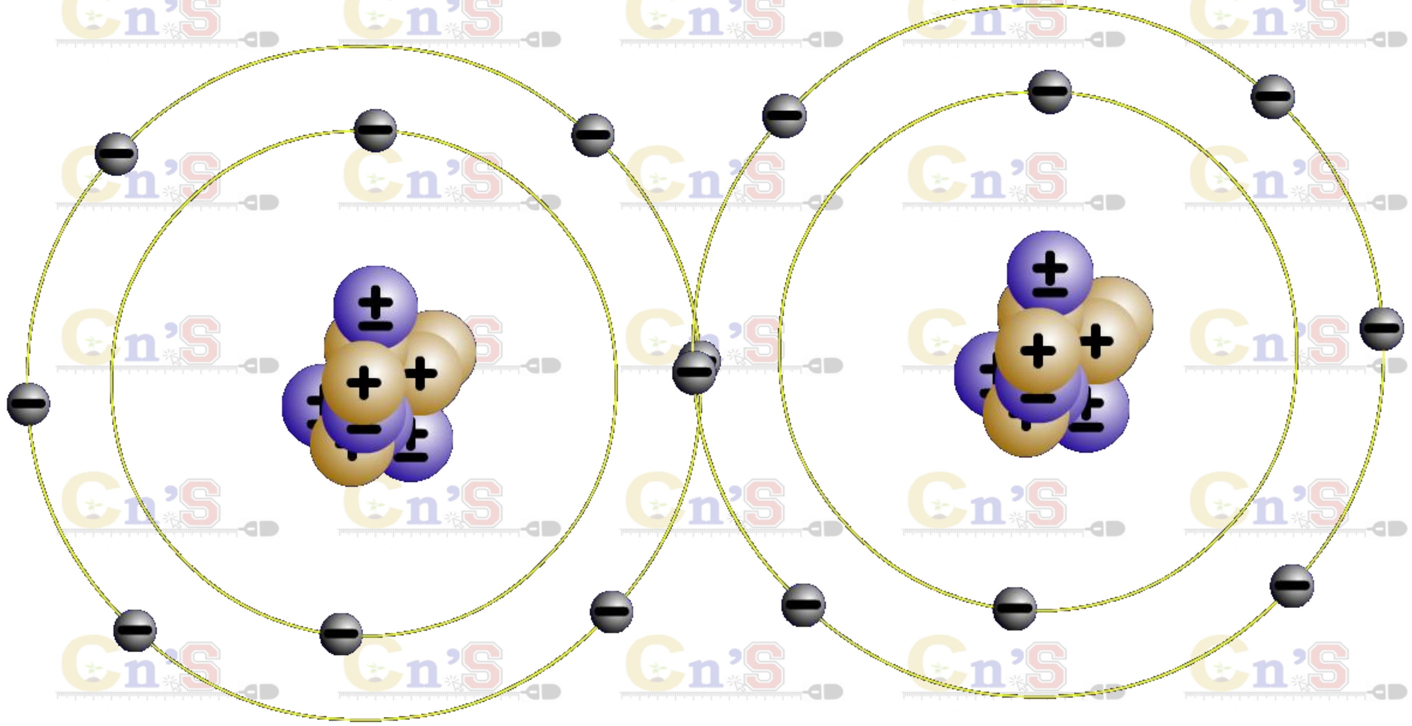
CHEMICAL BONDING



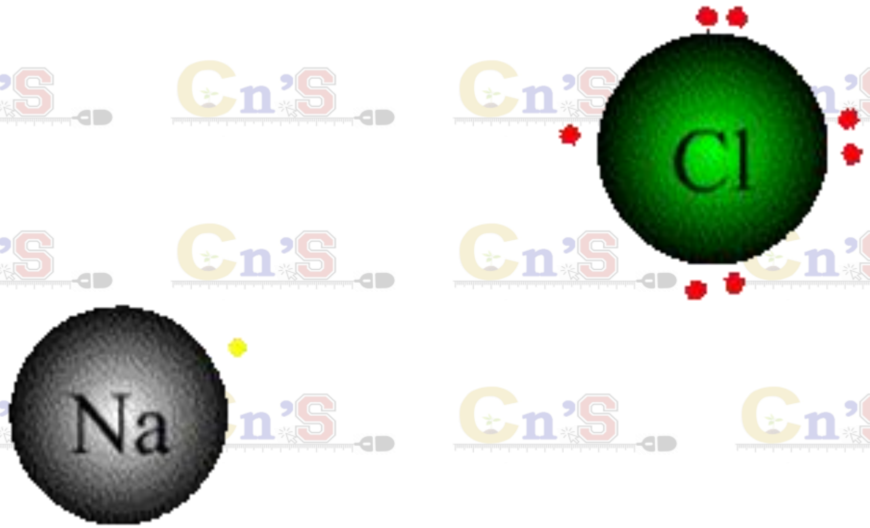
OBJECTIVES

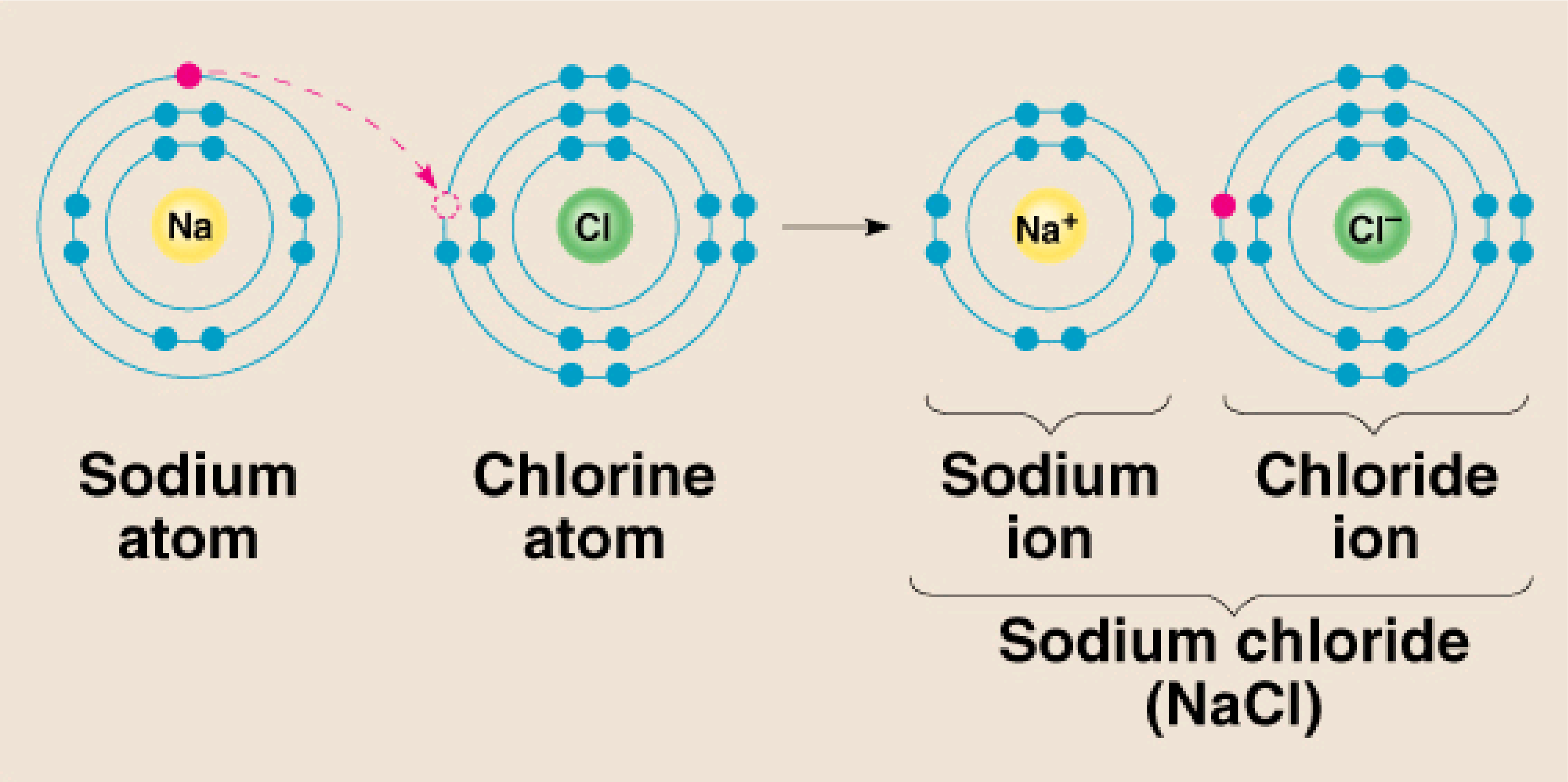
- overview chemical bonds to understand the participation of valence electrons by sharing electrons.
- explain the formation of covalent bonds by sharing electrons.
- identifiy the single bonds and multiple bonds.
- describe the rules regarding the drawing of Lewis structures.
- draw Lewis structures of covalent molecules and groups of ions.
- compare the nature of non- polar covalent bonds, polar covalent bonds and ionic bonds depending on the difference of electronegativity of the atoms involved in the bond.

- describe the polar covalent nature of bond and molecules using the concepts of polarization and dipole moment giving suitable examples.
- explain the formation of the dative-covalent bond.
- explain the formation of ionic bonds.
- explains the structure and physical properties of ionic lattice using sodium chloride.
- explain the covalent character of ionic bonds based on the polarizing power of cations and polarizability of anions.
- compare the ionic character and covalent character of compound
- explain the structure of the metallic bond.
- states the covalent, ionic and metallic bonding as primary interactions.

