UNIT 3: The Octet Rule & Chemical Bonding

- Section 1: The Octet Rule
- Section 2: Ionic Bonding & Ionic Compounds
- Section 3: Covalent Bonding & Molecules
- Section 4: VSEPR Theory & Polarity
- Section 5: Intermolecular Forces
- Section 6: Periodic Table Trends
- Section 7: Nomenclature

UNIT 3 Synapsis

In the previous unit we looked at individual atoms and how they can change into other atoms. In this Unit we will look at how atoms combine to form molecules and compounds. Previously, we focused on the nucleus and now we will focus on the electrons because they are what interact to form chemical bonds.

In Section 1 we will look at one of the fundamental rules that govern chemical bond formation; the octet rule. The octet rule simplifies much of what we discussed regarding electron configuration. It generalizes how and why elements form bonds the way they do.

In Sections 2, 3, and 4 we will utilize the octet rule to explain the two major types of chemical bonding and the various ways the compounds they create can be modeled.

In Section 5 we will take what we learned about chemical bonds one step further and look at house it affects the forces between individual molecular units.

In Section 6 we will utilize the octet rule to get a better understanding of the periodic table and some of it’s trends.

In Section 7 we will look at the rules for naming substances.

Section 1: The Octet Rule

This part of the Unit is covered on page 192 in your textbook

Section 1: The Octet Rule / Objectives

- Students Will Be Able too…
  - …*identify* **metals, semi-metals, and non-metals** on the periodic table."
  - …*identify* where the **main group** is on the periodic table and *determine* the number of **valence electrons** main group elements have."
  - …*explain* the **octet rule**."
  - …*use* the octet rule to *determine* if an element wants to gain or lose electrons and how many electrons it wants to gain or lose."
  - …*determine* what noble gas an element will resemble and what **ion** it will form after obeying the octet rule."
  - …*explain* why metals form **cations** and non-metals form **anions**."

Metals, Semi-metals, Non-metals

- The periodic table can be divided into many different groups. For example you have already become familiar with the noble gases.
- First and foremost, the periodic table should be divided into **metals, semi-metals, and non-metals**.
- The semi-metals splits the non-metals and metals.
- There are six semi-metals. Remember these six semi-metals and identifying metals and non-metals is easy. This is because all the metals are to the left of the semi-metals and the non-metals are to the right (except Hydrogen).
- Remember: “Bacon So Good At All Times” for the six semi-metals (Boron, Silicon, Germanium, Arsenic, Antimony, Tellurium)

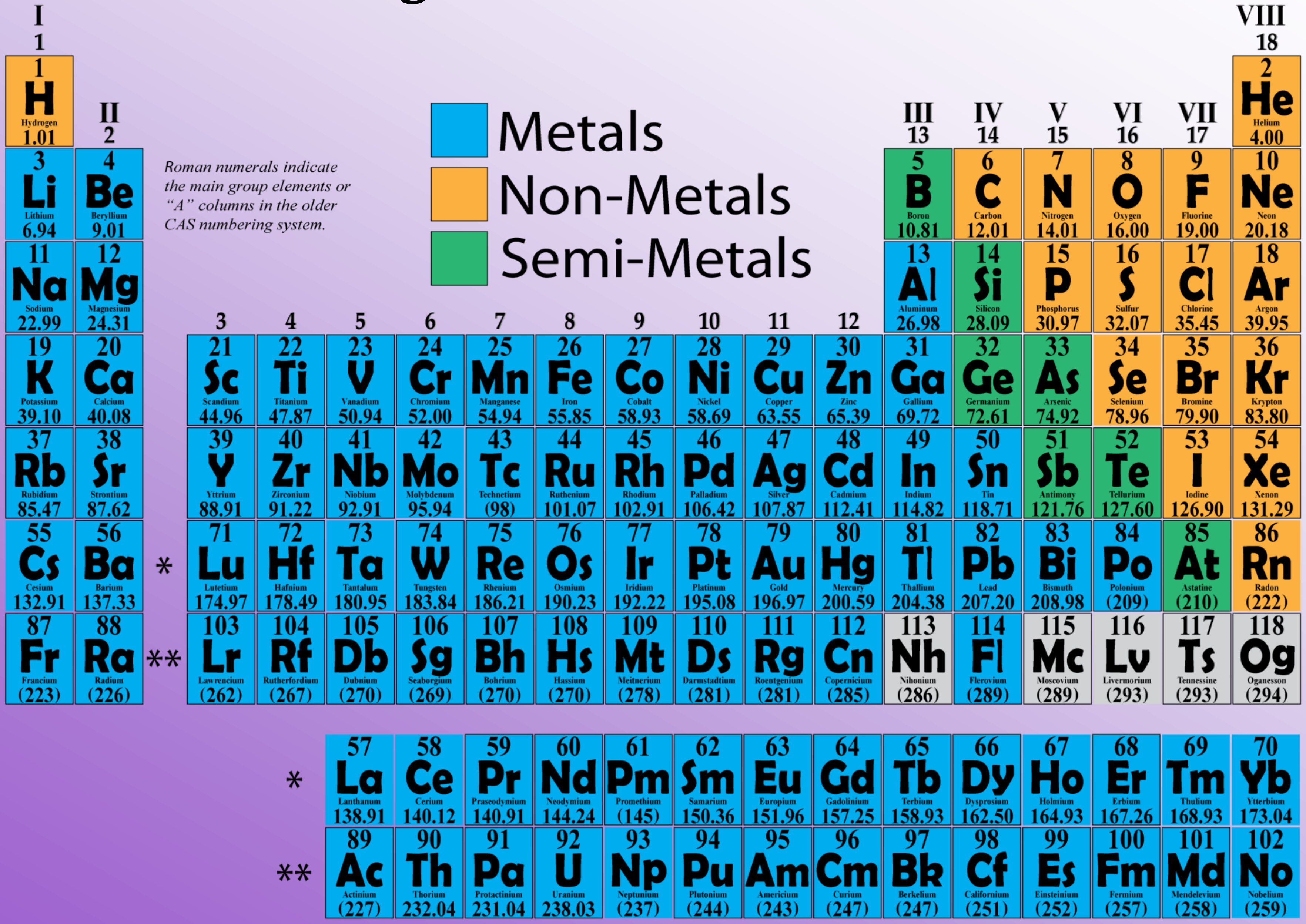
Periodic Table Showing Metals, Semi-metals, & Non-metals

🟦 Metals

🟧 Non-Metals

🟩 Semi-Metals

Roman numerals indicate the main group elements or "A" columns in the older CAS numbering system.



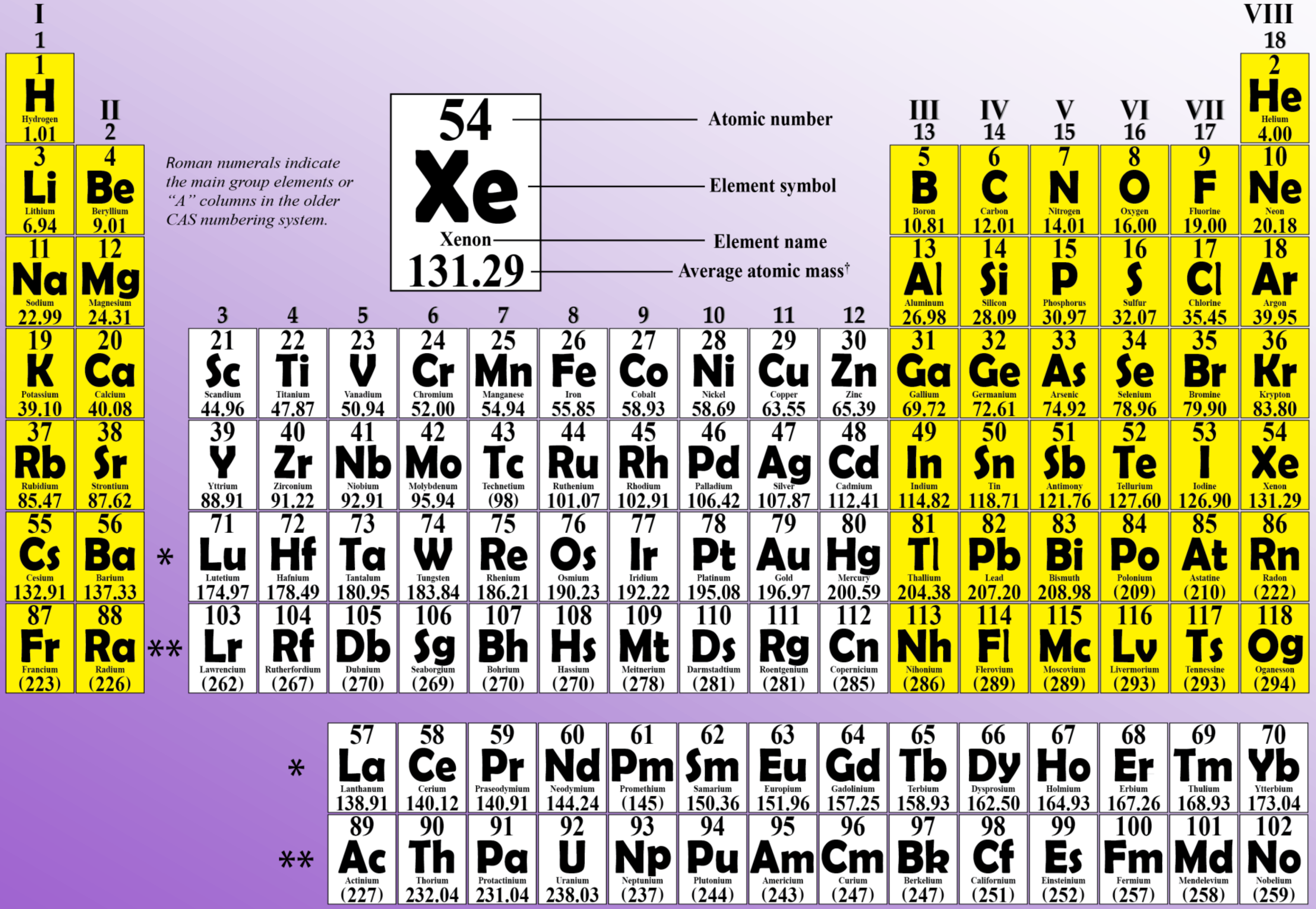
I 1	II 2		3	4	5	6	7	8	9	10	11	12	III 13	IV 14	V 15	VI 16	VII 17	VIII 18
1 **H** Hydrogen 1.01																		2 **He** Helium 4.00
3 **Li** Lithium 6.94	4 **Be** Beryllium 9.01												5 **B** Boron 10.81	6 **C** Carbon 12.01	7 **N** Nitrogen 14.01	8 **O** Oxygen 16.00	9 **F** Fluorine 19.00	10 **Ne** Neon 20.18
11 **Na** Sodium 22.99	12 **Mg** Magnesium 24.31												13 **Al** Aluminum 26.98	14 **Si** Silicon 28.09	15 **P** Phosphorus 30.97	16 **S** Sulfur 32.07	17 **Cl** Chlorine 35.45	18 **Ar** Argon 39.95
19 **K** Potassium 39.10	20 **Ca** Calcium 40.08		21 **Sc** Scandium 44.96	22 **Ti** Titanium 47.87	23 **V** Vanadium 50.94	24 **Cr** Chromium 52.00	25 **Mn** Manganese 54.94	26 **Fe** Iron 55.85	27 **Co** Cobalt 58.93	28 **Ni** Nickel 58.69	29 **Cu** Copper 63.55	30 **Zn** Zinc 65.39	31 **Ga** Gallium 69.72	32 **Ge** Germanium 72.61	33 **As** Arsenic 74.92	34 **Se** Selenium 78.96	35 **Br** Bromine 79.90	36 **Kr** Krypton 83.80
37 **Rb** Rubidium 85.47	38 **Sr** Strontium 87.62		39 **Y** Yttrium 88.91	40 **Zr** Zirconium 91.22	41 **Nb** Niobium 92.91	42 **Mo** Molybdenum 95.94	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.91	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.87	48 **Cd** Cadmium 112.41	49 **In** Indium 114.82	50 **Sn** Tin 118.71	51 **Sb** Antimony 121.76	52 **Te** Tellurium 127.60	53 **I** Iodine 126.90	54 **Xe** Xenon 131.29
55 **Cs** Cesium 132.91	56 **Ba** Barium 137.33	*	71 **Lu** Lutetium 174.97	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.95	74 **W** Tungsten 183.84	75 **Re** Rhenium 186.21	76 **Os** Osmium 190.23	77 **Ir** Iridium 192.22	78 **Pt** Platinum 195.08	79 **Au** Gold 196.97	80 **Hg** Mercury 200.59	81 **Tl** Thallium 204.38	82 **Pb** Lead 207.20	83 **Bi** Bismuth 208.98	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
87 **Fr** Francium (223)	88 **Ra** Radium (226)	**	103 **Lr** Lawrencium (262)	104 **Rf** Rutherfordium (267)	105 **Db** Dubnium (270)	106 **Sg** Seaborgium (269)	107 **Bh** Bohrium (270)	108 **Hs** Hassium (270)	109 **Mt** Meitnerium (278)	110 **Ds** Darmstadtium (281)	111 **Rg** Roentgenium (281)	112 **Cn** Copernicium (285)	113 **Nh** Nihonium (286)	114 **Fl** Flerovium (289)	115 **Mc** Moscovium (289)	116 **Lv** Livermorium (293)	117 **Ts** Tennessine (293)	118 **Og** Oganesson (294)
				*	57 **La** Lanthanum 138.91	58 **Ce** Cerium 140.12	59 **Pr** Praseodymium 140.91	60 **Nd** Neodymium 144.24	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.96	64 **Gd** Gadolinium 157.25	65 **Tb** Terbium 158.93	66 **Dy** Dysprosium 162.50	67 **Ho** Holmium 164.93	68 **Er** Erbium 167.26	69 **Tm** Thulium 168.93	70 **Yb** Ytterbium 173.04
				**	89 **Ac** Actinium (227)	90 **Th** Thorium 232.04	91 **Pa** Protactinium 231.04	92 **U** Uranium 238.03	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)

†Elements that have their average atomic mass in parentheses have no stable isotopes. The number given is the mass number for the longest lived isotope according to the 2013 IUPAC technical report.

The Main Group

- Another important way to divide the periodic table is into the **Main Group Elements**.
- The main group is usually indicated using roman numerals. This is how it is done on the periodic table you get to use on your assessments.
- The main group includes all the semi-metals & non-metals, as well as several metals.
- The main group contains the elements that make up 99% of living things, 99% of all the matter in the universe, plastics, and many, many other things.
- It is where we will focus the vast majority of our attention in Chemistry A & Chemistry B.

Periodic Table Showing the main Group in Yellow



Roman numerals indicate the main group elements or "A" columns in the older CAS numbering system.

54 — Atomic number
Xe — Element symbol
Xenon — Element name
131.29 — Average atomic mass†

I 1	II 2		3	4	5	6	7	8	9	10	11	12	III 13	IV 14	V 15	VI 16	VII 17	VIII 18
1 H Hydrogen 1.01																		2 He Helium 4.00
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19 K Potassium 39.10	20 Ca Calcium 40.08		21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62		39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
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*	57 La Lanthanum 138.91	58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04
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†Elements that have their average atomic mass in parentheses have no stable isotopes. The number given is the mass number for the longest lived isotope according to the 2013 IUPAC technical report.

Valence Electrons

- **Valence electrons** can be looked at in a couple of different ways:
    - The outermost electrons available for forming chemical bonds
    - The electrons in the s & p orbitals. Plenty of main group elements have electrons in d & f subshells, but recall that d & f subshells are always a shell (or energy level) below the electrons in the s & p subshells for a given element.
- You don't have to write out the electron configuration to figure out the number of valence electrons a main group element has though! It is simply the column number the main group element is in (so long as you are looking at the roman numerals from the old-school column numbering system)
- For example
    - Hydrogen has 1 valence electron
    - Neon has 8 valence electrons
    - Carbon has 4 valence electrons
    - Bromine has 7 valence electrons
- The concept of valence electrons doesn't apply well to elements outside the main group but recall it's only the main group we are focused on.

Practice Problems: Indentifying element groups & Valence Electrons

Directions: (1) Determine if an element is a metal, semi-metal, or non-metals. (2) Determine if it is part of the main group or not. (3) If it is part of the main group, determine how many valence electrons it has.

1) Al
2) P
3) Na
4) Mn
5) Ca
6) Li
7) Fe
8) O
9) N
10) Cl
11) Fe

The Octet Rule

- All of the elements on the Periodic table "want" to be like the noble gases because they have completely filled s & p orbitals. Having completely filled s & p orbitals makes elements stable. In other words it makes them happy.
- Main group elements react and form chemical bonds in a way that brings them closer to having the same electron configuration as the noble gas they are closest too.
- What this means is that they will *gain, lose, or share their valence electrons until they have 8 valence electrons or none at all*.
- Thus, during bond formation elements in columns 1,2,3 will lose all their valence electrons until they have 0.
- Likewise, during bond formation elements in columns 5,6,7 will gain valence electrons to get to 8.
- Elements in column 8 already have 8 valence electrons so they don't form chemical bonds or react. Elements in column 4 we will discuss in section 3.
- Once an element gets to 8 or 0 valence electrons they will have the same electron configuration as a noble gas and this makes them more stable. Since the noble gases are in column 8, this rule of chemical bonding is called **The Octet Rule.**

The Octet Rule Examples

- Phosphorus…
  - Has 15 total electrons
  - Has 5 valence electrons
  - Wants to gain 3 valance electrons to get to 8
  - When it gains 3 valance electrons it will have 18 total electrons and be like Argon
- Selenium…
  - Has 34 total electrons
  - Has 6 valence electrons
  - Wants to gain 2 valance electrons to get to 8
  - When it gains 2 valence electrons it will have 36 total electrons and be like Krypton
- Aluminum…
  - Has 13 total electrons
  - Has 3 valence electrons
  - Wants to lose 3 valence electrons to get to 0
  - When it loses 3 valence electron it will have 10 total electrons and be like Neon

Practice Problems: The Octet Rule

Directions: (1) Determine if an element is a metal, semi-metal, or non-metals. (2) Write out the total number of electrons. (3) Write out the number of valence electrons. (4) determine if the element wants to gain or lose electrons based on the octet rule and how many electrons it wants to gain or lose (5) Write down what noble gas it will resemble after obeying the octet rule

1) Br

2) K

3) O

4) Ca

5) Sr

6) N

7) Ba

8) Na

Ion Formation

- As we learned in the previous unit atoms form **ions** when they gain or lose electrons.
- During chemical bond formation and chemical reactions, electrons get exchanged or shared but the nucleus remains completely unchanged. This means the number of protons remains unchanged.
- So when gaining or losing electrons based on the octet rule atoms form ions.
- Elements in columns 1,2, and 3 lose electrons and then form positive ions called **cations**
  - Elements in column 1 form +1 cations
  - Elements in column 2 form +2 cations
  - Elements in column 3 form +3 cations
- Elements in columns 5,6, and 7 gain electrons and thus form negative ions called **anions**
  - Elements in column 5 form -3 anions
  - Elements in column 6 form -2 anions
  - Elements in column 7 form -1 anions
- You might have noticed metals form cations and non-metals form anions (except Hydrogen).

Video Time !!!

Fuse School Video: “What are Ions?”

Section 1 Additional Resources & Links…

- Science Post Video: “The Octet Rule – Clear and Simple”
- Tyler Dewitt’s Video: “What’s an Ion?”

Section 2: Ionic Bonding & Ionic Compounds

This part of the Unit is covered on page 205-217 in your textbook

Section 2: Ionic Bonding & Ionic Compounds / Objectives

- Students Will Be Able too…
- …*recall* that **Ionic bonding** occurs when electrons are exchanged between atoms.
- …*write* the chemical formulas for **ionic compounds** when presented with two main group elements or a cation and an anion (including polyatomic ions.)"
- ...*recognize* the metals outside the main group form positive ions of varying charges and most are multivalent."
- …*define* and *identify* **polyatomic ions**."
- …*identify* the **parent cation** & **parent anion** in an ionic compound."
- …*give* some of the basic properties of ionic compounds and *explain* those properties by discussing their **crystal lattice** structure."

