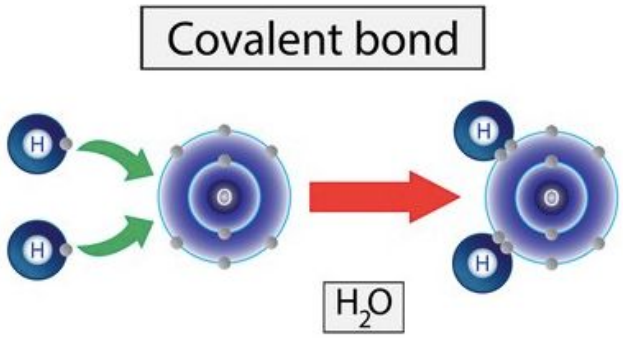
Chemical Bonds: Forces within Molecules

"O LORD, truly I *am* Your servant; I *am* Your servant, the son of Your maidservant; You have loosed my bonds." Psalm 116:16

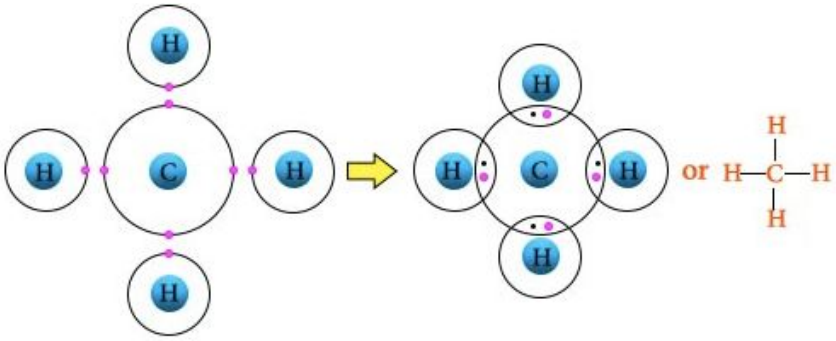
- Chemical bonds: attractive forces that link atoms
- Determined by the number of e- in valence shell → atoms tend to either share or transfer e- **to acquire eight e- in valence shell (octet rule)**
- H exhibits same tendency but with two valence e-

Covalent bonds

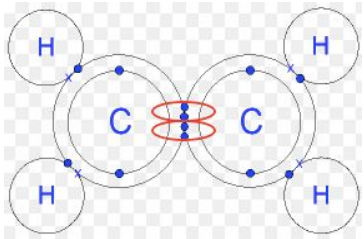
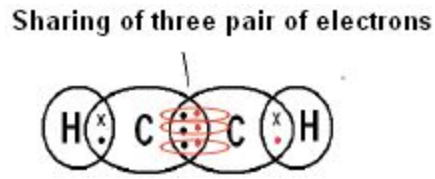


- Covalent bond: chemical bond resulting from the sharing of valence e- between atoms; atoms do not gain or lose e- so their electrical charges do not change; also number of valence shell does not change
  Ex: Cl - Cl

- Single covalent bond: two atoms share a single pair of e-
  Ex: CH_4 (methane)



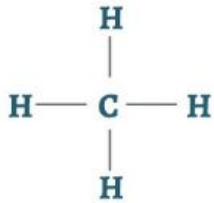
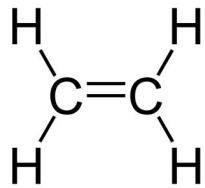
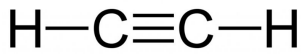
- Multiple covalent bond: atoms share two or three pairs of e-

Ex: double bond C_2H_4 (ethene)	
Ex: triple bond C_2H_2 (ethyne)	

- Lewis structures: diagram showing the locations of all the atoms and valence e- in a compound; each atom represented by its chemical symbol; atom's valence e- indicated as dots around the symbol; shared e- indicated as dots or lines [Table p.142]