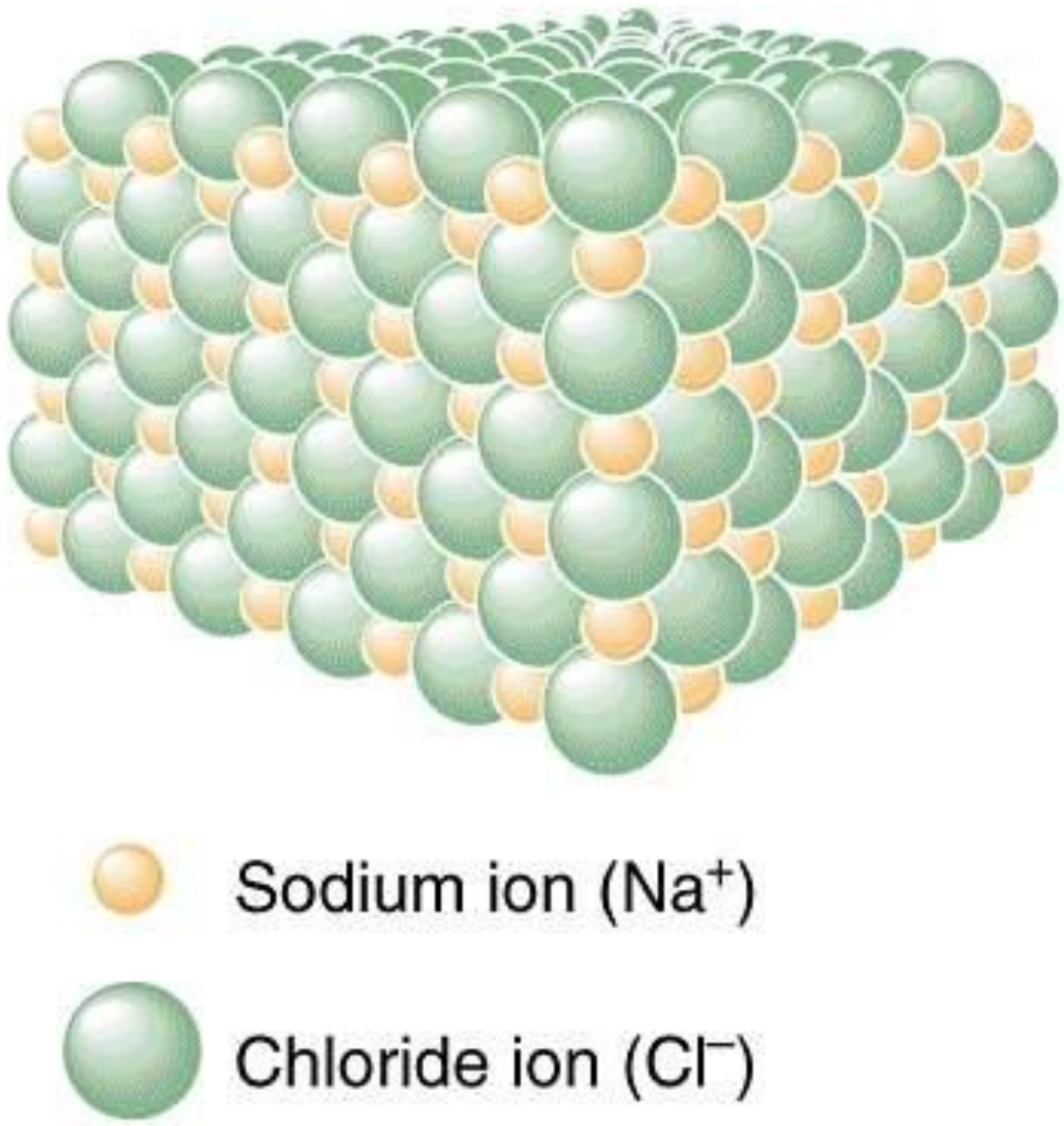
Ionic Bond

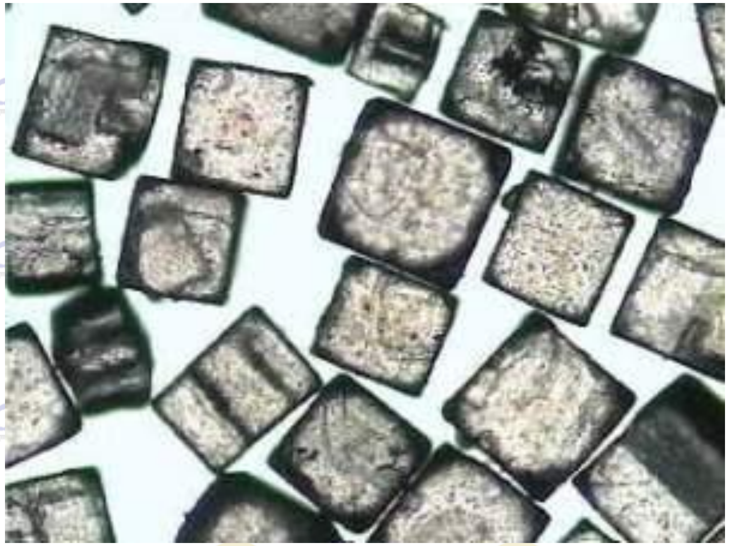
- Between atoms of metals and nonmetals with very different electronegativity
- Bond formed by transfer of electrons
- Produce charged ions all states. Conductors and have high melting point.
- Examples; NaCl, $CaCl_2$, K_2O



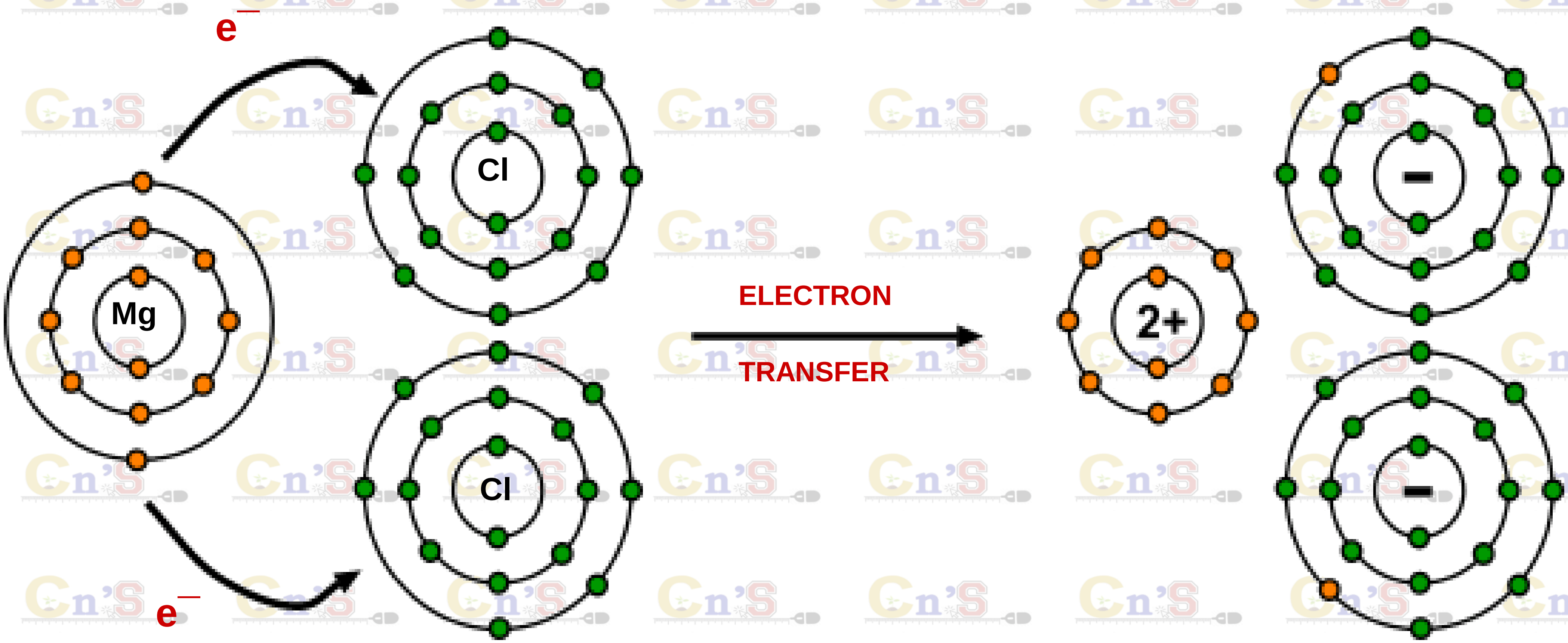


1). Ionic bond – electron from Na is transferred to Cl, this causes a charge imbalance in each atom. The Na becomes (Na^+) and the Cl becomes (Cl^-), charged particles or ions.


Sodium ion (Na^+)
Chloride ion (Cl^-)

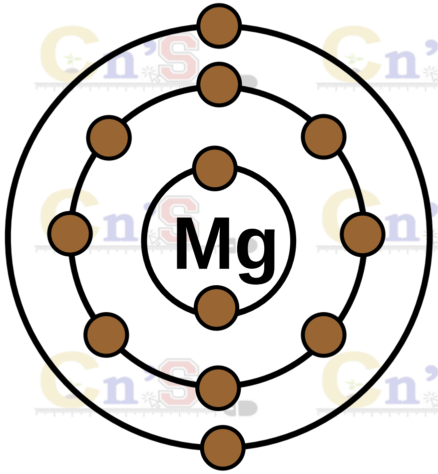


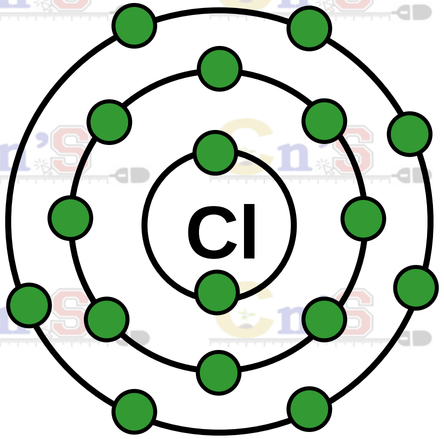
FORMATION OF MAGNESIUM CHLORIDE



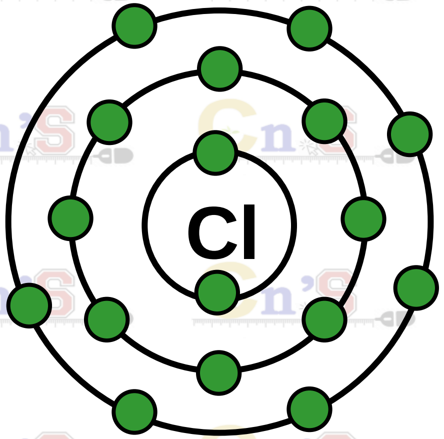
$Mg \longrightarrow Mg^{2+} + 2e^{-}$ and $2Cl + 2e^{-} \longrightarrow 2\,Cl^{-}$

MAGNESIUM CHLORIDE

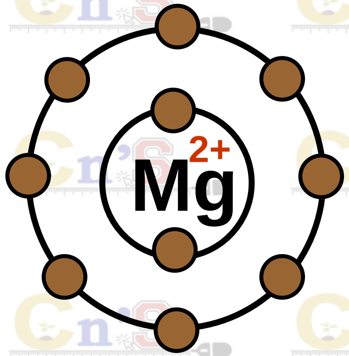


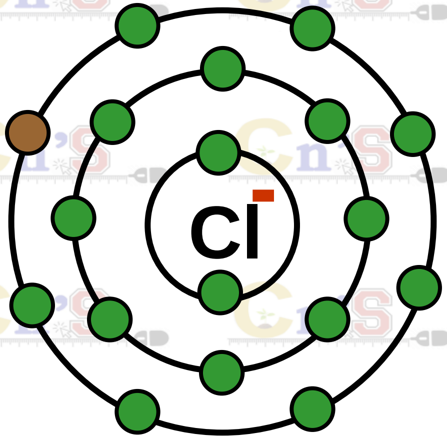


CHLORINE ATOMS
2,8,7

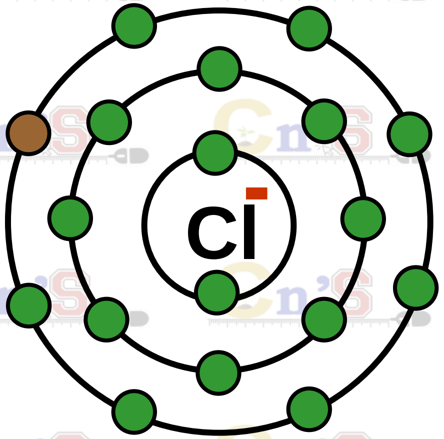


MAGNESIUM CHLORIDE





CHLORIDE IONS
2,8,8

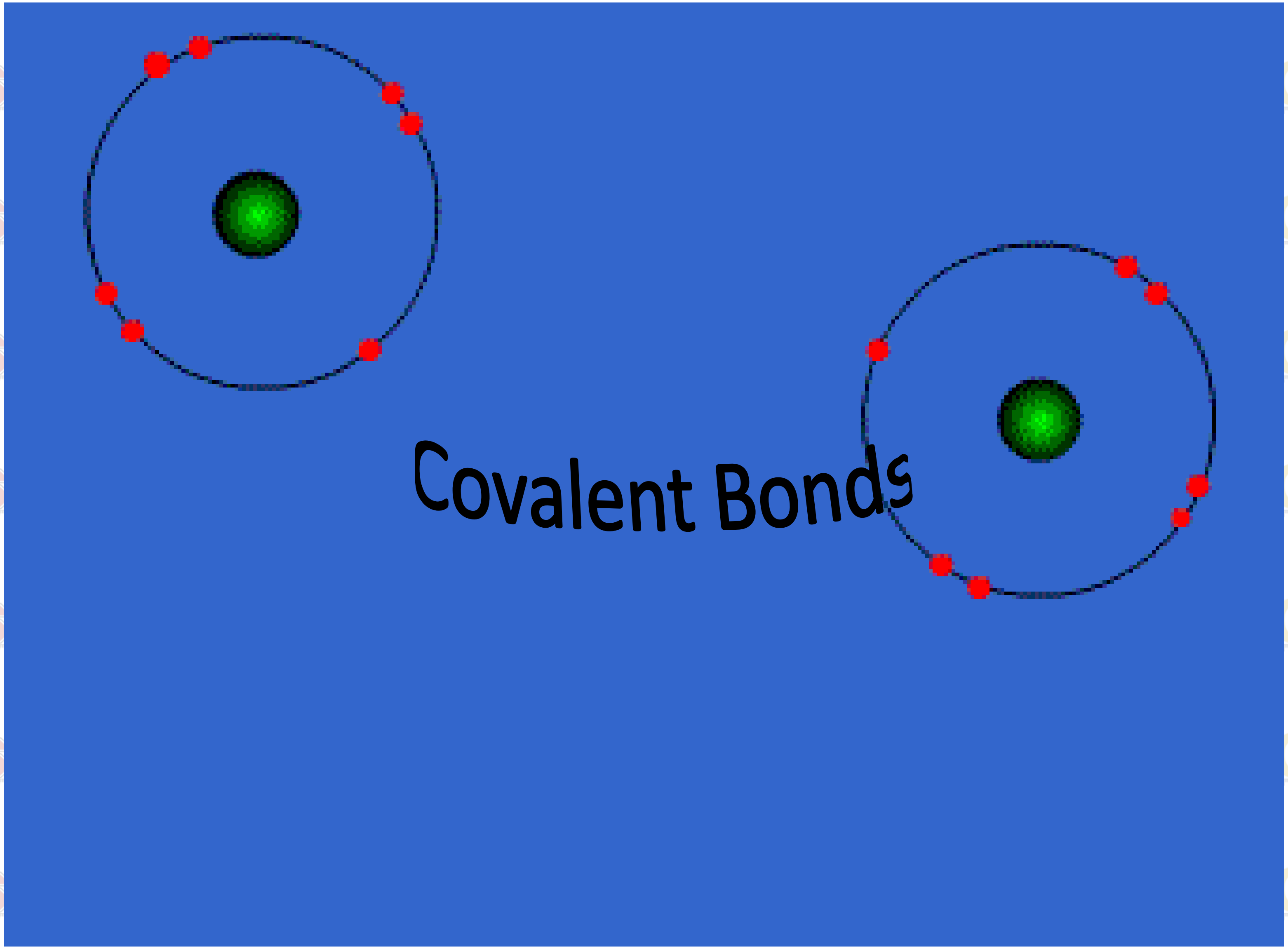


<u>COVALENT BOND</u>

bond formed by the *sharing* of electrons

Covalent Bond

- Between nonmetallic elements of similar electronegativity.
- Formed by sharing electron pairs
- Stable non-ionizing particles, they are not conductors at any state
- Examples; O_2, CO_2, C_2H_6, H_2O, SiC