Ionic Bonding

- Ionic bonding is basically what you get when a metal that wants to lose it's electrons get's together with a non-metal that wants to gain electrons.
- When this happens one or more electron is transferred between the two atoms and an ionic bond is formed.
- The resulting substance is called an ionic compound.
- Some of the simplest ionic compounds you are familiar with include:
  - Table Salt: $NaCl$ (a Na^+ ion & a Cl^- ion)
  - Sodium Fluoride: NaF (a Na^+ ion & a F^- ion)
  - Limestone: CaO (a Ca^{2+} ion and a O^{2-} ion)

Ionic Bonding

- In the previous examples of ionic compounds the ratio was always 1:1. 1 anion for every 1 cation. The ratio is not always 1 to 1 though. Sometimes the ratio is 3:1 or 2:1 or 2:3. There are actually several possible ratios.
- Some examples of those compounds include:
  - Lithium Oxide: Li_2O (2 Li^+ ions for every 1 O^{2-} ion)
  - Calcium Chloride: $CaCl_2$ (1 Ca^{2+} ion for every 2 Cl^- ions)
  - Aluminum oxide: Al_2O_3 (2 Al^{3+} ions for every 3 O^{2-} ions)
- Any cation can get to together with any anion to make an ionic compound. The only rule is the charges balance each other out so often times you will need more than 1 ion. You never have charges left over.

Video Time !!!

Fuse School Video: “What are Ionic Bonds?”

Practice Problems: Writing Formulas for Ionic Compounds

Directions: (1) Determine what ion the elements below will form. (2) write the resulting formula when these elements combine to form a compound.

1) Al & P

2) Na & O

3) Mg & N

4) Ca & Br

5) K & S

6) Rb & N

7) Sr & F

8) Li & Cl

Practice Problems: Writing Formulas for Ionic Compounds

Directions: (1) Determine what ion the elements below will form. (2) write the resulting formula when these elements combine to form a compound.

1) Ca & P

2) Na & N

3) Mg & Se

4) Ca & O

5) In & I

6) Al & S

7) Li & Br

8) Al & Cl

Ionic Compounds made from Metals Outside the Main Group

- As mentioned earlier we are mostly concerned with the main group elements as we move forward with our exciting exploration of chemistry.
- We do need to take a second to mention some of the elements outside the main group when talking about ionic compounds though:
  - All the elements outside the main group are metals
  - Like metals in the main group they always form positive ions
  - The ions they form cannot be easily prediction and you will not be expected to know what ion non-main group elements form.
  - Interestingly, most of the metals outside the main group can form a couple of different ions depending on the situation. For example Iron is known to form Fe^{2+} , Fe^{3+} and occasionally even Fe^{6+}. Elements capable of doing this are called **multivalent**, and almost all the transition metals are multivalent.
- Besides remembering the things mentioned above, the only other objective for elements outside the main group is that you can write formulas for ionic compounds if you are given the ion.

Practice Problems: Writing Formulas for Ionic Compounds

Directions: write the resulting formula from the ionic compound formed from the pair of cations and anions below.

1) V^{5+} & Cl^{-}

2) Fe^{3+} & O^{2-}

3) Cu^{+} & N^{3-}

4) Ni^{2+} & Br^{-}

5) Tc^{7+} & S^{2-}

6) Ti^{4+} & N^{3-}

7) Mn^{2+} & F^{-}

8) Co^{3+} & Cl^{-}

Polyatomic Ions

- Some ions are not just a single atom but a group of atoms. These ions are almost always anions. They are called **polyatomic ions.** You will find a table of these ion on the back of your periodic table.
- Polyatomic ions follow the same rules as mono-atomic ions when it comes to forming ionic compounds, but they can be a bit trickier to write formulas for. This is because you sometimes do - and sometimes do not - need to use parenthesis when writing the formula for the resulting compound.
- The rule is simple: If you have more than 1 polyatomic ion in an ionic compound, you have to put it in parenthesis and put the number of polyatomic ions outside the parenthesis.
- Examples:
- Al^{3+} & CN^- forms…
  - $Al(CN)_3$
- Al^{3+} & CO_3^{2-} forms…
  - $Al_2(CO_3)_3$
- Na^+ & SO_4^{2-} forms…
  - Na_2SO_4

Table of Polyatomic Ions

Polyatomic Ions			
-1		**-2 and -3**	
Acetate	$C_2H_3O_2^-$	Carbonate	CO_3^{2-}
Bicarbonate	HCO_3^-	Chromate	CrO_4^{2-}
Chlorite	ClO_2^-	Dichromate	$Cr_2O_7^{2-}$
Chlorate	ClO_3^-	Oxalate	$C_2O_4^{2-}$
Cyanide	CN^-	Sulfite	SO_3^{2-}
Hydroxide	OH^-	Sulfate	SO_4^{2-}
Hypochlorite	ClO^-	Phosphate	PO_4^{3-}
Iodate	IO_3^-	Phosphite	PO_3^{3-}
Nitrite	NO_2^-		
Nitrate	NO_3^-		
Perchlorate	ClO_4^-		
Permanganate	MnO_4^-		
Thiocyanate	SCN^-		

Practice Problems: Polyatomic Ions & Ionic Compounds

Directions: write the resulting formula for the ionic compound formed from the pairs of cations and anions below.

1) Cs^+ & OH^-
2) Ca^{2+} & CrO_4^{3-}
3) Al^{3+} & PO_4^{3-}
4) Fe^{2+} & NO_3^-
5) Fe^{3+} & SCN^-
6) Mg^{2+} & PO_3^{3-}
7) Ni^+ & CO_3^{2-}
8) Cr^{2+} & HCO_3^-
9) K^+ & NO_2^-

Practice Problems: Polyatomic Ions & Ionic Compounds

Directions: write the resulting formula for the ionic compound formed from the pairs of cations and anions below.

1) Ca^{2+} & OH^{-}
2) Ca^{2+} & PO_4^{3-}
3) V^{5+} & SO_4^{2-}
4) Ba^{2+} & ClO_3^{-}
5) Ti^{4+} & OH^{-}
6) Sc^{+} & ClO^{-}
7) K^{+} & CO_3^{2-}
8) Na^{+} & HCO_3^{-}
9) Rb^{+} & NO_2^{-}

Identifying Parent Ions

- Now that you know how cations and anions combine to form ionic compounds, lets you see if you can do the opposite and pull them apart.
- As you become more familiar with these ions, you may find yourself knowing the cation and anion just by looking at the formula.
- Until then, follow these steps:
    1) Identify the metal first. The metal is always a mono-atomic ion
    2) If the metal is main group, you should know the charge immediately. If it's not main group element, move on to the anion.
    3) Identify the anion next. If it is a mono-atomic anion, it will be apart of the main group and you should know the charge immediately.
    4) If the anion is polyatomic, find it on the back of your periodic table
    5) Based on the overall charge of the anion go back to the multivalent cation and see if you can figure out it's charge.
- When identifying the cation and anion in an ionic compound you may sometimes hear them referred to as "**parent cation**" and "**parent anion**"

Practice Problems: Decoding Ionic Compounds

Directions: (1) Identify the parent cation & parent anion in the compound below. (2) Identify how many of the parent cation and parent anion there are.

1) Cs_2O
2) Rb_3N
3) $Al(NO_3)_3$
4) MgS
5) K_2CO_3
6) $MgCl_2$
7) $Mg_3(PO_4)_2$
8) LiF
9) $FeCl_3$

Practice Problems: Decoding Ionic Compounds

Directions: (1) Identify the parent cation & parent anion in the compound below. (2) Identify how many of the parent cation and parent anion there are.

1) Na_2SO_4
2) $Ti(ClO_3)_4$
3) $SrCO_3$
4) MgF_2
5) $Al_2(SO_4)_3$
6) Ni_3PO_4
7) $Cu(NO_3)_2$
8) $CuNO_3$
9) $CaCl_2$

The Crystal Lattice Structure of Ionic Compounds

- The formula for an ionic compound shows the parent cation, the parent anion and the ratio between the two. However, what you need to understand is that ionic compounds do not exist in isolation as a group of just a few atoms or "units".
- The ions are actually "packed together" in a repeating pattern of ions we call a **crystal lattice**. They have a very ordered arrangement.
- The main reason ionic compounds have the crystal lattice structure is because while the charges in ionic compounds *balance out*, they do not *cancel out*.
- For example Table Salt is an overall neutral compound, but each sodium atom in a grain of table salt maintains a +1 charge, while each chlorine atom maintains a -1 charge.
- Thus, each chlorine ion is surrounded by 6 sodium ions and each sodium ion is surrounded by 6 chlorine ions. This pattern of ions repeats itself over and over again for millions or more ions (depending on how big your crystal is).

Video Time !!!

Fuse School Video: “Ionic Compounds and Their Properties”
TEDed Video: “How Atoms Bond”
Cosmos Clip Video: “Democritus & Bonding”

Model of the Crystal Lattice Structure in Table Salt

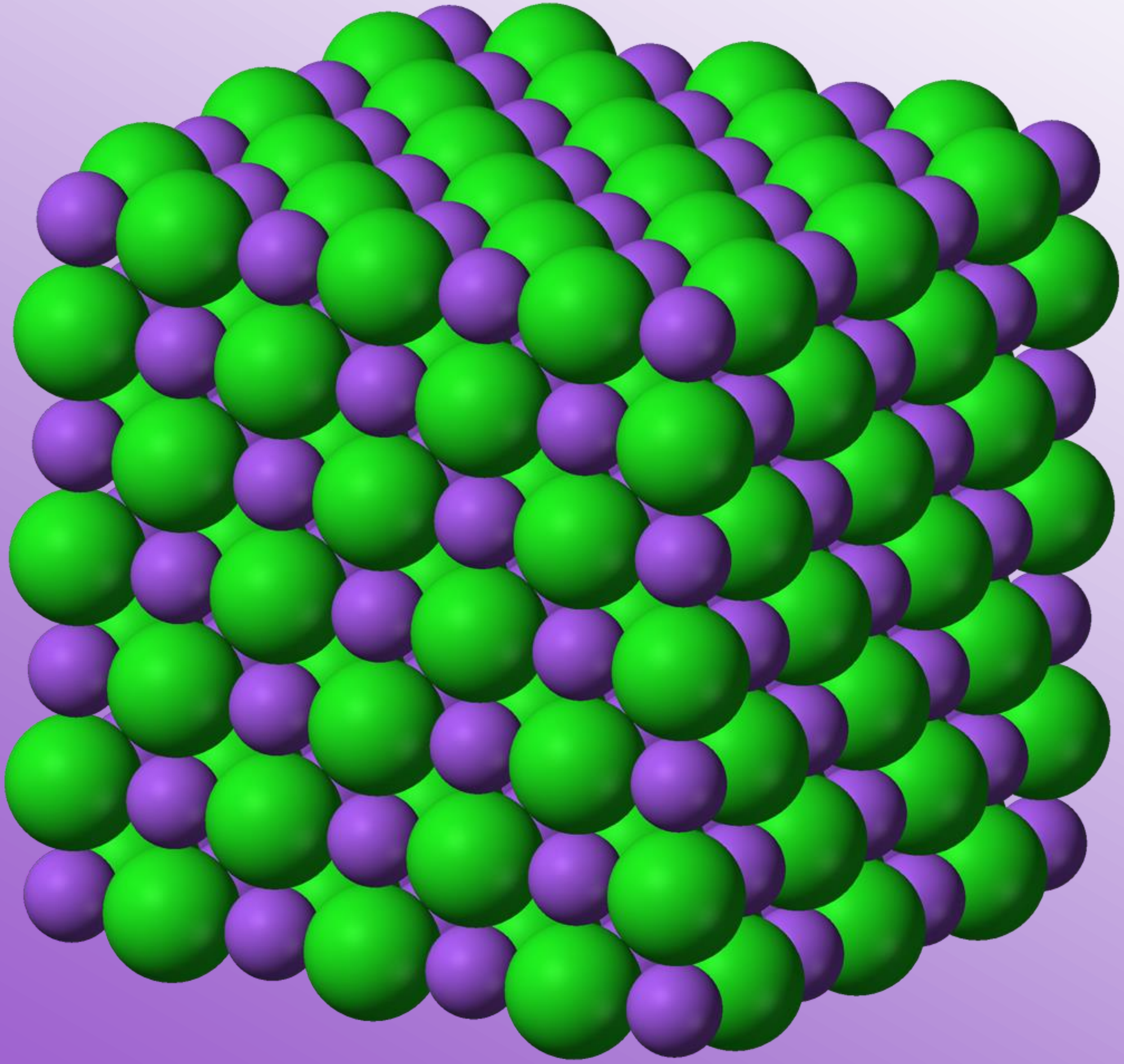


Image Credit: wikipedia

Model of the Crystal Lattice Structure in Calcium Oxide

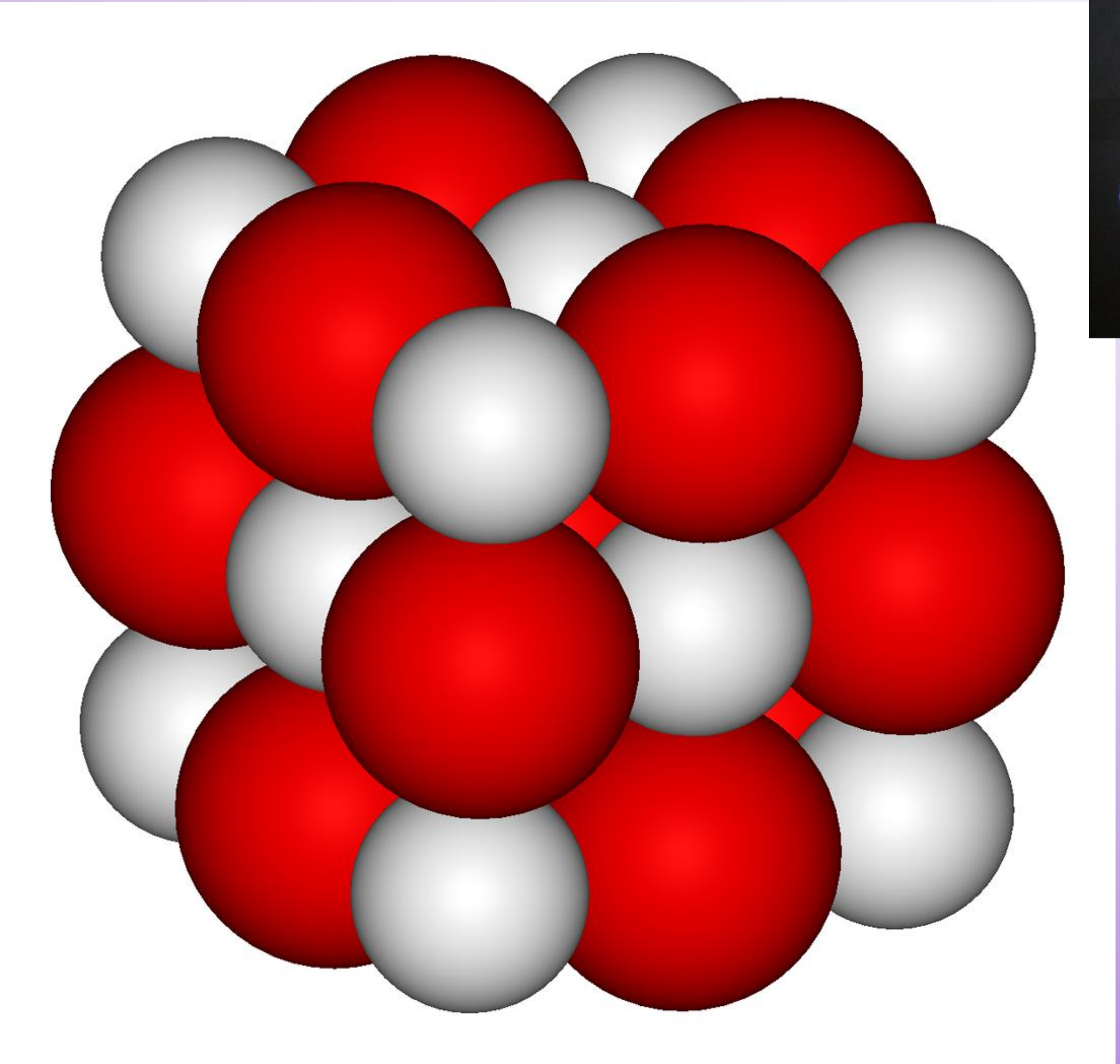


Image Credit: wikipedia

Model of the Crystal Lattice Structure in Potassium Iodide

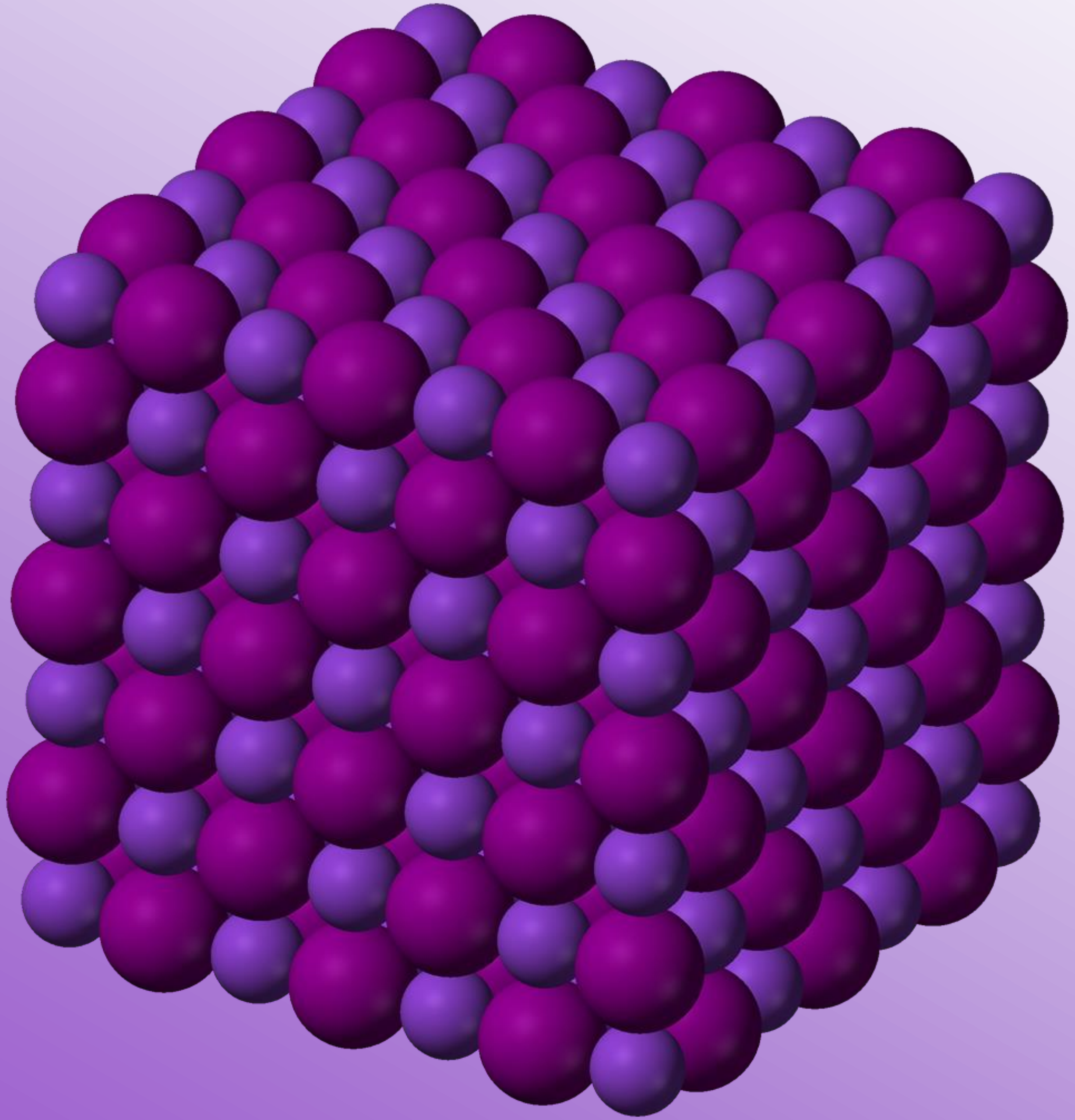


Image Credit: wikipedia

Model of the Crystal Lattice Structure in $CuSO_4$

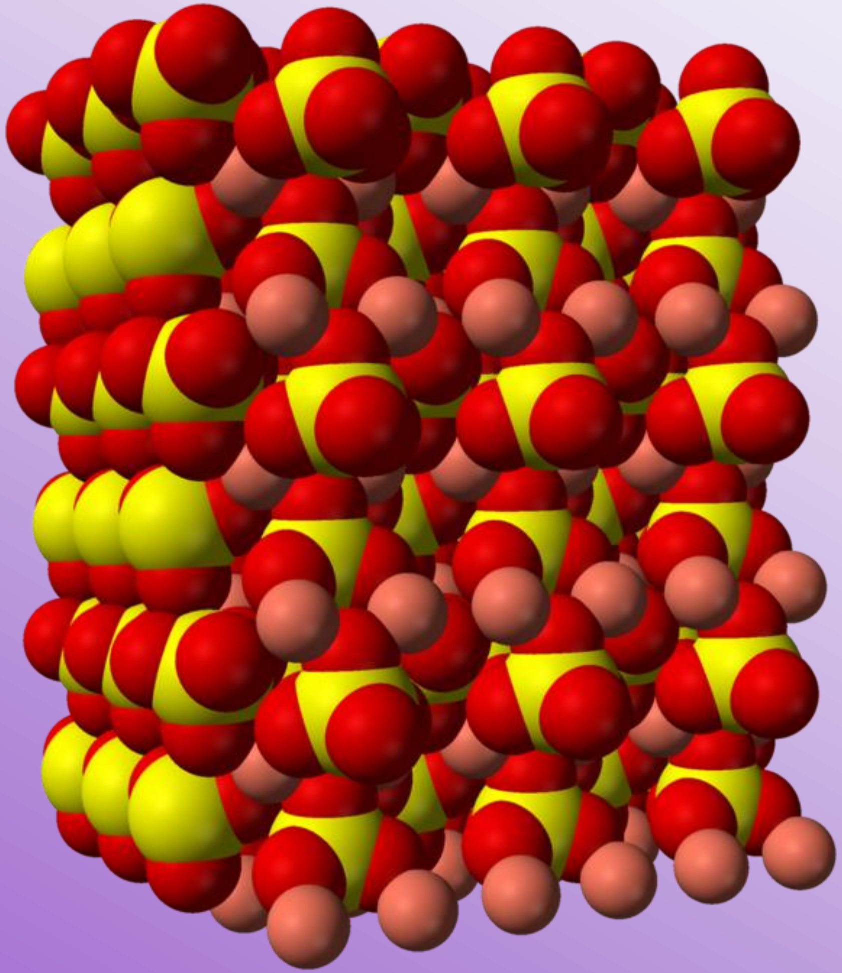


Image Credit: wikipedia

Model of the Crystal Lattice Structure in Calcium Carbonate (Chalk)

