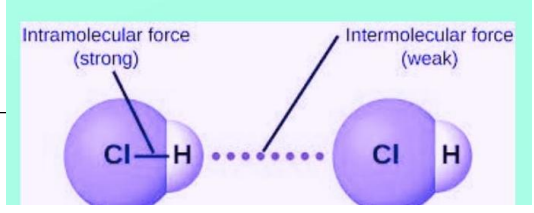
AP Chemistry Unit 2

2.1 Types of Chemical Bonds

Overview



- **Intramolecular forces** are the forces that hold ATOMS together with a molecule: Types- Ionic, Covalent, Metallic
- **Intermolecular forces (IMFs)** - this involves TEMPORARY ATTRACTIONS between separate molecules. Unit 3
- Bonding within a molecule will heavily influence the types of attractions between separate molecules. Therefore, it makes sense to start with bonding first

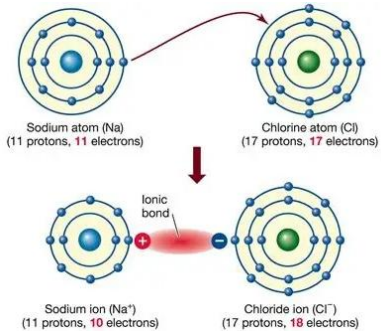
Why do bonds form?

- This goes back to the concept of what drives chemical reactions.
- Bonds form due to increased stability.
- What drives increased stability? Getting to a lower energy level.
- **In order for bonds to form (unless it is forced) - the drive is that the bonded atoms are at a lower energy when together than when separated.**

Review

- **Octet rule:** atoms tend to gain, lose, or share electrons in order to acquire eight valence electrons
- **3 Types of chemical bonds**
  - **Ionic bond** – transfer of electron – should be review
  - **Covalent bond** – sharing of electrons – should be review
  - **Metallic bond** – electron sea model – new info

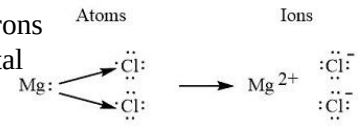
Ionic Bonds

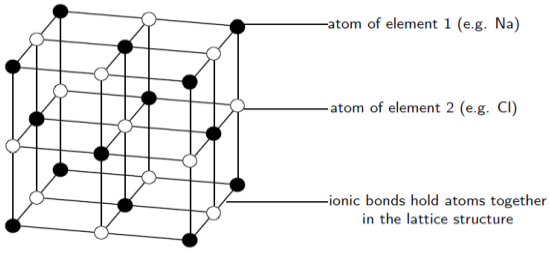


- **Definition**: **Ionic bonds** form when one atom transfers electrons to another atom, resulting in the formation of positively and negatively charged ions. These opposite charges attract each other, creating a **strong electrostatic force**.
  - Coulomb's Law applies here
- **Formation**: Typically occurs between metals (which lose electrons) and non-metals (which gain electrons).

Ionic Bonds- Characteristics

- **Formation of Ions**: Metal atoms lose electrons to form positive ions (cations), and non-metal atoms gain electrons to form negative ions (anions).

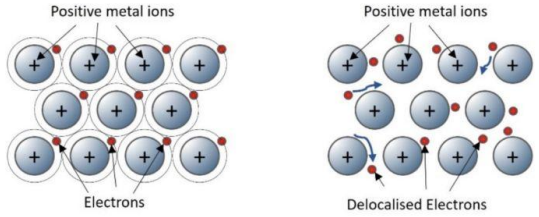




- **Crystal Lattice Structure**: Ionic compounds form a regular repeating pattern of ions.

Metallic Bonds

- **Definition**: Metallic bonds form between metal atoms, where electrons are free to move throughout the structure, often described as a "sea of electrons."
- **Formation**: Between metal atoms.



Examples:

- **Aluminum (Al)**: Aluminum atoms share a pool of electrons.
- **Gold (Au)**: Gold atoms share a sea of electrons.