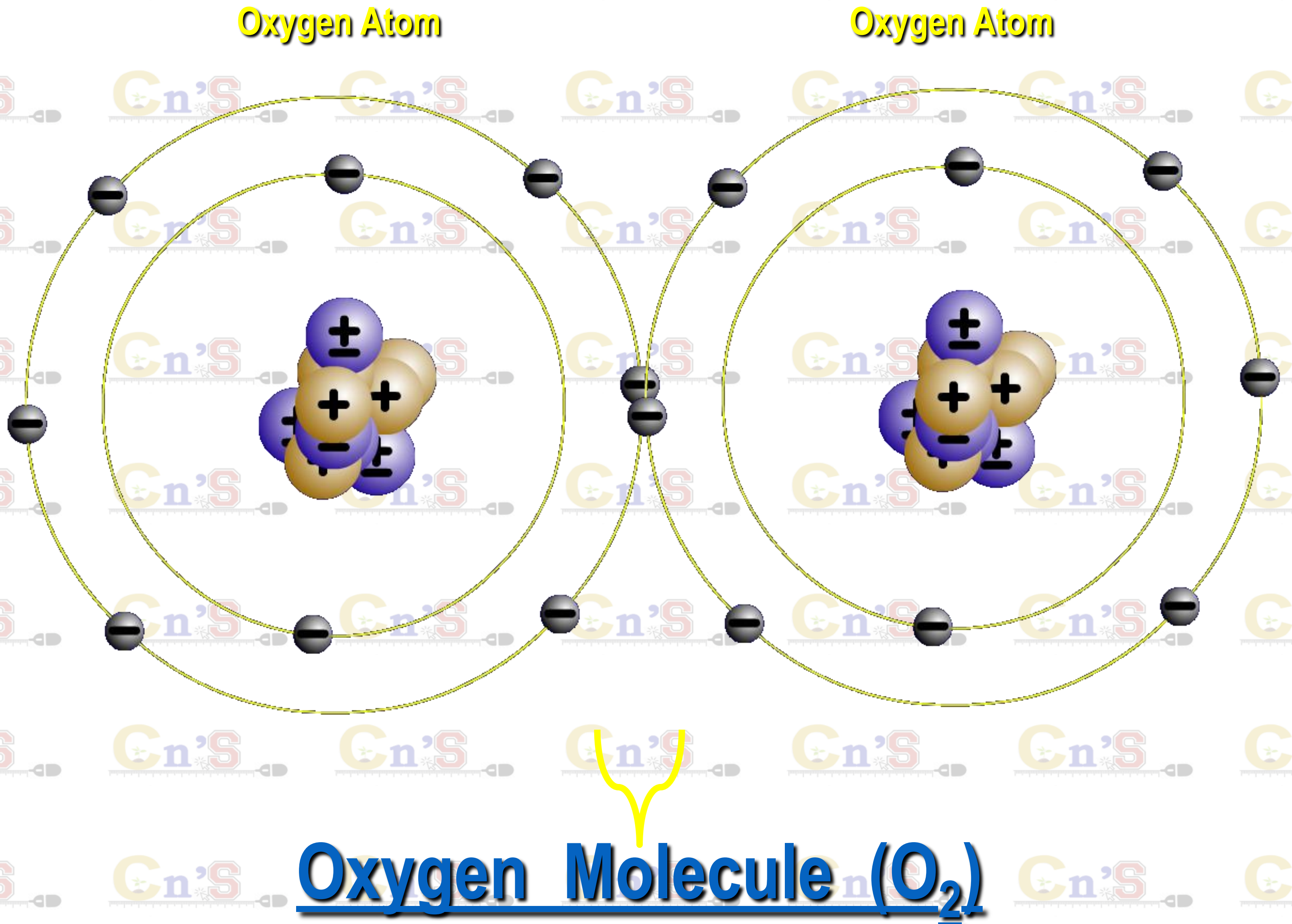

Covalent Bonds

NONPOLAR COVALENT BONDS

when electrons are shared *equally*

H_2 or Cl_2

2. Covalent bonds- Two atoms share one or more pairs of outer-shell electrons.



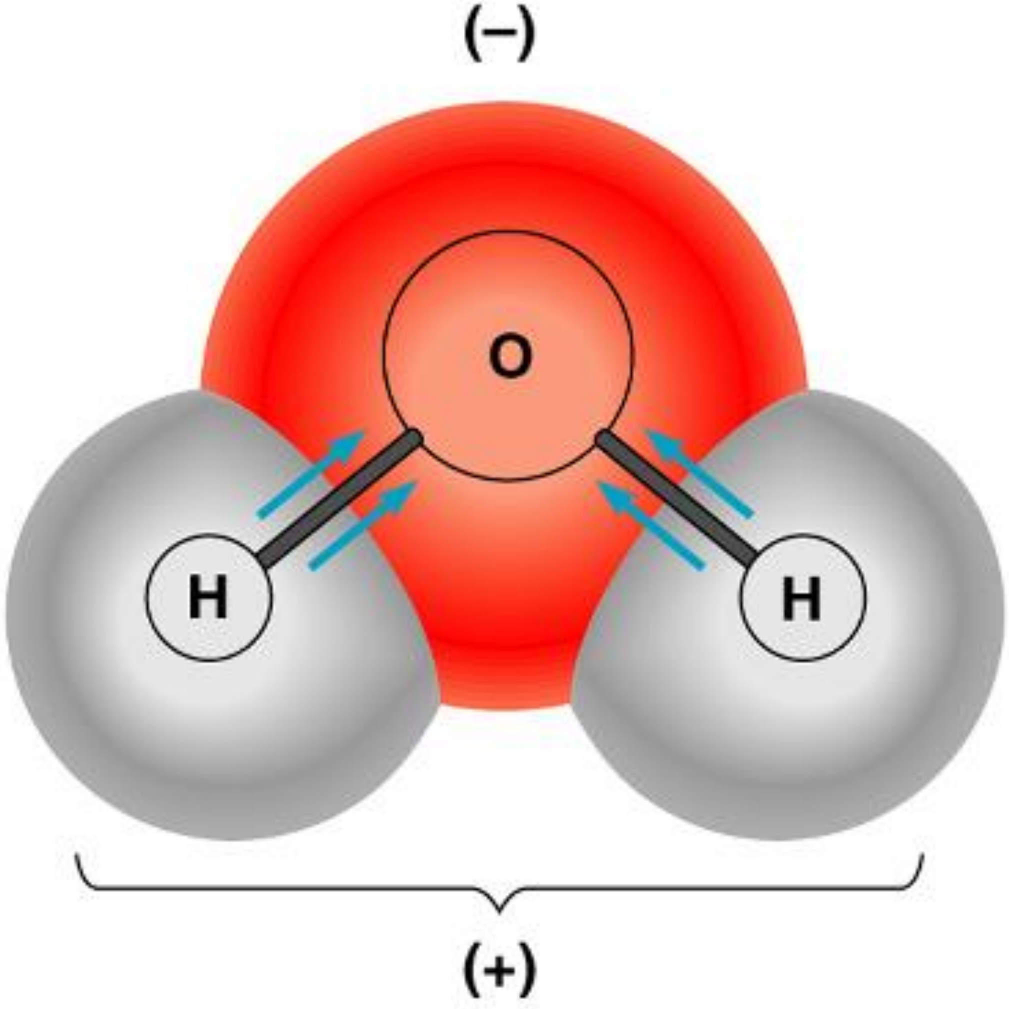
$H-H$ $\quad$ $:\underset{\cdot\cdot}{\ddot{F}}-\underset{\cdot\cdot}{\ddot{F}}:$ $\quad$ $:N\equiv N:$

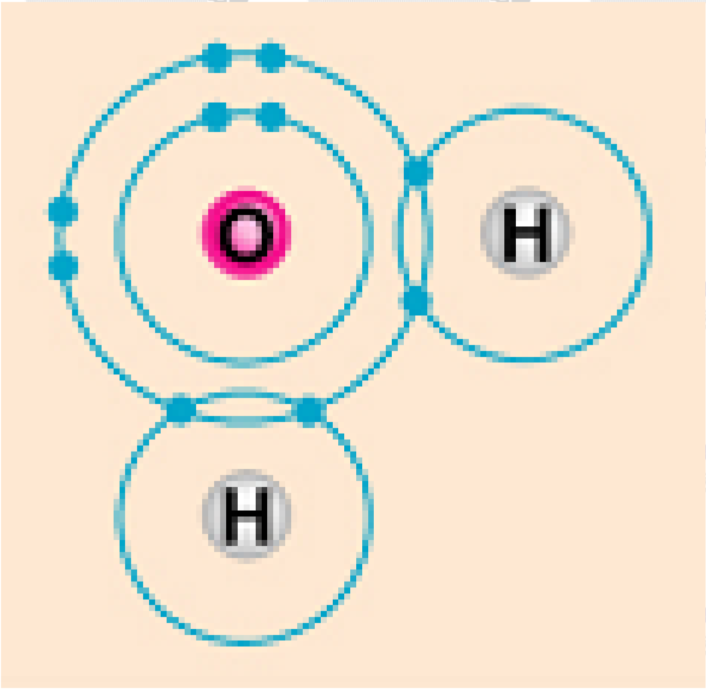
nonpolar bonds connect atoms of the same electronegativity

POLAR COVALENT BONDS

when electrons are shared but shared *unequally*

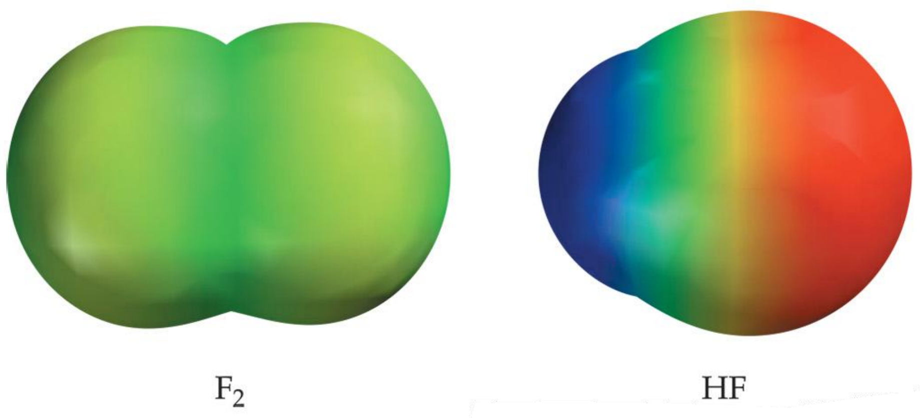
H_2O





- water is a *polar molecule* because oxygen is more electronegative than hydrogen, and therefore electrons are pulled closer to oxygen.

