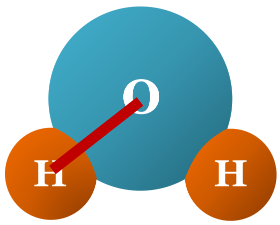
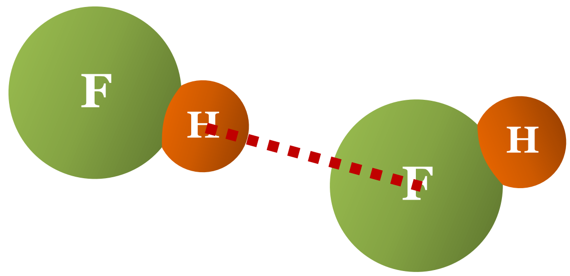
Chemical Bonding
Unit 4

Chemical Bonding

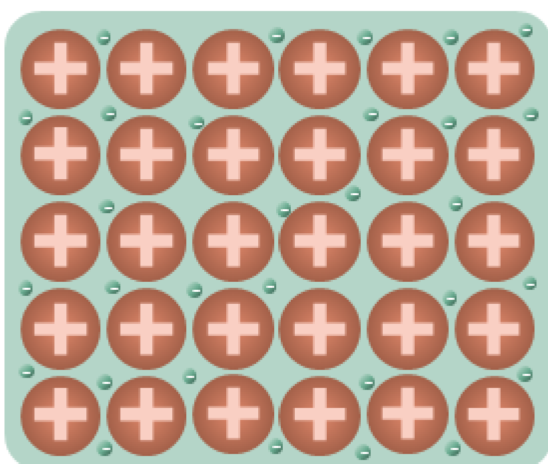
Bond Formation

- All chemical bonds occur when the protons (+) of one atom attract the electrons (–) of another atom
  - Bond formation is *always* exothermic (energy is released)
  - Bond breaking is *always* endothermic (energy is absorbed)
  - All chemical bonding is a form of potential energy (PE)
- There are two main categories of chemical bonds:

Regular Chemical Bonds	Intermolecular Forces (IMFs)
	
(an O-H bond)	**(a hydrogen bond)**
bonded atoms are in contact with each other	an atom of one molecule bonds to an atom on a nearby molecule

Types of Chemical Bonds (from strongest to weakest)

- Metallic Bonds: characterized by a sea of freely moving electrons
  - Valence electrons move off the kernel and move freely throughout the metal crystal in empty d sublevels

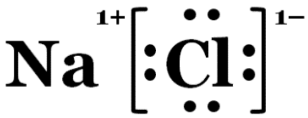


  - Compare: metallic bonds are like other chemical bonds because + nuclei attract e^- between
  - Contrast: because of the sea of free e^-, all the particles are cations which are very small making metallic bonds the shortest bonds (and therefore the strongest bonds)
    - Example: the melting point of tungsten is 3422 °C

Ionic Bonds: characterized by the transfer of electrons

Metals lose e^- and nonmetals gain e^-

metal + nonmetal → ionic compound



Compare: ionic bonds are like other chemical bonds because + nuclei attract e^- between
Contrast: the extra attraction of the + and – ions adds extra bond strength

Example: the melting point of NaCl is 801 °C or 786 kJ/mol lattice energy

Covalent Bonds: electrons are shared between atoms

Polar covalent bonds: e^- are shared *unevenly*
Nonpolar covalent bonds: e^- are shared *evenly*
Coordinate covalent bonds: one atom donates *both* the shared e^-

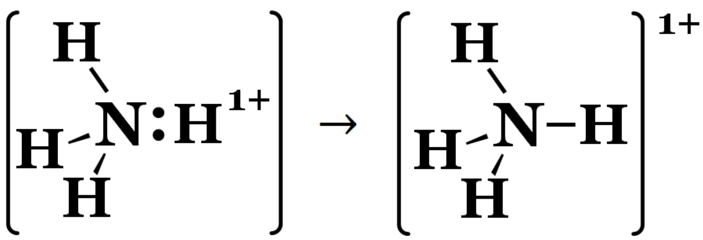
Nonmetals can share e^-

nonmetal + nonmetal → covalent or molecular compound

polar covalent

nonpolar covalent

coordinate covalent



the e^- pair from the N atom forms a bond to the H^+ ion

once the bond has formed all the N–H bonds are equal

Compare: covalent bonds are like other chemical bonds because + nuclei attract e^- between
Contrast: covalent bonds have no added bond strength