Rules/Guidelines for Lewis structures

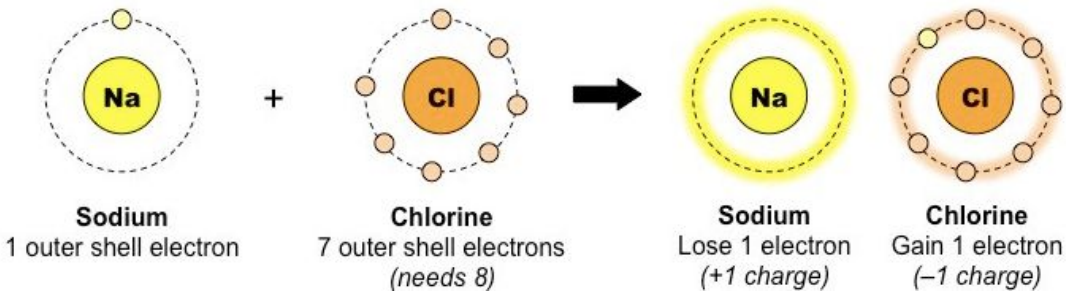
1. Write the number of valence e- for each atom in the molecule 2. Total the number of valence e- for the molecule 3. Arrange the atoms to form a "skeletal structure" for the molecule -draw only the symbols of the element -the first element listed is usually the central atom (except for **hydrogen**) 4. Draw two dots (⍰⍰) in between every two atoms (single bonds) 5. Draw two dots (⍰⍰) around all the outside atoms until you have reached the total number of valence electrons 6. Check each atom to make sure each has 8 electrons (or 2 for H and He) 7. If you need more electrons, you probably need to rearrange your dots to form **double** or **triple** bonds 8. Double check: number of dots = total number of valence electrons each atom follows the octet rule (except for H or He) 9. Rewrite all the shared electrons (single/double/triple) as lines (instead of dots)	Cl_2 N_2 H_2O CO_2

Ex: CH_4 (methane)	
Ex: double bond C_2H_4 (ethene)	
Ex: triple bond C_2H_2 (ethyne)	

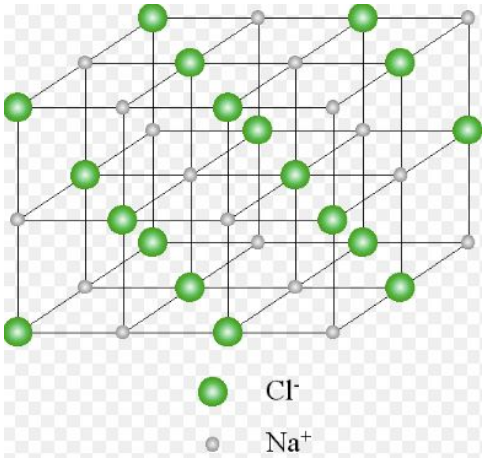
- Electronegativity: how strongly atoms of that element pull on the e- involved in chemical bonds; higher electronegativity → stronger pull
  -As the number of protons in the nucleus increases, the electronegativity or attraction will increase. Therefore **electronegativity increases from left to right in a row in the periodic table**.
  -From top to bottom down a group, electronegativity decreases. This is because atomic number increases down a group, and thus there is an increased distance between the valence electrons and nucleus, or a greater atomic radius.
- Nonpolar: covalent bond in which both atoms share e- equally
- Polar: one of the atoms has greater electronegativity and that atom has greater pull of shared e- toward itself; formation of dipole: set of oppositely charged poles with partial charges (**δ**)
  Ex: H-Cl

- Covalent networks: atoms covalently bonded into large structures that do not contain a specific number of atoms but do contain a fixed ratio of elements
  Formula unit: simplest ratio of atoms in a covalent network
  Ex: SiO_2; also graphite v. diamond

Ionic bonds

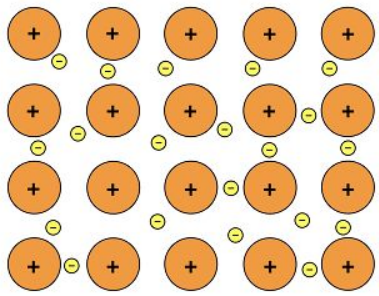


- Ions: atoms that develop electric charges due to a loss or gain of e-
  Anions: atoms gain e- and become (-) charged
  Cations: atoms lose e- and become (+) charged
- Ionic bond: attraction between oppositely charged ions
  Ex: NaCl
  -Na has 1 valence e-, Cl has 7 valence e-
  -Na valence e- completely transferred to valence shell of Cl
  -both atoms gain complete outer shells
  -Na loses e- so now is (+1) also Na^+, Cl gains and e- and is now (-1) also Cl^-
- Ionic bonds also form between groups of atoms
  Ex: NH_3 (ammonia) can covalently bond to H^+ to become NH_4^+; NH_4^+ can be attracted to any anion to form an ionic compound such as NH_4Cl (ammonium chloride)
- Polyatomic ion: ion formed from a group of atoms that includes ionic and covalent bonds
  Ex: NH_4Cl
- each ion is attracted to as many oppositely charged ions as can crowd around it
  → Ionic crystals: ions all bonded by electrical attraction (salt crystals)



