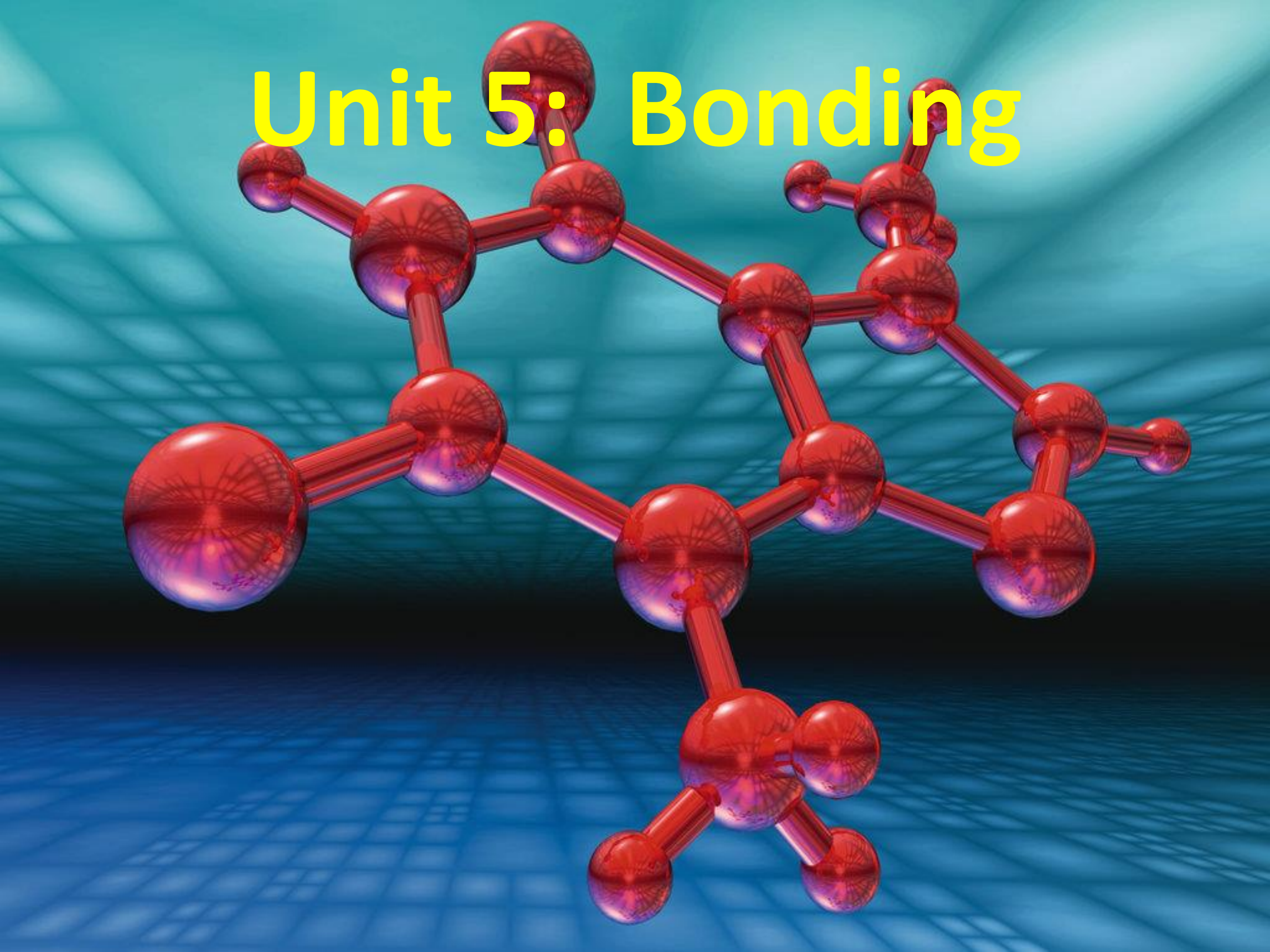
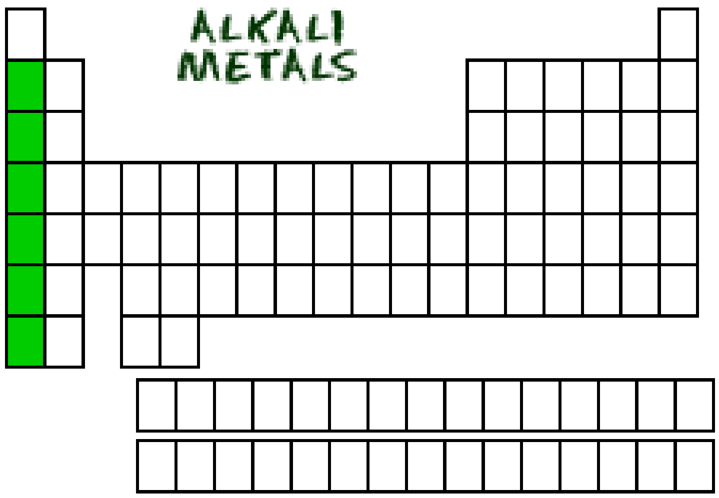
Unit 5: Bonding

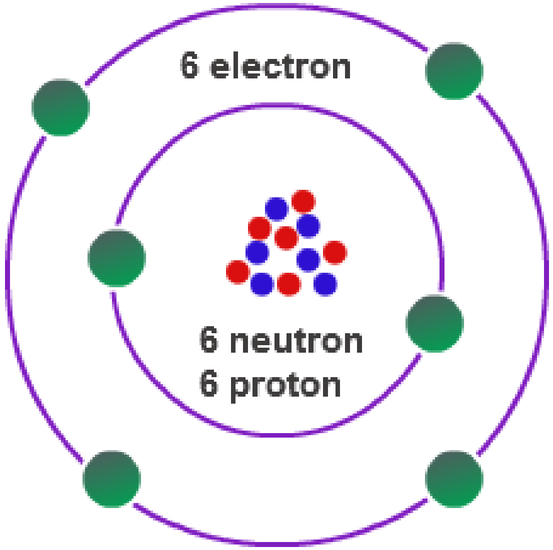


Scientists learned that elements in same group on PT react in a similar way



Why?

They all have the same number of valence electrons....which are electrons in the highest occupied energy level (outer shell)

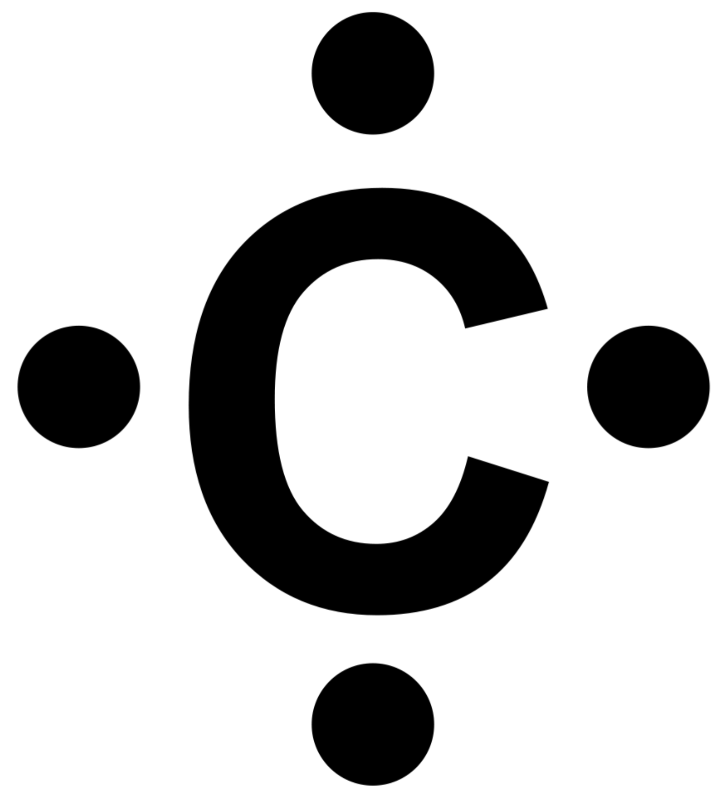


How do you determine the number of valence electrons in an atom?

Electron dot structures

Valence electrons are usually the only electrons involved in bonding

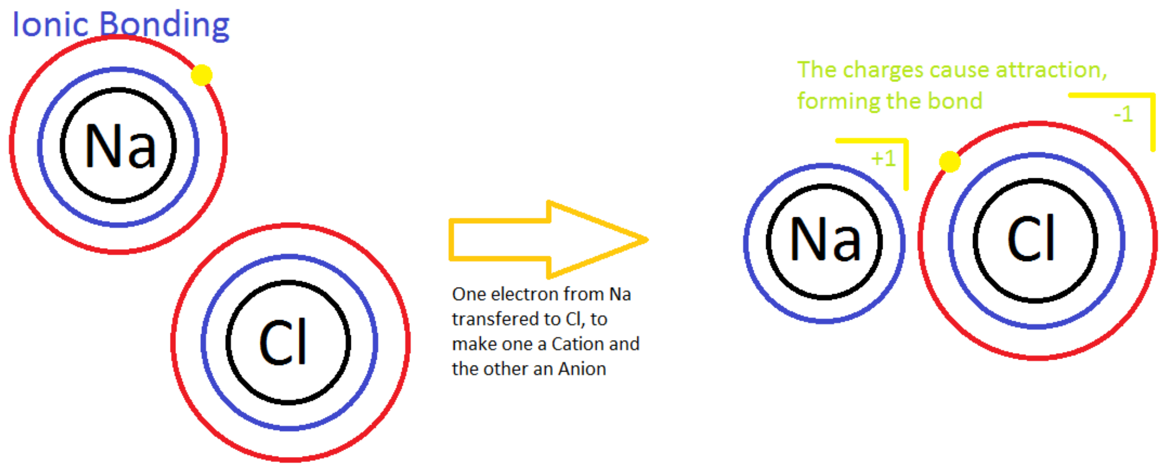
Represented with **Lewis dot structures**