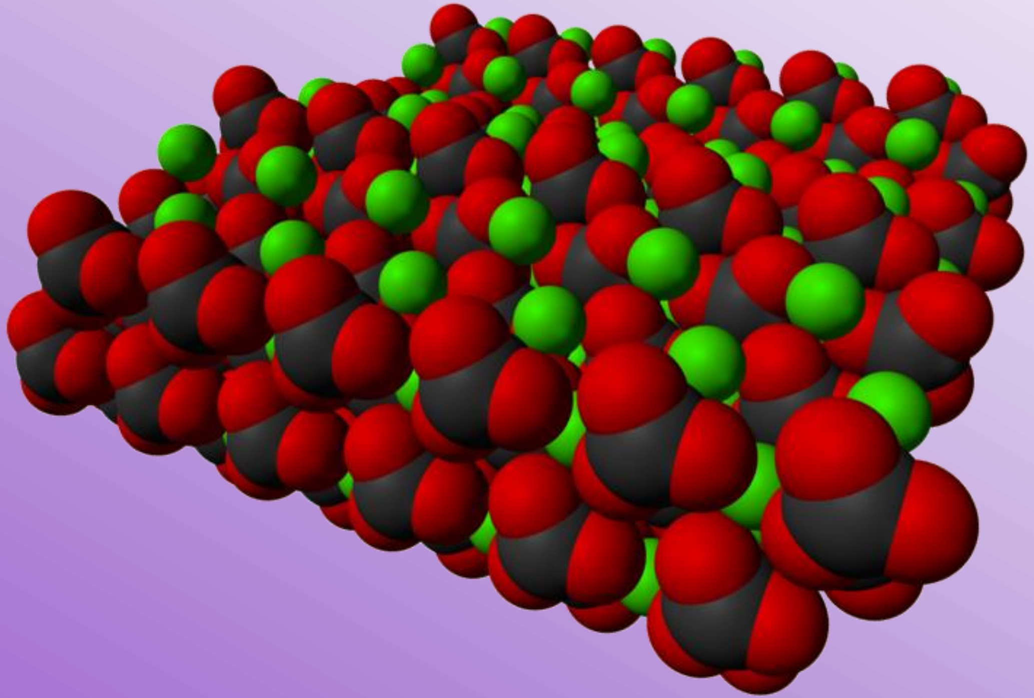


Image Credit: wikipedia

Model of the Crystal Lattice Structure in Lithium Chloride

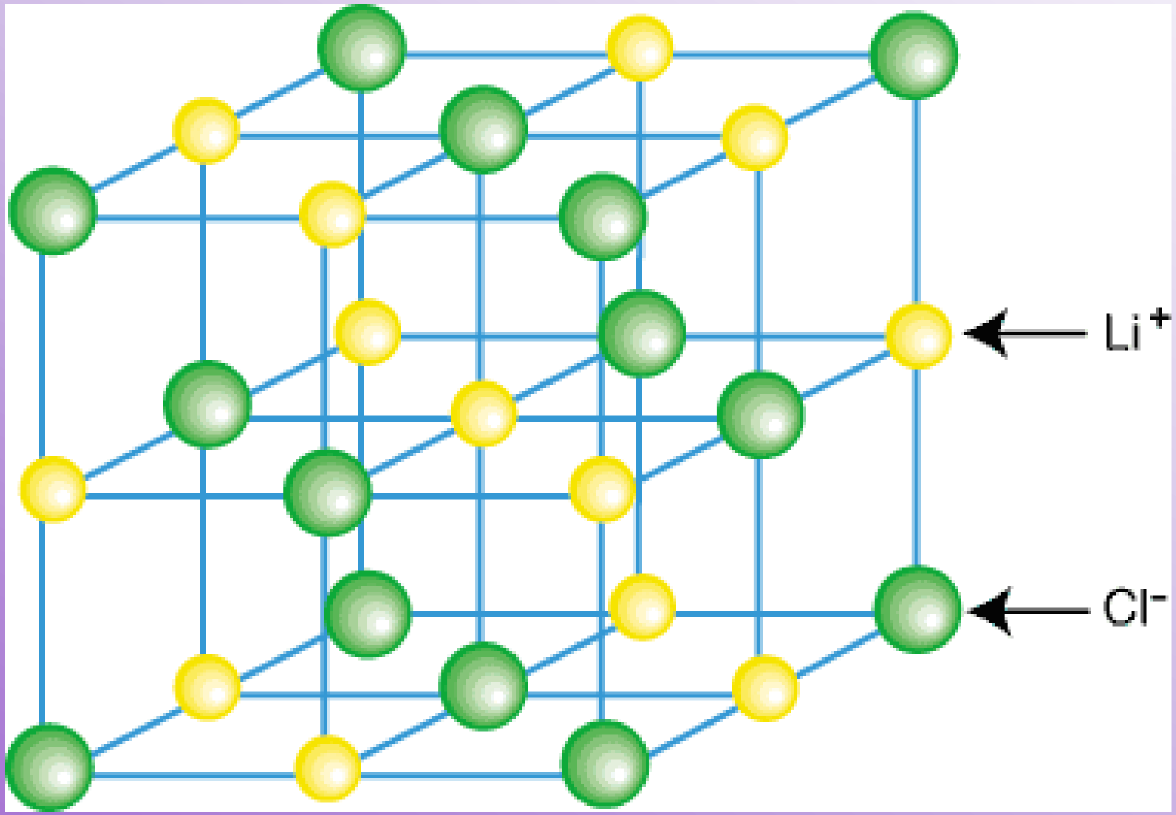


Image Credit: wikipedia

The Crystal Lattice Structure of Ionic Compounds

- As mentioned in the previous videos, the crystal lattice structure of ionic compounds is responsible for many of the observed physical properties of ionic compounds:
  - *Very high melting points*. This is because it takes a lot of energy to break the **electrostatic attraction** between the ions in the lattice and get them to flow freely like the molecules in liquid water do.
  - *Ionic compounds are brittle*. This is because applying a force to the ionic compound causes it to break or cleave at a certain point within the lattice
- Ionic compounds are solids substances that make up rocks and minerals. As you know, rocks don’t bend they crack of break. Rocks also do not melt or become molten unless you have extremely high temperatures.

Section 2 Additional Resources & Links…

- Tyler Dewitt's Video: "Writing Ionic Formulas: Introduction"
- Tyler Dewitt's Video: "Writing Formulas for Polyatomic Ions"
- Tyler Dewitt's Video: "Ionic Bonding: An Introduction"
- Tyler Dewitt's Video: "Ionic Bonding: Part 2"
- Tyler Dewitt's Video: "Ionic Bonding: Part 3"
- Fuse School Video: "What are Ionic Bonds"
- Fuse School Video: "Ionic Compounds and their Properties"

Section 3: Covalent Bonding & Molecules

This part of the Unit is covered on page 240-243 in your textbook

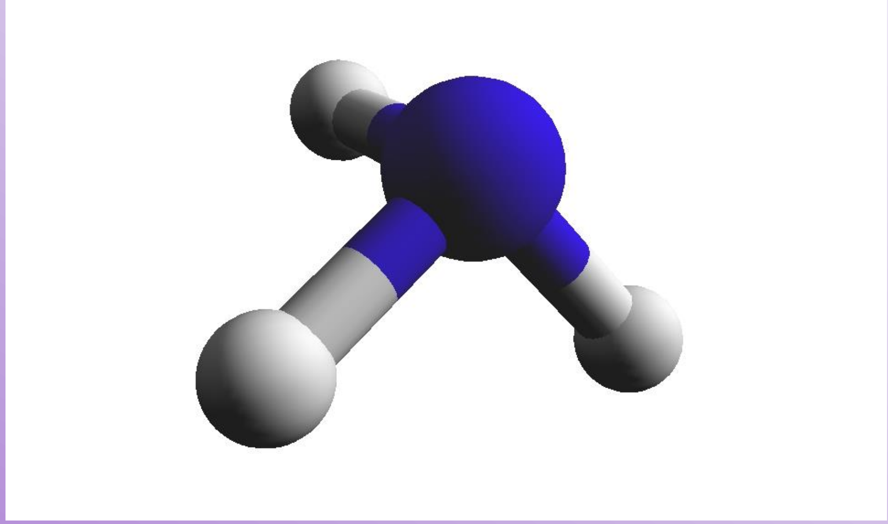
Section 3: Covalent Bonding & Molecules / Objectives

- Students Will Be Able too…
  - …*use* the octet rule to *determine* how many **covalent bonds** an element will usually make."
  - …*draw* **Lewis structures."**
  - …*draw* Lewis structures for molecules with **double** and **triple bonds**."
  - …*recall* the three types of **chemical bonding** and the key words for what happens with the electrons."
  - …*compare and contrast* Ionic & Covalent bonding."

Covalent Bonding

- **Covalent bonding** is the opposite of Ionic bonding:
  - In ionic bonding electrons are *transferred* but in covalent bonding electrons are *shared.*
  - In ionic bonding metals combine with non-metals but in covalent bonding non-metals combine with other non-metals.
  - Ionic bonding forms Ionic compounds but covalent bonding forms **molecules**.
  - Ionic compounds are always solids but molecules are gases, liquids, and occasionally solids.
  - Ionic compounds have high melting points and molecules have very low melting points (so low that they have already melted into liquids and even boiled into gases at room temperature!)
  - Ionic compounds have a crystal lattice of repeating ions in fixed positions. Molecules are individual "clumps of atoms" that sometimes chain or group themselves together but are not rigid like a crystal lattice is.

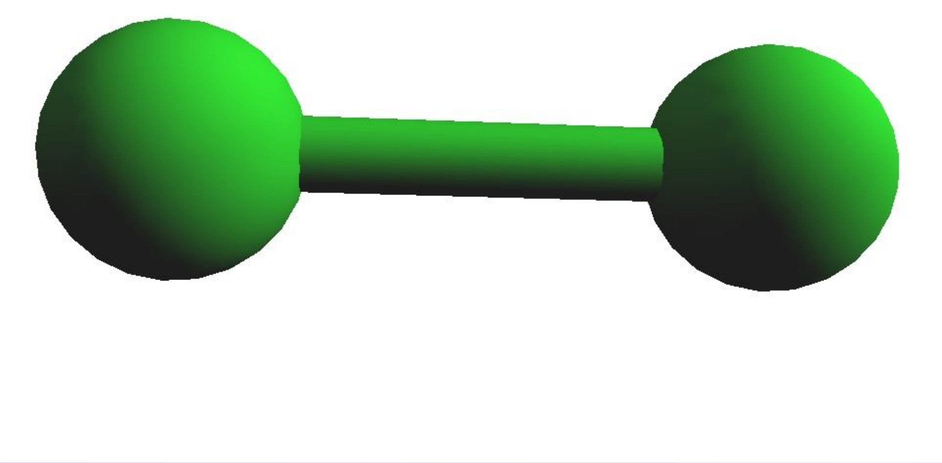
3D Ball and Stick Models for Some Simple Molecules



Ammonia

“NH_3”

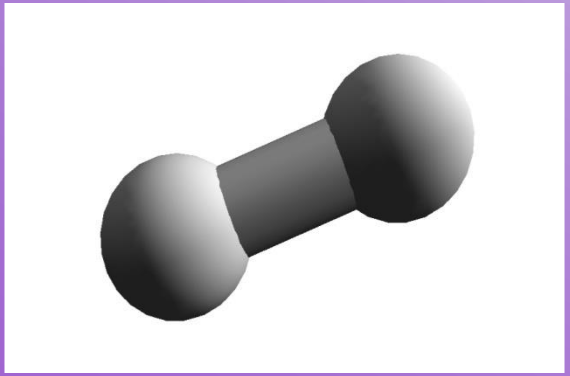
- Household Cleaner
- Industrial chemical (fertilizers)



Chlorine gas

“Cl_2”

- Chemical Weapon used in WWI
- Industrial chemical

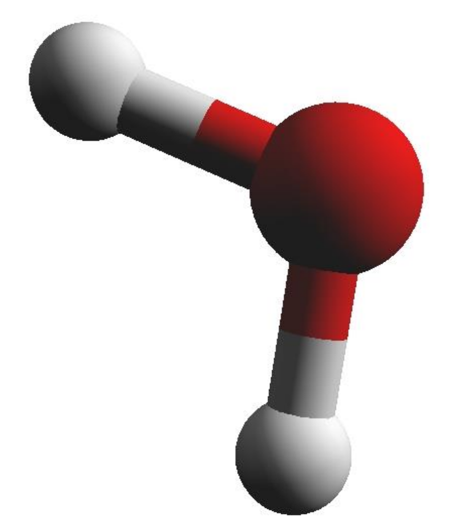


Hydrogen Gas

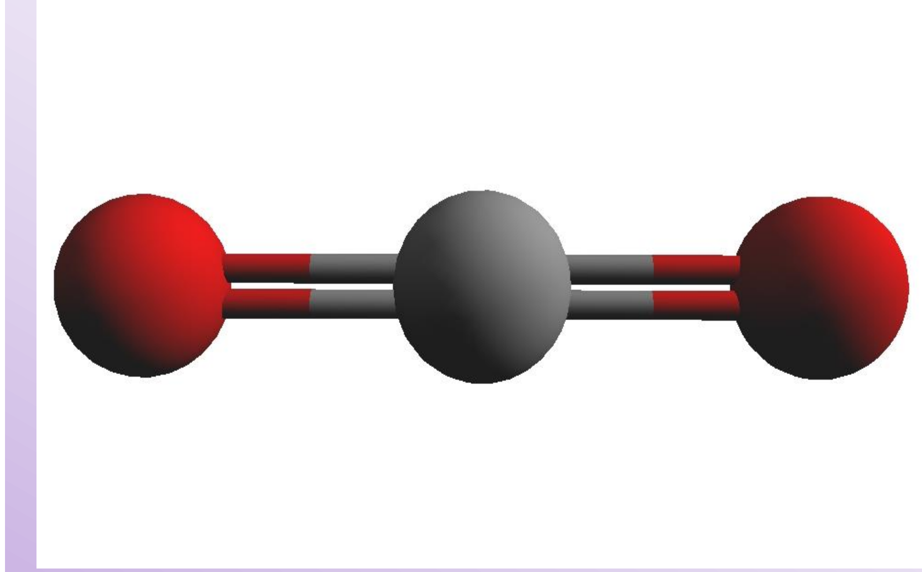
“H_2”

- Clean burning Fuel

3D Ball and Stick Models for Some Simple Molecules



Water
"H_2O"



Carbon Dioxide
"CO_2"

- Bi-product of essentially anything that burns (except hydrogen)
- Greenhouse gas

Video Time !!!

- Cosmos Clip Video: Democritus and Bonding
- SciShow Video: "Why Do Atom's Bond?"
- TEDed Video: "How atoms Bond"
- Some Random Teacher's Video: "Ionic and Covalent Bonding Animiation"
- TEDed Video: "The Science of Macaroni Salad: What's in A Molecule"

The Octet Rule & Covalent Bonding

- One thing ionic and covalent bonding do share in common though, is that the octet rule can be used to explain things about both!
  - In ionic bonding, the octet rule explains and determines what ion a main-group element will form.
  - In covalent bonding, the octet rule explains and determines how many covalent bonds non-metals will make (most of the time).

The Octet Rule & Covalent Bonding

- Recall that the octet says elements want to get to 8 or 0 valence electrons depending on on what column they are in.
- Since covalent bonding only happens between non- metals, we only need to look at the elements in columns 4,5,6, and 7 because those are the columns non-metal elements are in (and hydrogen).
- Remember the goal for elements in those columns (except hydrogen) is to get to 8 valance electrons. The number of covalent bonds formed is equal to the number of electrons an element wants to gain.
- Column 5 = 5 valence electrons = wants to gain 3 = 3 covalent bonds.
- Column 6 = 6 valence electrons = wants to gain 2 = 2 covalent bonds.
- Column 7 = 7 valence electrons = wants to gain 1 = 1 covalent bond.
- Carbon forms 4 covalent bonds.
- Hydrogen forms 1 covalent bond.
- Covalent bonding creates **molecules**. Molecules are modeled in a couple of different ways. One way we will model them is by using Lewis structures.
- The first step with Lewis structures is Lewis Dot Diagrams; basically the element’s symbol surrounded by it’s number of valence electrons.

Lewis Dot Diagrams for the First 20 Elements

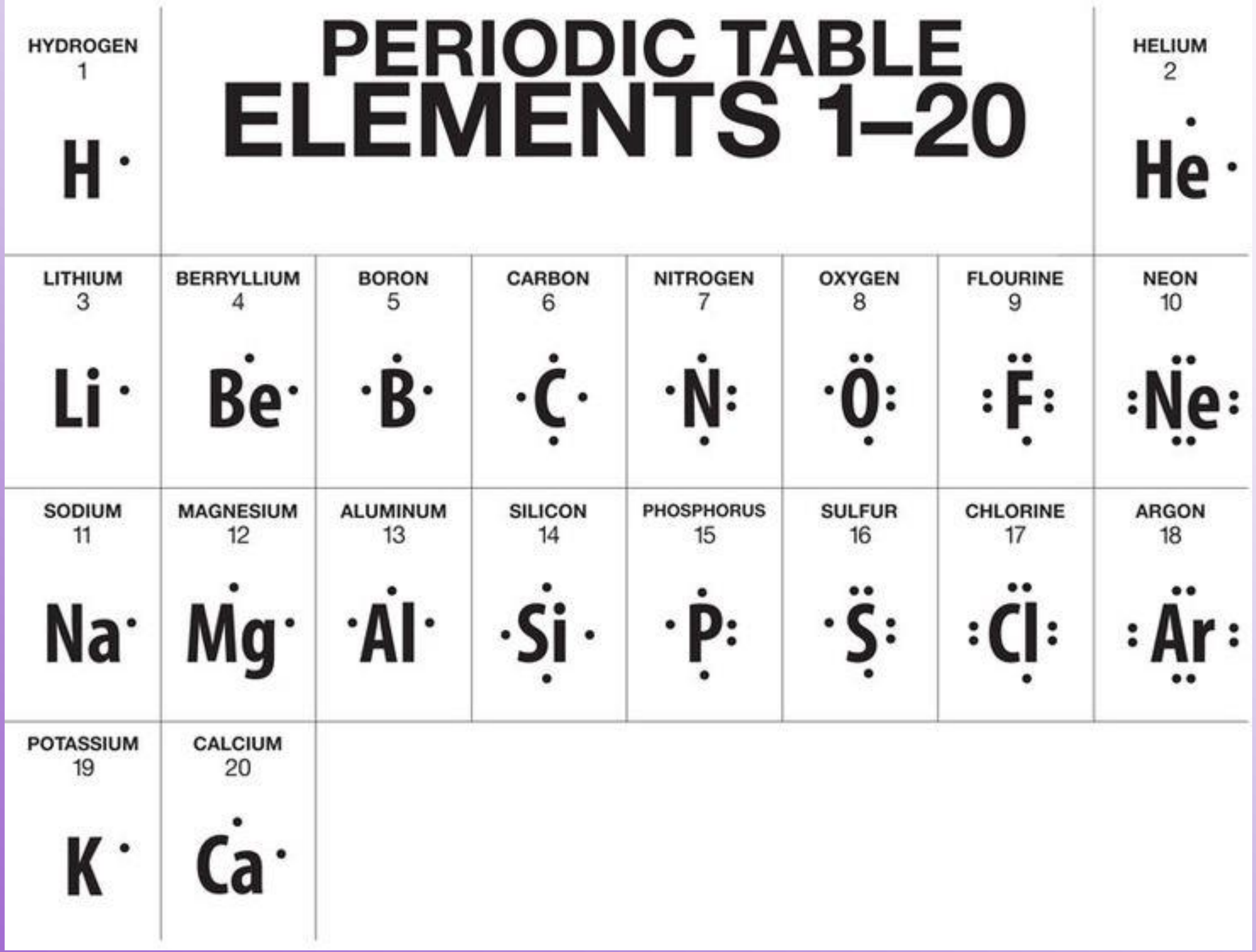


Image Credit: i.pinimg.com

Drawing Lewis Structures

- The first thing you need to realize about drawing Lewis structures is that it's like a puzzle, the pieces will only fit together in a certain way
- The second thing you need to realize about drawing Lewis structures is that they can take a lot of practice so don't get frustrated if you don't get the correct answer the first time. Sometimes you have to move atoms around before you can get all of the pieces to fit correctly.
- We will look at one set of steps to follow for beginners. Those of you who move on to chemistry in college or AP chemistry will probably use a different set of steps for drawing Lewis structures.