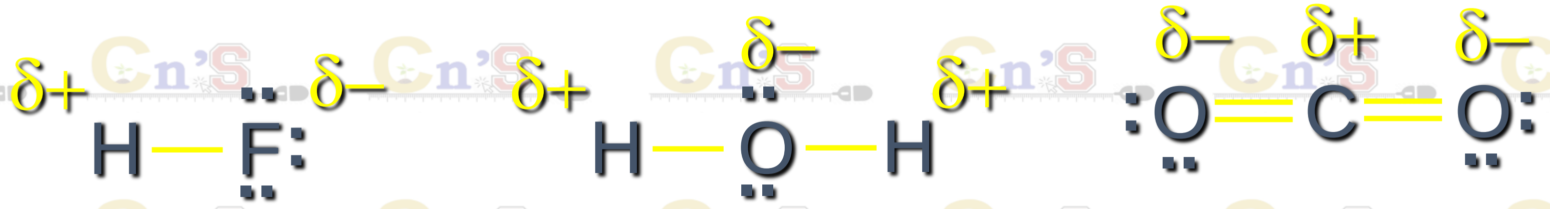
Polar Covalent Bonds



- Covalent bond, sharing electrons,
- But electron sharing not always equal.

- Fluorine pulls harder on the shared electrons than hydrogen does.
- Therefore, the fluorine end has more electron density than the hydrogen end.
- But how do you know who pulls hardest?

- The greater the difference in electronegativity between two bonded atoms; the more polar the bond.

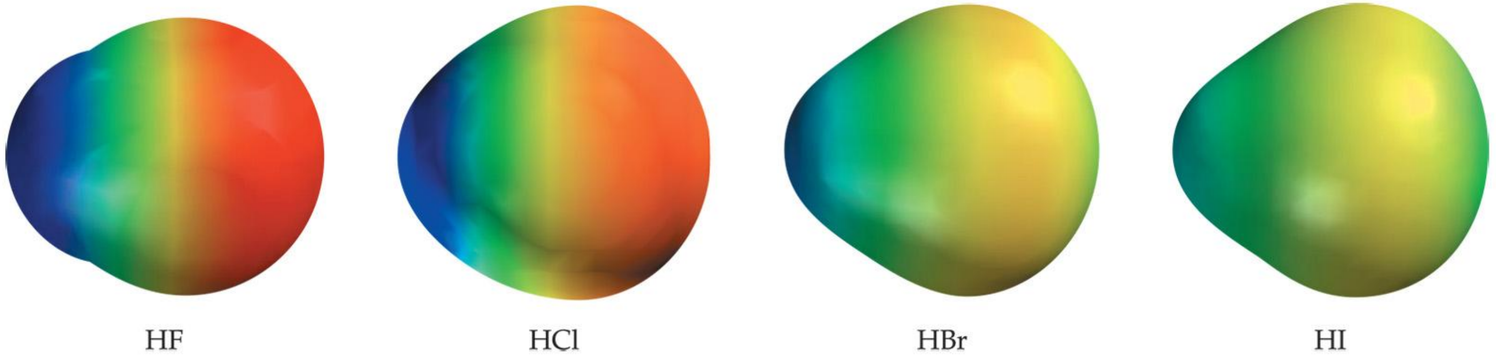


polar bonds connect atoms of different electronegativity

Polar Covalent Bonds

Compound	Bond Length (Å)	Electronegativity Difference	Dipole Moment (D)
HF	0.92	1.9	1.82
HCl	1.27	0.9	1.08
HBr	1.41	0.7	0.82
HI	1.61	0.4	0.44

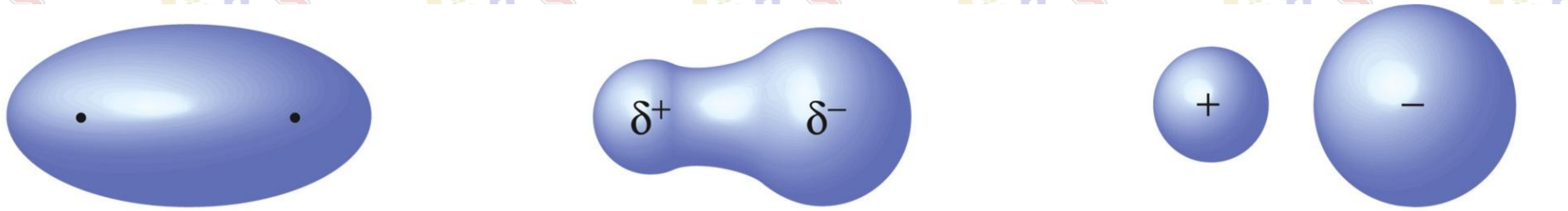
The greater the difference in electronegativity, the more polar is the bond.



Polar Covalent Bonds

- The absolute value of the difference in electronegativity of two bonded atoms gives a rough measure of the **polarity** of the bond.
  - When this difference is small (less than 0.5), the bond is **nonpolar**.
  - When this difference is large (greater than 0.5), the bond is considered **polar**.
  - If the difference exceeds approximately 1.8, sharing of electrons is no longer possible and the bond becomes **ionic**.

- The polarity of a bond depends on the difference between the electronegativity values of the atoms forming the bond.



a
A covalent bond formed between identical atoms.

b
A polar covalent bond, with both ionic and covalent components.

c
An ionic bond, with no electron sharing.

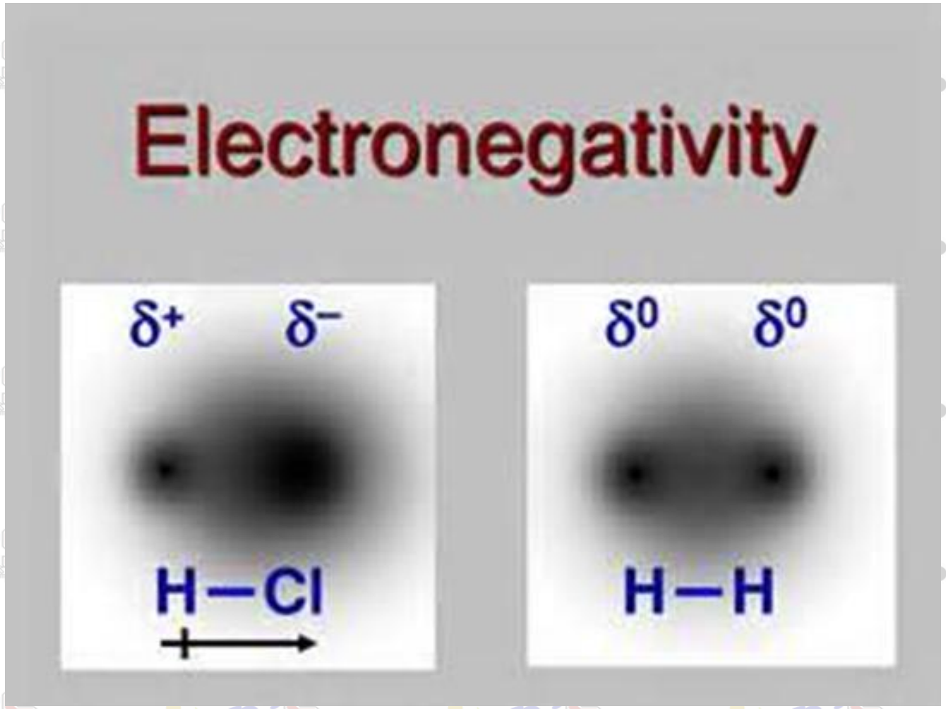
Polar Covalent Bonds and Electronegativity

Electronegativity

Electronegativity is a measure of an element to attract electrons toward itself when bonded to another element.

•An electronegative element attracts electrons.

•An electropositive element releases electrons.