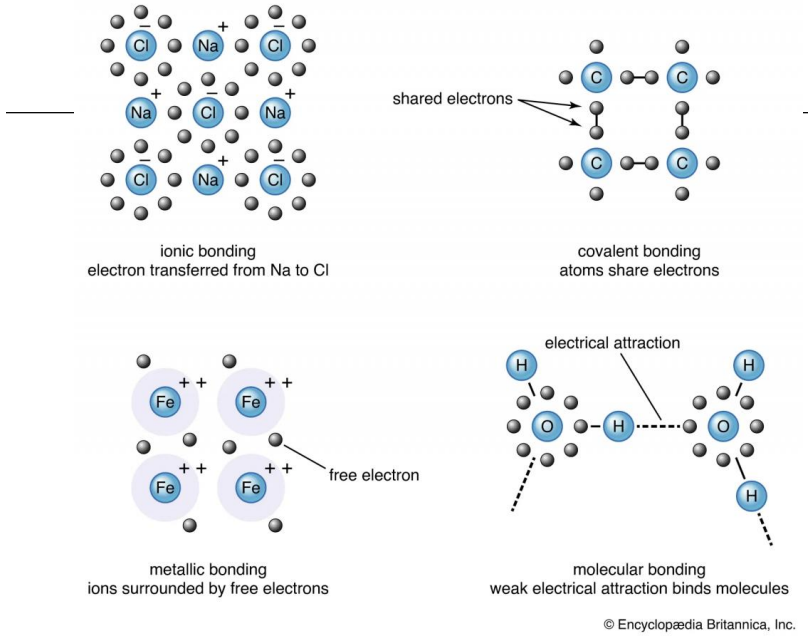
Metallic Bonds

Characteristics

- **Delocalized Electrons**: Electrons are not associated with any specific atom and can move freely.
- **Conductivity**: Metals are good conductors of electricity and heat due to free-moving electrons.

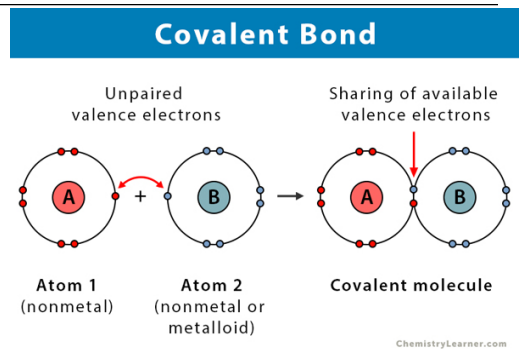
Properties:

- **Malleability and Ductility**: Metals can be hammered into sheets and drawn into wires.
- **Luster**: Metals have a shiny appearance.
- **High Melting and Boiling Points**: Due to strong metallic bonding.



Covalent Bonds

- **Definition**: Covalent bonds form when two atoms share one or more pairs of electrons. This bond typically occurs between non-metal atoms with similar electronegativities.



- **Formation**: Between non-metals.

Covalent Bonds

Characteristics:

- **Electron Sharing**: Electrons are shared between atoms to achieve a stable electron configuration.
- **Bond Strength**: Can vary depending on the number of shared electron pairs (single, double, or triple bonds).

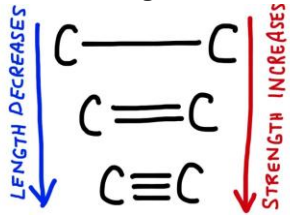
Properties:

- **Lower Melting and Boiling Points**: Compared to ionic compounds.
- **Variable Solubility**: Solubility in water can vary; many covalent compounds are not soluble.
- **Poor Electrical Conductivity**: Do not conduct electricity in solid or liquid states.

Strength of Covalent Bonds

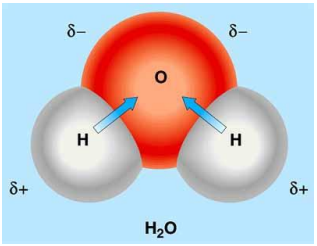
- The **strength depends** on the distance between the two nuclei, or **bond length**
  - As length increases, strength decreases!!!

- The amount of energy required to break a bond is called the **bond dissociation energy**
  - The shorter the bond length, the greater the energy required to break it
  - As the number of shared **electron pairs increase** the **bond length decreases**
    - Shorter the bond, the stronger the bond.
      - Example thinking of breaking a pencil



Polar & Non-polar covalent bonds

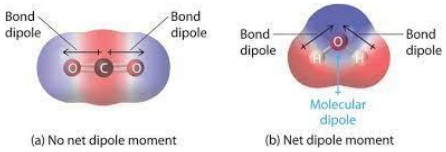
- Covalently bonded molecules are either **polar** or **non-polar**
- Unequal sharing of electrons results in a **Polar covalent bond**
  - Polar covalent bonds form when atoms pull on electrons in a molecule unequally
  - charges at the ends of the bonds called a **dipole**
  - A molecule is **POLAR** if: the central atom as a lone pair
  - Polar molecules align with an electric field
    - Example: Water
- **Non-polar molecules** are not attracted by an electric field



Chemical Bonding

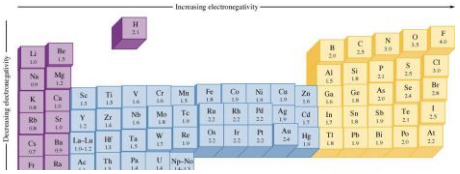
Dipole Moment has a **Magnitude** and a **Direction**