Metallic bonds

- Valence electrons detach from their original atomic owners and float around in the "sea" of metal atoms; the metal atoms become positive ions. The result is an orderly structure of positive metal atoms surrounded by a sea of negative electrons that hold the ions together like glue.



- Bonding in metals; valence e- are shared by all the atoms in the metal; shared e- are free to move among the atoms
- e- are not restricted to specific bonds, so metal atoms can combine in almost any ratio resulting in vast number of alloys (metal mixtures)
  brass (Cu + Zn), bronze (Cu + Sn), steel (Fe + C)

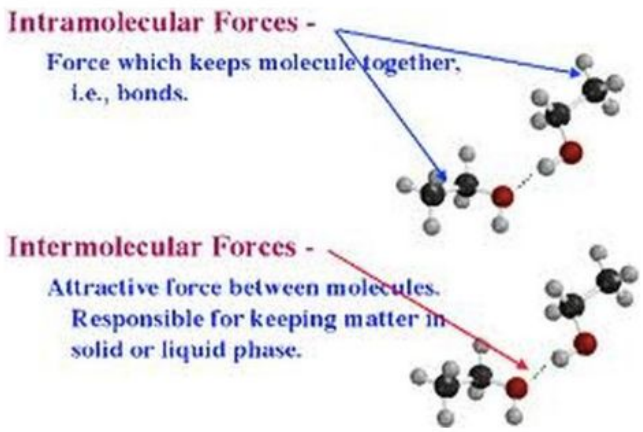
Naming compounds

- Name of a chemical based on its molecular composition, naming each element and giving its prefix based on its abundance; name of last element altered to end in *-ide*
  [Prefix table p.147]
- Prefix mono- used only when there is another compound composed of same types of atoms:
  CO (carbon monoxide), CO_2 (carbon dioxide)
  H_2O (ask students)

Chemical formulas

- Combinations of chemical symbols and numbers showing the type and number of atoms in each molecule
- Molecular formula: number of each type of atom in a molecule; does not show arrangement of atoms within a molecule
- Structural formula: general arrangement of atoms in a molecule
- Empirical formula: shows only simplest ratio of atoms in a compound; same as formula unit; rarely used
  [Chemical formulas table p.149]

Forces between Molecules



- Intermolecular forces: forces that bind molecules together to form larger structures
  Responsible for cohesion and adhesion
- Dipole-dipole: force of attraction <u>between</u> polar molecules (picture p.150); present in all polar molecules (where there is unequal sharing of valence e-)
- London forces: as e- move around in orbitals, temporary concentration in one area of e- results in a deformed cloud and temporary dipole → generates partial charges in a neighboring molecule
  -individual London attractions are very weak, but cumulative effect can be significant
  -strength of London forces is proportional to the size of the molecule
- Hydrogen bonds: unusually strong intermolecular force present when H covalently bonds to a more electronegative atom
  -strongest of all intermolecular forces
  -responsible for cohesion, adhesion of water molecules and crystalline structure of ice
  [penny experiment]

<u>Effects of Intermolecular Forces</u>

- Stronger intermolecular forces between the molecules of a solid results in higher melting points
- Chemical bonds are much stronger than intermolecular forces → nonmolecular compounds (covalent and ionic) typically have higher melting points than molecular compounds (held together by dipole-dipole/London/H bonds)
- Vapor pressure and boiling point of a molecular substance depend primarily on the molecular mass of the molecules and the types of intermolecular forces between them
- Polar and ionic solutes are dissolved by polar solvents, while nonpolar solvents are dissolved by nonpolar solvents; "like dissolves like"
  Ex: sugar in water (polar-polar)

Iodine in oil