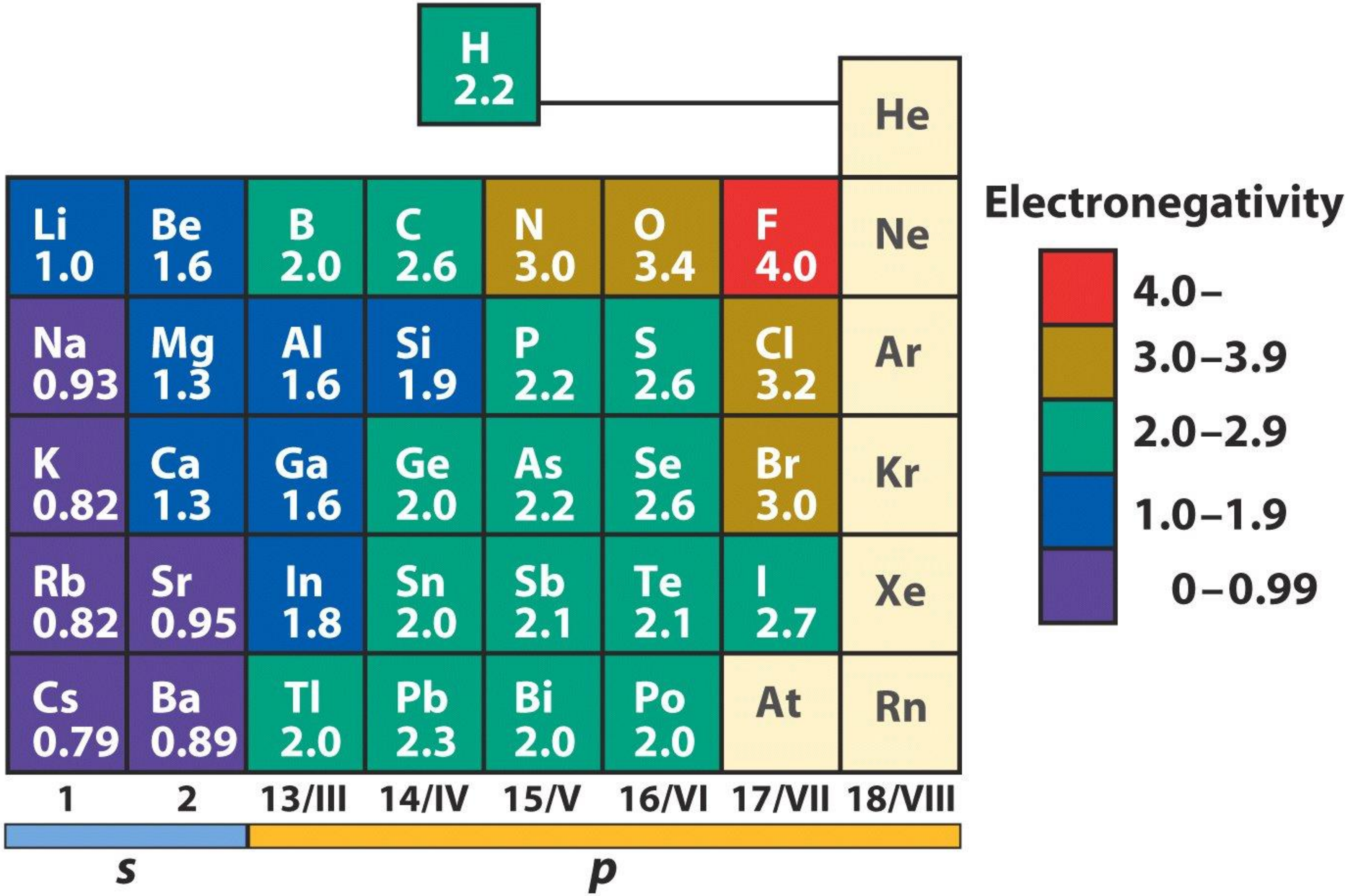


Pauling Electronegativity Scale

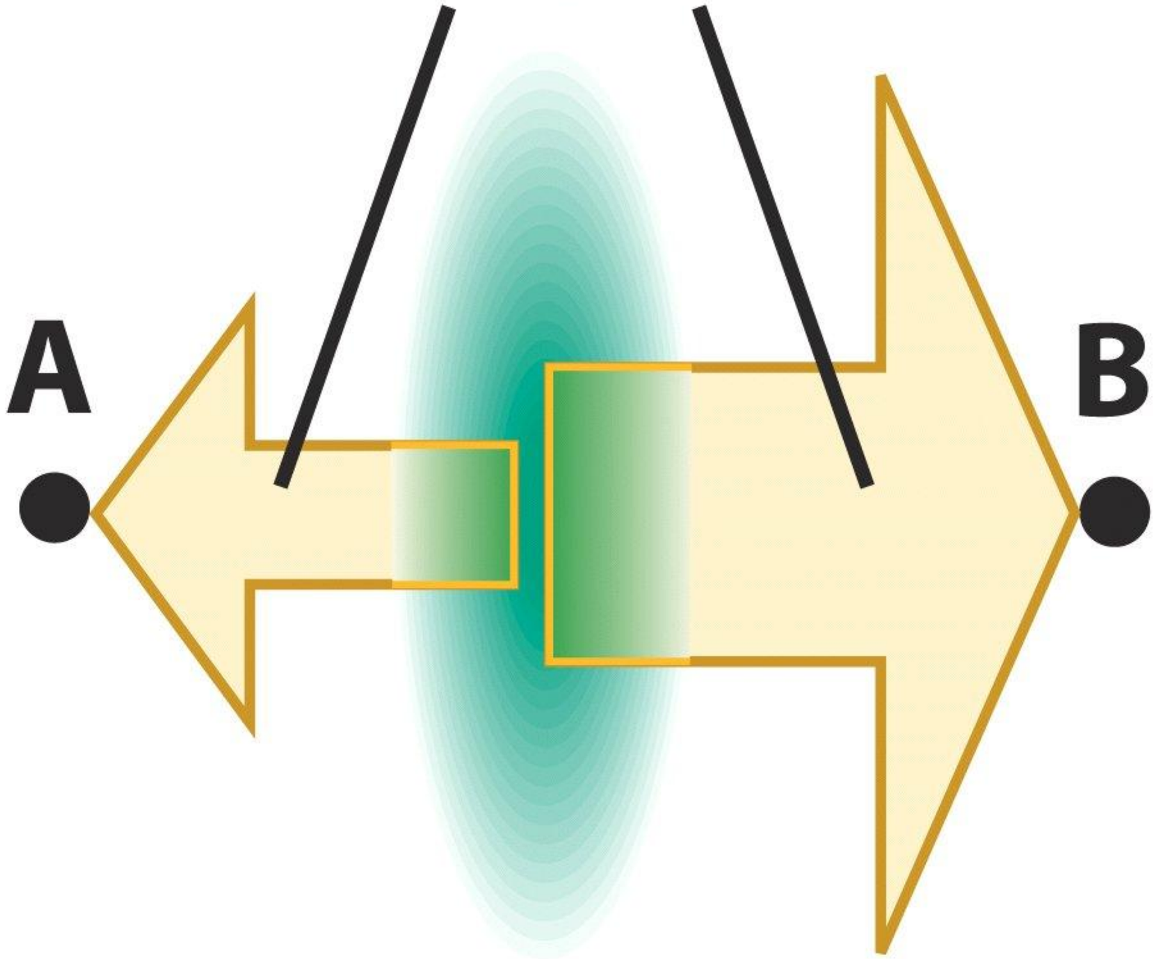
Li	**Be**	**B**	**C**	**N**	**O**	**F**
1.0	**1.5**	**2.0**	**2.5**	**3.0**	**3.5**	**4.0**
Na	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**
0.9	**1.2**	**1.5**	**1.8**	**2.1**	**2.5**	**3.0**

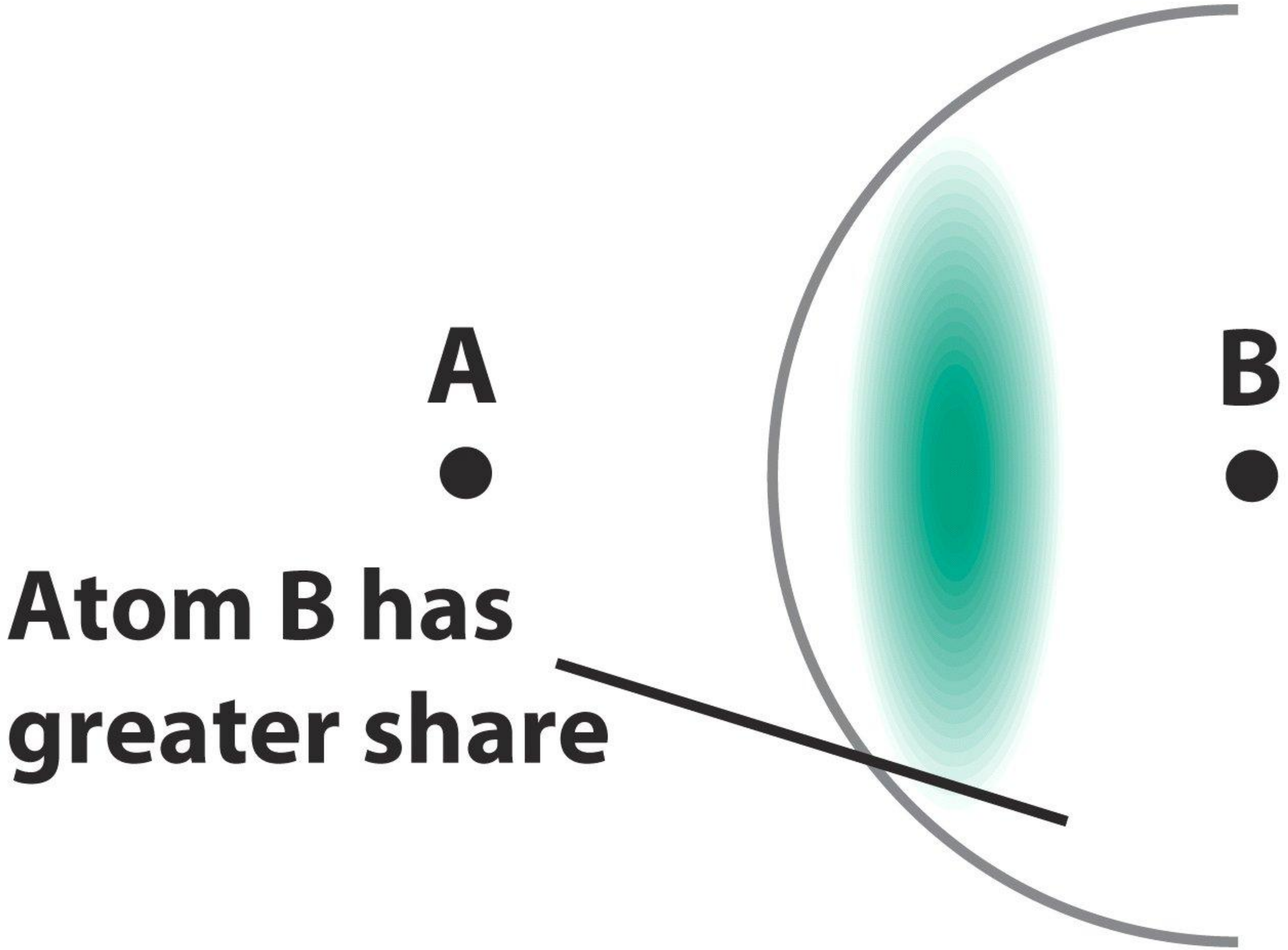
- Electronegativity increases from left to right in the periodic table.

- Electronegativity decreases going down a group.


H 2.2
He
Li 1.0
Be 1.6
B 2.0
C 2.6
N 3.0
O 3.4
F 4.0
Ne
Na 0.93
Mg 1.3
Al 1.6
Si 1.9
P 2.2
S 2.6
Cl 3.2
Ar
K 0.82
Ca 1.3
Ga 1.6
Ge 2.0
As 2.2
Se 2.6
Br 3.0
Kr
Rb 0.82
Sr 0.95
In 1.8
Sn 2.0
Sb 2.1
Te 2.1
I 2.7
Xe
Cs 0.79
Ba 0.89
Tl 2.0
Pb 2.3
Bi 2.0
Po 2.0
At
Rn
1
2
13/III
14/IV
15/V
16/VI
17/VII
18/VIII
s
p
Electronegativity
4.0–
3.0–3.9
2.0–2.9
1.0–1.9
0–0.99

Relative pulling power of atom




A
B
Atom B has greater share

Exercise

Arrange the following bonds from most to least polar:

a)	N–F	O–F	C–F
b)	C–F	N–O	Si–F
c)	Cl–Cl	B–Cl	S–Cl
a)	C–F,	N–F,	O–F
b)	Si–F,	C–F,	N–O
c)	B–Cl, S–Cl,	Cl–Cl	

Which of the following bonds would be the least polar yet still be considered polar covalent?

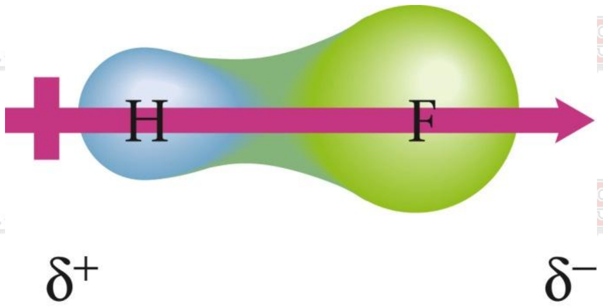
Mg–O   C–O   O–O   Si–O $\boxed{\text{N–O}}$

Which of the following bonds would be the most polar without being considered ionic?

Mg–O   C–O   O–O $\boxed{\text{Si–O}}$ N–O

Dipole Moment

- Property of a molecule whose charge distribution can be represented by a center of positive charge and a center of negative charge.
- Use an arrow to represent a dipole moment.
  - Point to the negative charge center with the tail of the arrow indicating the positive center of charge.

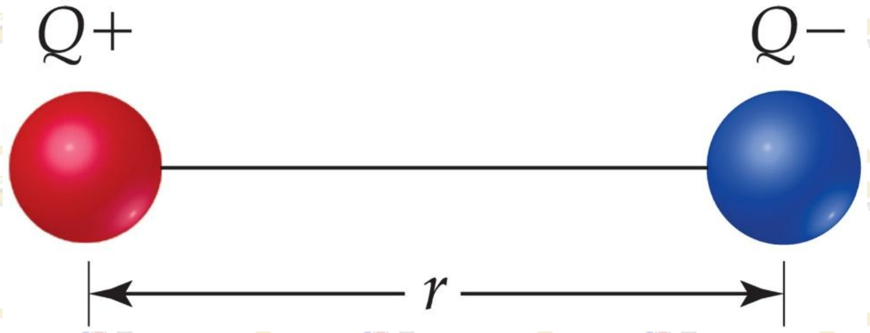


Dipole Moment

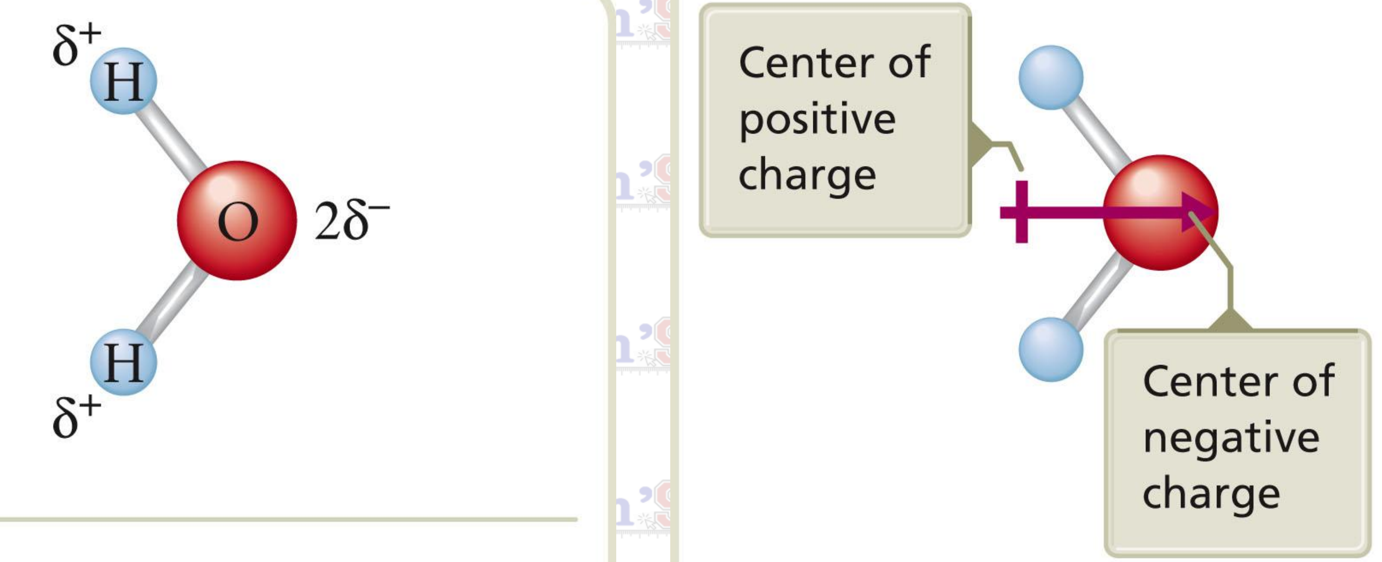
- When two atoms share electrons unequally, a bond dipole results.
- The dipole moment, μ, produced by two equal but opposite charges separated by a distance, r, is calculated:

$$\mu = Qr$$

- It is measured in debyes (D).