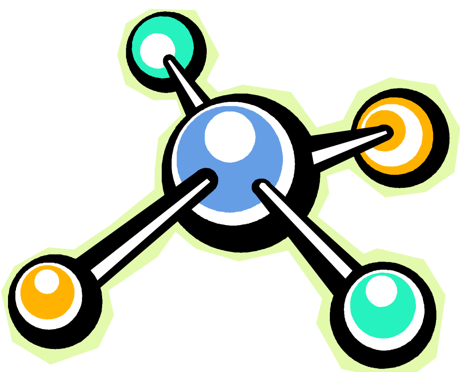


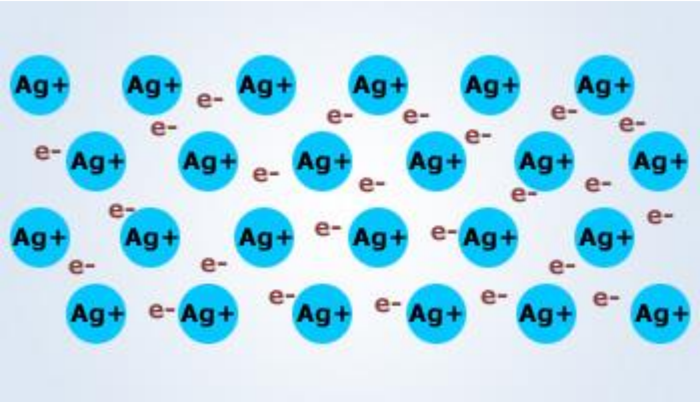
Types of bonds

Intramolecular bonds **(atom to atom bonds)**

- **Ionic, Covalent or Metallic**



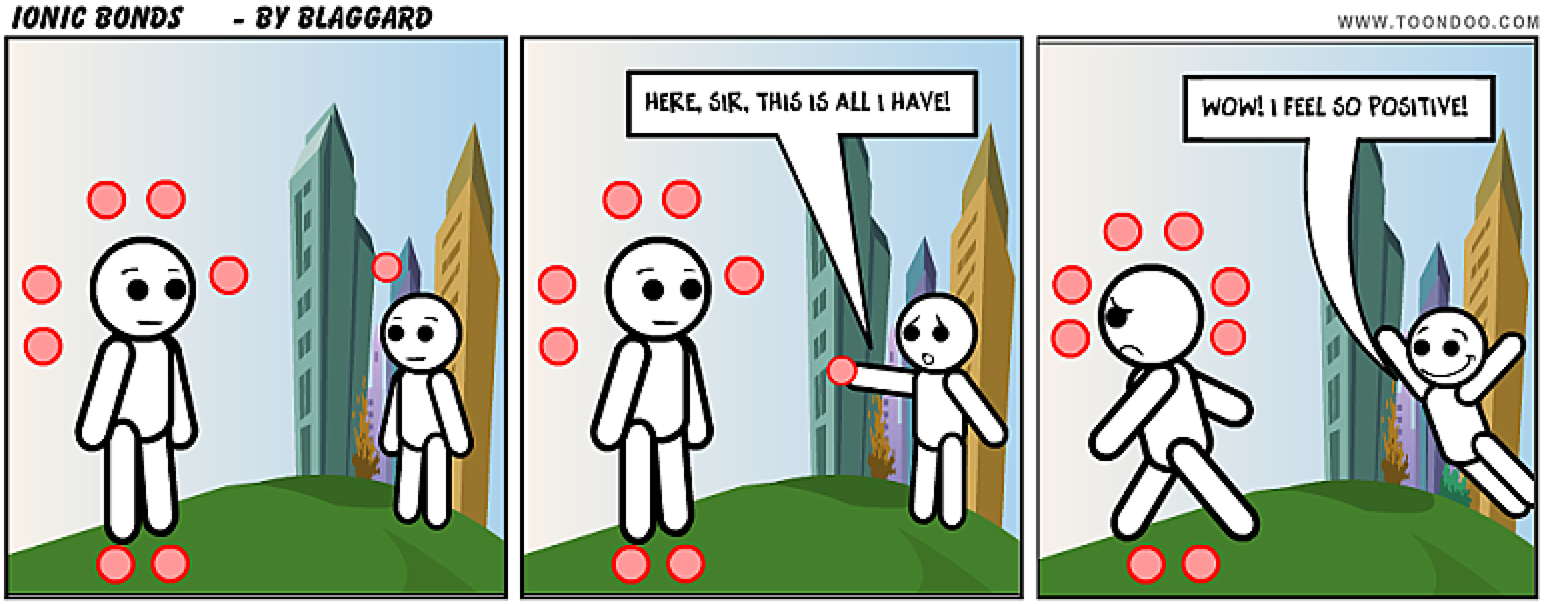




What is Chemical Bonding?

A **chemical bond** (net attractive force) is formed when electrons are shared_or given between two or more atoms.

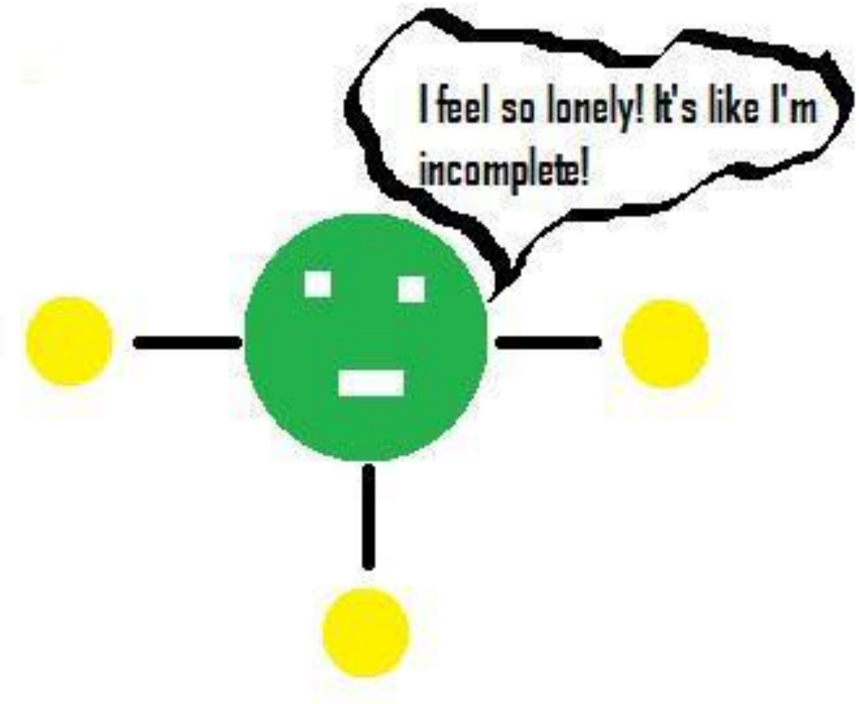
The electrons involved are only the outermost electrons – the valence electrons.



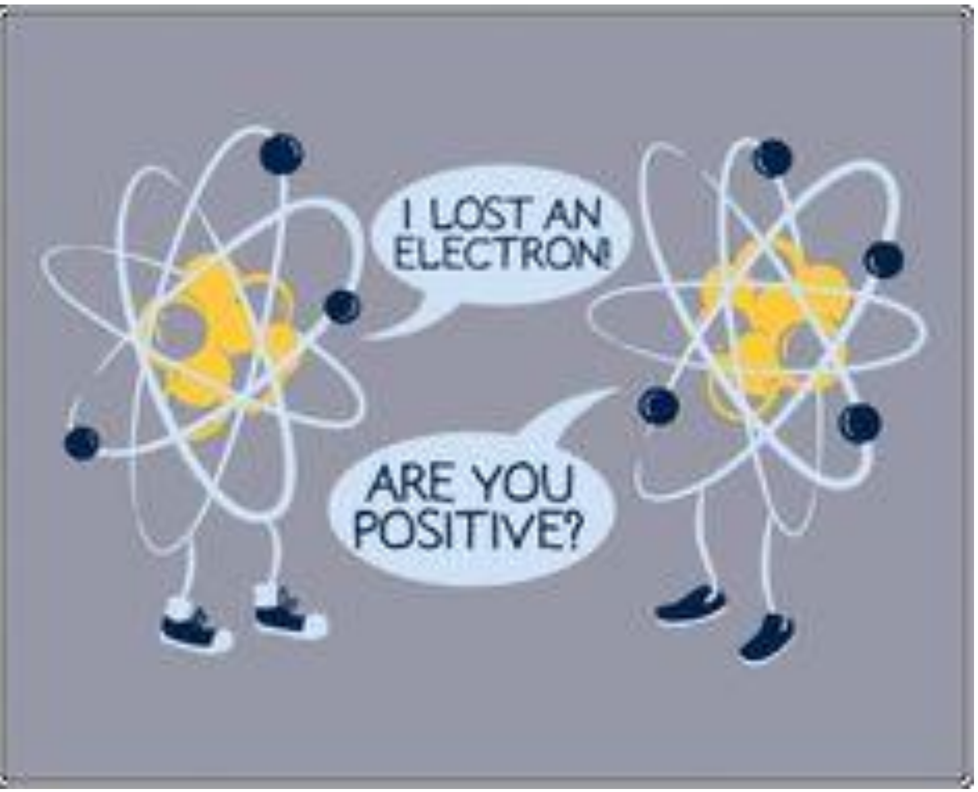
The Octet Rule

Atoms typically bond to form an **octet** (a group of 8 electrons) in their valence level.

All atoms want this stability (electron configuration of a Noble gas)



Ions and ionic bonding



Atoms are electrically neutral

If they lose or gain an electron they become a charged particle- an **ion**

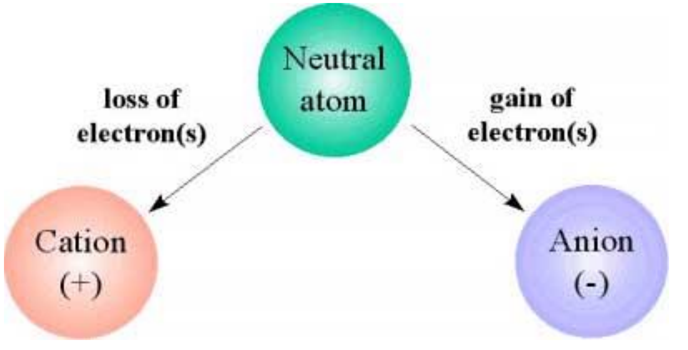
Cations= positively charged (lose e^-)

Anions= negatively charged (gain e^-)

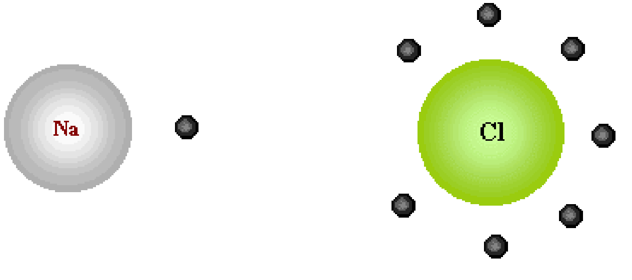
Ionic Compound (aka salts)

An ionic bond is formed when **electrons are transferred** from one atom to another (both atoms end up with an octet of electrons in their outer shell).

Metals lose electrons (cation) and non-metals gain electrons (anion). This ***creates positive and negative ions*.**

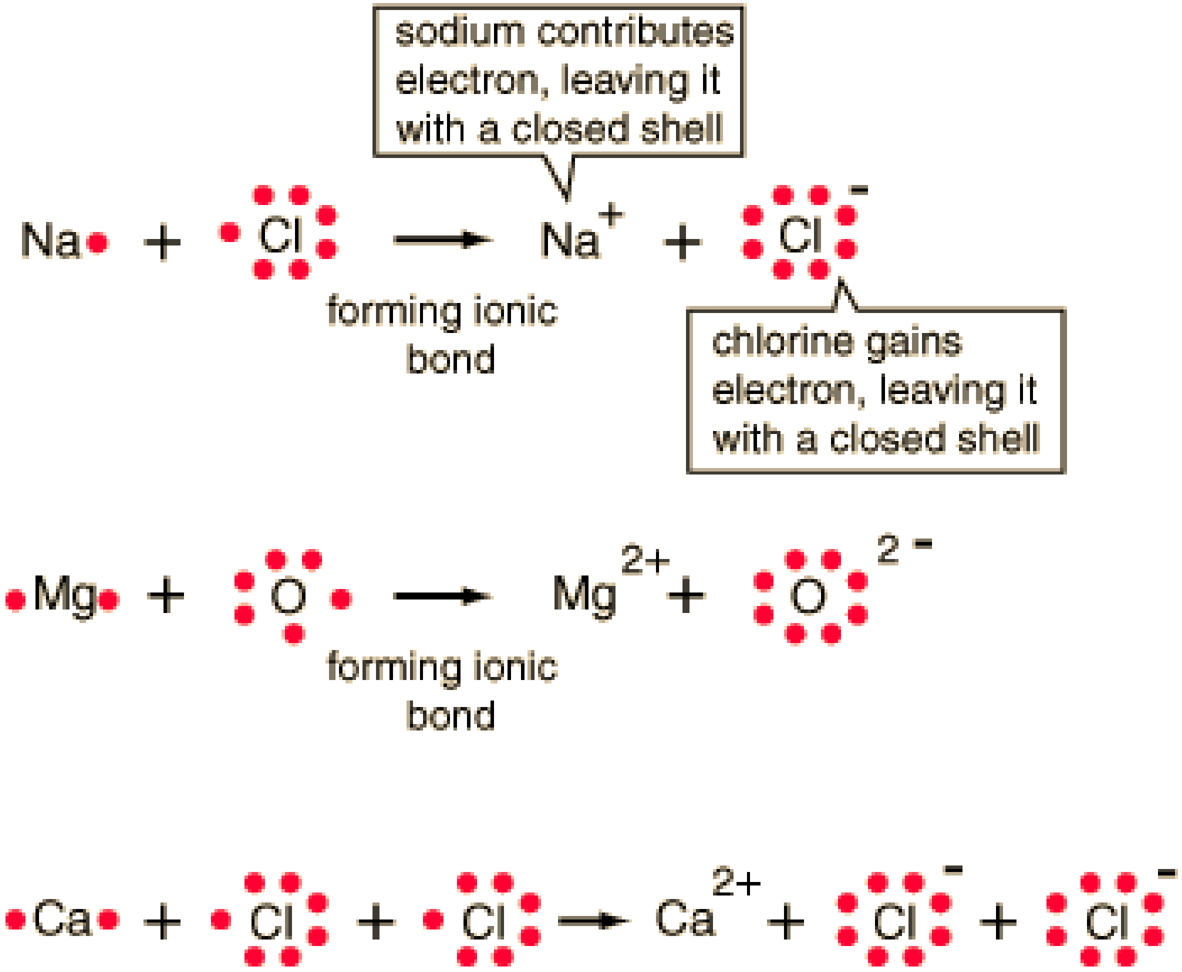


Since they have opposite charges, they are **attracted to each other through electrostatic force.**



Ionic Bond

Ionic Compounds can be visualized with Lewis dot notation.



