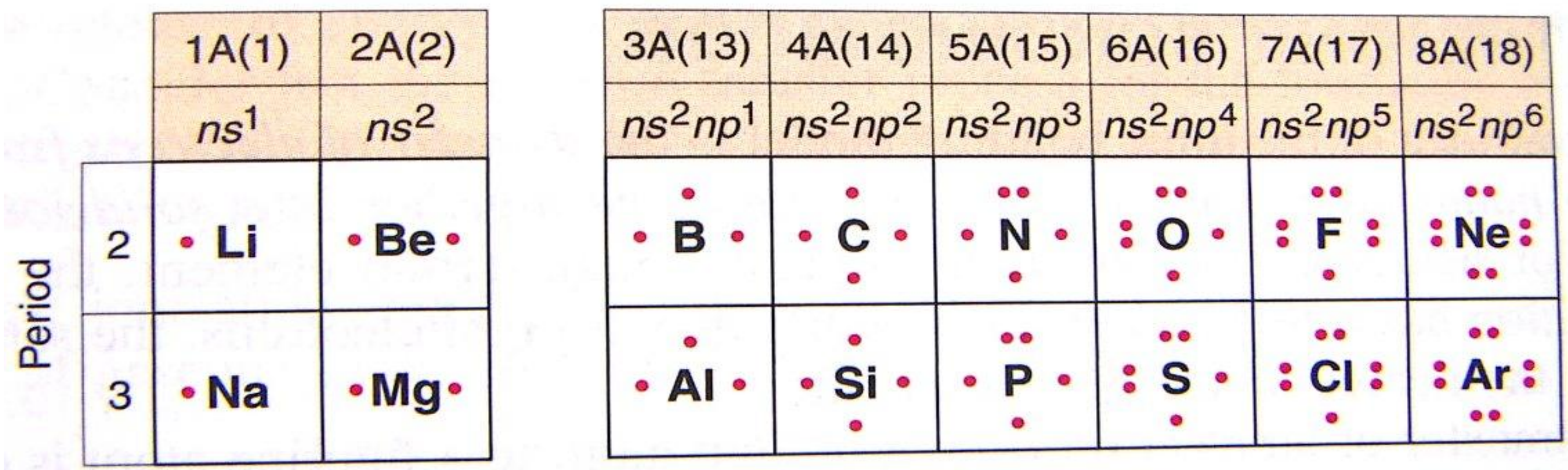
Notes

- You cannot always predict the bond type solely from the elements' positions in the periodic table
- All binary compounds contain a metal and a nonmetal, but all metals do not form binary ionic compounds with all nonmetals
- Example: metal beryllium (Group 2A(2)) combines with the nonmetal chlorine (Group 7A(17)), the bonding fits the covalent model better than the ionic model.

LEWIS ELECTRON-DOT SYMBOLS



- The element symbol represents the nucleus and the inner electrons
- The dots around it represent valence electrons, either paired or unpaired
- The number of unpaired dots indicates the number of electrons a metal atom loses, or the number of nonmetal atom gains, or the number of covalent bonds a nonmetal atom usually forms

Lewis Symbol .....

- Note its A-group number (1A to 8A), which gives the number of valence electrons
- Place one dot at a time on the four sides (top, right, bottom, left) of the element symbol
- Keep adding dots, pairing the dots until all are used up
- The placement of dots is not important
- The number and pairing of dots provide information about an element's bonding behavior:
  - For a metal, the total number of dots is the maximum number of electrons an atom loses to form a cation
  - For a nonmetal, the number of unpaired dots is the number of electrons that become paired either through electron gain or through electron sharing.
  - Thus the number of unpaired dots equals either the number of electrons a nonmetal atom gains in becoming an anion or the number of covalent bonds it usually forms

Lewis : ==> OCTET RULE

- when atoms bond, they lose, gain, or share electrons to attain a filled outer level of eight (or two) electrons
- The octet rule holds for nearly all of the compounds of Period 2 elements and a large number of others as well