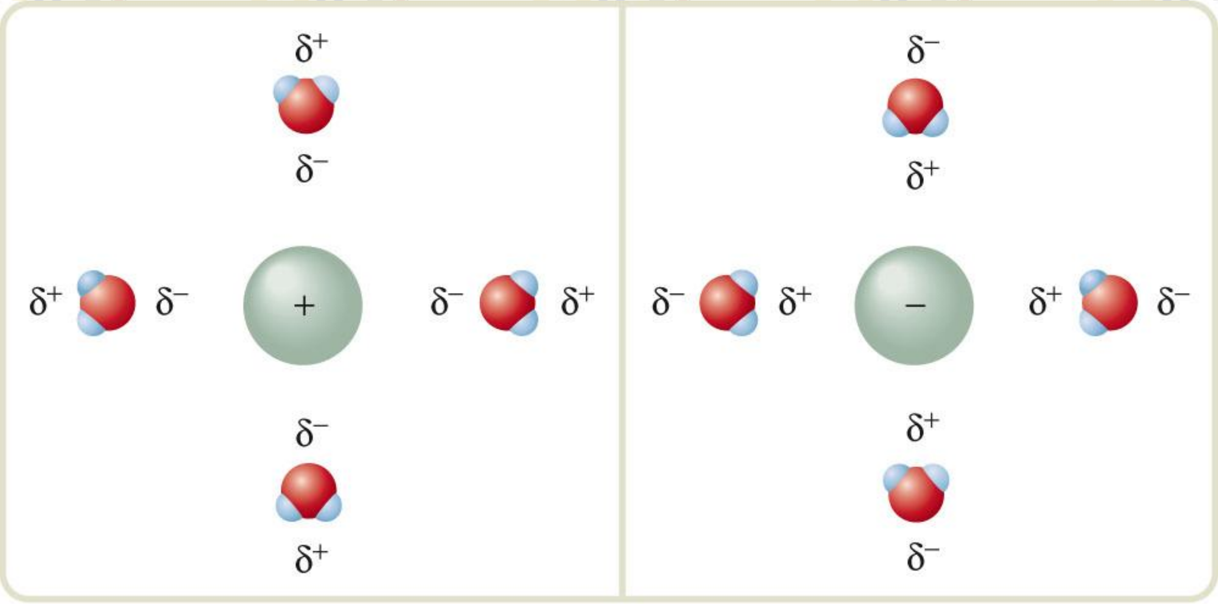
Dipole Moment in a Water Molecule



The charge distribution in the water molecule. The oxygen has a charge of $2\delta^-$ because it pulls δ^- of charge from each hydrogen atom ($\delta^- + \delta^- = 2\delta^-$).

The water molecule behaves as if it had a positive end and a negative end, as indicated by the arrow.

- The polarity of water affects its properties.
  - Permits ionic compounds to dissolve in it.



a *Polar water molecules are strongly attracted to positive ions by their negative ends.*

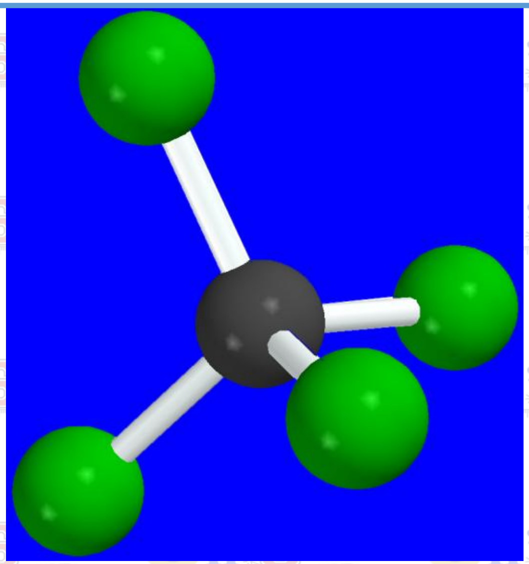
b *They are also strongly attracted to negative ions by their positive ends.*

  - Ca ure.

Molecular Dipole Moments

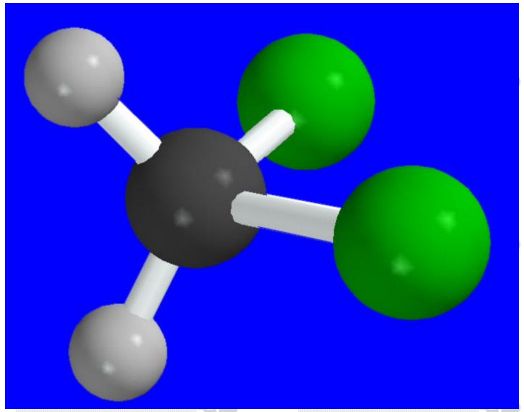
Carbon dioxide has no dipole moment; $\mu = 0$ D

Comparison of Dipole Moments



Carbon tetrachloride

$\mu = 0$ D

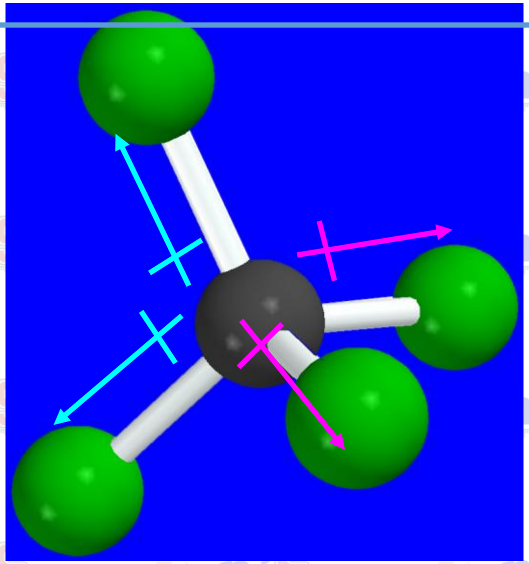


Dichloromethane

$\mu = 1.62$ D

Carbon tetrachloride

Resultant of these two bond dipoles is

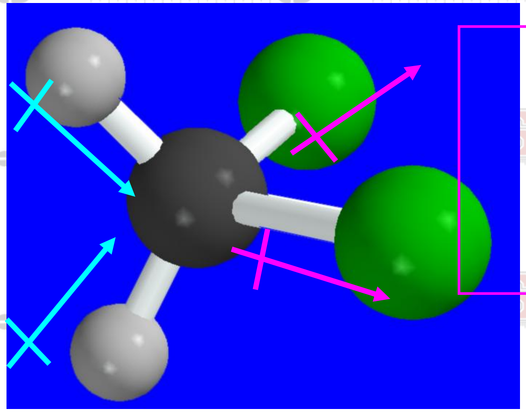


Resultant of these two bond dipoles is

$\mu = 0$ D

Carbon tetrachloride has no dipole moment because all of the individual bond dipoles cancel.

Dichloromethane

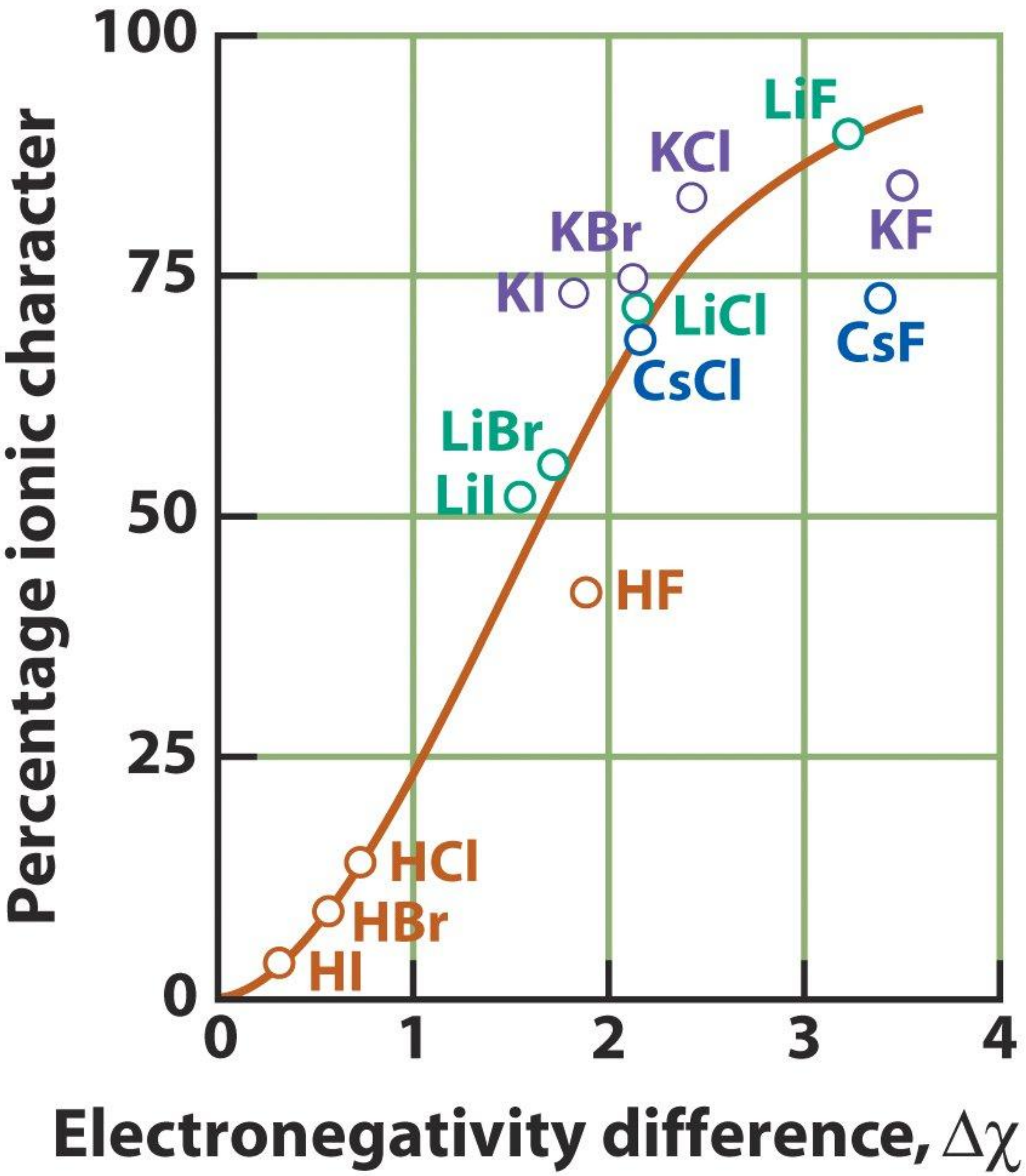


Resultant of these two bond dipoles is

Resultant of these two bond dipoles is

μ = 1.62 D

The individual bond dipoles do not cancel in dichloromethane; it has a dipole moment.