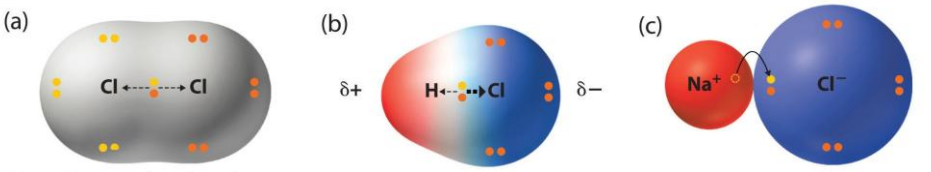
- **Covalent Bonds:**
  - Polar and non-polar
  - **Dipole moments** – separation of the charge in a molecule; product of the size of the charge and the distance of separation.
    - Molecules with a somewhat negative and a somewhat positive end have dipole moments.
    - Arrows are used to indicate DPM from positive toward negative.
    - If octet rule is obeyed AND all the surrounding bonds are the same all DPM should cancel our and the molecule is NON polar.



Bond Characterization

- Based on electronegativity differences
  - < .3 – Non polar covalent
  - .4-1.7 – Polar covalent
  - >1.7 – Ionic





Nonpolar covalent bond
Bonding electrons shared equally between two atoms. No charges on atoms.

Polar covalent bond
Bonding electrons shared unequally between two atoms. Partial charges on atoms.

Ionic bond
Complete transfer of one or more valence electrons. Full charges on resulting ions.

Determining Bond Character with Electronegativity

- The character of a bond depends on how strongly each of the bonded atoms attracts electrons
  - Bonding is often not clearly ionic or covalent
- You can **subtract the electronegativities** of the elements in the compound to determine the **character and bond type**

H_20
Electronegativity O = 3.44
Electronegativity H = 2.20
Difference = **1.24**
Polar covalent !

Table 8.7	**EN Difference and Bond Character**
Electronegativity Difference	Bond Character
> 1.7	mostly ionic
0.4 – 1.7	polar covalent
< 0.4	mostly covalent
0	nonpolar covalent

Since Oxygen electronegativity is higher than that of hydrogen, the shared pair of electrons are closer to Oxygen. Making the Oxygen end of the compound more negative

Electron Affinity, Electronegativity, and Bond Character

- **Electron affinity** measure the tendency of an atom to accept an electron
  - Noble gases are not listed because they do not form compounds

Electronegativity Values for Selected Elements

1 **H** 2.20																	
3 **Li** 0.98	4 **Be** 1.57											5 **B** 2.04	6 **C** 2.55	7 **N** 3.04	8 **O** 3.44	9 **F** 3.98	
11 **Na** 0.93	12 **Mg** 1.31											13 **Al** 1.61	14 **Si** 1.90	15 **P** 2.19	16 **S** 2.58	17 **Cl** 3.16	
19 **K** 0.82	20 **Ca** 1.00	21 **Sc** 1.36	22 **Ti** 1.54	23 **V** 1.63	24 **Cr** 1.66	25 **Mn** 1.55	26 **Fe** 1.83	27 **Co** 1.88	28 **Ni** 1.91	29 **Cu** 1.90	30 **Zn** 1.65	31 **Ga** 1.81	32 **Ge** 2.01	33 **As** 2.18	34 **Se** 2.55	35 **Br** 2.96	
37 **Rb** 0.82	38 **Sr** 0.95	39 **Y** 1.22	40 **Zr** 1.33	41 **Nb** 1.6	42 **Mo** 2.16	43 **Tc** 2.10	44 **Ru** 2.2	45 **Rh** 2.28	46 **Pd** 2.20	47 **Ag** 1.93	48 **Cd** 1.69	49 **In** 1.78	50 **Sn** 1.96	51 **Sb** 2.05	52 **Te** 2.1	53 **I** 2.66	
55 **Cs** 0.79	56 **Ba** 0.89	57 **La** 1.10	72 **Hf** 1.3	73 **Ta** 1.5	74 **W** 1.7	75 **Re** 1.9	76 **Os** 2.2	77 **Ir** 2.2	78 **Pt** 2.2	79 **Au** 2.4	80 **Hg** 1.9	81 **Tl** 1.8	82 **Pb** 1.8	83 **Bi** 1.9	84 **Po** 2.0	85 **At** 2.2	
87 **Fr** 0.7	88 **Ra** 0.9	89 **Ac** 1.1															

Metal
Metalloid
Nonmetal

Extra Information follows

Review

- **Chemical Bonding**
  - Nature is striving for a lower energy state
  - For atoms approach to form bonds:
    - 2 "bad" things happen – E/E repulsion and P/P repulsion
    - 1 "good" thing happens – electron/proton attraction.
  - When the attractive forces offset the repulsive forces the energy of the atoms decreases and a bond forms.