Use Lewis dot structures to predict formulas for

1. Potassium and oxygen

2. Magnesium and Nitrogen

Properties of Ionic Compounds

High melting points / boiling points – it takes a lot of energy to break strong bonds.

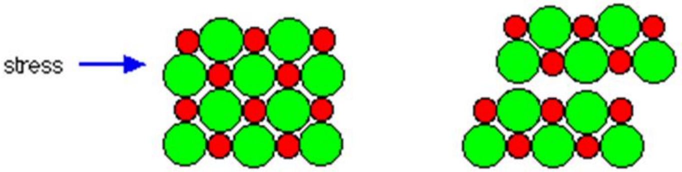
Hard, brittle solids

Many are **soluble** in water

When dissolved, free ions float and conduct electricity (they are **electrolytes**)

Form crystalline solids – have a regular/repeating structure

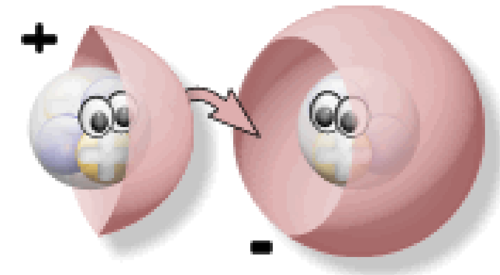
Brittleness is the tendency to break, snap, or crack, when subjected to pressure.



Ions of the same charge are brought side-by-side and so the crystal repels itself to pieces!

Held together by lattice energy

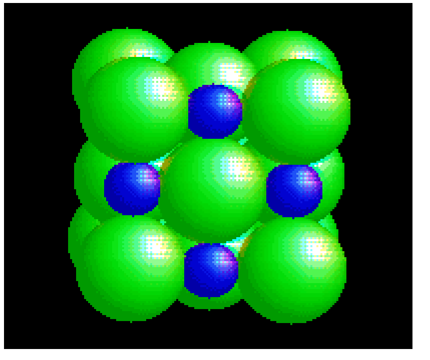
Ionic Bond



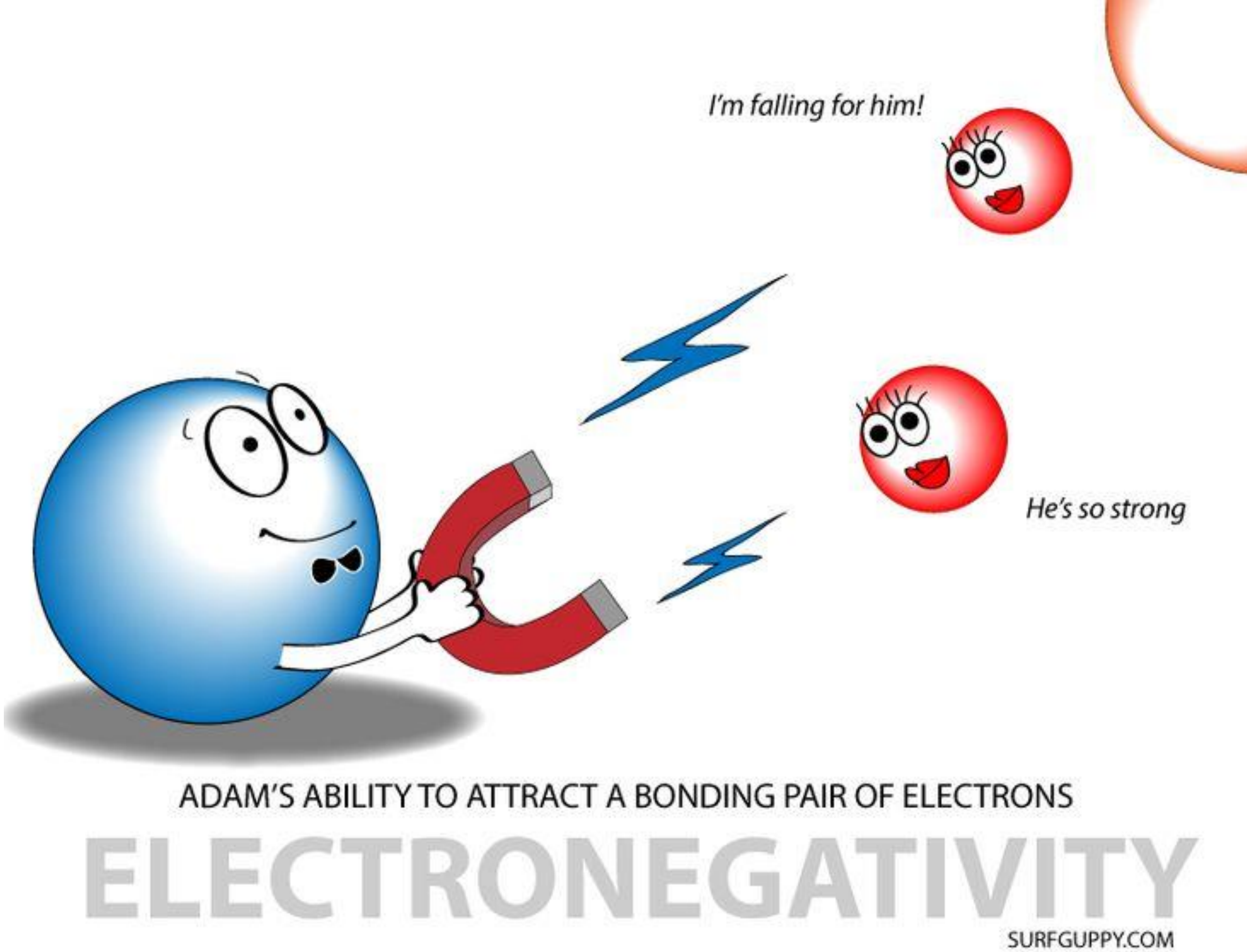
Overall, **salts are neutral**. They have equal amounts of + and - charge.

Ionic compounds do not exist as single discrete units but as collections of + and -ly charged ions arranged in repeating patterns

A **formula unit** is a formula that tells the ratio of ions in an ionic compound. (The **smallest part of an ionic compound** that still has the composition of the compound.)



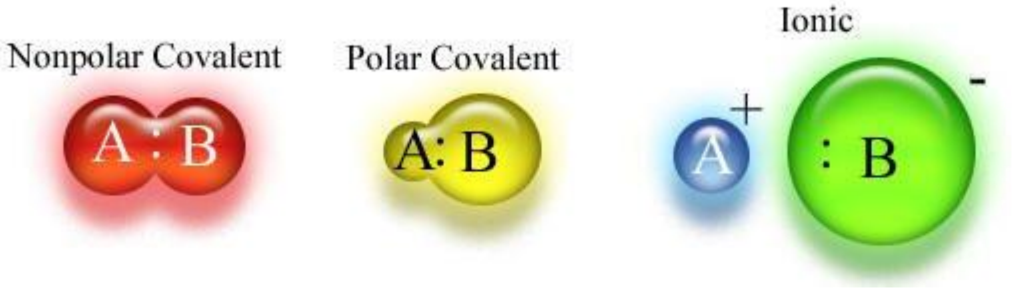
$Na_{12}Cl_{12}$ Simpler ratio is NaCl


I'm falling for him!
He's so strong
ADAM'S ABILITY TO ATTRACT A BONDING PAIR OF ELECTRONS
ELECTRONEGATIVITY
SURFGUPPY.COM

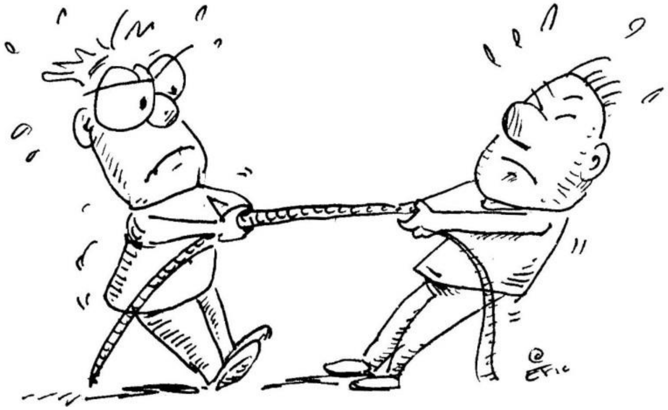
Electronegativity

In order for electrons to be transferred, one element must be able to take electrons from the other (more electronegative)

For ionic bonds to form there must be a large difference in electronegativity **(difference greater than 1.7)**



Covalent (Molecular) Bonds



A bond that results from the sharing of electrons.

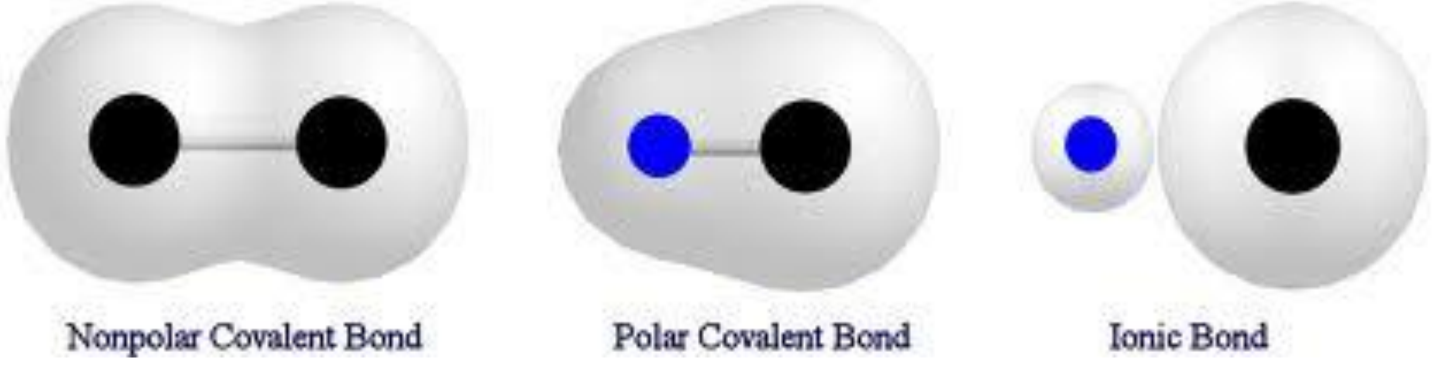
Atoms must stay together to share so molecules are formed

Usually occurs between two nonmetals.

Covalent Bonds & Electronegativity

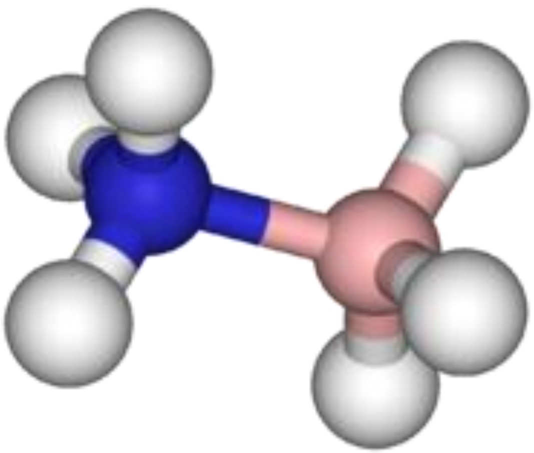
When two atoms do not have a big difference in Electronegativity (EN), they will share electrons. There are varying degrees of how electrons can be shared.

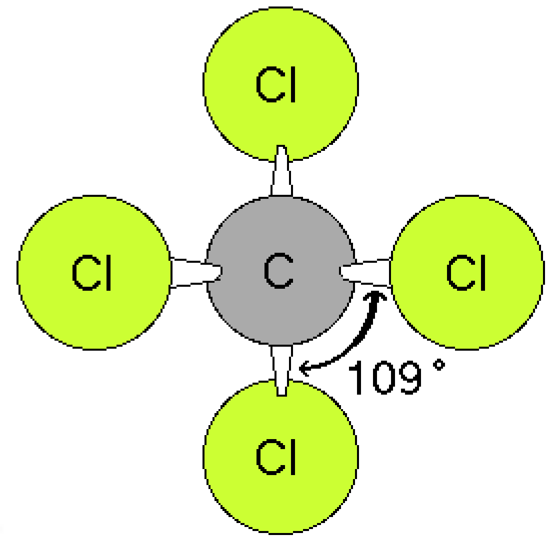
Shared equally: **nonpolar covalent bond**.
The EN values are almost equal.
Shared unequally: **polar covalent bond.**
The EN values are not equal, but not different enough to form ionic.



Nonpolar Covalent Bonds

If electrons are shared equally, the molecules overall charge is neutral.