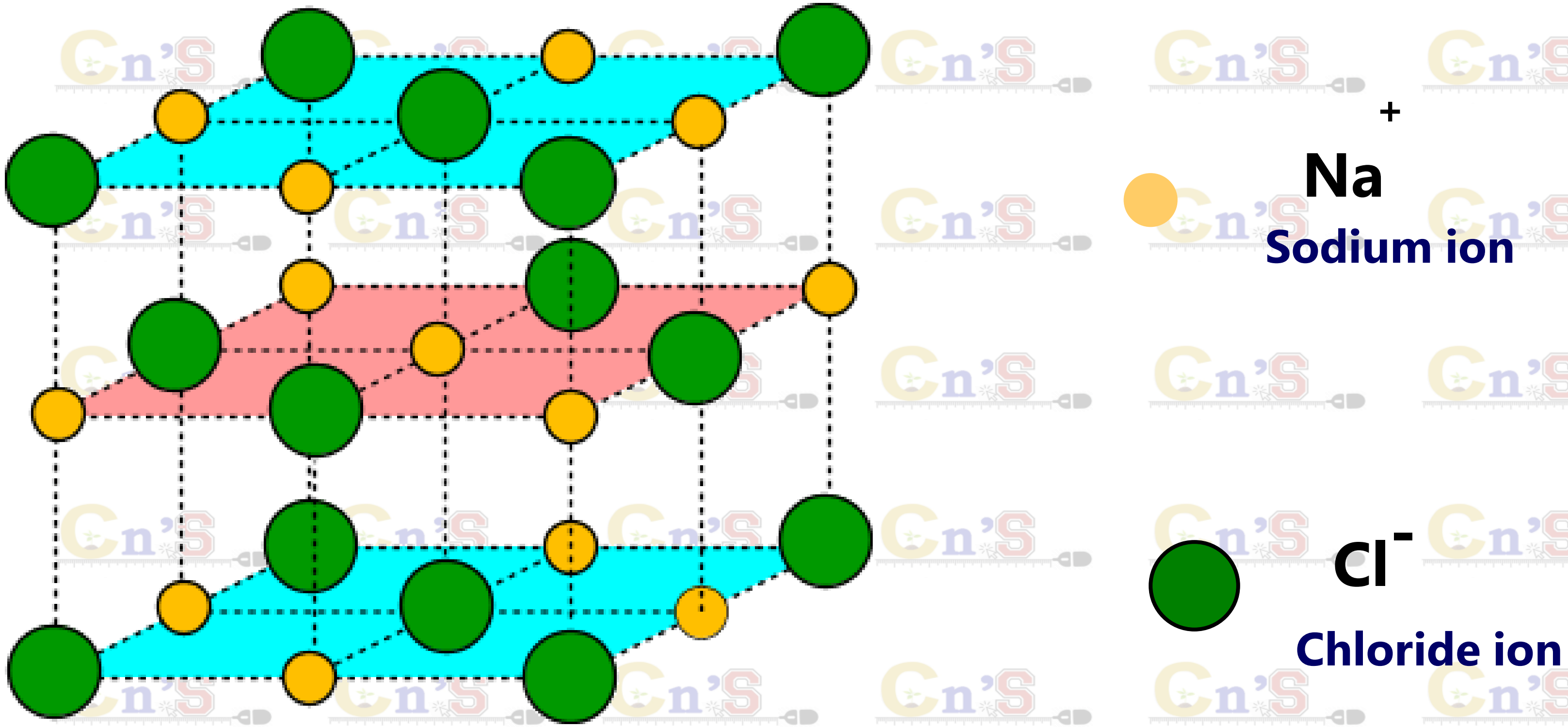


100
75
50
25
0
Percentage ionic character
LiF
KCl
KF
KBr
KI
LiCl
CsF
CsCl
LiBr
LiI
HF
HCl
HBr
HI
0
1
2
3
4
Electronegativity difference, $\Delta\chi$

IONIC COMPOUNDS - CRYSTAL LATTICE STRUCTURE

Oppositely charged ions held in a regular 3-dimensional lattice by electrostatic attraction

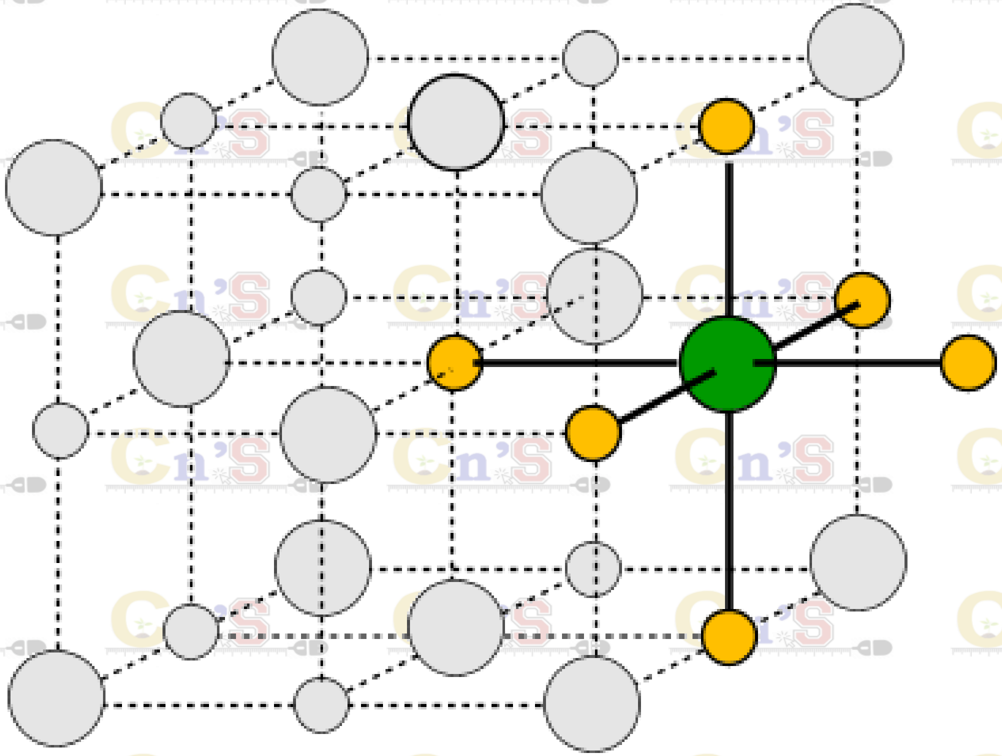
The arrangement of ions in a crystal lattice depends on the relative sizes of the ions



The Na ion is small enough relative to a Cl^{-} ion to fit in the spaces so that both ions occur in every plane.

Oppositely charged ions held in a regular 3-dimensional lattice by electrostatic attraction:

The arrangement of ions in a crystal lattice depends on the relative sizes of the ions



Each Na^{+} is surrounded by 6 Cl^{-}

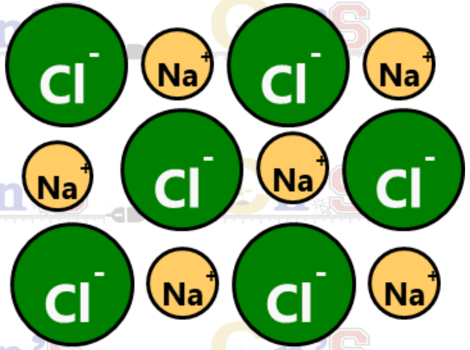
and each Cl^{-} is surrounded by 6 Na^{+}

Properties of Ionic compounds

- Ionic compound exist in solid state
- The network of ions have a definite geometric pattern which depends on the size and charge of ions
- Posses high melting and boiling points due to strong electrostatic force of attraction between the ions
- Good conductor of electricity in molten or dissolved state
- Does not conduct electricity in solid state as ions are not free to move
- Are soluble in polar solvent like water as solvent interacts with the ions of ionic solid

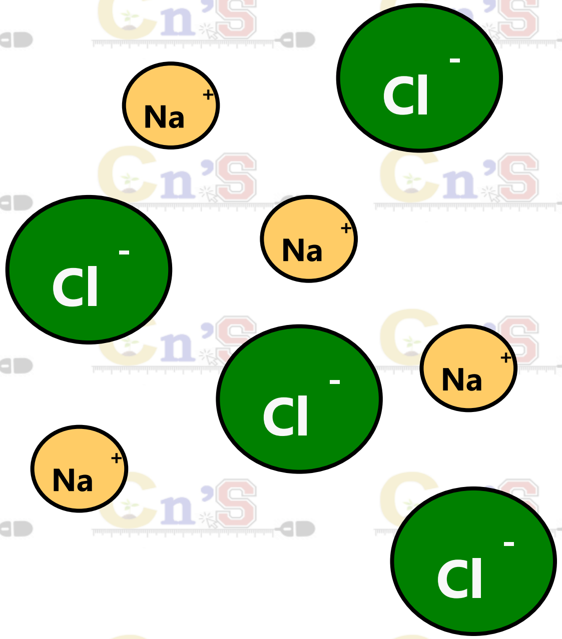
SOLID IONIC COMPOUNDS DO NOT CONDUCT ELECTRICITY

IONS ARE HELD STRONGLY TOGETHER
+ IONS CAN'T MOVE
- IONS CAN'T MOVE



MOLTEN IONIC COMPOUNDS DO CONDUCT ELECTRICITY

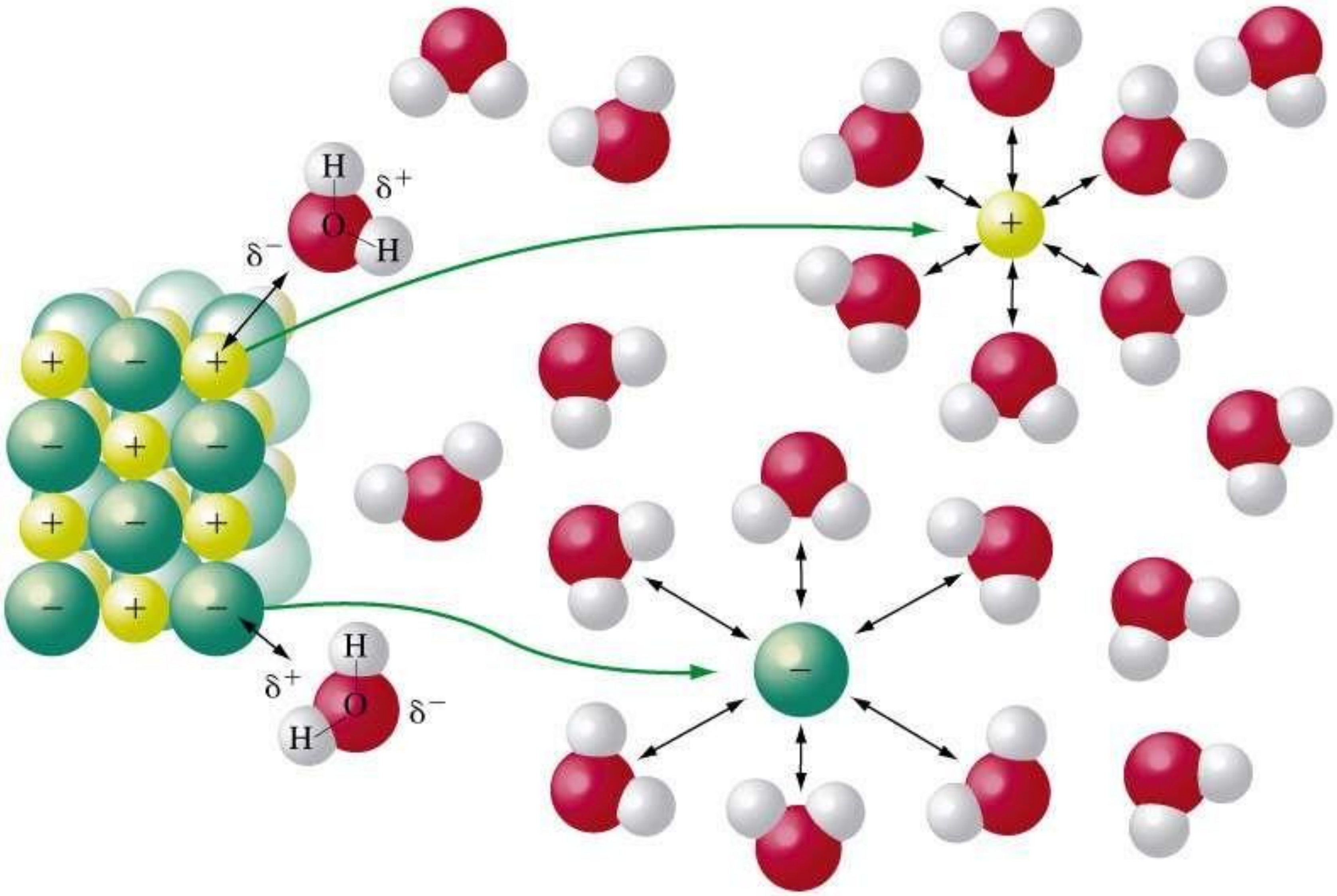
IONS HAVE MORE FREEDOM IN A LIQUID SO CAN MOVE TO THE ELECTRODES



SOLUTIONS OF IONIC COMPOUNDS IN WATER DO CONDUCT ELECTRICITY

DISSOLVING AN IONIC COMPOUND IN WATER BREAKS UP THE STRUCTURE SO IONS ARE FREE TO MOVE TO THE ELECTRODES

Many ionic compounds are soluble in water - dissolving



The formations of ionic bond governed by the following factors:

- 1. Ionization energy:
- Formation of ionic bond metal atom loses electron to form cation
- Energy required for this equal to ionization energy
- Alkali metals have lowest ionization energy, thus have more tendency to form cation
- 2. Electron gain enthalpy:
- • Electron released in the formation of cation are to be accepted by the other atom taking part in the ionic bond formation
- • Electron accepting tendencies depend on upon the electron gain enthalpy
- • Defined as energy released when isolated gaseous atom takes up an electron to form anion.
- • Greater the negative enthalpy, easier the formation of anion