Steps for Drawing Lewis Structures

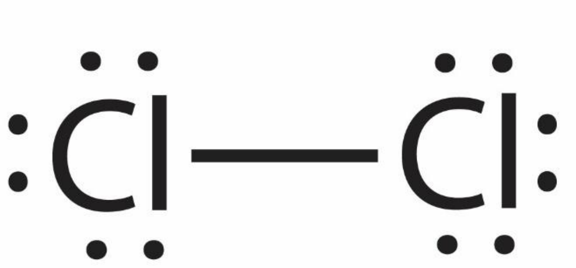
1) Draw the atom that makes the largest number of bonds first because it will be your central atom. If there is more than one of those atoms, make a "chain" of them.
2) Move on to the element that makes the 2nd largest number of bonds and attach it to the end of your first atom or the chain of atoms you created in step 1.
3) Save Hydrogen and Halogens (column 7) for last because they only make 1 bond.
4) When your structure is finished, there should never be any single dots left over, and all elements except for hydrogen should be surrounded by 8 electrons (1 bond = 2 electrons). If you have single dots left over, there must be double or triple bonds in the structure somewhere.
5) When there are no more single dots left and all atoms have the correct number of bonds and lone pairs, your structure is complete.

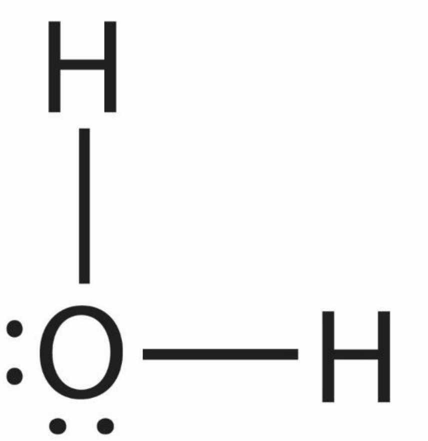
Practice Problems: Drawing Lewis Structures

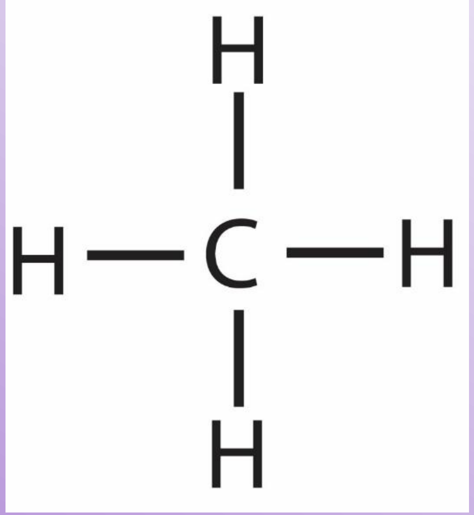
Directions: Draw the Lewis Structures for the following molecules

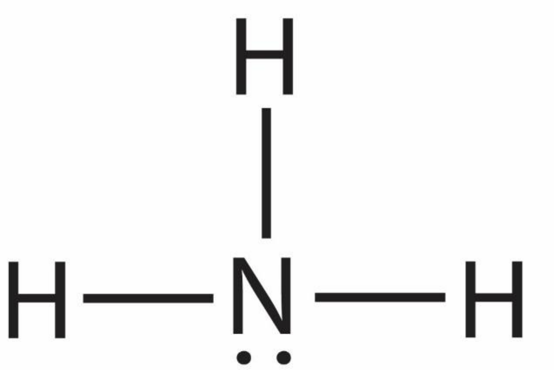
1) H_2O
2) CH_4
3) NH_3
4) Br_2
5) C_2H_6
6) N_2H_4
7) CH_4O
8) C_2H_6O
9) PF_3

Lewis Structure Examples for H_2, Cl_2, H_2O, CH_4, & NH_3









Drawing Lewis Structures for Molecules w/ Double or Triple Bonds

- Many elements are capable of forming more than 1 bond to the same atom. Including an atom of the same element.
- Elements in column 6 like oxygen and sulfur, can form **double covalent bonds.**
  - O_2 which accounts for 21% of the air we breathe has a double bond.
- Elements in column 5 like nitrogen and Phosphorus can form double & **triple covalent bonds.**
  - N_2, which accounts for 78% of air has a triple bond
- Carbon can also form double and triple bonds.
- If you run out of atoms when drawing your Lewis structure and you still have single dots left, that indicates the compound contains double and/or triple bonds.

Practice Problems: Lewis Structures w/ Double & Triple Bonds

Directions: Draw the Lewis Structures for the following molecules

1) O_2

2) N_2

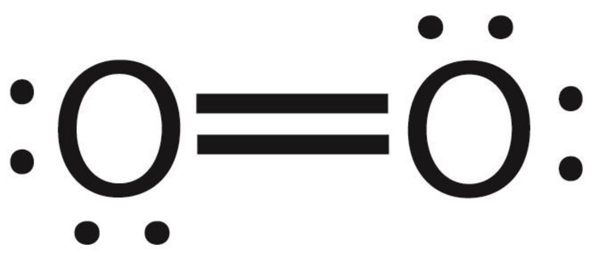
3) C_2H_4

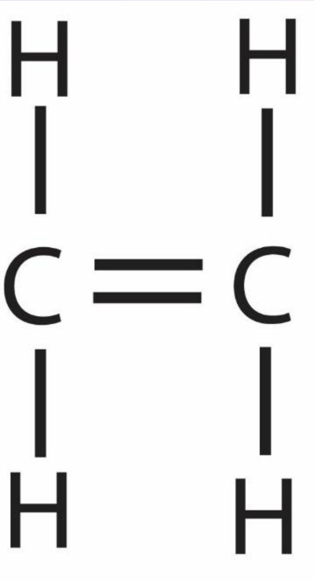
4) C_2H_2

5) CH_2O

6) N_2F_2

Lewis Structure Examples for O_2, N_2, & C_2H_4

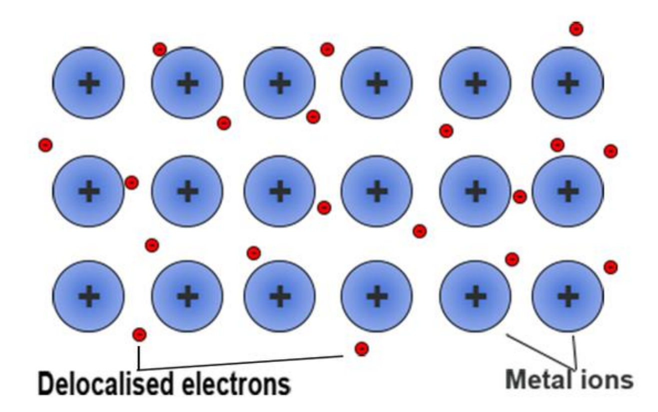


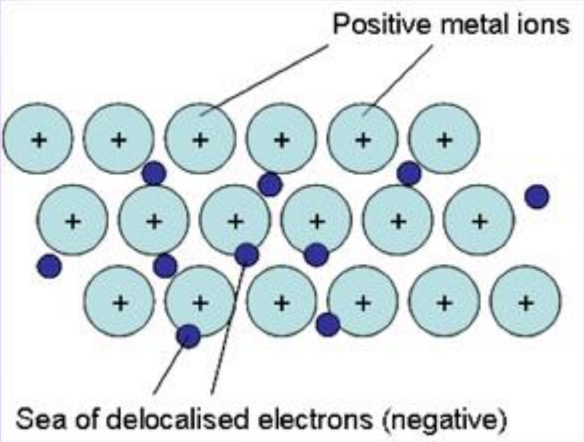


Just 1 slide on Metallic Bonding

- A final type of chemical bonding is metallic bonding. **Metallic Bonding** is what you gets between the atoms in pure metals or alloys. **Alloys** are mixtures of metals, like 18 K gold is 75% gold atoms and 25% silver atoms.
- Metallic Bonding is generally not an area of focus for beginning chemistry students but there are a few things you should know:
  - Like Ionic Compounds, Solid metal objects have a lattice structure.
  - Unlike the crystal lattice in Ionic compounds, the electrons can move between the atoms in metals freely. They are not fixed on a single atom or between two atoms like in ionic and covalent bonding. The electrons in metals are described as being a “sea of freely flowing electrons”
  - In other words, the electrons can easily move between the atoms. This is why all metals conduct electricity. Electricity is essentially the movement of electrons between atoms, hence, the term “electrical current”

Some Models of the Lattice Structure in Metallic Bonding





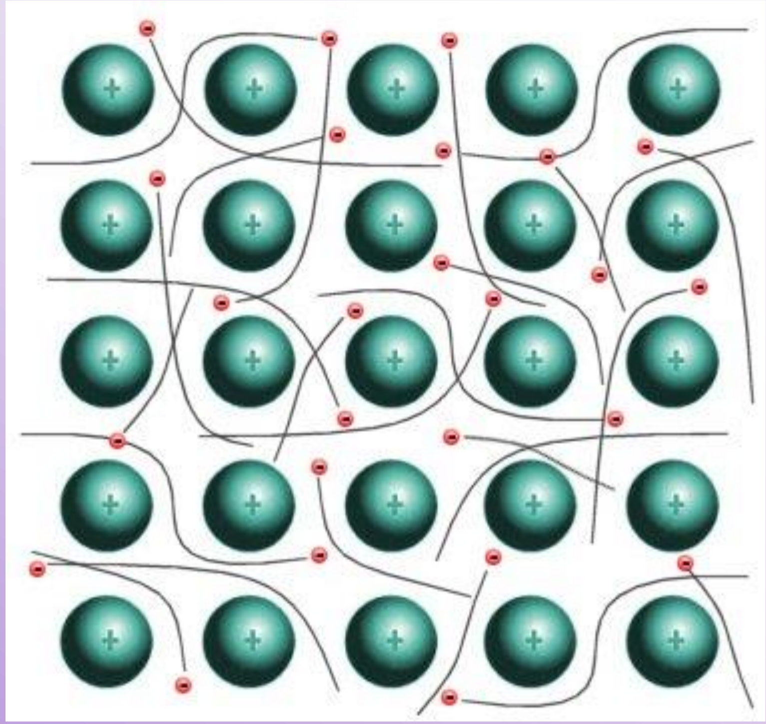


Image Credit: https://chemistryclinic.co.uk/bonding/
Image Credit: https://www.quora.com/What-is-the-nature-of-metallic-bonding
Image Credit: https://www.technologyuk.net/physics/matter/chemical-bonding.shtml

Section 3 Additional Resources & Links…

- Tyler Dewitt's Video: "Ionic vs. Molecular"
- CCC #24 Video: "Bonding Models and Lewis Structures"

Section 4: VSEPR Theory & Polarity

This part of the Unit is covered on page ??? - ??? in your textbook

Section 4: VSEPR Theory & Polarity / Objectives

Students Will Be Able too…

- …*define* **steric number** and *determine* the steric number for a given molecule.
- …*use* the Lewis structure of a molecule to determine it's **molecular geometry** (**VESPR chart** provided).
- …*determine* if a covalent bond is a **polar covalent bond** or a **non-polar covalent bond** and assign partial charges (on a Lewis structure) to atoms with polar bonds.
- …*distinguish* polar bonds for **polar molecules**.
- …*determine* if a molecule is polar or non-polar.

Steric Number

- Lewis structures are 2 dimensional models for molecules that are really 3 dimensional.
- They are very useful in two ways:
    - They show you how the atoms are bonded to one another.
    - They are the first step in determining the 3 dimensional shape.
- The second step in determining the 3 dimensional shape of a molecule is to find the steric number of the central atom.
- The **steric number** is the number of lone pairs + the number of atoms bonded to the central atom.
- It might be easier to think of it as the number of lone pairs + the number of bonds but, double bonds and triple bonds only count as 1 when it comes to steric number.
- We will only worry about molecules with a steric number of 2,3, or 4 in this class. Steric numbers of 5 and 6 (and higher) exist but almost always involve molecules that break the octet rule.
- BTW, steric number is sometimes referred to as the number of "bonding groups"

Practice Problems: Determining Steric Number

Directions: Draw the Lewis Structures for the following molecules

1) CH_2O
2) HCN
3) CH_4
4) NH_3
5) CCl_2S
6) HNO
7) CF_2S
8) CO_2
9) PF_3

VSEPR Theory & Molecular Shape

- **VSEPR Theory** is an acronym that stands for "Valence Shell Electron Pair Repulsion."
- It sounds way more complicated than it is.
- VSEPR Theory is basically a model we use to predict the three dimensional structure of a molecule.
- The easiest way to do this is with the VSEPR chart (on the next slide) but first it's worth mentioning the theory itself and how it explains molecular geometry.
- The main idea is that electrons want to be as far away from each other as possible because their like charges repel one another.
- Therefore, the pair of electrons that make up the covalent bond in a molecule will try and get *as far away as possible* from the pair of electrons in a neighboring covalent bond or a lone pair.
- This repulsion between electron pairs (bonds and lone pairs) causes the molecule to twist and bend until the electron pairs are at the maximum possible distances from each other (without the bonds breaking of coarse).

VSEPR Theory Chart

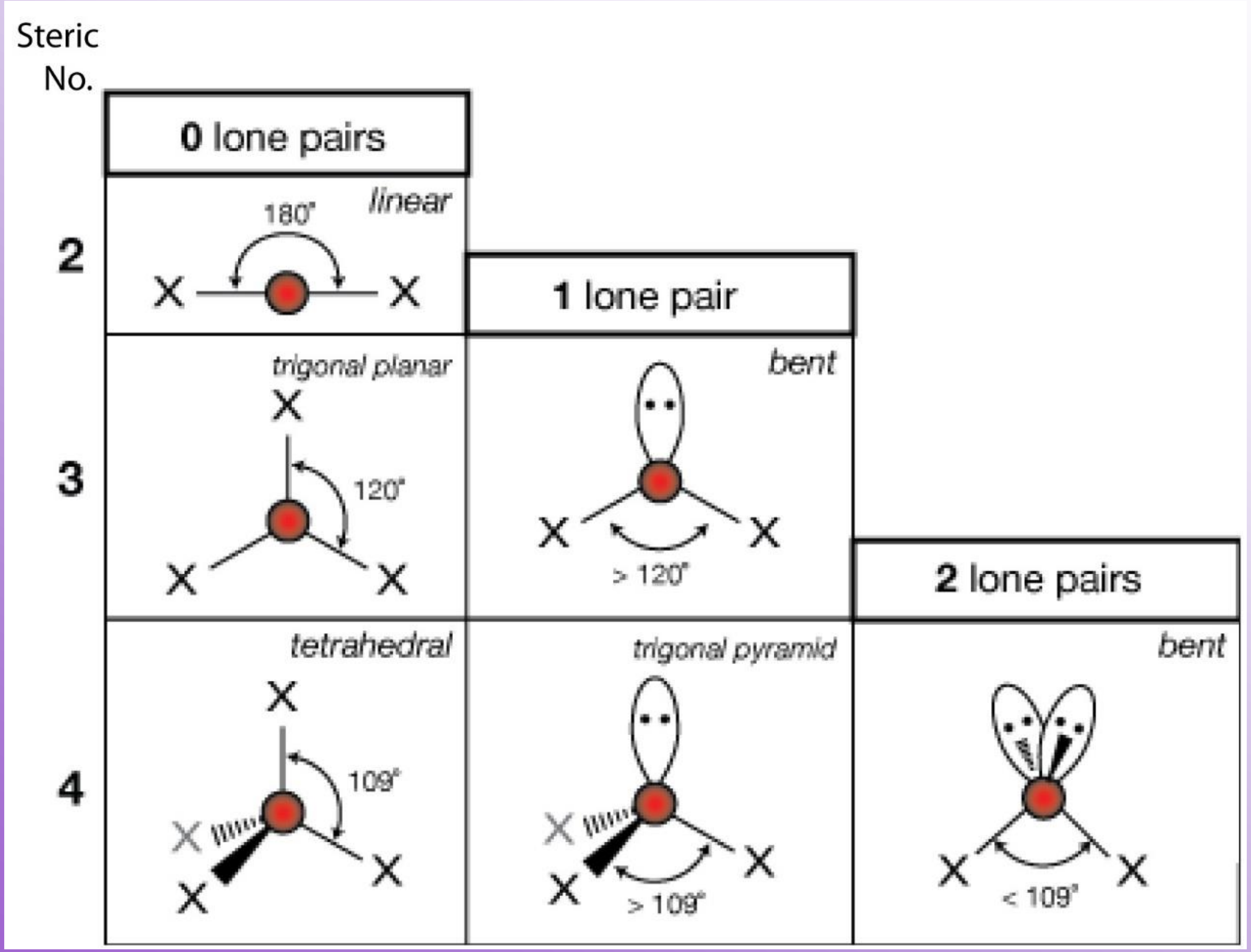


Image Credit: xactly.com

5 Common Molecular Shapes encountered in Chem A

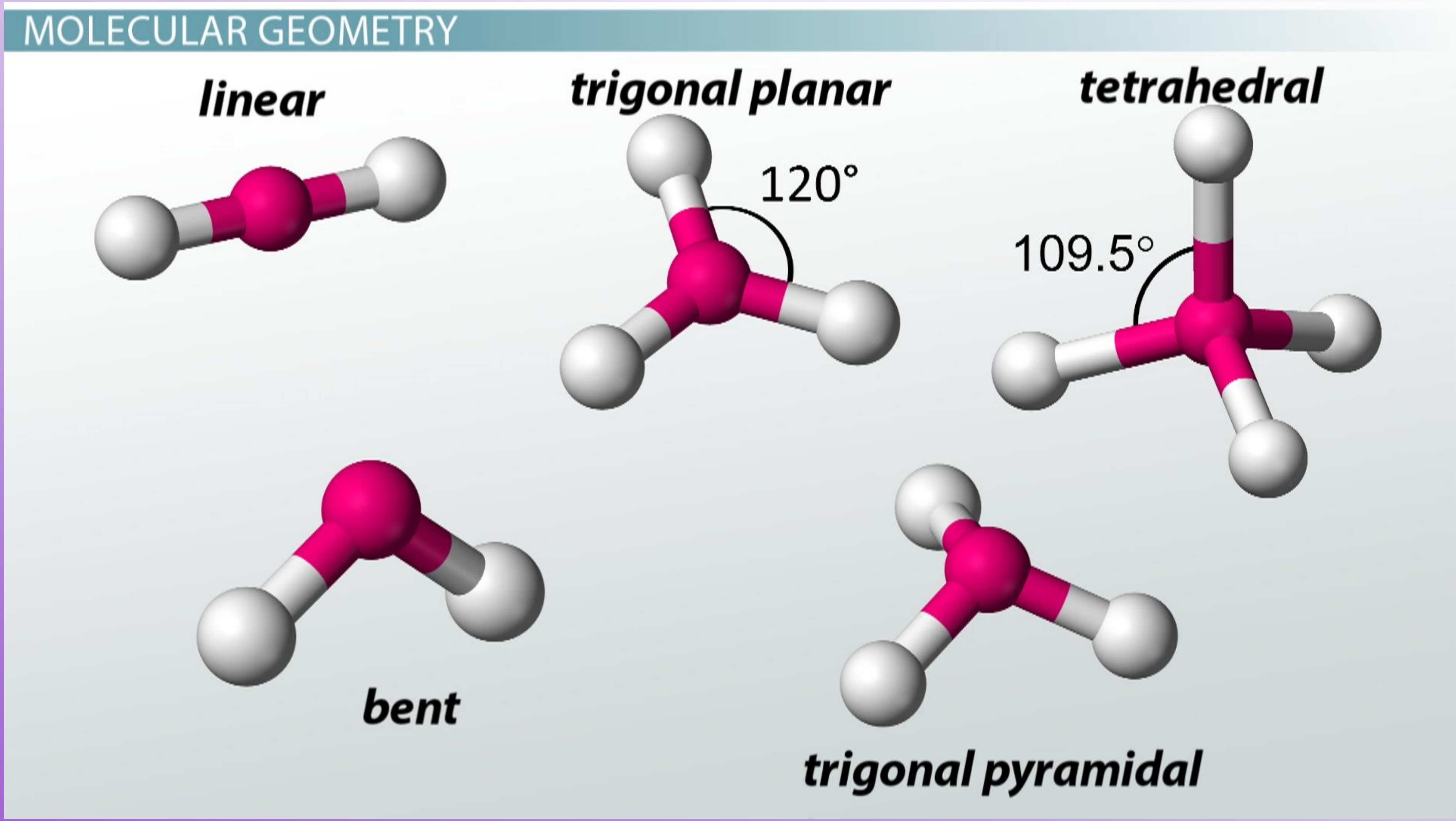
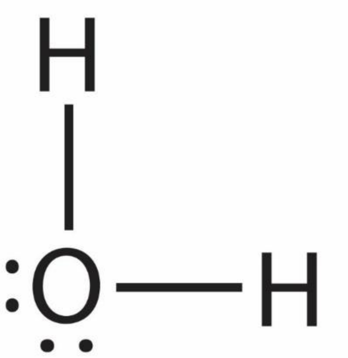
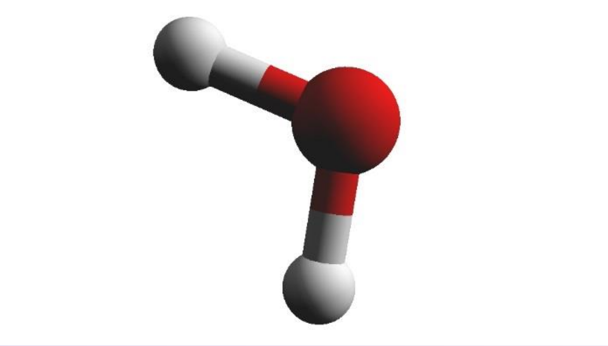
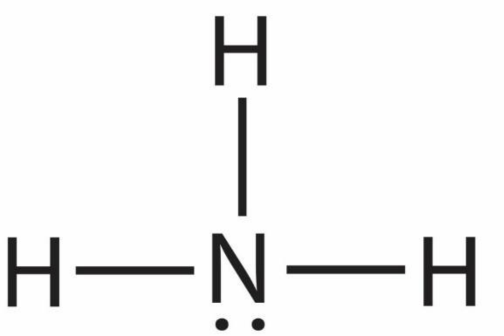