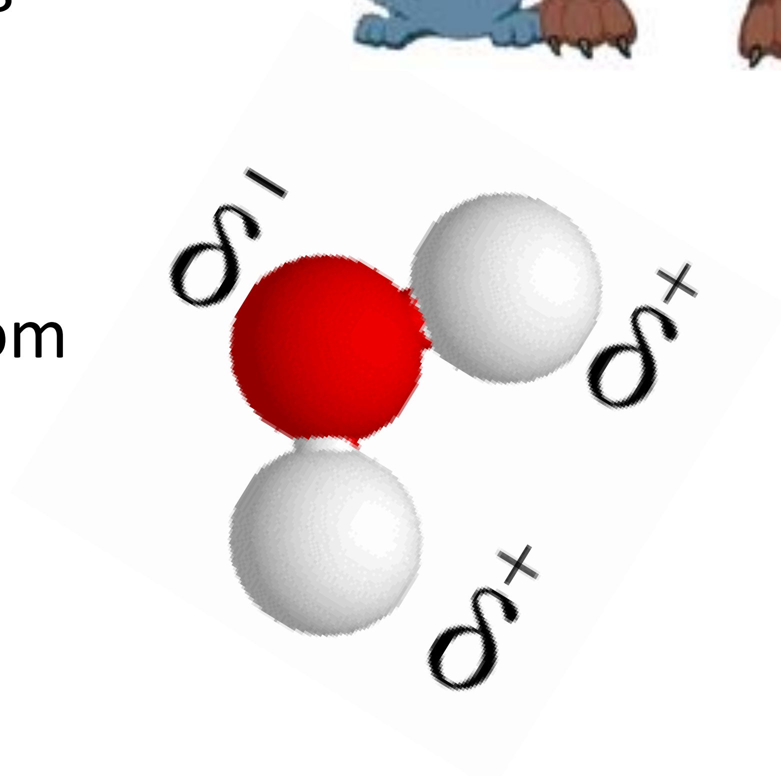


Nonpolar Covalent Bond

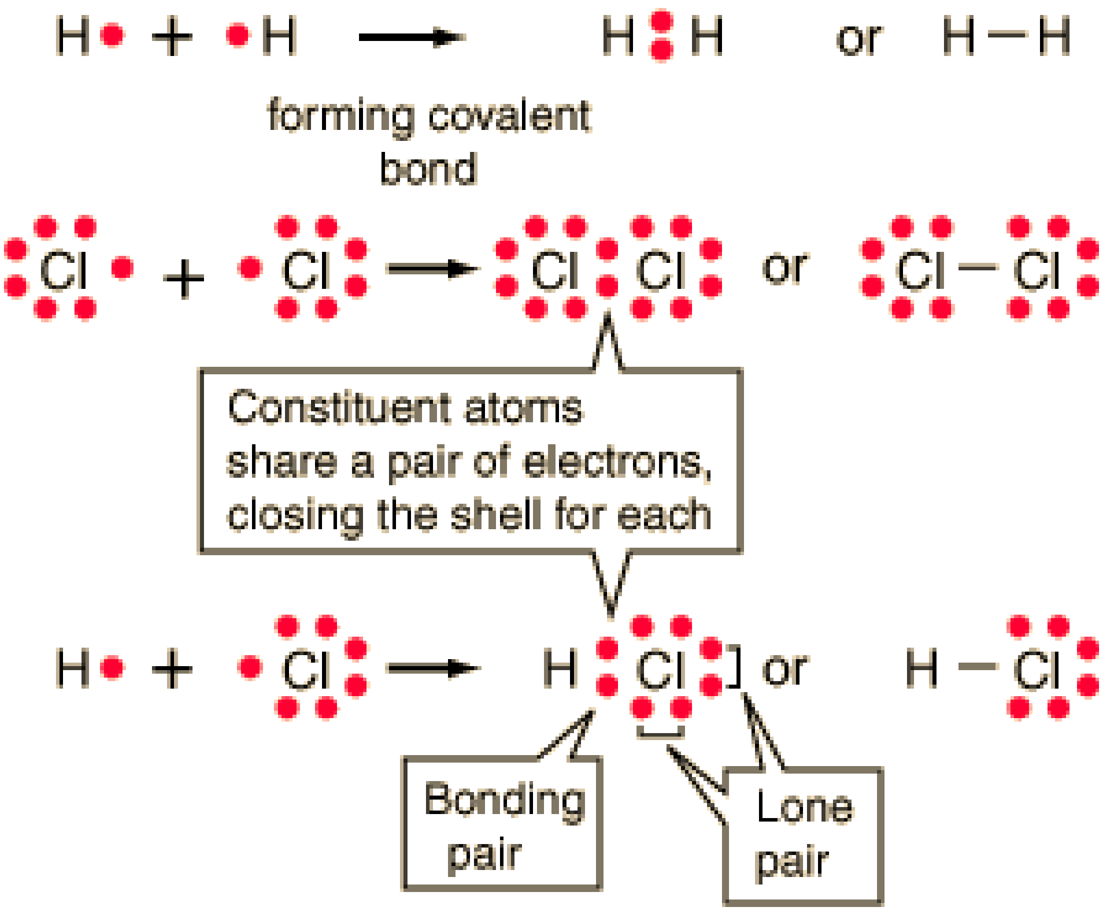
Polar Covalent Bonds

Electrons are not shared equally, the *more electronegative* atom attracts the electrons more, forming a partially negative region of the atom.

The less electronegative atom becomes partially positive.



Lewis Dot and Covalent Bonds



Properties of Covalent Bonds

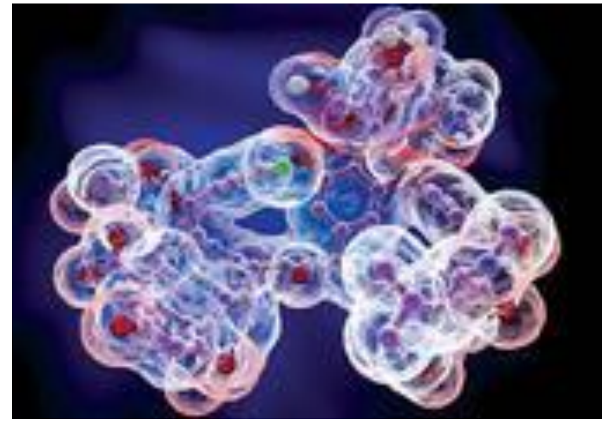
No ions, no charges, do not conduct electricity.

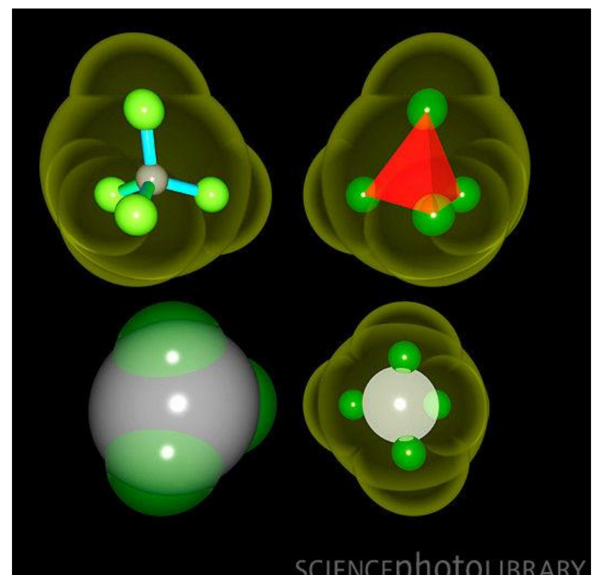
Weaker attraction between molecules than ionic bonds.

Usually liquids or gases at room temperature (if solid will have a low melting point)

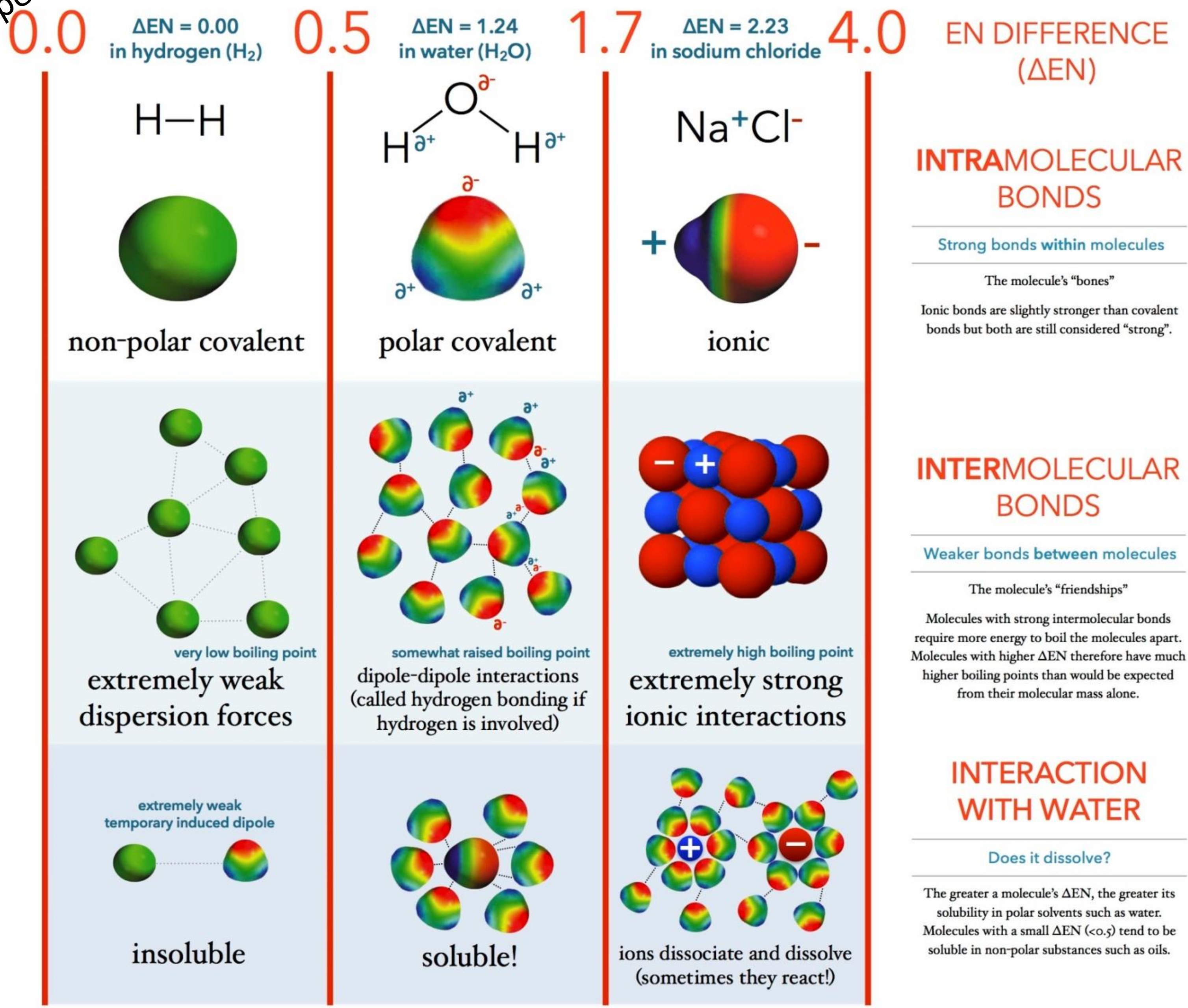
Insoluble in water (remain molecules)

Amorphous solids – do not have a regular/repeating structure





How electronegativity determines bond type:



Identify the type of bonds between each pair of elements

1. H and O
2. F and Zn
3. Na and Cl
4. P and O

What is the electronegativity difference, type of bond, and which is more electronegative atom?