Examples: the polar HCl is 432 kJ/mol
the nonpolar Cl_2 is 243 kJ/mol

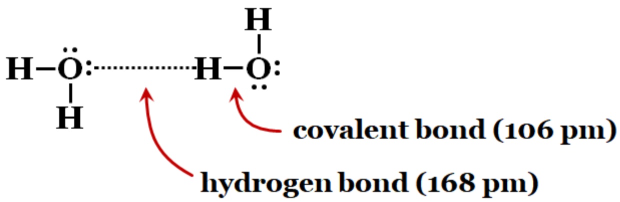
Molecules: a group of atoms held together by covalent bonds

- the bonds can be polar covalent, nonpolar covalent, or a mix of each
- molecules are neutral particles
- molecules can be polar or nonpolar
    - do not confuse polar molecules with polar bonds
        - CO_2 has polar bonds but is a nonpolar molecule
        - PH_3 has nonpolar bonds but is a polar molecule
    - polar molecules have a δ+ charge on one side and a δ– charge on the other side

Types of Intermolecular Forces (IMFs)

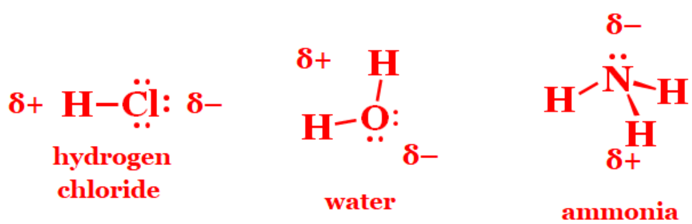
Inter: <*from Latin*> between, IMF is literally forces between molecules

Hydrogen bonds
- require a H covalently bonded to N, O, or F
- bridge *two* molecules
- are very strong dipole–dipole bonds

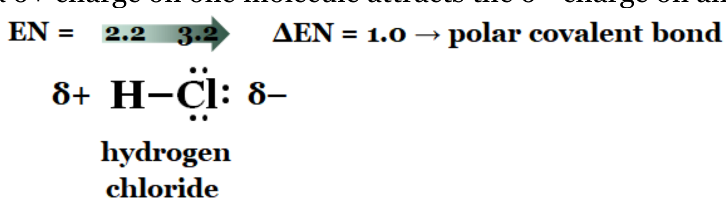


H
Examples: HF···HF, H_3N···H_3N, H_2O···H_2O, CH_3O···$HOCH_3$

Dipole–dipole bonds
- require that *both* molecules be polar
  - Polar molecule: any molecule that has a slight negative charge (δ–) on one end and a slight positive charge (δ+) on the other
    - ○ polar molecules must be asymmetrical in shape
    - ○ polar molecules must have an asymmetrical charge distribution
    - ○ polar molecules are usually (but not always) caused by polar bonding
      - ▪ polar bonds are detected by differences in electronegativity (ΔEN)
        - ▫ $0.3 \le \Delta EN \le 1.7$ will result in a polar covalent bond
      - ▪ Examples of polar molecules:



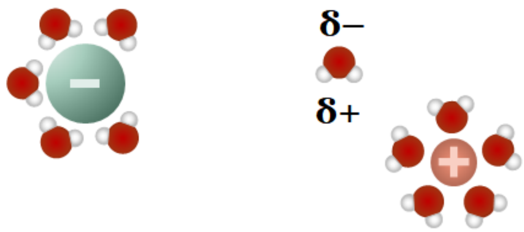
- bridge *two* molecules
- a δ+ charge on one molecule attracts the δ– charge on another molecule



Examples of molecules that dipole–dipole bond:
HBr, HI, SO_2, SO_3, CO_2, $HCCl_3$

Molecule–ion attractions

- occur with ions in polar solvents
- bridge *polar molecules* and *ions*



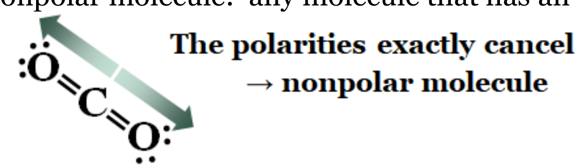
Examples of molecule–ion attractions:

$NaCl_{(aq)}$, $Na_2CO_{3(aq)}$, $KBr_{(aq)}$, $NaOH_{(aq)}$

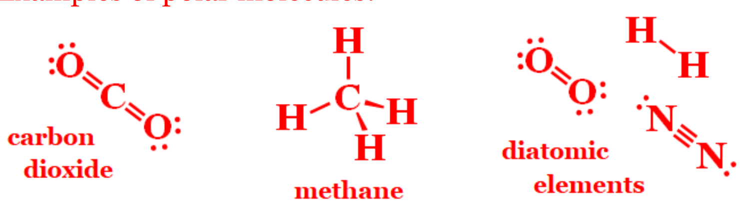
van der Waal’s forces (a.k.a. London dispersion forces or dispersion forces)

- require that *both* molecules be nonpolar

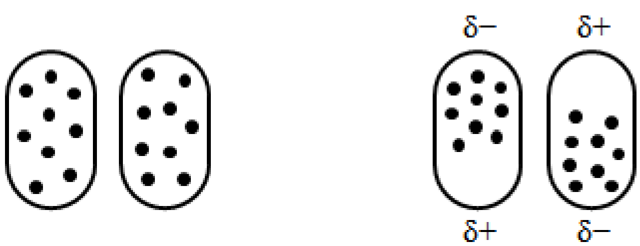
  Nonpolar molecule: any molecule that has an even charge distribution



- Examples of polar molecules:



- bridge *two* molecules
- occur because of concerted motion of e^- in polarizable atoms or molecules



random dispersion results in neutral charge dispersion

electron motion causes asymmetric charge distribution and momentary charges attract

Identifying Bond Type

Metallic Bonds: occur whenever all the atoms are metals

Examples: iron, $Fe_{(s)}$; silver, $Ag_{(s)}$; or brass, Cu_3Zn_2

Ionic, Polar Covalent, and Nonpolar Covalent Bonds: use ΔEN to calculate bond type

Electronegativity difference (ΔEN) and bond type table:

nonpolar	**0.3**	**polar**	**1.7**	**ionic**

Examples:

PH_3 molecule

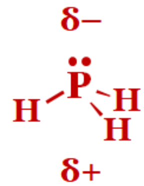
Determine the bond type:

Find the electronegativities on Table S:

P = 2.2

H = 2.2

ΔEN = 0.0 which are nonpolar covalent bonds



Determine the molecule polarity:

Electrons spend more time between the H and P atoms, so the H end is δ+

The electron pair above the P atom is δ–

Even though the bonds are nonpolar, the molecule is polar

Notice that bonds are nonpolar due to ΔEN

but molecule is polar due to asymmetry (uneven e^- distribution)

CO_2 molecule

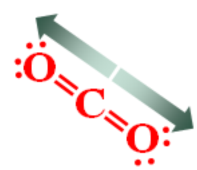
Determine the bond type:

Find the electronegativities on Table S:

C = 2.6

O = 3.4

ΔEN = 0.8 which are polar covalent bonds



Determine the molecule polarity:

Electrons are evenly distributed (the molecule is symmetrical)

The electron pair above the P atom is δ–

Even though the bonds are polar, the molecule is nonpolar

Notice that bonds are polar due to ΔEN

but molecule is polar due to symmetry (even e^- distribution)

HCl molecule

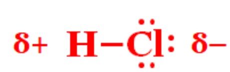
Determine the bond type:

Find the electronegativities on Table S:

H = 2.2

Cl = 3.2

ΔEN = 1.0 which are polar covalent bonds



Determine the molecule polarity:

Electrons are unevenly distributed (the molecule is asymmetrical)

The bonds are polar and the molecule is polar

Notice that bonds are polar due to ΔEN

and molecule is polar due to asymmetry (uneven e^- distribution)

NaCl

Determine the bond type:

Find the electronegativities on Table S:

Na = 0.9

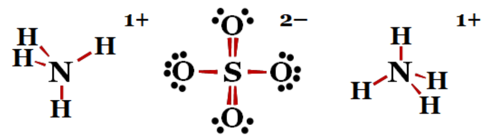
Cl = 3.2

ΔEN = 2.2 which is an ionic bond

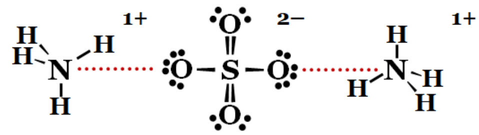
$Na^{+}Cl^{-}$

Polyatomic ions (e.g. $(NH_4)_2SO_4$)

Bonds inside the ions are covalent



Bonds between ions are ionic



Summary of identifying bond types:

Metals → Metallic Bonds

Metals + Nonmetals → Ionic Bonds (ΔEN > 1.7)