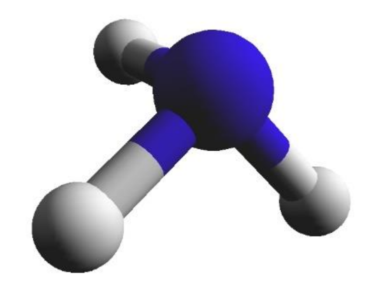
Image Credit: study.com

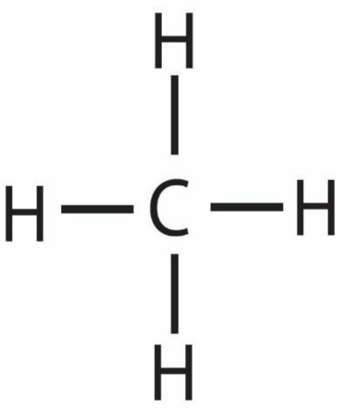
Lewis Structures Next to Their 3D Shapes

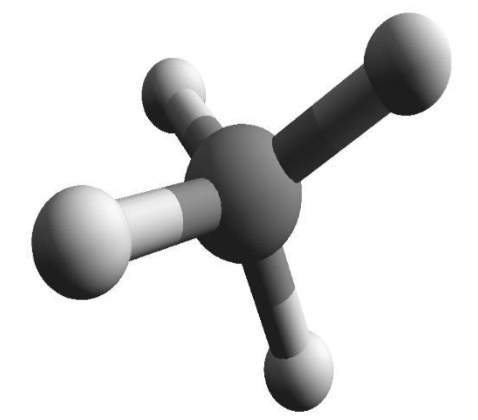












VSEPR Theory & Molecular Shape

- Utilizing the VSEPR theory matrix or chart you should be able to determine the **molecular shape (a.k.a molecular geometry)** for any molecule with a steric number of 2,3, or 4.
- As the chart shows there are only 5 shapes you have to worry about (**bent, trigonal planar, trigonal pyramidal, linear, and tetrahedral**).
- If the molecule is **diatomic**, meaning there is only 2 atoms, the shape is always linear.

Video Time !!!

- TEDed Video: “What’s the Shape of a Molecule”

Practice Problems: Determining Molecular Geometry

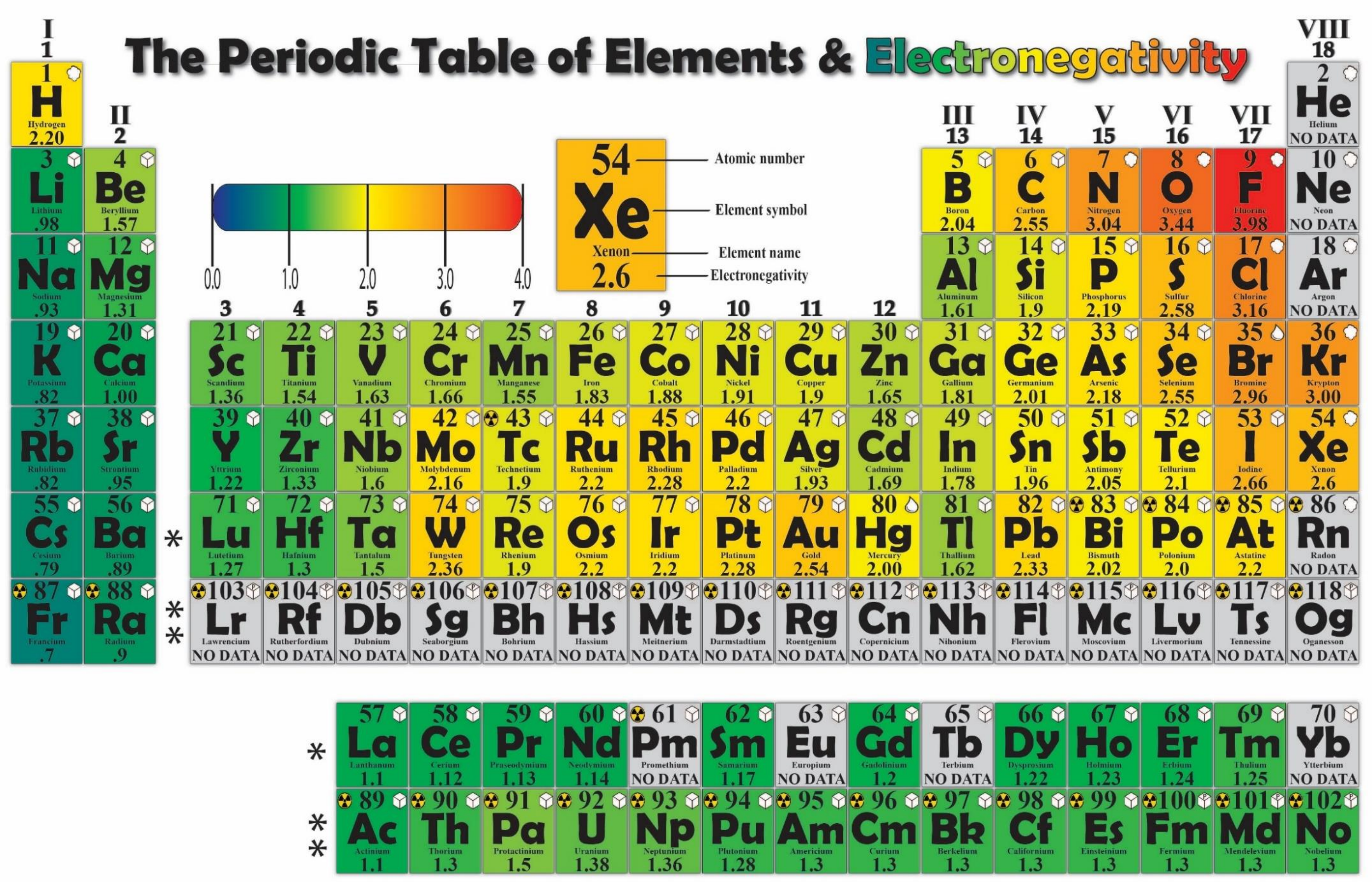
Directions: (1) Draw the Lewis Structures for the Following molecules. (2) write down the steric number (3) determine the molecular geometry

1) NBr_3
2) CH_3F
3) N_2
4) CF_2S
5) H_2S
6) NOF
7) CF_2Cl_2
8) OF_2
9) CCl_2O

Electronegativity

- The covalent bonds in molecules happen when electrons are shared between atoms trying to achieve noble gas electron configuration or octets.
- However, this sharing is not always an equal sharing.
- Usually the sharing of electrons between atoms is unequal, and this happens whenever the atoms bonded to one another are different elements.
- So in a C-C bond the sharing is totally equal. However, in a C-H bond, or a C-N bond, or a O-F bond, or any time the atoms are different elements, the sharing is unequal.
- How unequal depends on the boned atoms electronegativity values.
- **Electronegativity** is defined as the tendency of an atom to attract electrons towards itself.
- It can be thought of in more simpler terms as how much an element "wants" electrons.
- You can a find electronegativity values for non-metals listed on the back of your periodic table.

Periodic Table Showing Electronegativity



The Periodic Table of Elements & Electronegativity

54 Atomic number
Xe Element symbol
Xenon Element name
2.6 Electronegativity

I 1	II 2		3	4	5	6	7	8	9	10	11	12	III 13	IV 14	V 15	VI 16	VII 17	VIII 18
1 H Hydrogen 2.20																		2 He Helium NO DATA
3 Li Lithium .98	4 Be Beryllium 1.57												5 B Boron 2.04	6 C Carbon 2.55	7 N Nitrogen 3.04	8 O Oxygen 3.44	9 F Fluorine 3.98	10 Ne Neon NO DATA
11 Na Sodium .93	12 Mg Magnesium 1.31												13 Al Aluminum 1.61	14 Si Silicon 1.9	15 P Phosphorus 2.19	16 S Sulfur 2.58	17 Cl Chlorine 3.16	18 Ar Argon NO DATA
19 K Potassium .82	20 Ca Calcium 1.00		21 Sc Scandium 1.36	22 Ti Titanium 1.54	23 V Vanadium 1.63	24 Cr Chromium 1.66	25 Mn Manganese 1.55	26 Fe Iron 1.83	27 Co Cobalt 1.88	28 Ni Nickel 1.91	29 Cu Copper 1.9	30 Zn Zinc 1.65	31 Ga Gallium 1.81	32 Ge Germanium 2.01	33 As Arsenic 2.18	34 Se Selenium 2.55	35 Br Bromine 2.96	36 Kr Krypton 3.00
37 Rb Rubidium .82	38 Sr Strontium .95		39 Y Yttrium 1.22	40 Zr Zirconium 1.33	41 Nb Niobium 1.6	42 Mo Molybdenum 2.16	43 Tc Technetium 1.9	44 Ru Ruthenium 2.2	45 Rh Rhodium 2.28	46 Pd Palladium 2.2	47 Ag Silver 1.93	48 Cd Cadmium 1.69	49 In Indium 1.78	50 Sn Tin 1.96	51 Sb Antimony 2.05	52 Te Tellurium 2.1	53 I Iodine 2.66	54 Xe Xenon 2.6
55 Cs Cesium .79	56 Ba Barium .89	*	71 Lu Lutetium 1.27	72 Hf Hafnium 1.3	73 Ta Tantalum 1.5	74 W Tungsten 2.36	75 Re Rhenium 1.9	76 Os Osmium 2.2	77 Ir Iridium 2.2	78 Pt Platinum 2.28	79 Au Gold 2.54	80 Hg Mercury 2.00	81 Tl Thallium 1.62	82 Pb Lead 2.33	83 Bi Bismuth 2.02	84 Po Polonium 2.0	85 At Astatine 2.2	86 Rn Radon NO DATA
87 Fr Francium .7	88 Ra Radium .9	**	103 Lr Lawrencium NO DATA	104 Rf Rutherfordium NO DATA	105 Db Dubnium NO DATA	106 Sg Seaborgium NO DATA	107 Bh Bohrium NO DATA	108 Hs Hassium NO DATA	109 Mt Meitnerium NO DATA	110 Ds Darmstadtium NO DATA	111 Rg Roentgenium NO DATA	112 Cn Copernicium NO DATA	113 Nh Nihonium NO DATA	114 Fl Flerovium NO DATA	115 Mc Moscovium NO DATA	116 Lv Livermorium NO DATA	117 Ts Tennessine NO DATA	118 Og Oganesson NO DATA

*	57 La Lanthanum 1.1	58 Ce Cerium 1.12	59 Pr Praseodymium 1.13	60 Nd Neodymium 1.14	61 Pm Promethium NO DATA	62 Sm Samarium 1.17	63 Eu Europium NO DATA	64 Gd Gadolinium 1.2	65 Tb Terbium NO DATA	66 Dy Dysprosium 1.22	67 Ho Holmium 1.23	68 Er Erbium 1.24	69 Tm Thulium 1.25	70 Yb Ytterbium NO DATA
**	89 Ac Actinium 1.1	90 Th Thorium 1.3	91 Pa Protactinium 1.5	92 U Uranium 1.38	93 Np Neptunium 1.36	94 Pu Plutonium 1.28	95 Am Americium 1.3	96 Cm Curium 1.3	97 Bk Berkelium 1.3	98 Cf Californium 1.3	99 Es Einsteinium 1.3	100 Fm Fermium 1.3	101 Md Mendelevium 1.3	102 No Nobelium 1.3

Bond Polarity

- When the difference in electronegativity between two atoms in a chemical bond is great than .4, the bond is said to be a polar bond or **polar covalent bond**.
- If the difference is less than .4, the bond is described as being a non-polar bond or **non-polar covalent bond**.
- It's important to note that .4 is not a magic line that gets crossed and all of a sudden a chemical bond becomes polar. Even a bond with a difference of .1 will still exhibit a small degree of polarity. Bond polarity is a continuum or sliding scale, not a dividing line.
- However, for beginning chemistry students labels are important and so anytime the difference in electronegativity is less than .4 we will label the bond non-polar and anytime it is greater than .4 we will label the bond polar.

Bond Polarity Divisions

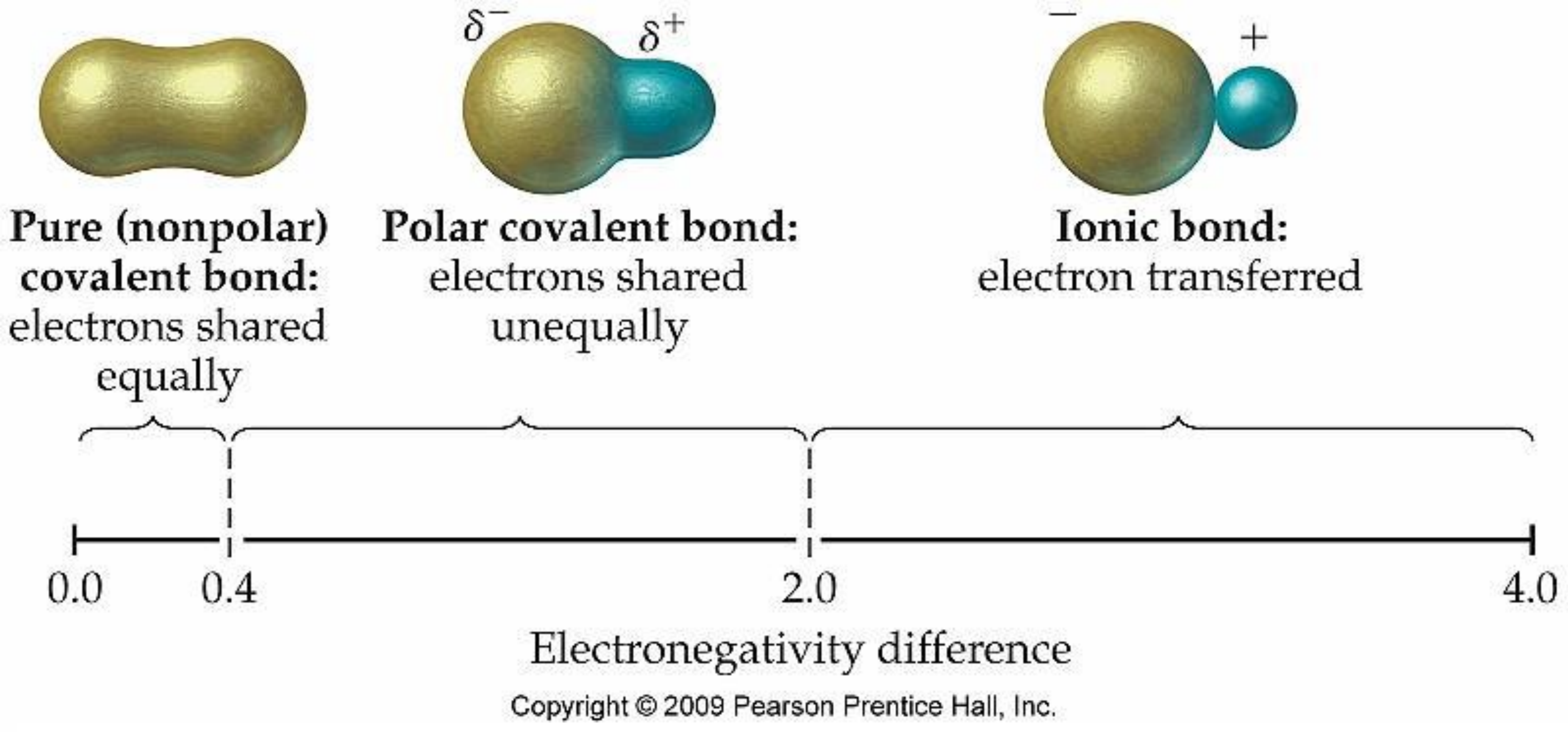


Image Credit: Pearson

Video Time !!!

- CCC #22 Video : “Atomic Hook-ups: Types of Chemical Bonds”
  Note: It does not hurt to pay attention to what is discussed about bond length & coulombs law in this video, but those concepts are not in our objectives and you will not be assessed on them.

Dipoles

- Polar bonds create dipoles between the bonded atoms.
- **Dipoles** are the positive and negative ends of the bond

- **Positive Dipole** or "partial positive charge" on an atom is indicated using the symbol: δ^+

- **Negative Dipole or** "partial negative charge" on an atom is indicated using the symbol: δ^-

- The negative dipole is always on the more electronegative atom.
- The positive dipole is always on the less electronegative atom.