Periodic Table of the Elements

Electronegativity

1.0 1.3 1.6 1.9 2.2 2.5 2.8 3.1

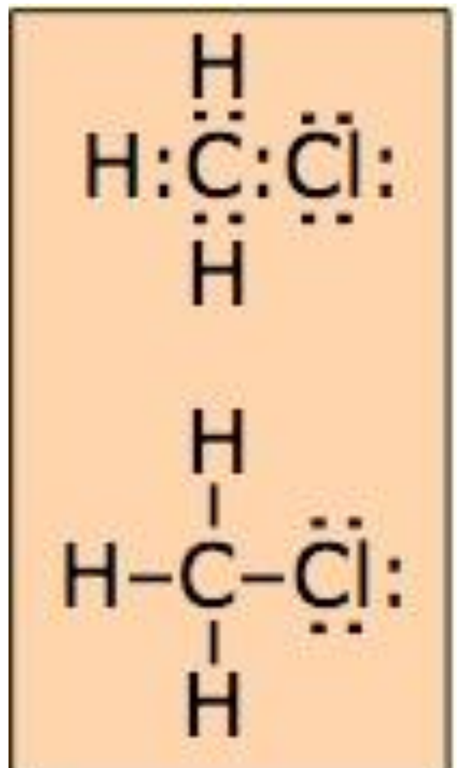
1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
1 **H** 2.20																	2 **He** no data
3 **Li** 0.98	4 **Be** 1.57											5 **B** 2.04	6 **C** 2.55	7 **N** 3.04	8 **O** 3.44	9 **F** 3.98	10 **Ne** no data
11 **Na** 0.93	12 **Mg** 1.31											13 **Al** 1.61	14 **Si** 1.90	15 **P** 2.19	16 **S** 2.58	17 **Cl** 3.16	18 **Ar** no data
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	36 **Kr** 3.00
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** 1.9	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	54 **Xe** 2.6
55 **Cs** 0.79	56 **Ba** 0.89	57-71 Lanthanides	72 **Hf** 1.3	73 **Ta** 1.5	74 **W** 2.36	75 **Re** 1.9	76 **Os** 2.2	77 **Ir** 2.20	78 **Pt** 2.28	79 **Au** 2.54	80 **Hg** 2.00	81 **Tl** 1.62	82 **Pb** 2.33	83 **Bi** 2.02	84 **Po** 2.0	85 **At** 2.2	86 **Rn** no data
87 **Fr** 0.7	88 **Ra** 0.89	89-103 Actinides															

***** Elements > 104 exist only for very short half-lifes and the data is unknown.*****

Lanthanides	57 **La** 1.10	58 **Ce** 1.12	59 **Pr** 1.13	60 **Nd** 1.14	61 **Pm** 1.13	62 **Sm** 1.17	63 **Eu** 1.2	64 **Gd** 1.2	65 **Tb** 1.2	66 **Dy** 1.22	67 **Ho** 1.23	68 **Er** 1.24	69 **Tm** 1.25	70 **Yb** 1.1	71 **Lu** 1.27
Actinides	89 **Ac** 1.1	90 **Th** 1.3	91 **Pa** 1.5	92 **U** 1.38	93 **Np** 1.36	94 **Pu** 1.28	95 **Am** 1.3	96 **Cm** 1.3	97 **Bk** 1.3	98 **Cf** 1.3	99 **Es** 1.3	100 **Fm** 1.3	101 **Md** 1.3	102 **No** 1.3	103 **Lr** no data

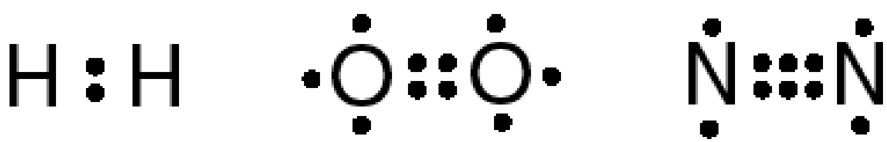
Drawing Lewis Structures of covalent compounds

Lewis structures can only be drawn for covalently bonded molecules (showing sharing of electrons)



Types of covalent bonds

Single bond	Double bond	Triple bond
H—H	O=O	N≡N



Triple bonds are stronger than double bonds; double bonds are stronger than single bonds (*it takes more energy to break a double bond than a single bond*)

Multiple bonds increase the electron density between 2 nuclei (the nuclei move closer and the bond length is shorter with triple bond being the shortest)

TABLE 7.5 Average Bond Lengths for some Single, Double and Triple Bonds

Bond	Bond Length (Å)	Bond	Bond Length (Å)
C—C	1.54	N—N	1.47
C=C	1.34	N=N	1.24
C≡C	1.20	N≡N	1.10
C—N	1.43	N—O	1.36
C=N	1.38	N=O	1.22
C≡N	1.16		
		O—O	1.48
C—O	1.43	O=O	1.21
C=O	1.23		
C≡O	1.13		

Rules for drawing Lewis dot structures for covalent compounds:

Step 1: Count total valence electrons in the compound.

Example: **SF_4:**

valence electrons: S=6

F=(7 x 4)

34 in total (17 electron pairs)

Rules for drawing Lewis dot structures for covalent compounds:

Step 2: Put LEAST electronegative element in the center (never hydrogen) and evenly space the other atoms around it. Connect central atom to each terminal atom with a single bond.

Rules for drawing Lewis dot structures for covalent compounds:

Step 3: Add remaining electrons, in pairs, to the terminal atoms first to satisfy the octet rule.

****Exceptions to the octet rule:**

- Hydrogen can only hold 2.
- Beryllium has a maximum of 4.
- Boron is stable with only 6.
- If the central atom is from periods 3-7, it can accommodate more than 8 electrons due to the presence of **d orbitals**

Rules for drawing Lewis dot structures for covalent compounds:

Step 4: Put any remaining pairs or single electrons around the central atom. These are called **lone pairs**.

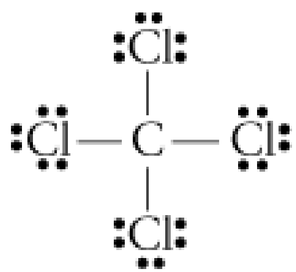
Step 5: If central atom does not yet have a full octet, convert a terminal lone pair to a multiple bond. (CNOPS)

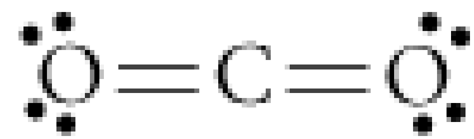
Draw Lewis structures for the following compounds:

1. **PH_3**
2. **CH_4**
3. **O_3 (resonance)**
4. **Sulfur ion**
5. **BF_3**
6. **N_2**

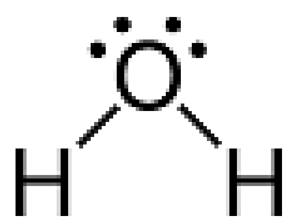
How can we tell if a molecule is polar or nonpolar?

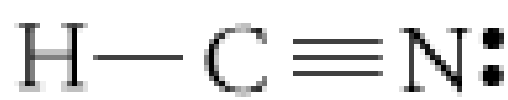
If the central atom has only one type of terminal atom and no **lone pairs** it is **nonpolar.**

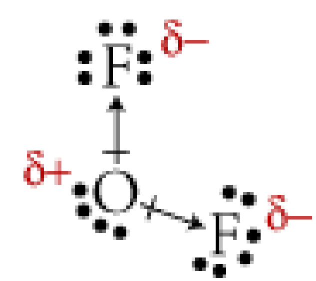




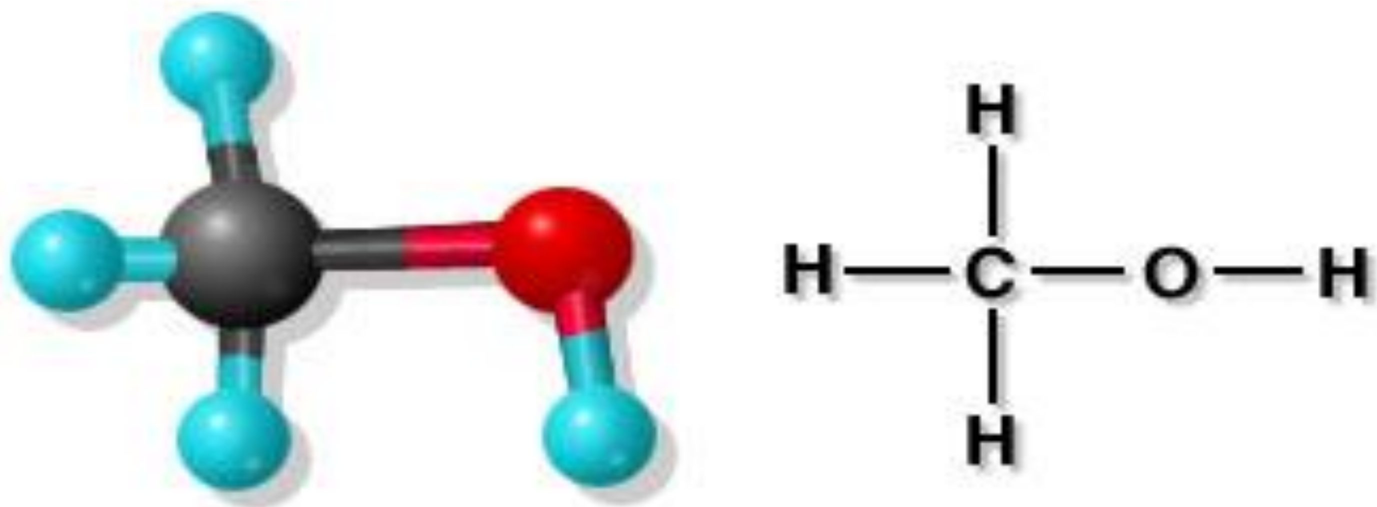
If the structure does not fit the description above, then the molecule is **polar.**







STRUCTURE OF MOLECULAR COMPOUNDS



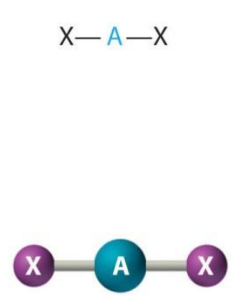
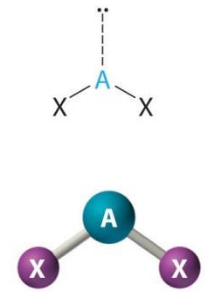
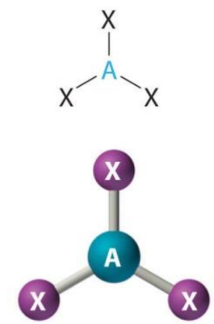
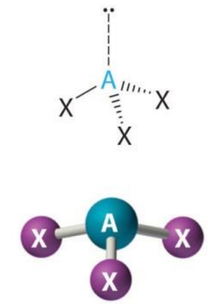
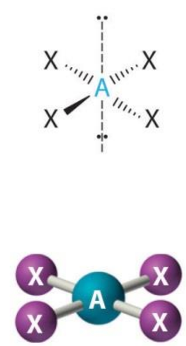
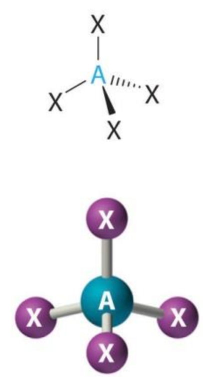
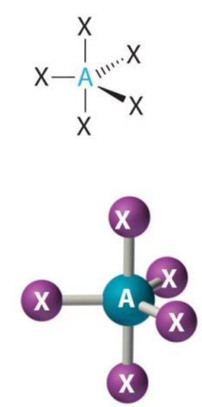
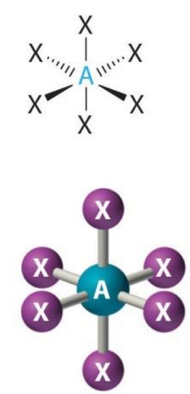
3D Representation

Lewis Structure

Valence Shell Electron Pair Repulsion theory (predicts 3-D shape of molecules or ions)

VSEPR theory

states electron pairs shared (in bonds) or unshared (lone pair) will adopt a geometry that maximizes the distance between the bonds or lone pairs.

