15 • Chemical Bonding
Three Types of Bonding
(1 of 8)

There are **three** general **classes** of **bonds** that form between atoms. You can predict which will form by classifying the atoms as **metals** or **nonmetals**:

metal + metal	metallic bond	Au-Ag alloy
metal + nonmetal	ionic bond	$MgCl_2$
nonmetal + nonmetal	covalent bond	SO_2 or CH_4

Some compounds can contain **both ionic** and **covalent** bonds such as K_2SO_4... the sulfate ion is held together with covalent bonds... the potassium ions are ionically bonded to the sulfate ions.

Acids are **exceptions**... they are **ionic** only when **dissolved**.

15 • Chemical Bonding
The Ionic Bond
(2 of 8)

Many ions can be explained because they have gained or lost electrons and attain a **noble gas configuration.**

For example: P^{3-} S^{2-} Cl^{-} **Ar** K^{+} Ca^{2+}

all have the same electron arrangement: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

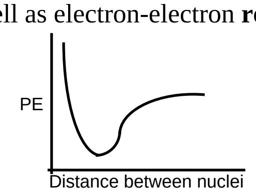
The importance of this configuration is that this is one reason why ions form. After these ions form, they stick together in a crystal lattice because **opposites attract**:

+ - + - + - + -
- + - + - + - +
+ - + - + - + -
- + - + - + - +
+ - + - + - + -

There are other reasons why some ions (ex: Cu^{+} or Zn^{2+}) form. One is the pseudonoble gas configuration.

15 • Chemical Bonding
The Covalent Bond
(3 of 8)

The covalent bond between two atoms depends on the **balance** of **attractions** between one atom's + nucleus and the other atom's – electrons and the proton-proton **repulsions** as well as electron-electron **repulstions**.



If two atoms have **half-filled orbitals**, the interactions balance at a **small enough distance** so the e^{-}'s can be **close to both nuclei** at the same time... this is a **covalent bond**.

15 • Chemical Bonding
Groves' Electron Dot System
Multiple & Extended Valence Bonds
(4 of 8)

Count up your <u>valence</u> electrons.

Give every atom who "wants" and octet an octet.
[the first 5 elements do not need octets... too small]
[Family I, II, and III do not form octets]

If you have drawn **too many** electrons... ex: C_2H_2

"Take away a lone pair... take away a lone pair... make these two atoms share"

If you have drawn **too few** electrons... ex: PCl_5
Place the extra electrons on the central atom **(extended valence shell)**

15 • Chemical Bonding
Lewis Electron Dot Structures
(5 of 8)

Lewis symbols conisist of the atomic symbol surrounded by valnece electrons. The four sides represent the four valence orbitals. Atoms are usually shown in their excited states (II, III, IV)

Li· ·Be· ·B· ·C· ·N· :O· :F: :Ne:

Ions include brackets. Positive ions show no valence electrons while negative ions usually have an octet.

$[Li]^{+}$ $[Mg]^{2+}$ $[:O:]^{2-}$

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15 • Chemical Bonding
Comparing Ionic and Molecular Substances
(6 of 8)

Compound	Molecular	Ionic
Conducts as Solid	NO	NO
Conducts as Liquid	NO	YES
Conducts in Solution	NO	YES
Conducts as Gas	NO	YES
Hardness	soft	hard
MP / BP	low	high
Bonding	covalent	ionic
Examples	He, CH_4, CO_2, $C_6H_{12}O_6$	NaCl, KI, $AgNO_3$

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15 • Chemical Bonding
Electronegativity and Polar Bonds
(7 of 8)

You will be given a chart of **electronegativity values.**
Memorize the most electronegative element s (F = 4.0) then oxygen (O = 3.5) and chlorine (Cl = 3.0). The noble gases have no electronegativity values… no bonds.
Trend is **large** electronegativity in the **upper right** of the per. table and small in the lower left portion of the table.
Classify the bond between any two atoms by subtracting their electronegativity values (Δe)

Non-polar covalent	$0 < \Delta e <$ **0.5**
Polar covalent	**0.5** $\leq \Delta e \leq$ **1.7**
Ionic	$\Delta e >$ **1.7**

The **more electronegative** atom is more **negative**.
Polar covalent bonds have **partial** charges δ+ and δ–

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15 • Chemical Bonding
Shapes and Polar Molecules
(8 of 8)

Use **VSEPR** theory to predict the shape of molecules. The **Steric Number** (the number of lone pairs + bonded atoms) corresponds to the shape of the electron pairs around a central atom.
[1=linear, 2=linear, 3=trigonal planar, 4=tetrahedral]

If each shape is **symmetrical**, the bond dipoles will cancel resulting in a **nonpolar** molecule.
If a shape has **lone pairs** of electrons on the central atom, the shape is often unsymmetrical, the molecule is **polar**.

Polar molecules and ions **dissolve** well in **polar** solvents while **nonpolar** molecules dissolve in **nonpolar** solvents.
"Like Dissolves Like"