Practice Problems: Assigning Partial Charges

- *Directions: (1) Draw the Lewis Structures for the Following molecules. (2) use the electronegativity values on the back of your periodic table to determine if there are polar bonds. If there are, assign partial positive and partial negative charges to the bonds.*

1) NF_3

2) CH_4

3) N_2

4) CF_2O

5) H_2S

6) CO_2

7) NH_2Cl

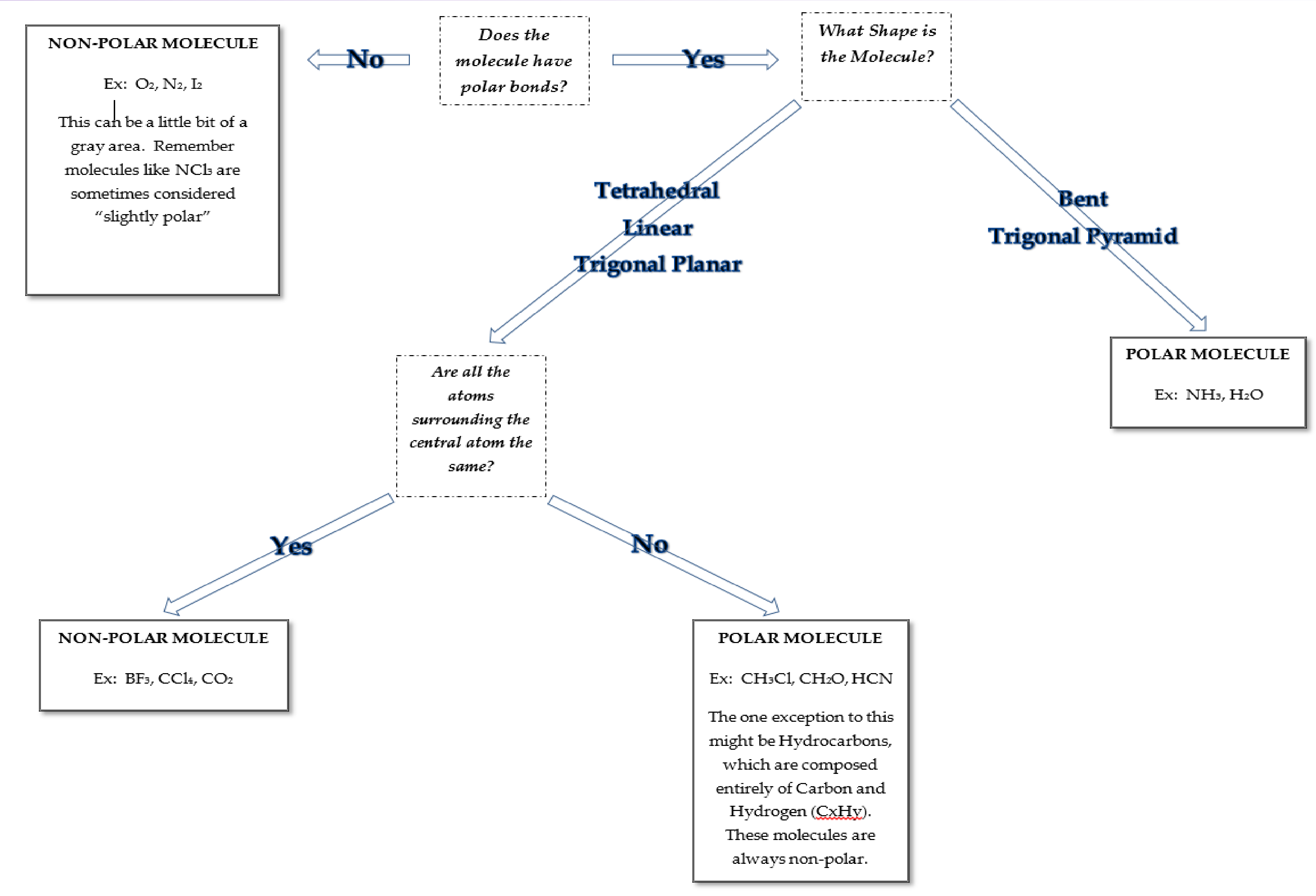
8) SF_2

9) CH_2O

Molecular Polarity

- Just as a bond can have a partial positive and partial negative end, and entire molecule can have a partial positive and partial negative sides to it.
- The partial positive and partial negative sides of a molecule is caused by the combination of two things: polar bonds and molecular shape.
- A molecule absent of polar bonds will be a non-polar molecule.
- If a molecule has polar bonds, it may or may not be a polar molecule depending on the shape and how that atoms are distributed in a molecule. In other words, polar bonds does not always mean polar molecule. Sometimes you can have a non-polar molecule with polar bonds.
- Polar molecules have polar bonds and are asymmetrical.
- Determining if a molecule is polar or non-polar will be one of the more difficult objectives this trimester but the flow chart on the next slide and many practice problems will hopefully be enough.

Polarity Flow Chart



Does the molecule have polar bonds?

No

NON-POLAR MOLECULE

Ex: O_2, N_2, I_2

This can be a little bit of a gray area. Remember molecules like NCl_3 are sometimes considered "slightly polar"

Yes

What Shape is the Molecule?

Tetrahedral
Linear
Trigonal Planar

Are all the atoms surrounding the central atom the same?

Yes

NON-POLAR MOLECULE

Ex: BF_3, CCl_4, CO_2

No

POLAR MOLECULE

Ex: CH_3Cl, CH_2O, HCN

The one exception to this might be Hydrocarbons, which are composed entirely of Carbon and Hydrogen (CxHy). These molecules are always non-polar.

Bent
Trigonal Pyramid

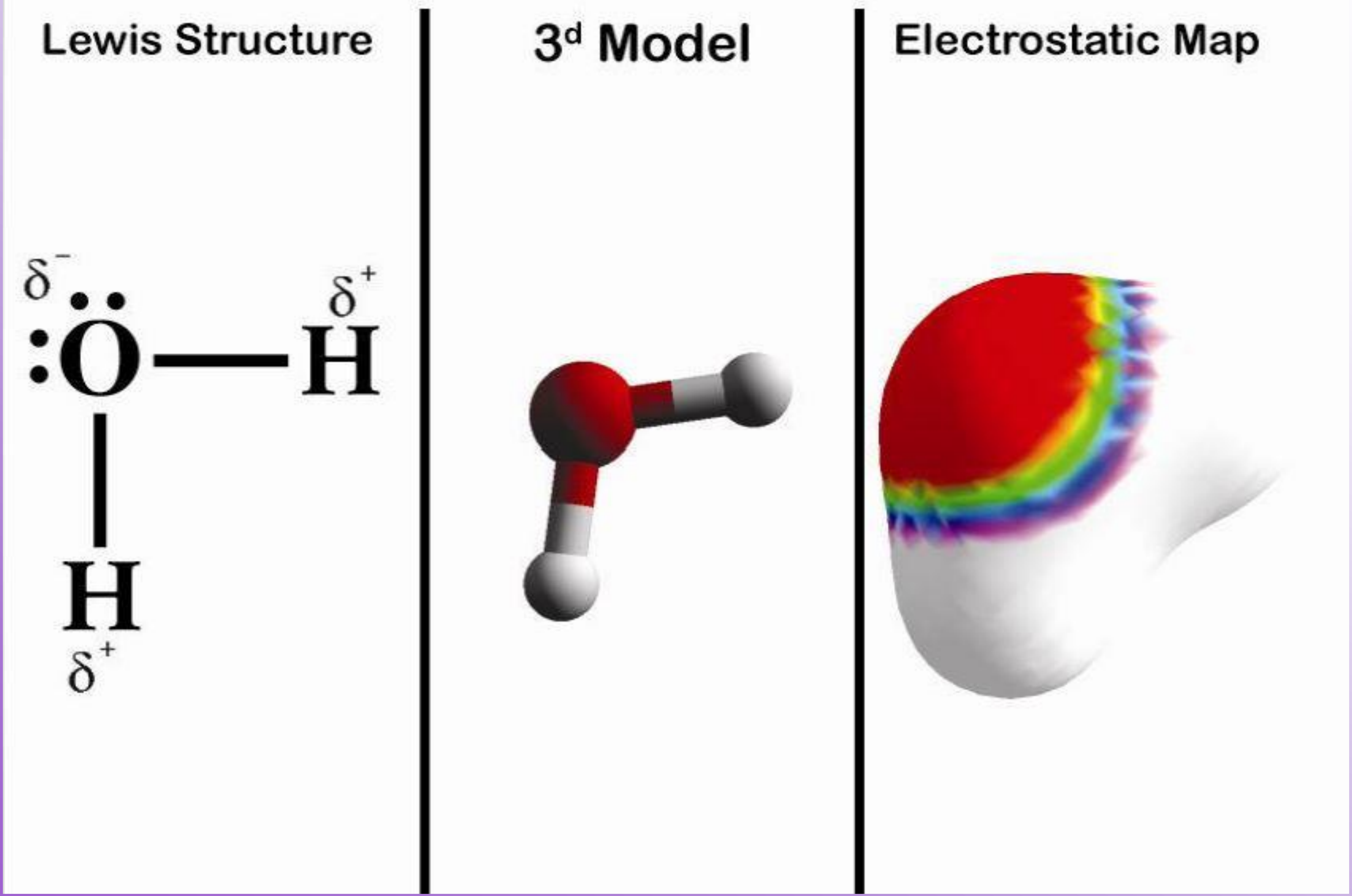
POLAR MOLECULE

Ex: NH_3, H_2O

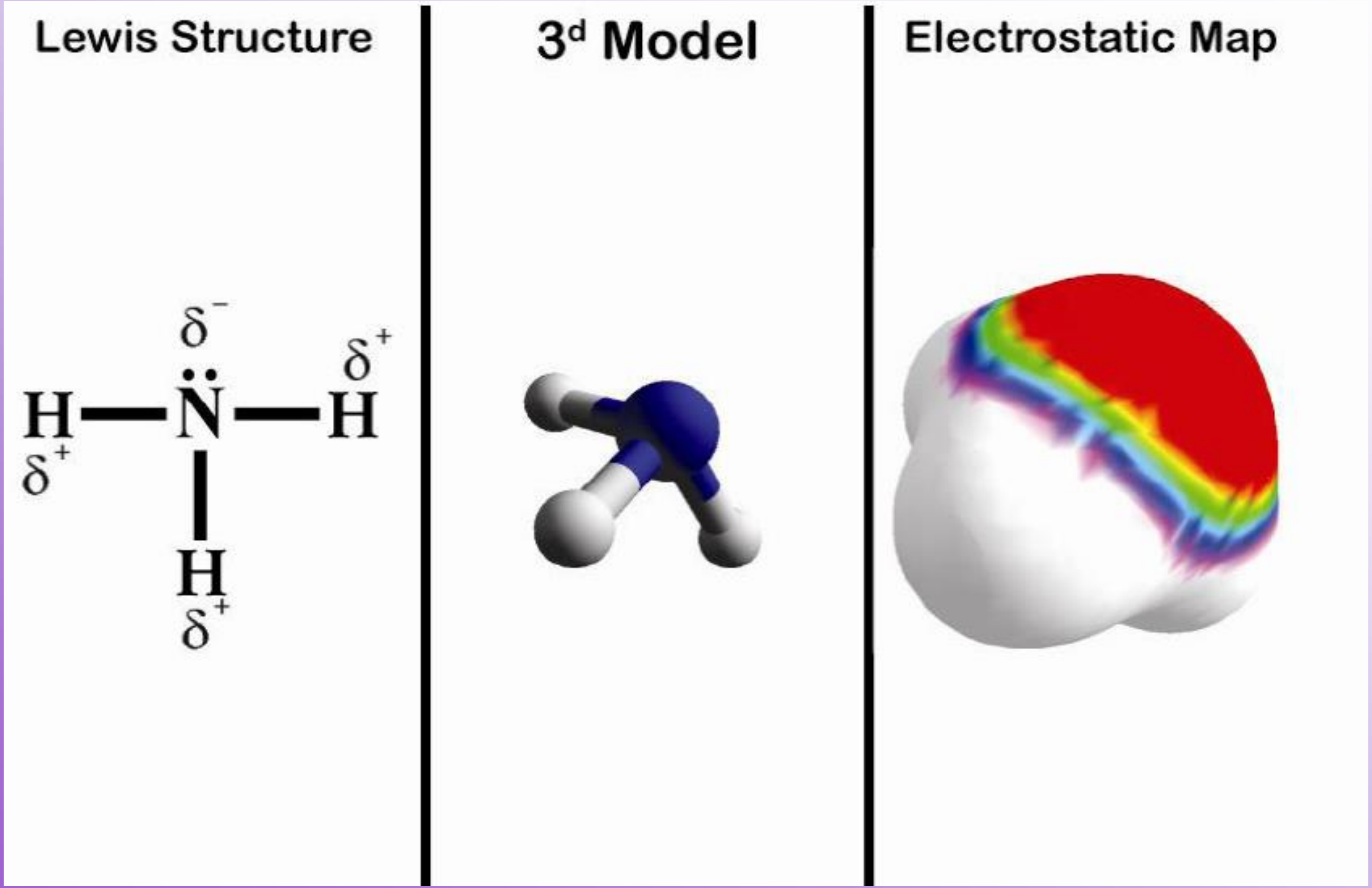
Video Time !!!

- CCC #23 Video: “Polar and Non-polar Molecules”

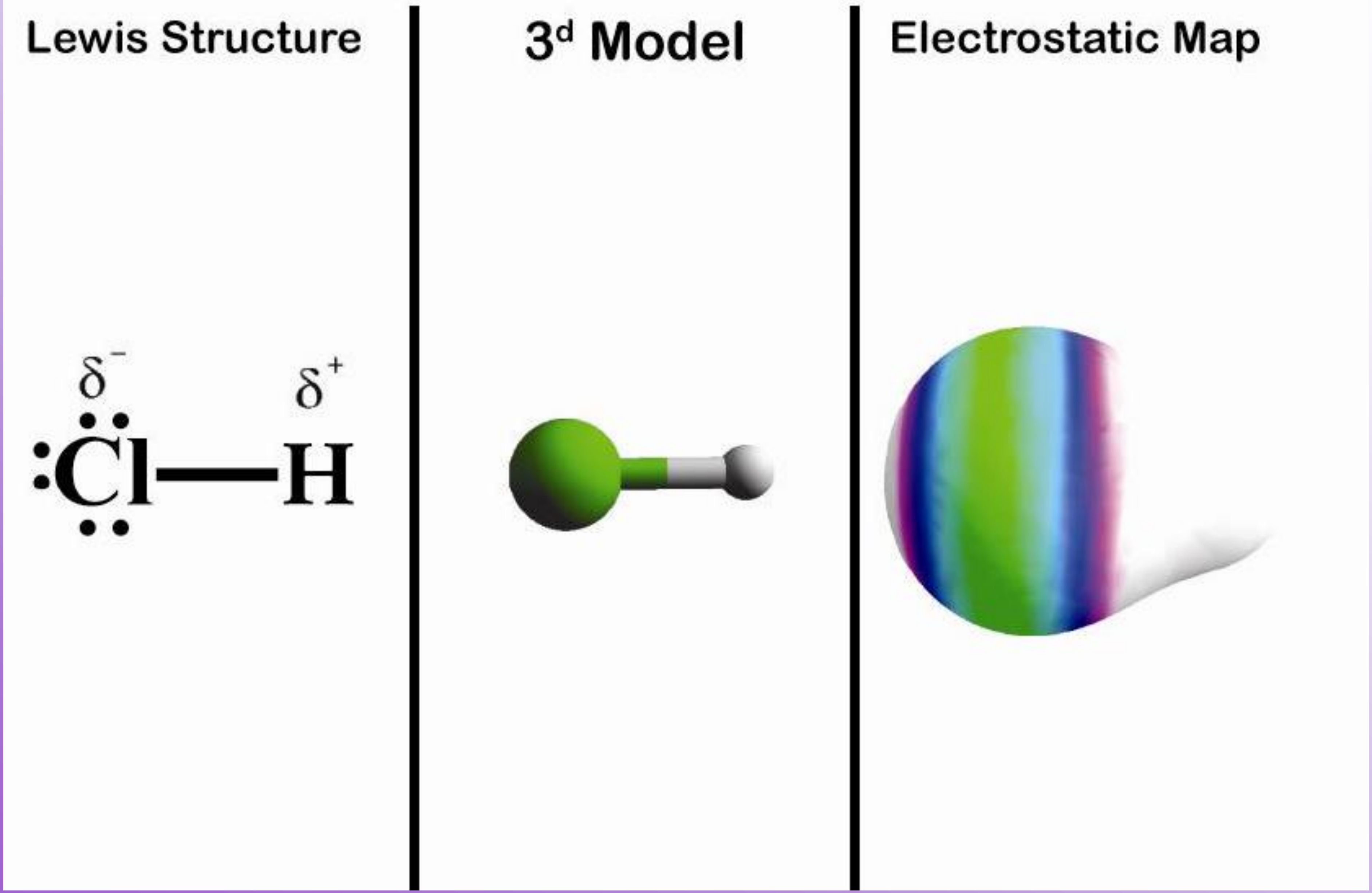
Polar Molecule Example: Water



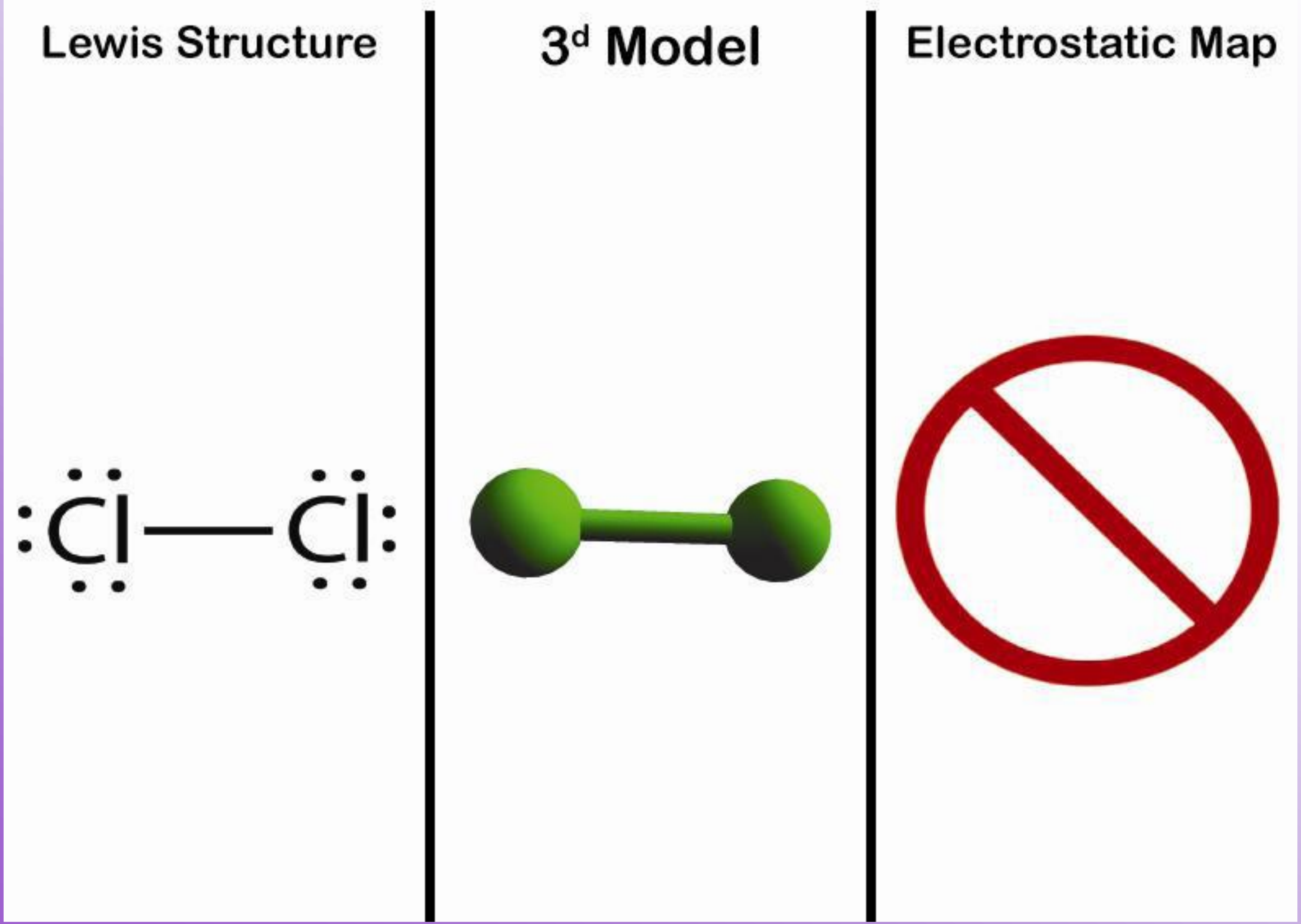
Polar Molecule Example: Ammonia



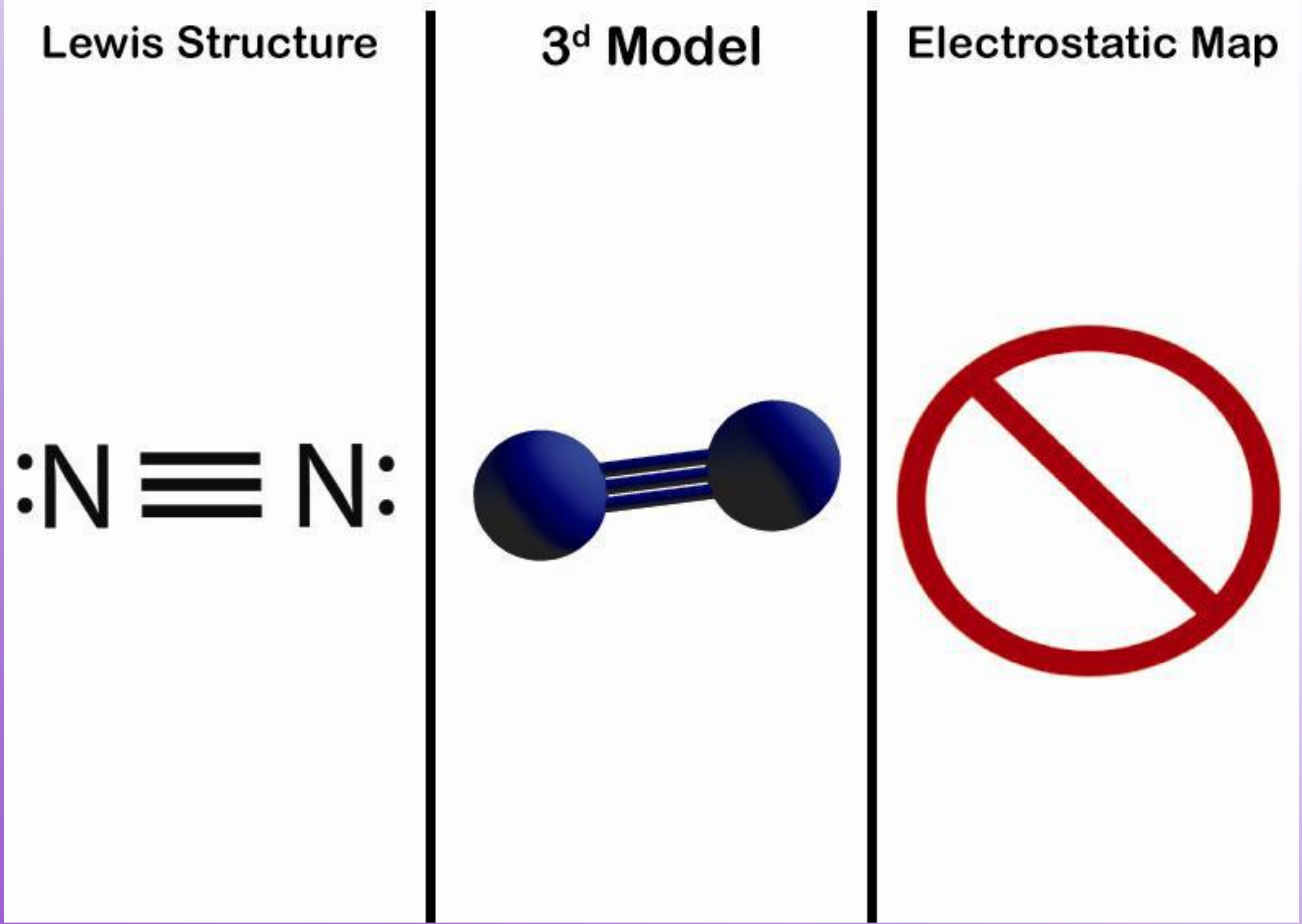
Polar Molecule Example: Hydrochloric Acid