AX_mE_n Notation	AX_2	AX_2E	AX_3	AX_3E
Geometry	Linear 	Bent (V-shaped) 	Trigonal planar 	Trigonal pyramidal 
Idealized Bond Angles	180°	<180°	120°	<120°
AX_mE_n Notation	AX_4E_2	AX_4	AX_5	AX_6
Geometry	Square planar 	Tetrahedral 	Trigonal bipyramidal 	Octahedral 
Idealized Bond Angles	90°	109.5°	90°, 120°	90°

Based on areas of electron density around the central atom

Bonded electrons on central atom (single, double and triple all count as 1 region of electron density)

OR

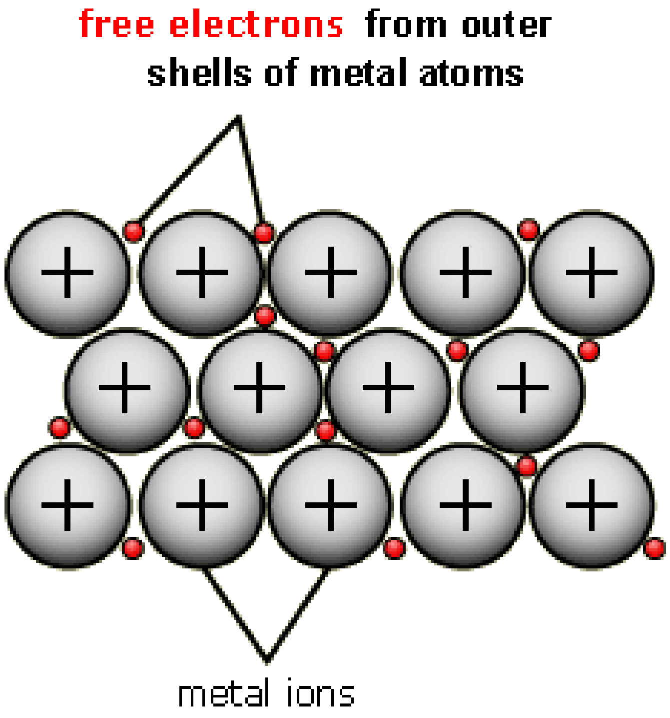
Unbonded electrons on central atom (lone pairs)

Phet Simulation

https://phet.colorado.edu/en/simulation/molecule-shapes

Metallic Bonds

- Occur between metal atoms.
- Bonding due to a "sea of electrons" – electrons that are not bound to one specific atom, they are *able to move around the substance* from atom to atom.

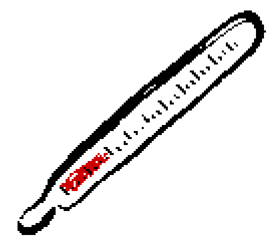


Physical Properties of Metals

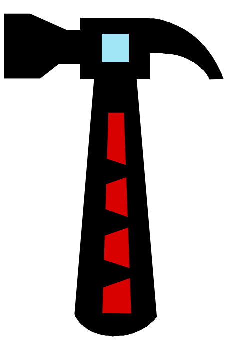
Good at conducting heat and electricity

Ductile

High melting and boiling points because of the strength of the metallic bond



Malleable

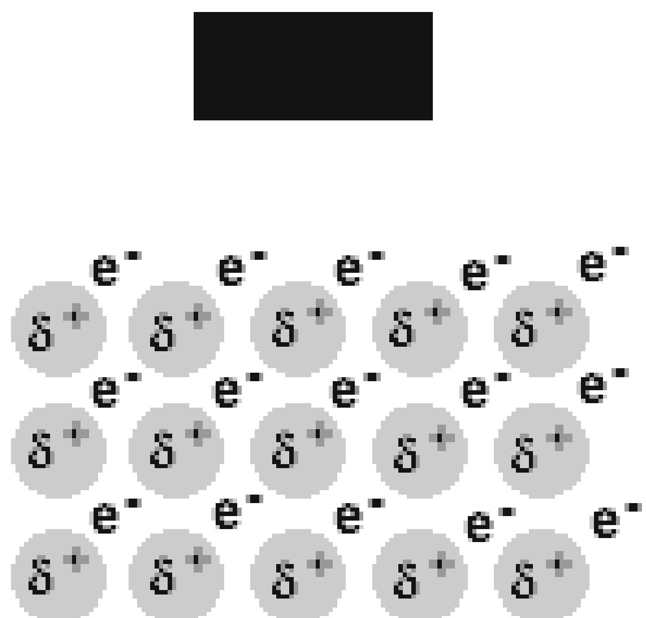


Metallic luster

**Properties are due to the free-floating (delocalized) electrons.

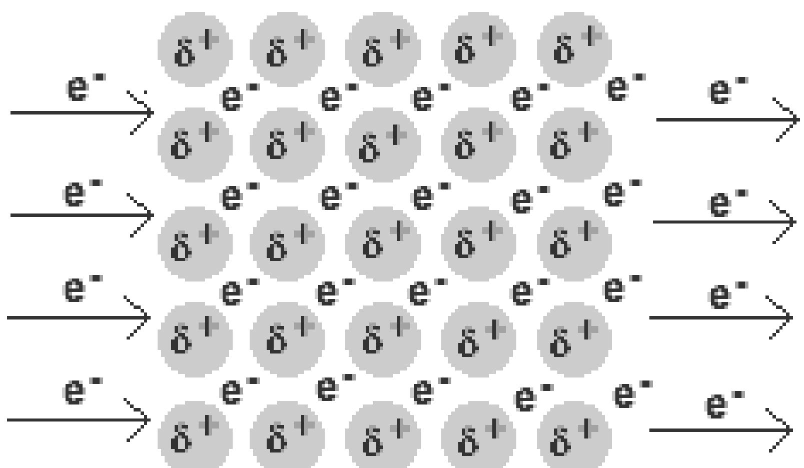
Metals are malleable and ductile

The 'sea' of electrons enable the metal atoms to roll over each other when a stress is applied.



Metals are good conductors of heat and electricity

The electrons are free to move and can act as charge carriers in the conduction of electricity or as energy conductors in the conduction of heat.



Ionic or Covalent?

1. H_2O
2. KCl
3. LiF
4. CO_2
5. $MgCl_2$
6. NO_2
7. H_2
8. Al_2O_3

Ionic, covalent or metallic?

1. A compound has a melting point of 750 °C and can conduct electricity when dissolved in water.
2. A compound contains N, O, C and H.
3. A compound has a high melting point, is ductile and malleable.
4. The compound crushes into pieces with a small amount of force, but will not melt in a Bunsen burner.
5. Liquid at room temperature and is insoluble in water.

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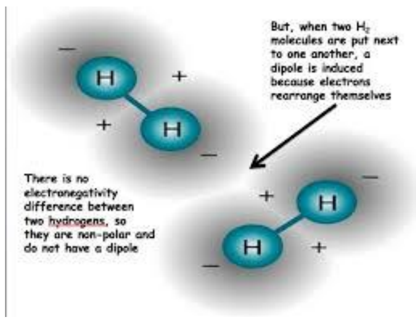
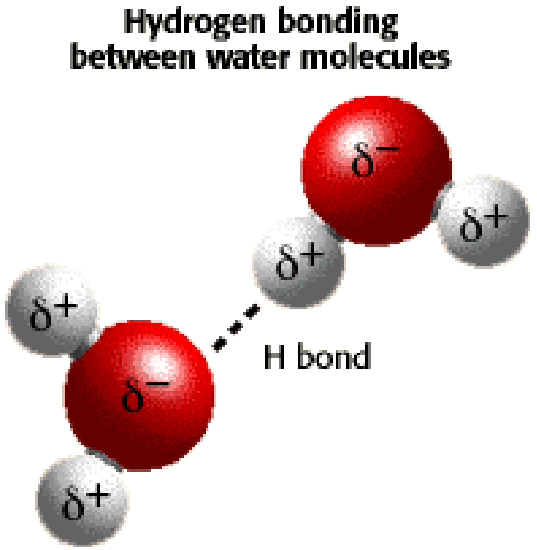
Types of bonds

Intermolecular Bonds **(hold two or more molecules together)**

- **Dipole-dipole, hydrogen bonds, London dispersion forces and ion-dipole forces**

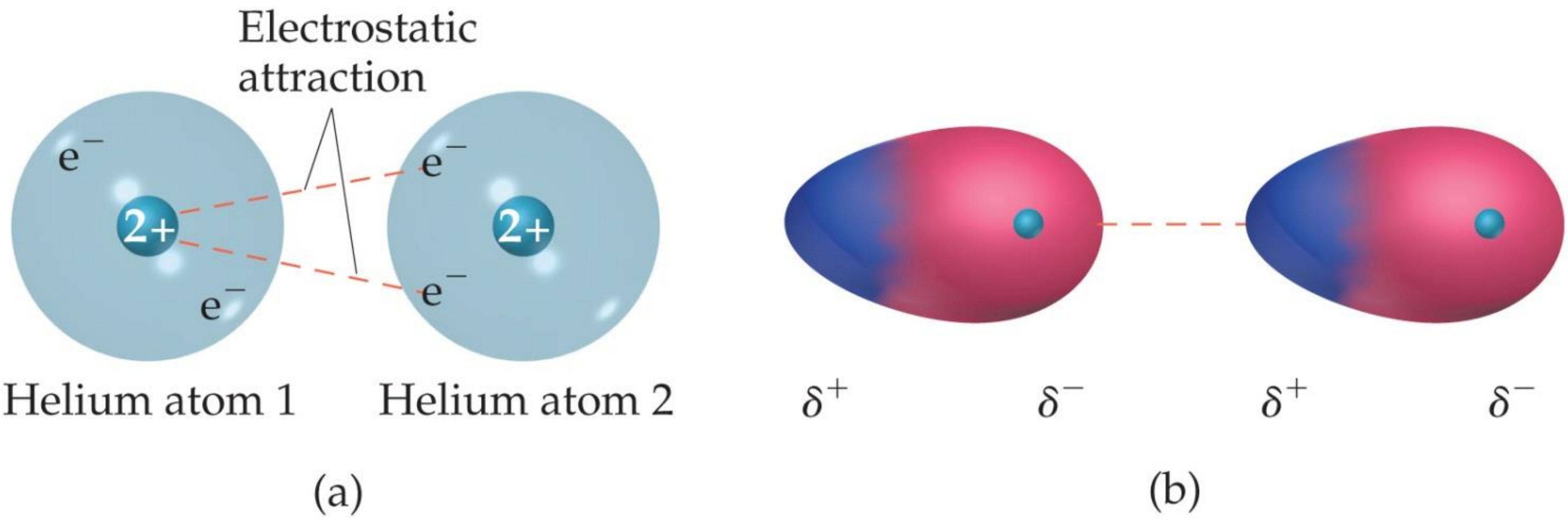




- **Weaker than ionic or covalent bonds**

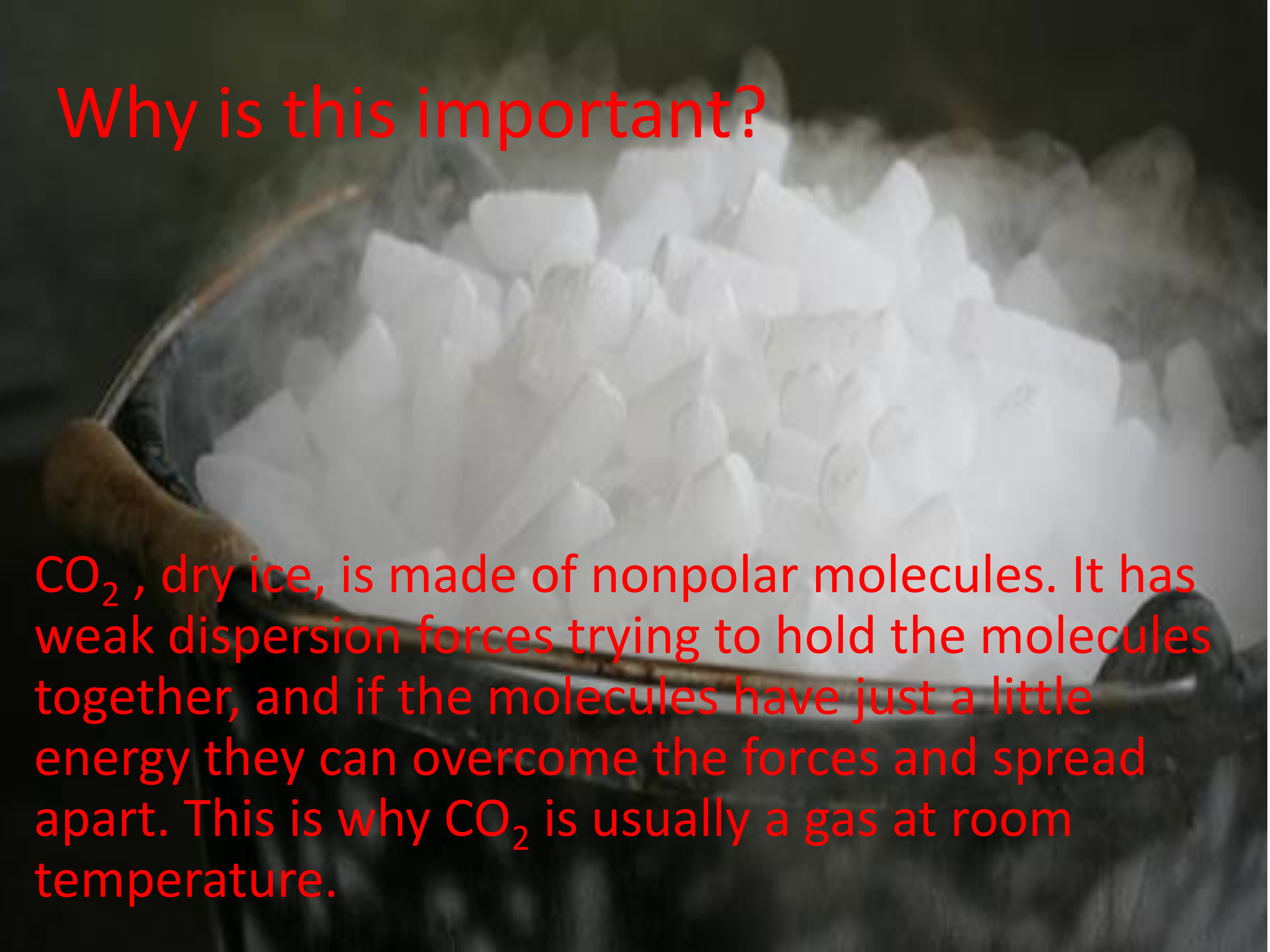
$$\overset{\delta+}{H}—\overset{\delta-}{Cl}\text{- - - - -}\overset{\delta+}{H}—\overset{\delta-}{Cl}$$

Intermolecular Forces (IMF's)



- Caused by an instantaneous dipole formation in which electron cloud becomes asymmetrical, and the molecules are slightly attracted to each other.