Nonmetals + Nonmetals → Covalent Bonds

(1.7 ≥ ΔEN ≥ 0.3 is a polar covalent bond)

(0.3 > ΔEN is a nonpolar covalent bond)

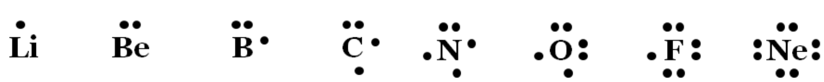
Polyatomic Ions → Ionic and Covalent Bonds

Lewis Electron Dot (LED) Diagrams and Bonding

Rules for elements: first two pair, Hund's Rule the remaining electrons

Period 2 elements by Group

1	2	13	14	15	16	17	18



Rules for ions:

- metals lose valence e⁻ forming cations
- nonmetals gain e⁻ forming anions with a stable octet

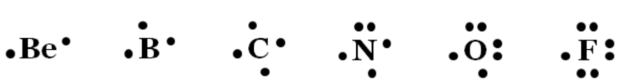
Period 3 elements by Group

1	2	13	14	15	16	17	18

Rules for covalent bonds: Hund's Rule all the valence electrons

Period 2 elements by Group

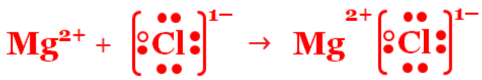
1	2	13	14	15	16	17	18



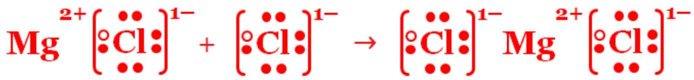
Constructing ionic compounds using LEDs:

1. Draw LED models of both ions
2. Start with the cation
3. Add an anion
4. Add ions to balance charge until a neutral species is formed

Example: magnesium and chlorine



charge = +2 – 1 = +1, add Cl^{1-}

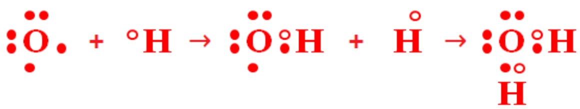


charges balance, formula = $MgCl_2$

Constructing molecular compounds (covalently bonded) using LEDs:

1. Draw LED models of both elements
2. Overlap valence shells
3. Share until stable octets form (H, Be, B, and Al do *not* form octets)

Example: hydrogen and oxygen



formula for water = H_2O

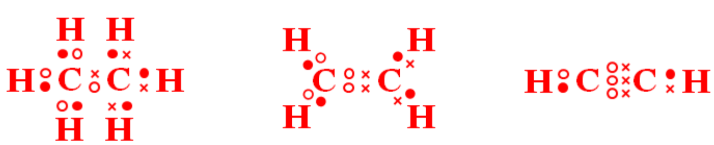
Example: oxygen molecule

7 e^- each   8 e^- each

formula for oxygen = O_2

Example: bond order and saturation in carbon – carbon bonds

ETHANE   ETHENE   ETHYNE



single bond saturated   double bond unsaturated   triple bond unsaturated

Four Solid Types, Bonding, and Properties

Metallic solid: characterized by a lattice of cations and freely moving electrons
- Good conductors of heat and electricity in solid and liquid states (due to sea of e^-)
- Malleable and ductile (due to sea of e^-)
- Hard with high MP and BP (due to very short bonds when all particles are cations)
- Not soluble in water

Ionic solid: characterized by a lattice of cations and anions (+ and – ions)
- Poor conductors of heat and electricity as solids (due to ions being trapped in the lattice) but good electrical conductors in the liquid state and solution (due movement of ions)
- Brittle (due to like ions repelling during crystalline dislocation)
- Hard with high MP and BP (due to strong ionic bonds)
- Usually are soluble in water (due to strong molecule – ion interactions) some ionic solids are not soluble in water (see Table F)

Network solid: characterized by a continuous web of covalent bonds
Generally poor conductors of heat and electricity in solid and liquid states (due strong lattice)
Brittle (due to opening up like a zipper along lattice bonds)
Very hard with high MP and BP (due to strong lattice of bonds, usually decompose with heat)
Generally insoluble in water

Molecular solid: characterized by discrete particles held together by IMFs
Poor conductors of heat and electricity in solid state (due to neutral nature of molecules)
Brittle (dipole – dipole molecules) to soft like wax (due to van der Waal's forces)
Soft with low MP and BP (most are liquids or gases at STP due to weak nature of IMFs)
Generally soluble in appropriate solvents (like dissolves like rule, polar with polar and nonpolar with nonpolar)