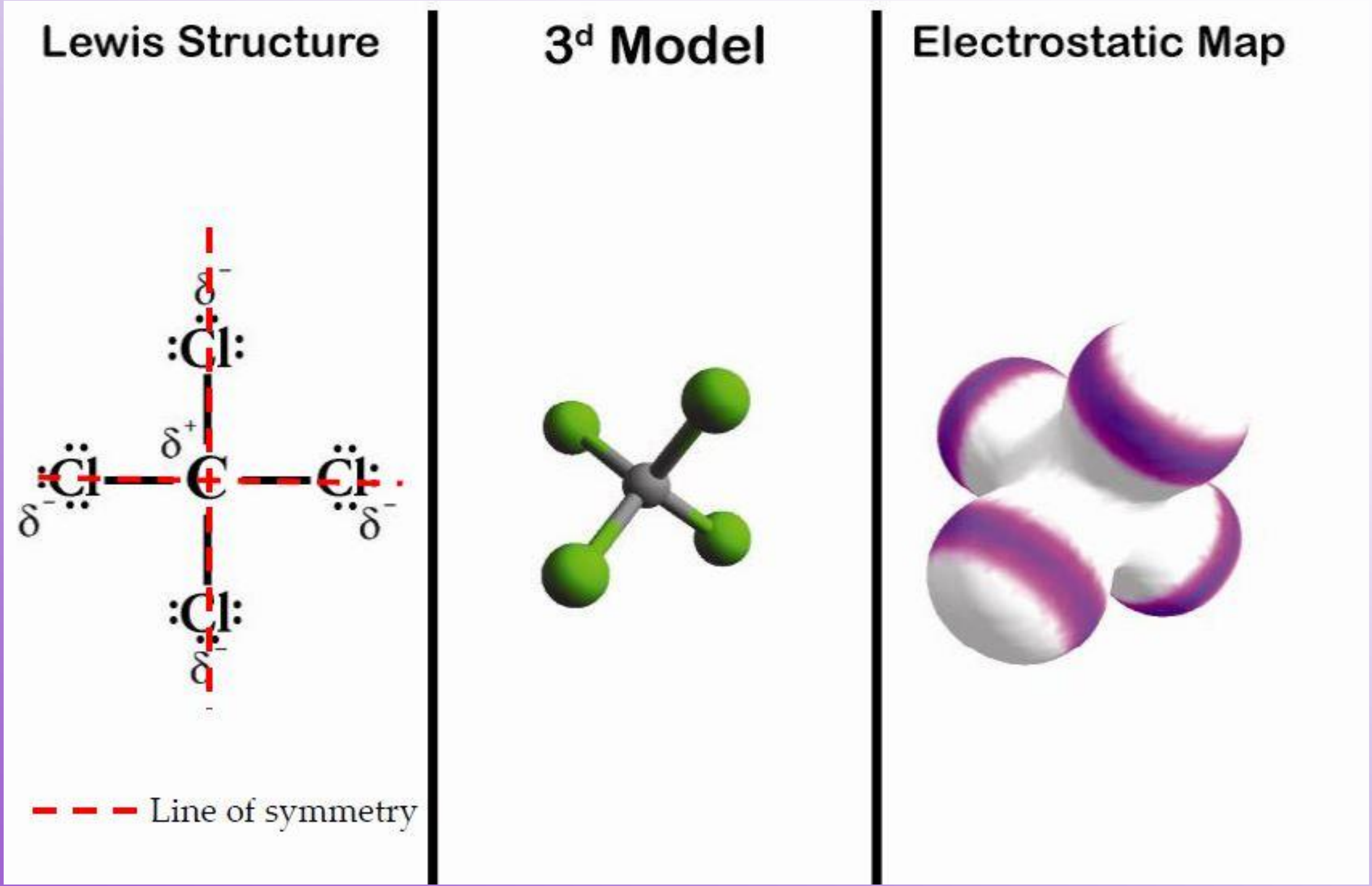
Non-Polar Molecule Example: Chlorine Gas



Non-polar Molecule Example: Nitrogen Gas



Non-polar Molecule Example: Carbon Tetrachloride



Practice Problems: Determining Molecular Polarity

- *Directions: (1) Draw the Lewis (2) label polar bonds if there are any. (3) determine molecular shape. (4) label the molecule as polar or non-polar*

1) PF_3

2) CO_2

3) CH_2O

4) CF_4

5) CI_4

6) HCN

7) OF_2

8) CH_2F_2

Section 4 Additional Resources & Links…

- TEDed Video: “What’s the Shape of a Molecule”
- Tyler Dewitt’s Video: “VSEPR: an Introduction”
- Tyler Dewitt’s Video: “VSEPR: Practice Problems”
- Tyler Dewitt’s Video: “VSEPR: Common Mistakes”
- Khan Academy Video: “Electronegativity and Bonding”
- The Organic Chemistry Tutor Video: “How to Tell if a Molecule is Polar or Non-polar”
- The Organic Chemistry Tutor Video: “Polar or Non-Polar”

Section 5: Polarity & Intermolecular Forces

This part of the Unit is covered on page ??? - ??? in your textbook

Section 5: Polarity & I.M.F. / Objectives

Students Will Be Able too…

- …*define* **intermolecular forces** and *give* at least two ways they can be evaluated.
- …*recall* the 5 types of IMF's discussed in class (**dipole-dipole, hydrogen bonding, dispersion, electrostatic attraction, and Ion-dipole**)
- …*recall* that only O, F, & N can form hydrogen bonds.
- …*recall* the 4 facts about London Dispersion Forces mentioned in class.
- …*Summarize* **Van Der Walls Forces** and *explain* the effect of molar mass on London Dispersion forces.
- …*Determine* the type(s) of intermolecular forces in a given compound or molecule
- …*generalize* the relative strengths of IMF and *make a prediction* about how the boiling points of a short list of substances will compare with one another based on their IMF's.

Intermolecular Forces

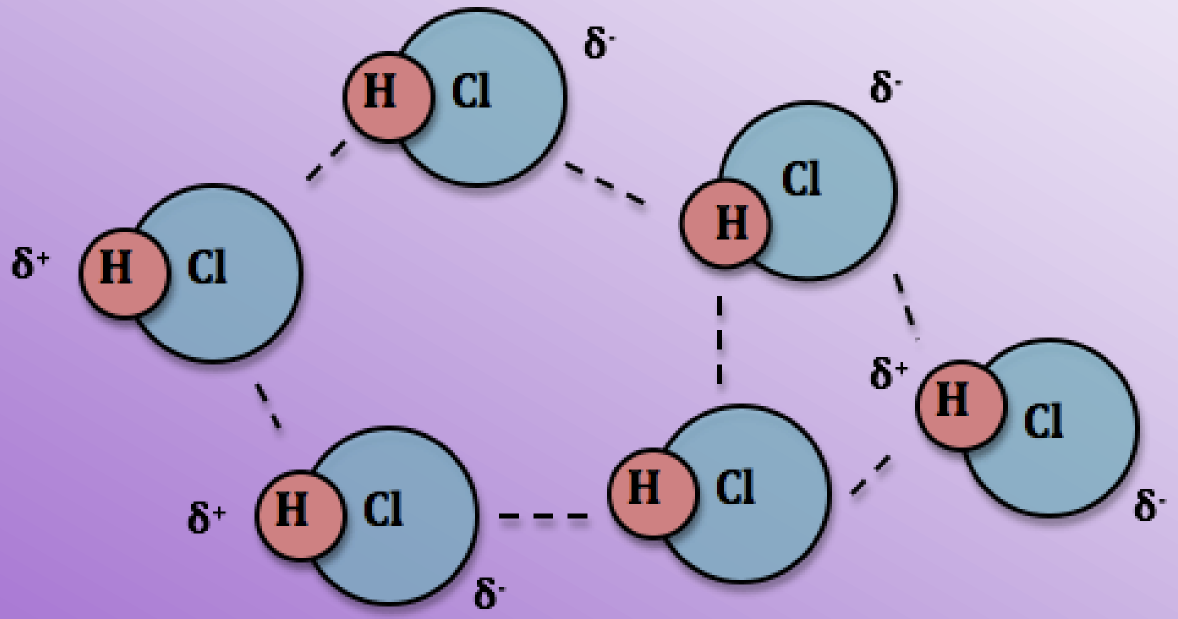
- **Intermolecular Forces** are defined as the forces of attraction between particles (particles could be the units of an ionic compound, molecules, or even atoms)
- There are a couple of ways to evaluate how strongly the particles in a pure substances are attracted to one another but Melting Point and Boiling Point are the most common. Others include, surface tension, viscosity, and vapor pressure.
- Boling Point is probably best because when a substance boils to a vapor it's as if the attraction have been completely overcome. This is because the particles of a gas or vapor are not really attracted to each other at all.

The Five "Types" of Intermolecular Forces

- There are five types of intermolecular forces we will discuss. All five types of intermolecular forces ultimately come down to electrostatic attraction. **Electrostatic attraction** is just a fancy word for what you already know: positive charges attract negative charges.
- The **Van Der Waal's Forces** are found in molecules and include:
  - **Dipole-Dipole interactions**
  - **Hydrogen Bonding**
  - **London dispersion forces**
- Next there is the attractions between the units in an ionic compound. (remember we don't refer to NaCl, Na_3PO_4 or $CaCl_2$ as molecules. we call them units because they are almost always locked in a crystal lattice)
- Lastly**, Ion-dipole interactions**, which is best understood as what happens when salt or other ionic compounds dissolve in a polar molecule like water.

IMF #1: Dipole – Dipole Interactions

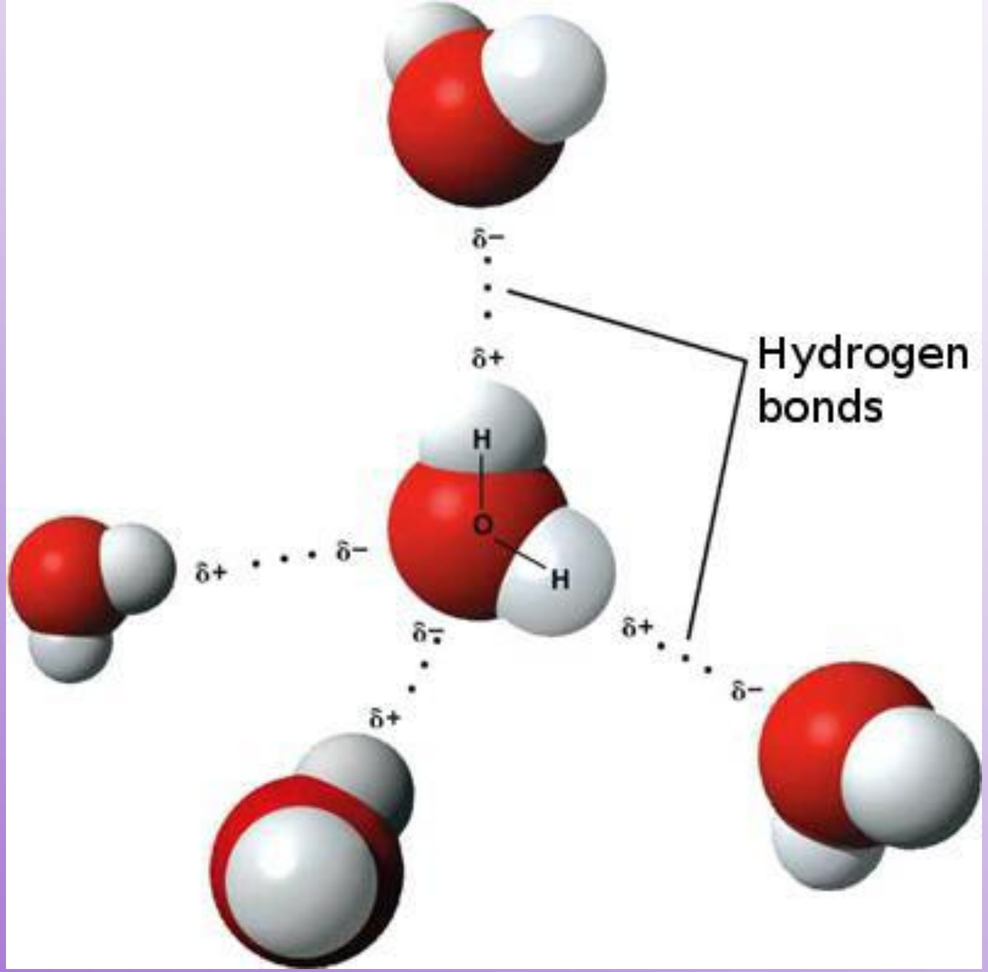
- **Dipole-Dipole interactions** are when the partial positive charge (δ^+) on one molecule is attracted to the partial negative charge (δ^-) on a nearby molecule.
- In order to have dipole-dipole interactions, a molecule needs to have a net dipole. In order words, it has to be a polar molecule with a positive and negative side.



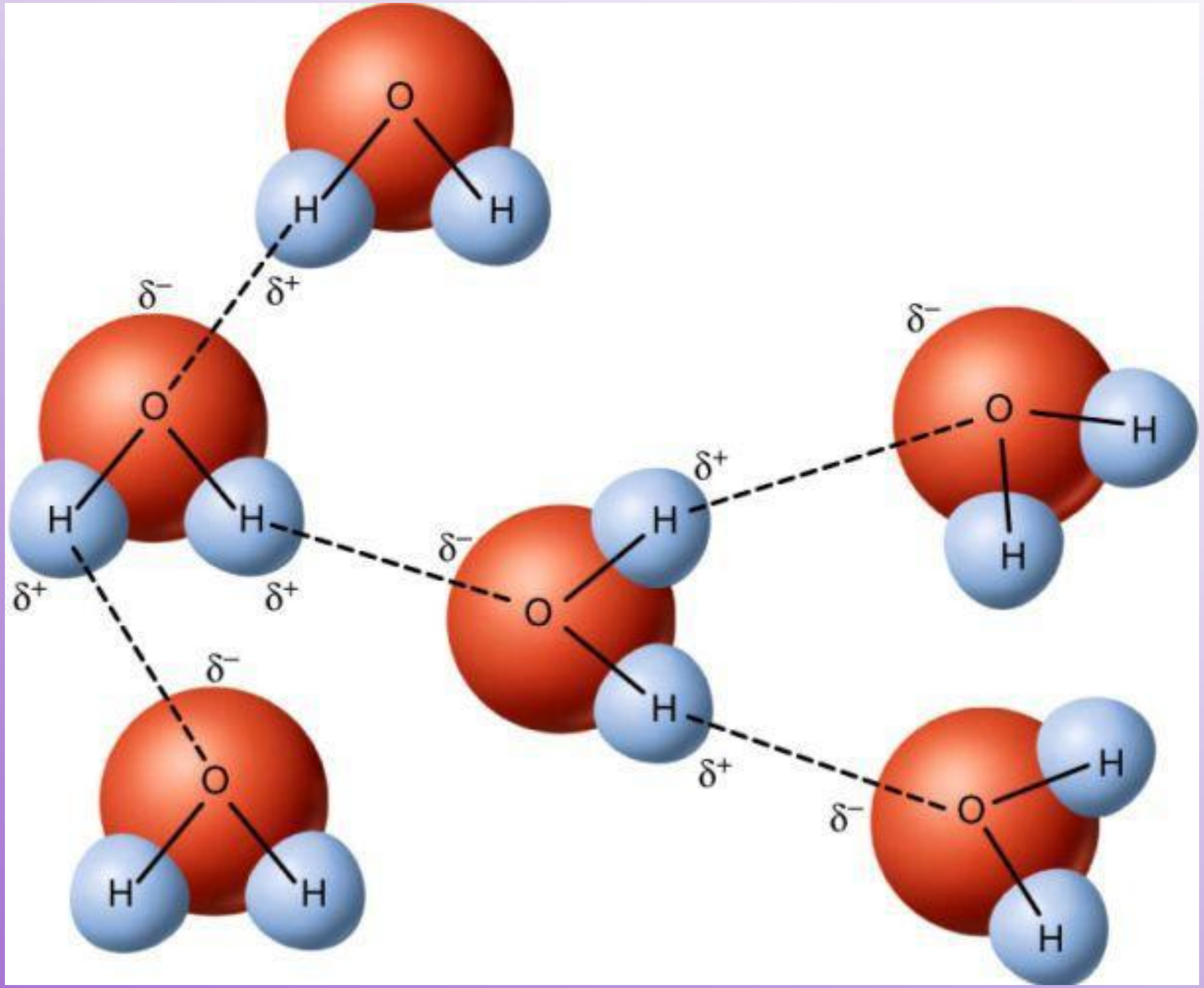
IMF #2: Hydrogen Bonding

- When a O, F, or N is bonded to a Hydrogen atom, the lone pairs of electrons and the very large difference in electronegativity creates very strong dipole-dipole interactions called **Hydrogen Bonding.**
- Hydrogen Bonding is just a stronger form of Dipole-Dipole Interactions.
- Remember that despite it's name, Hydrogen bonding is an I.M.F. and not a chemical bond.

Model of Hydrogen Bonding in Water



Model of Hydrogen Bonding in Water



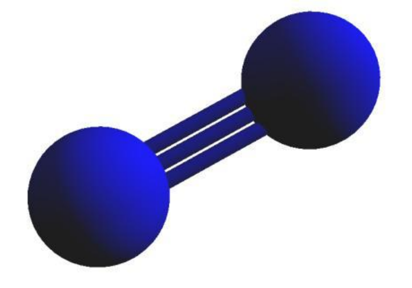
Video Time !!!

- TedEd Video: “How Polarity Makes Water Behave Strangely”
- TedEd Video: “How do Geckos Defy Gravity”

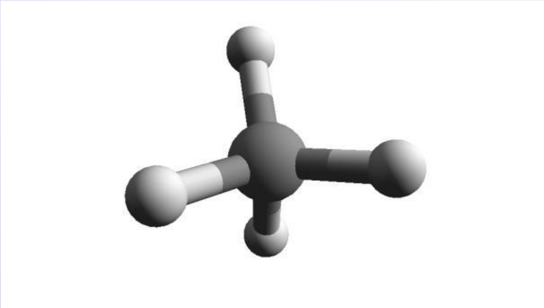
IMF #3: Dispersions Forces

- Dispersion Forces go by several names:
  - London Dispersion Forces
  - London Forces
  - Induced Dipole Forces
- Here are the things you need to know about Dispersion Forces:
  - They are present in all molecules. So even polar molecules that have Dipole-Dipole interactions have Dispersion Forces too.
  - They are the only IMF in non-polar molecules.
  - Generally, they are the weakest of all the intermolecular forces. This is why non-polar molecules all tend to be gases at STP.
  - They get stronger as the molar mass of a molecule increases. So some non-polar molecules with very large molar masses like candle wax ($C_{25}H_{52}$) has a higher boiling point than water despite the fact that Candle Wax's only IMF is dispersion forces. This fact is experimentally verified by looking at the boiling points of several substances.