Larger nonpolar molecules have stronger dispersion forces…
bigger e^- cloud and more polarizable (squashy)

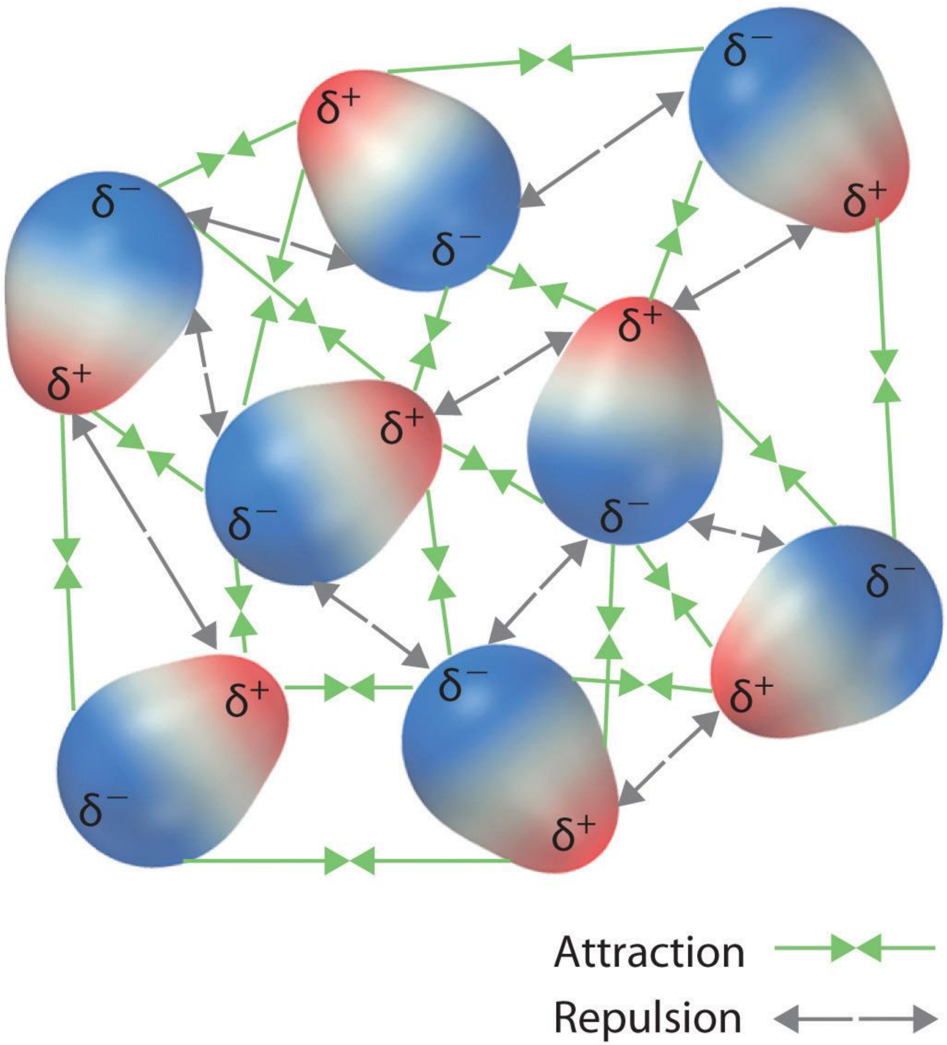


Why is this important?

CO_2 , dry ice, is made of nonpolar molecules. It has weak dispersion forces trying to hold the molecules together, and if the molecules have just a little energy they can overcome the forces and spread apart. This is why CO_2 is usually a gas at room temperature.

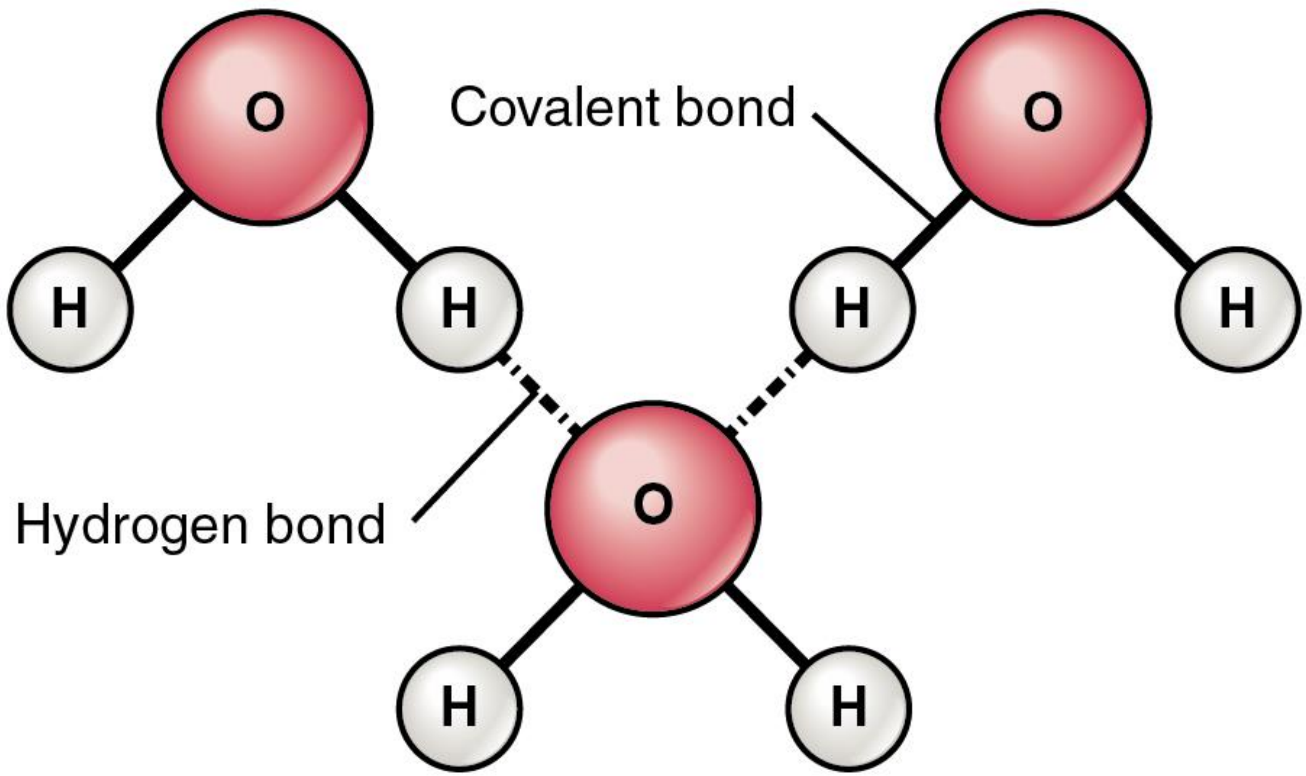
Dipole-Dipole

- These forces are found between molecules that are polar. These molecules have regions of + and – charge, and therefore experience attractive forces from one molecule to another.
- Stronger than dispersion forces because polarity is permanent.



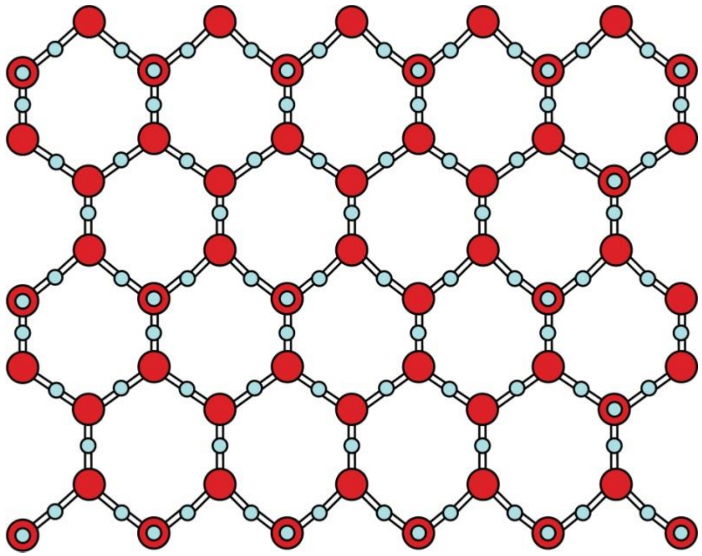
Hydrogen bonding

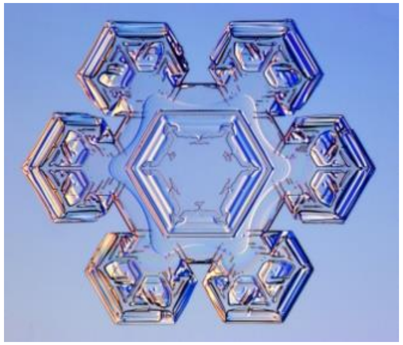
Ability of H atom already involved in a polar covalent bond with (N, O or F) to also be attracted to nearby molecules.

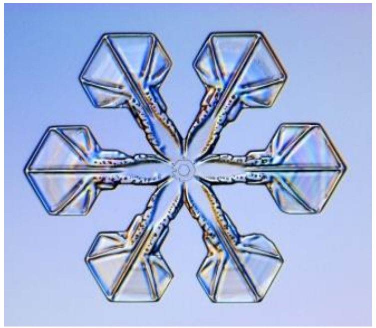


Hydrogen bonding is the reason water freezes into ice crystals of a certain repeating shape (hexagons)