- 3. Lattice energy:

- Combination of oppositely charged ions to form ionic crystal, with release of energy is referred as lattice energy
- Higher value of lattice energy, greater will be the stability of compound
- Magnitude of lattice energy gives idea about the strength of interionic forces
- Size of ions:
- In case of similar ions inter-nuclear distance is lesser due to which inter-ionic attraction is greater and hence the magnitude of lattice energy will be larger
- Charge on the ions:
- Ions have higher charge exerts stronger forces of attraction and hence larger amount of energy is released. Thus value of lattice energy is higher

Properties of covalent compounds

- Compounds formed exist as discrete molecules
- •Weak intermolecular force due to small molecular size
- •Mainly exist in liquid or gaseous state
- •Sugar, urea, starch etc. exist in solid state
- •Low melting and Boiling points due to weak attractive forces
- •Poor conductor of electricity in fused or dissolved state
- •Less soluble in water
- •Gives molecular reaction

Properties of covalent compounds

- Bonding Atoms are joined together within the molecule by covalent bonds.
- Don't conduct electricity as they have no mobile ions or electrons
- Solubility - more soluble in organic solvents than in water; some are hydrolysed
- Boiling point are low - compared to ionic compounds. As the Intermolecular forces are weak, little energy is required to separate molecules from each other. BP increases as molecules get a larger surface area.e.g. CH_4 -161 C C_2H_6 - 88 C C_3H_8 -42 C

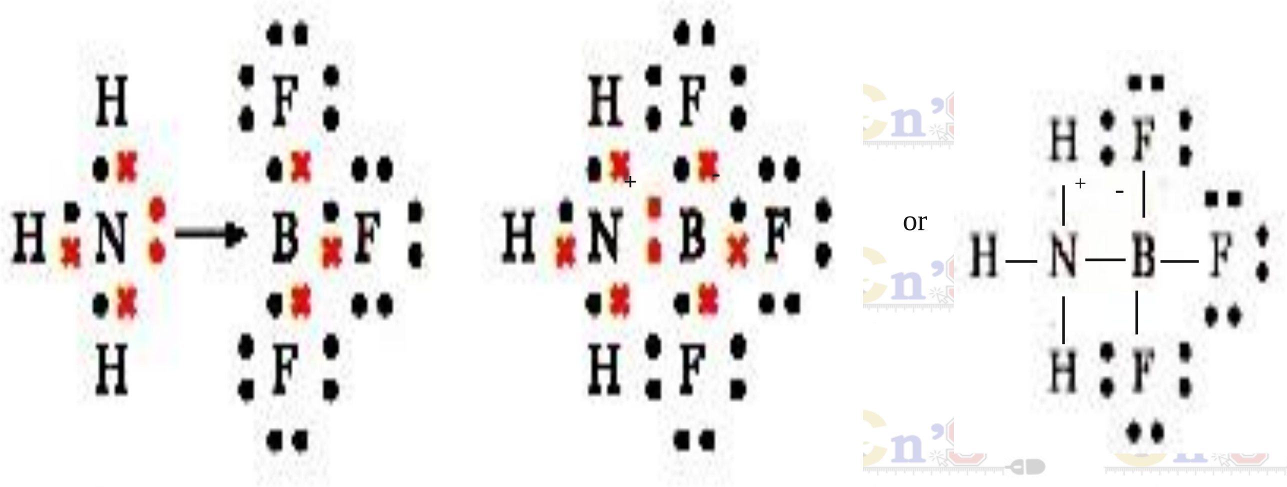
BOND STRENGTH

- Covalent bond strength depends on the number of electron pairs shared by the atoms.
- single bond < double bond < triple bond

Coordinate bond

A co -ordinate bond is a covalent bond (a shared pair of electrons) in which both electrons come from the same atom. A bond can be formed by the overlapping of an orbital containing lone pair of electrons with the empty orbital of the valance shell of another atom. The bond formed in this way is called the dative bond. There, the species which give the lone pair of electrons is called the donor group (Lewis base) and the species that receives the electrons to form the bond is called the acceptor group (Lewis acid).

eg.



Properties of Coordinate bond :

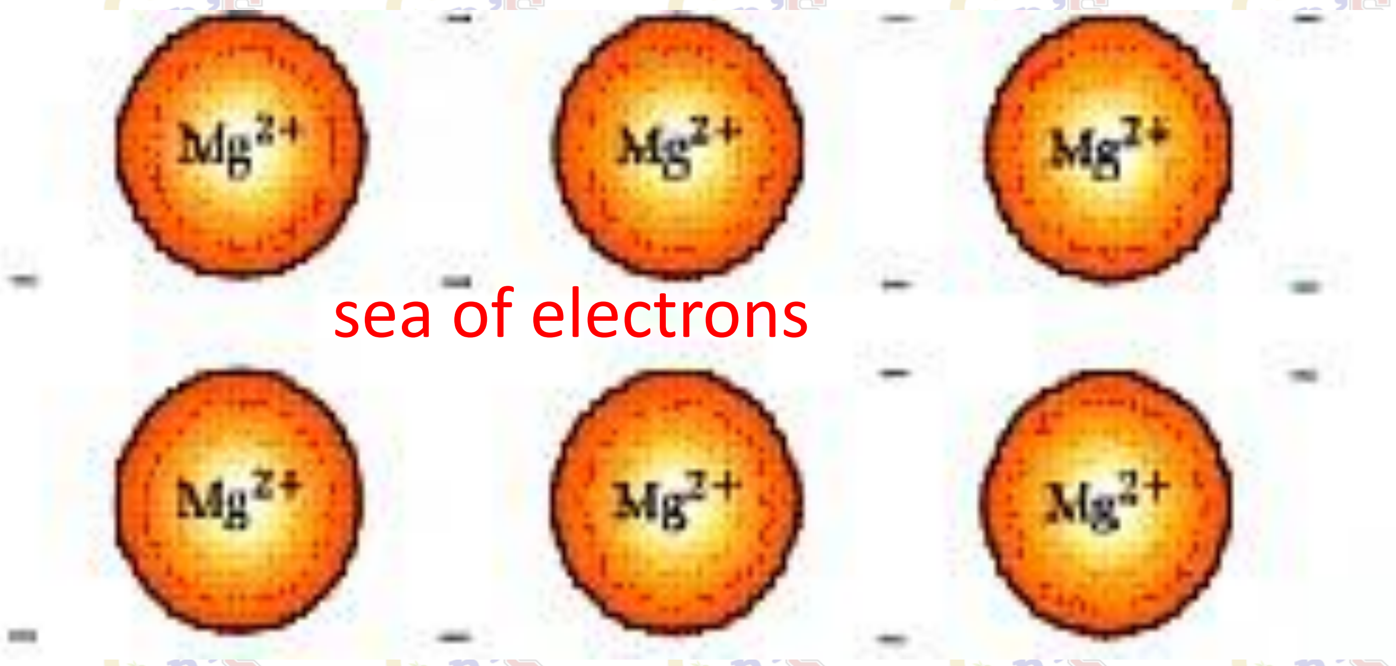
- 1. Are generally soluble in water and organic solvents
- 2. Boiling and melting points of these compounds are less than electrovalent compounds but are higher than covalent compounds
- 3. Compounds ionize in aqueous solution giving simple and complex ions
- 4. These bonds are also directional and stereoisomerism is also found
- 5. Molecules possess definite shape and definite bond angles, thus have definite geometry

METALLIC BOND

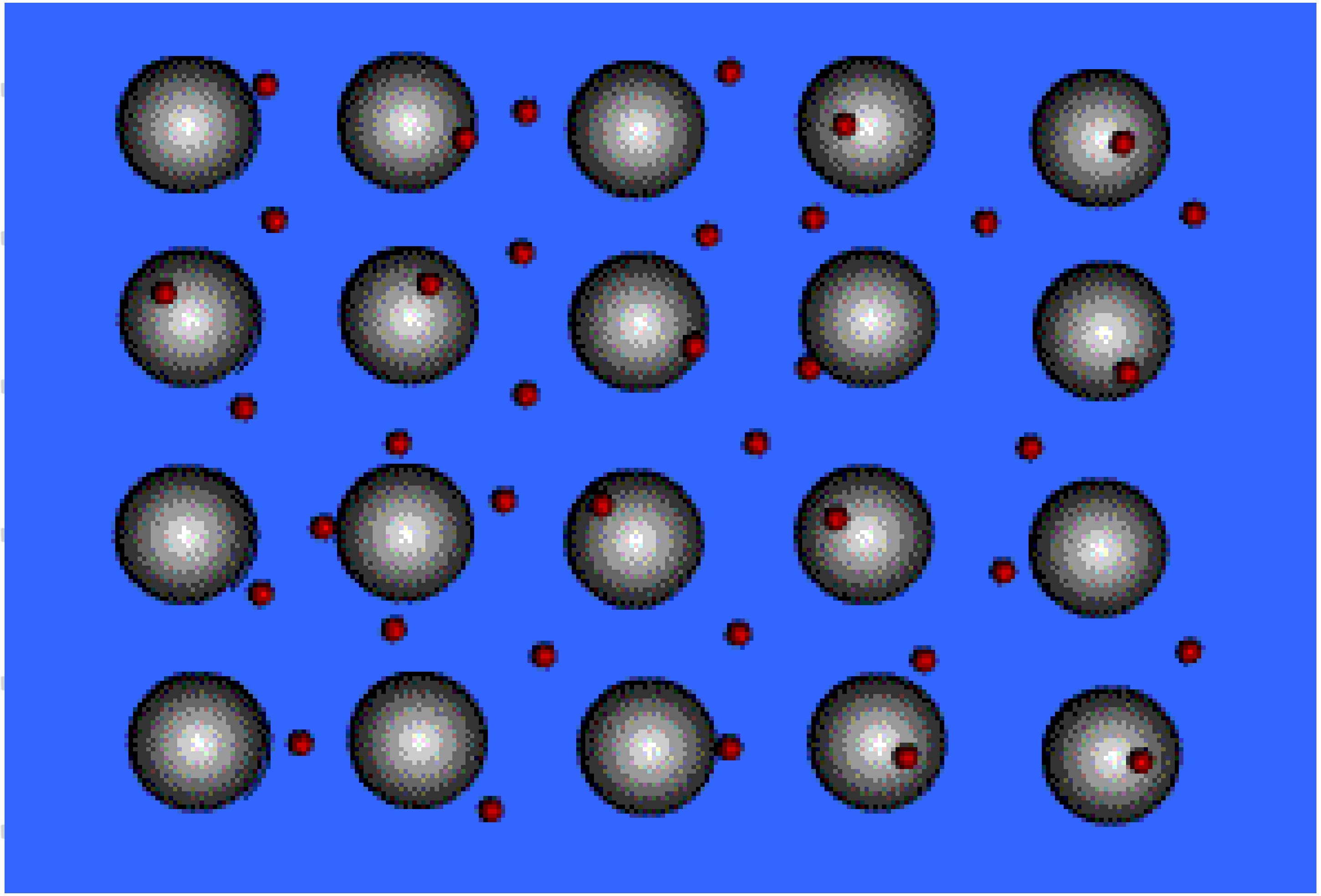
bond found in metals; holds metal atoms together very strongly

Metallic bonds

The electrons in the valence shell of metallic atoms are loosely bonded to the atom. Therefore, there is a tendency for the metallic atoms to release the electrons in valence shell and exist as positive ions. As a result a system is formed in which positive ions are immersed in a sea of electrons which were released from the metal atoms. The positive ions and the sea of electrons get attracted electrostatically to form metallic bonds.



Metalic Bond, A Sea of Electrons



- Metals conduct electricity due to the presence of free electrons.
  - Strong metallic bonds are formed when the size of the metal ion decreases, when the charge of the metal ion increases and when the number of electrons contributing to the metallic bond increases.
  - When the strength of metallic bond increases melting point of the metal also increases.
  - Attractions that exist in covalently bonded molecules or in ionic compounds or in metalic latices are referred as primary interactions.

Examples; Na, Fe, Al, Au, Co
Metallic bonding is the electrostatic attraction between the positively charged atomic nuclei of metal atoms and the delocalised electrons in the metal. In the solid state, both metallic and ionic compounds possess ordered arrays of atoms or ions and form crystalline materials with lattice structures.

Metals Form Alloys

Metals do not combine with metals. They form Alloys which is a solution of a metal in a metal. Examples are steel, brass, bronze and pewter.