- **3 types Covalent Bonds**
  - Single – share one pair of electrons H:H or H-H
  - Double – share two pairs of electrons O :: O or O=O
  - Triple – share three pair of electrons N::: N or $\equiv$ N

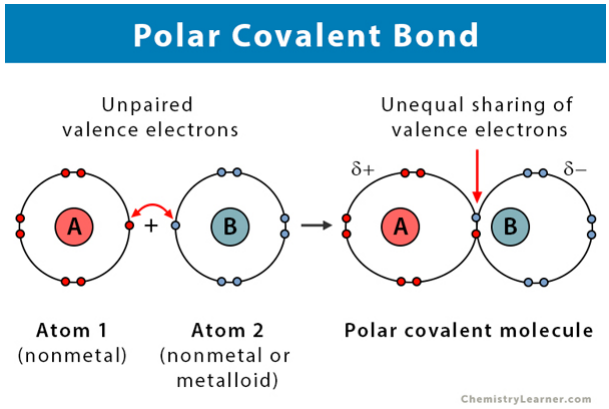
Review

- **Covalent bonds** – sharing of electrons between atoms
  - Atoms gain stability when they share electrons and form covalent bonds (full octet, or duet for H)
  - Sharing valence electrons with other atoms also results in noble-gas electron configurations - octet rule
  - Shared electrons are counted as owned by both atoms
  - A **molecule** is composed of two or more elements that are covalently bonded.
    - Ionic bonds form **formula units** or **ionic compounds**

- Covalent bonds can occur between same elements
  - **Diatomic elements** – H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2

Review – Polar Covalent bonds

Definition: A type of covalent bond where the electrons are shared unequally between the two atoms, resulting in a molecule with a partial positive charge on one end and a partial negative charge on the other

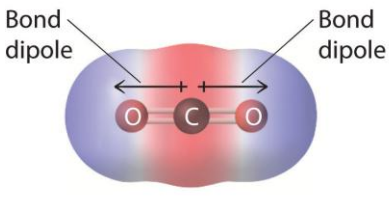


- **Formation**: Between atoms with different electronegativities.

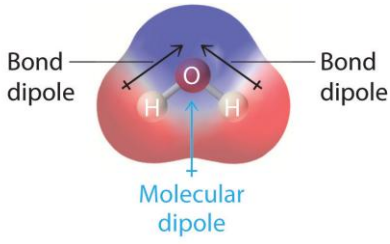
Review – Polar Covalent bonds

Characteristics:

- **Unequal Electron Sharing**: Electrons are more attracted to the atom with higher electronegativity.
- **Dipole Moment**: Creates a partial charge difference within the molecule.



(a) No net dipole moment



(b) Net dipole moment

Review – Polar Covalent bonds

Properties:

- **Dipole-Dipole Interactions**: Intermolecular forces between polar molecules.
- **Solubility in Polar Solvents**: Often soluble in water and other polar solvents.
- **Higher Melting and Boiling Points**: Compared to nonpolar covalent compounds.

Examples:

- **Water (H_2O)**: Oxygen is more electronegative than hydrogen, creating partial charges.
- **Hydrochloric acid (HCl)**: Chlorine is more electronegative than hydrogen.

Properties of Covalent Compounds

- Covalent bonds between atoms are strong, but attraction forces between molecules are weak.
- The weak attraction forces are known as van der Waals forces.
- The forces vary in strength but are weaker than the bonds in a molecule or ions in an ionic compound
- Non-polar molecules exhibit a weak dispersion force, or induced dipole. The force between two oppositely charged ends of two polar molecules is a dipole-dipole force.
- A hydrogen bond is an especially strong dipole-dipole force between a hydrogen end of one dipole and a fluorine, oxygen, or nitrogen atom on another dipole.
- Weak forces result in the relatively low melting and boiling points of molecular substances.
- Many covalent molecules are relatively soft solids.

Ionic Bonding and Coulomb's Law

So how strongly "bonded" are things in ionic solids?

It depends on:

1. How + or - things are (the greater the difference in charge, the stronger the attraction)
2. How close together things are (the lower the radius, the stronger the attraction)

You can predict everything you need to know about an ionic solid based off this.

Chemical Bonding

- Ionic Bonding results from cations and anions in a regular array called a crystal lattice.
  - Water must overcome the ion-ion attractions to dissolve ionic substance.
    - Size and charge affect this.
  - Lattice energy is the energy contained in the crystal lattice.
    - Can be calculated by:

$$\text{Lattice Energy} = k\left(\frac{Q_1 Q_2}{r}\right)$$

    - K is a proportionality constant that depends on the structure of the solid and the electron configurations of the ions.