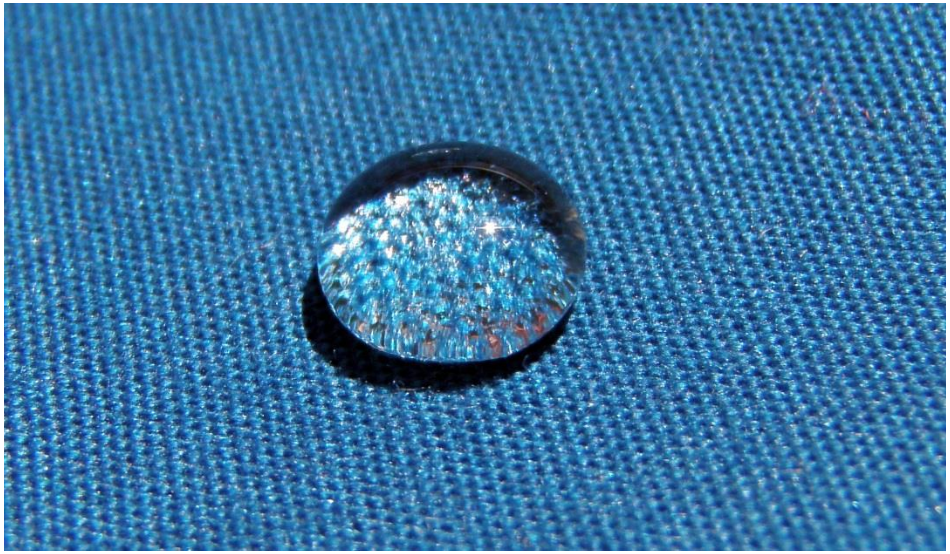
***Volume increases!**



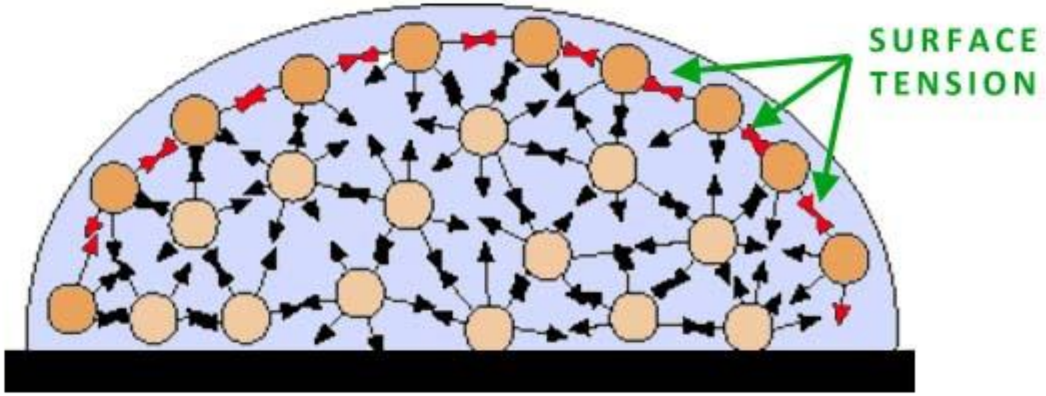






Unique Properties of Water due to Hydrogen Bonds



High Boiling Point

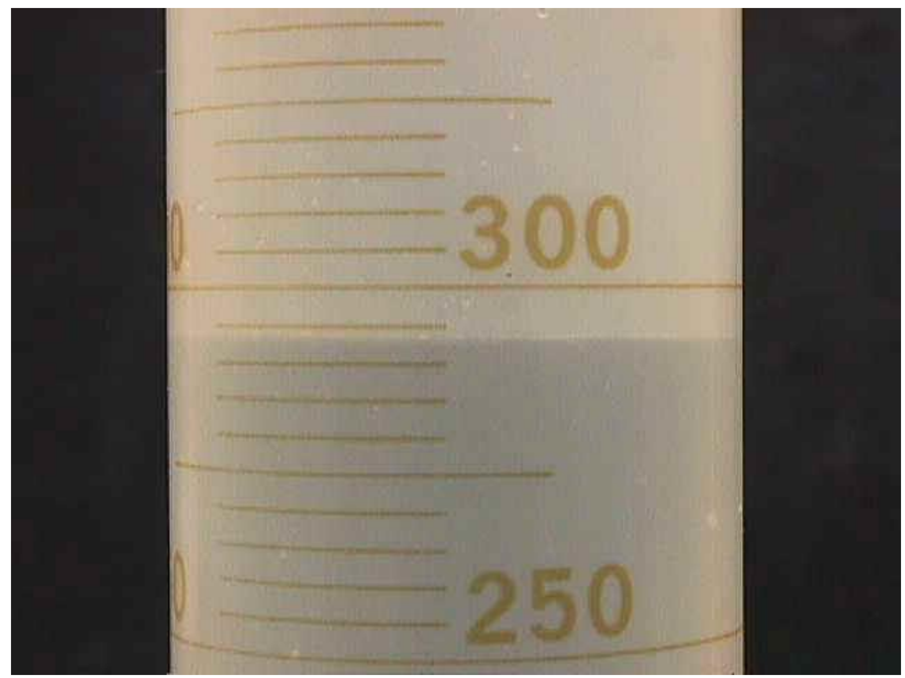


High Surface Tension (cohesion)

Adhesion

Formation of a Meniscus

Water molecules are attracted to the glass because glass molecules are also polar.

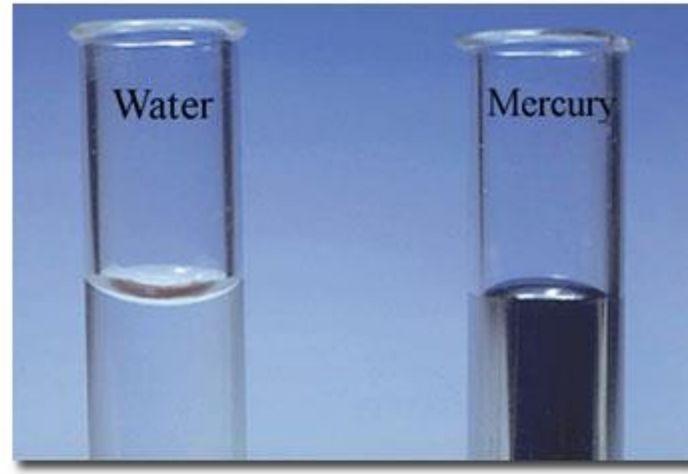


Surfactant

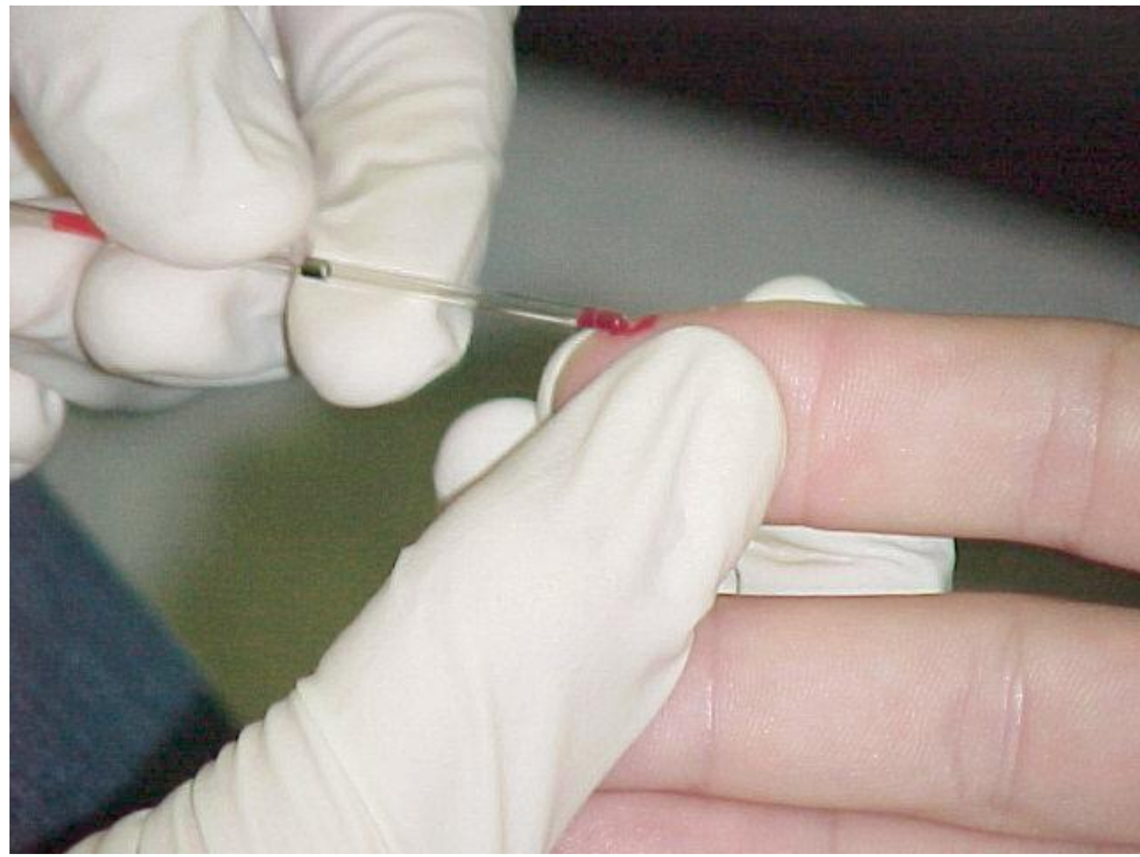
Surface Tension can be decreased by adding a **surfactant** – this type of substance interferes with hydrogen bonding.

Surfactants are used in soaps, detergents, paints, adhesives, inks…

Capillary Action

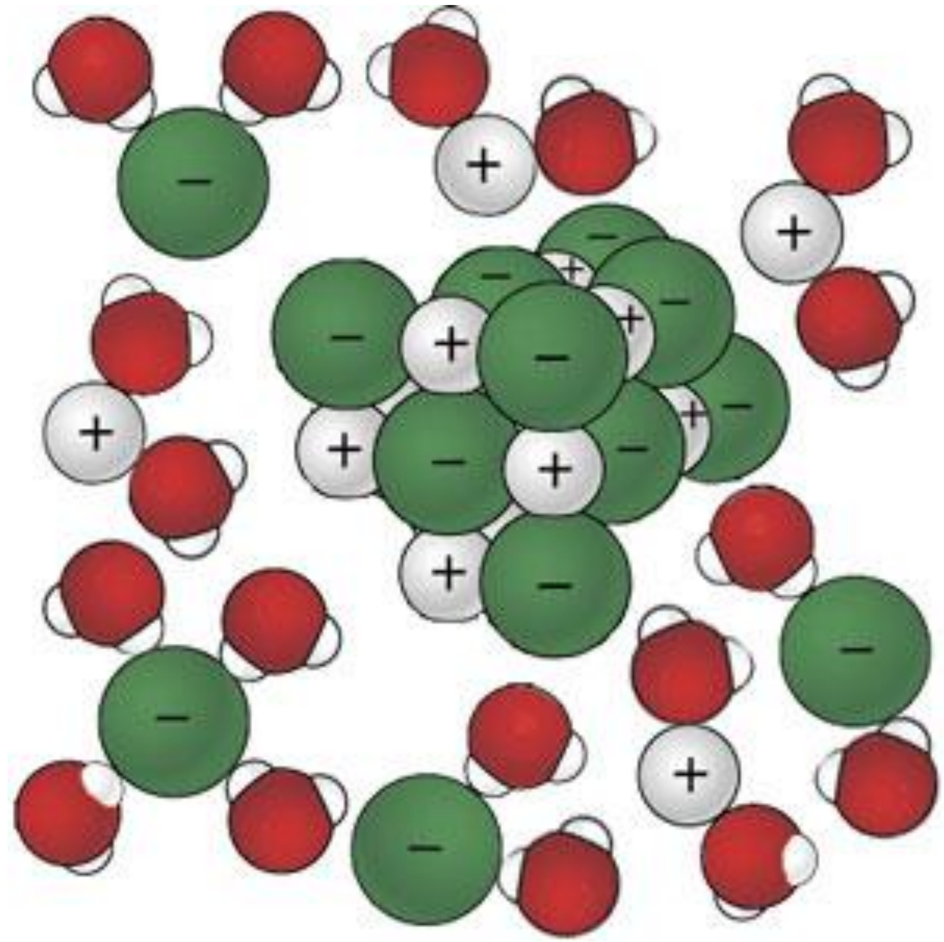


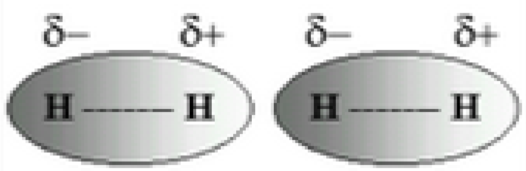
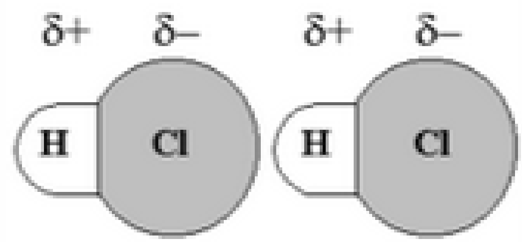
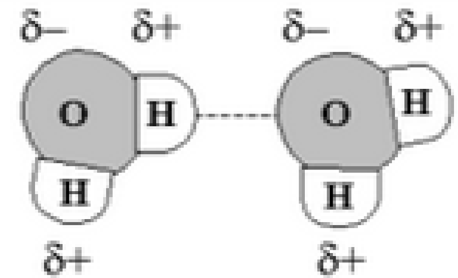
Water can be drawn up into a thin glass tube with no effort because of the attraction between the water and glass molecules (adhesive property of water).



Ion-dipole forces

Attraction that helps ionic compounds dissolve in a polar substance.



	LONDON DISPERSION FORCES	DIPOLE-DIPOLE FORCES	HYDROGEN BONDING
Definition	• Attraction between 2 instantaneous dipoles. • Asymmetrical electron distribution. • All atoms & molecules.	• Attraction between 2 permanent dipoles. • Polar molecules.	• Attraction between molecules with N-H, O-H, & F-H bonds. • Extremely polar bonds ⇒ very strong dipole-dipole force.
Diagram			
Relative Strength	• weakest	• medium strength	• strongest
Other Information	• Increase in strength as molar mass increases (more electrons).	• Stronger when molecules are closer together	• *Not* chemical bonding.