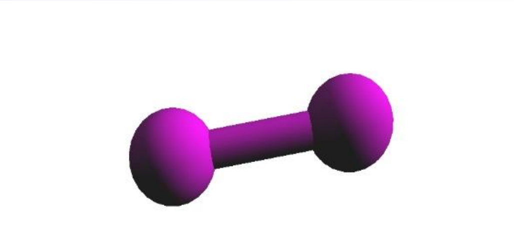
Molecules with Dispersion Force Only (All Gases @ STP)



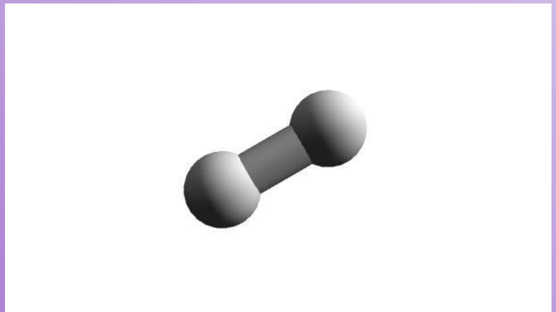
N_2



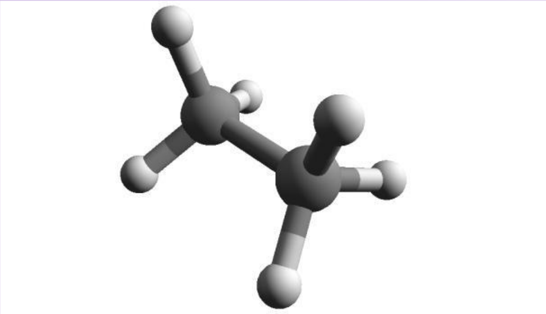
CH_4



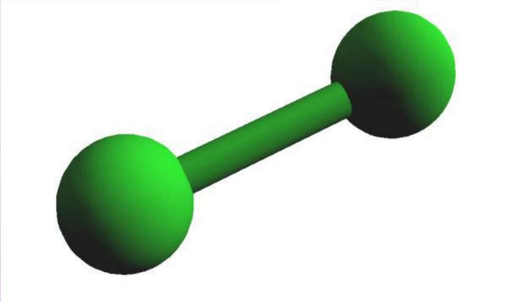
F_2



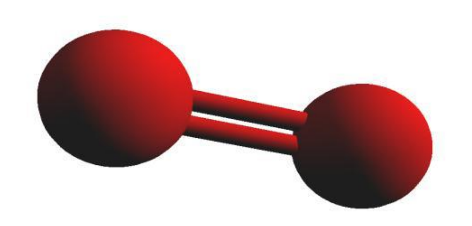
H_2



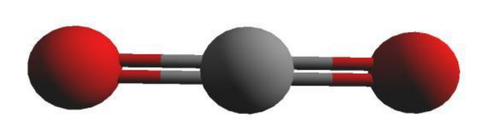
C_2H_6



Cl_2



O_2



CO_2

He

Boiling Point of Molecules with Identical Polarities but Different Molar Masses

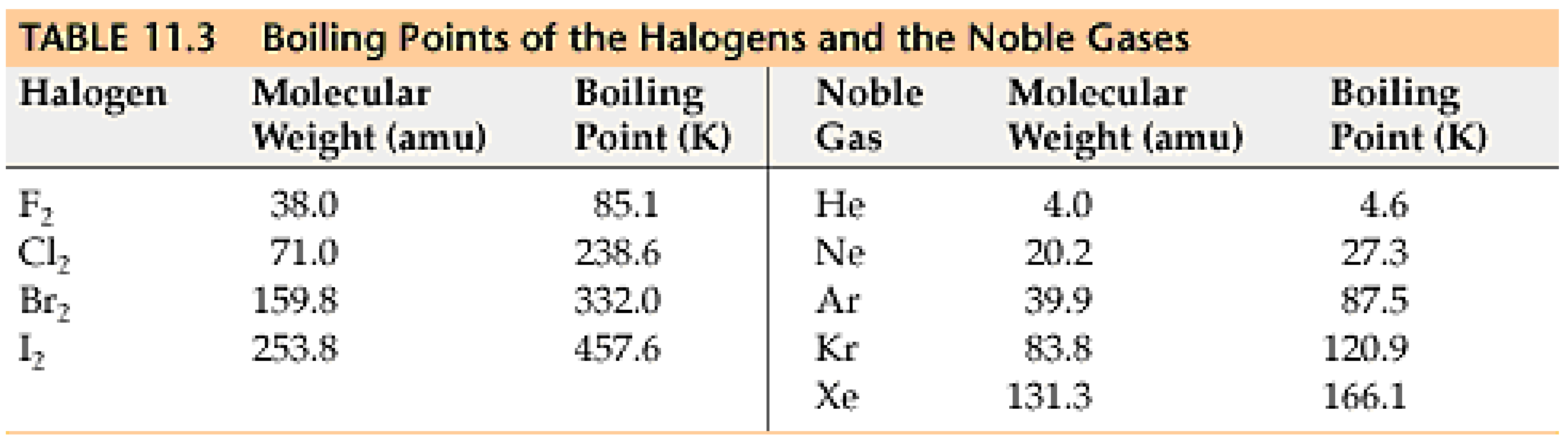


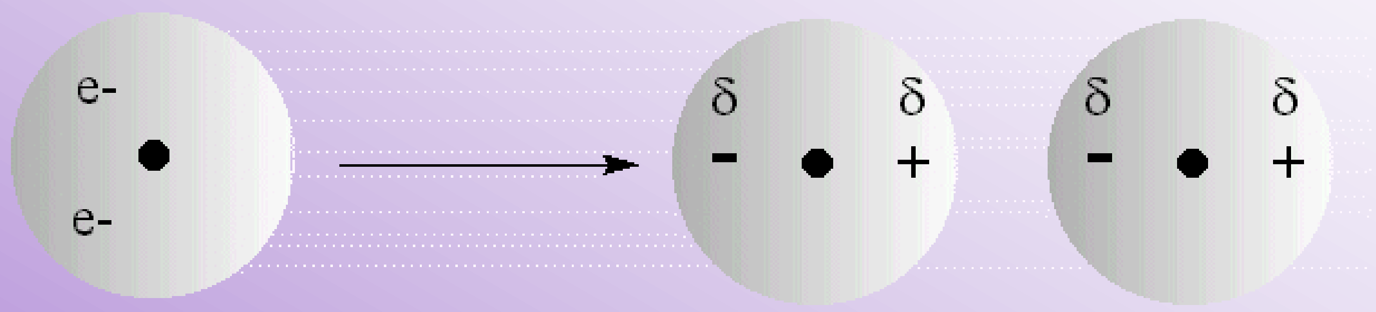
TABLE 11.3 Boiling Points of the Halogens and the Noble Gases

Halogen	Molecular Weight (amu)	Boiling Point (K)	Noble Gas	Molecular Weight (amu)	Boiling Point (K)
F_2	38.0	85.1	He	4.0	4.6
Cl_2	71.0	238.6	Ne	20.2	27.3
Br_2	159.8	332.0	Ar	39.9	87.5
I_2	253.8	457.6	Kr	83.8	120.9
			Xe	131.3	166.1

IMF #3: Dispersions Forces, What Causes Them?

- Dispersion Forces are caused by dipoles just like dipole-dipole interactions, because non-polar molecules can form temporary dipoles. This happens when the distribution of electrons around one of the atoms in the molecule becomes uneven at random.
- It's best to think of it as a series of steps using a simple substances like Neon.
  1) The electrons surrounding the Neon nucleus are moving. At any given point you may have more of the electrons on the right side of the nucleus compared to the left.
  2) The right side of the atom will therefore get a partial negative charge
  3) The left side of the atom will therefore get a partial positive charge
  4) If that neon atom is close to another neon atom, and those atoms are moving slowly enough, it's partial negative charge can induce a dipole in the near by neon atom.

Model of London Dispersion Forces

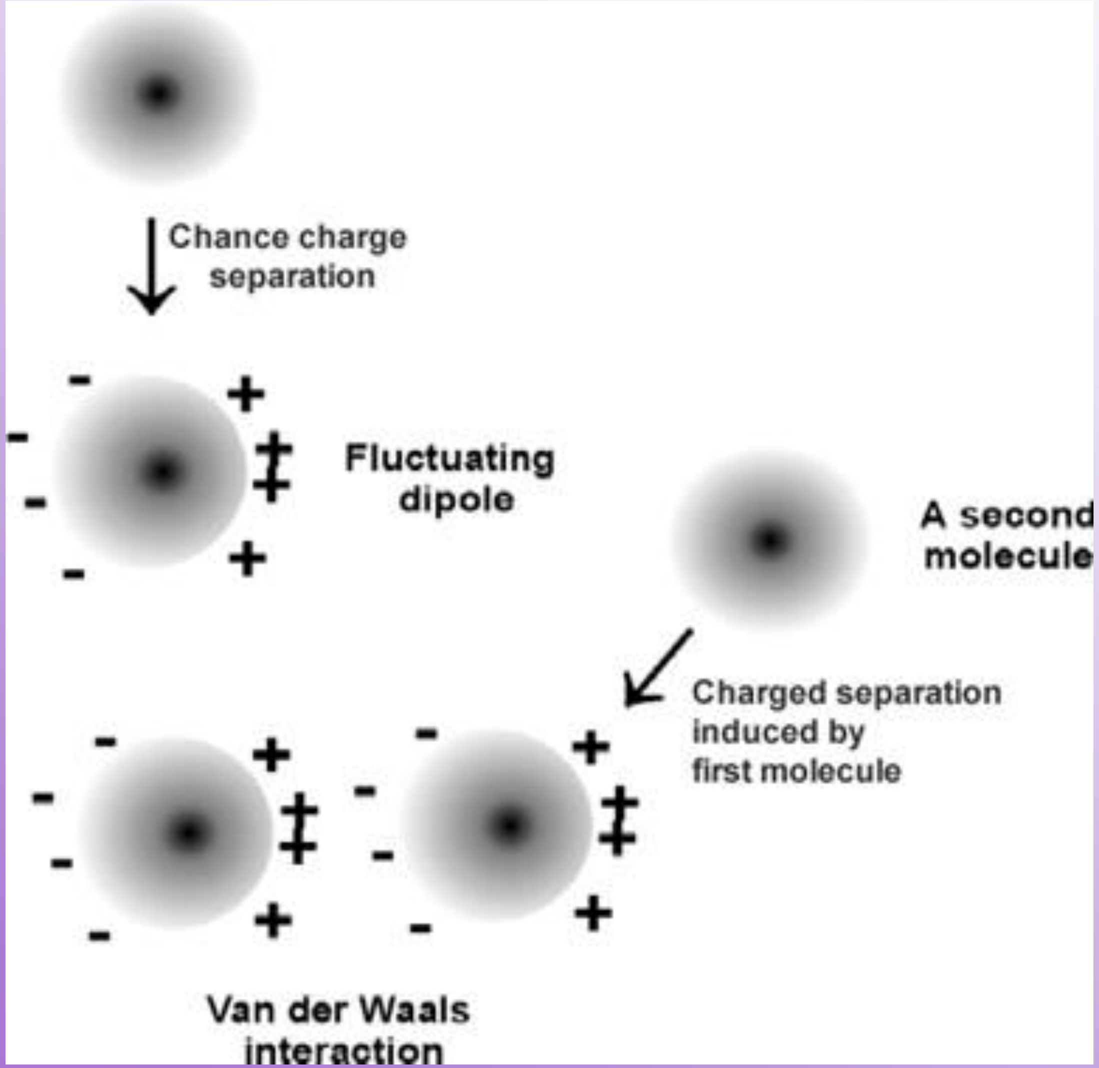


Instantaneous
uneven distribution
of electrons in He atom

Instantaneous
dipole

Induced dipole
on neighboring
He atom.
Resultant attractive
force

Model of London Dispersion Forces



Video Time !!!

- CCC #26 Video: ″Liquids”

Practice Problems: Determining a Molecules IMF's

Directions: Do the following for each molecule (1) Draw the Lewis Structure, (2) Determine if the molecule is Polar or Non-polar, (3) Determine the types of IMF's in each molecule. (4) Predict highest to lowest B.P. (1 highest, 6 lowest)

1) PF_3
2) NH_3
3) CO_2
4) CH_4
5) HCN
6) OF_2

IMF #4: Ionic Bonding

- Unlike the previous three types of intermolecular forces, Ionic bonding is not a van Der Wall's force.
- The idea is the same though, attractions between the positive sides of one unit attract the negative side of a nearby unit.
- Since ionic compounds have full charges and not partial charges, their attractions are much stronger than the attractions in Van Der Walls Forces.
- This is why Ionic compounds have extremely high melting and boiling points. It is also why ionic compounds are locked in a crystal lattice at anywhere close to normal temperature and pressures.
- Molecules can form a crystal lattice too, but only at low enough temperatures. Water for example has a crystal lattice structure when it forms ice. The Hydrogen bonding between water molecules is not very strong, but if the temperature gets low enough it is sufficient to "lock" the molecules in place.

Model of the Crystal Lattice Structure in Table Salt

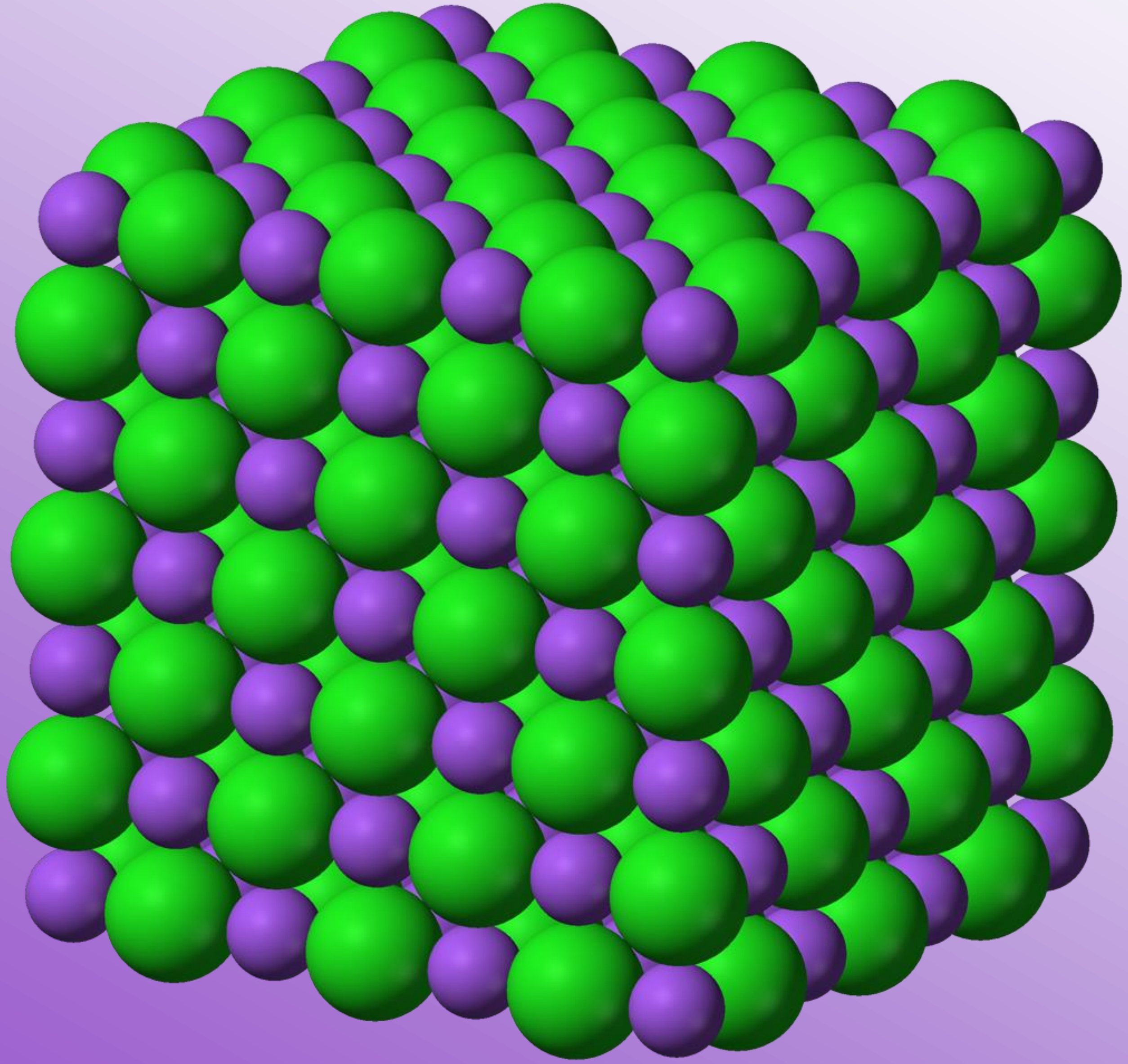


Image Credit: wikipedia

Model of the Crystal Lattice Structure in Calcium Oxide

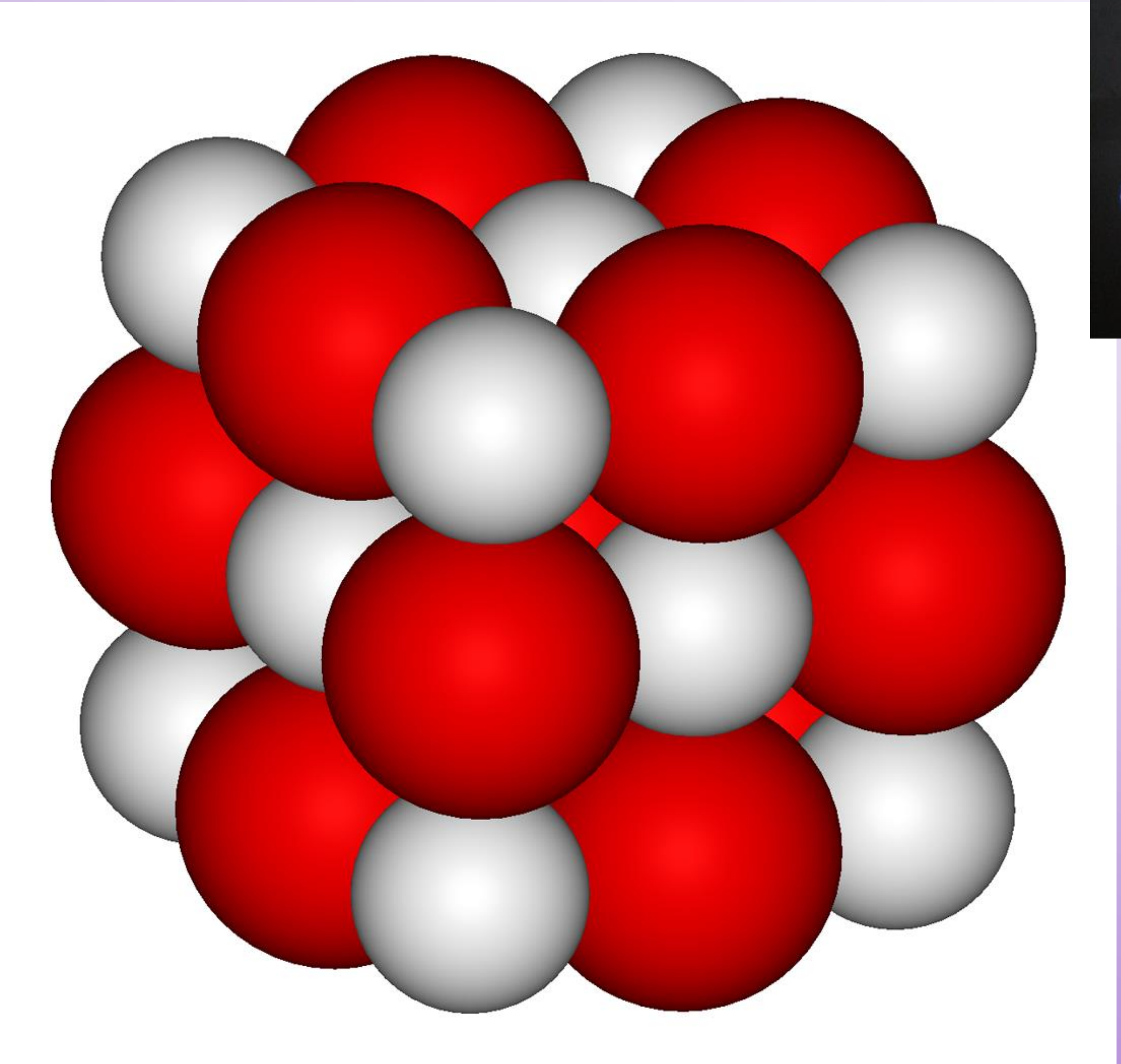


Image Credit: wikipedia

Model of the Crystal Lattice Structure in $CuSO_4$

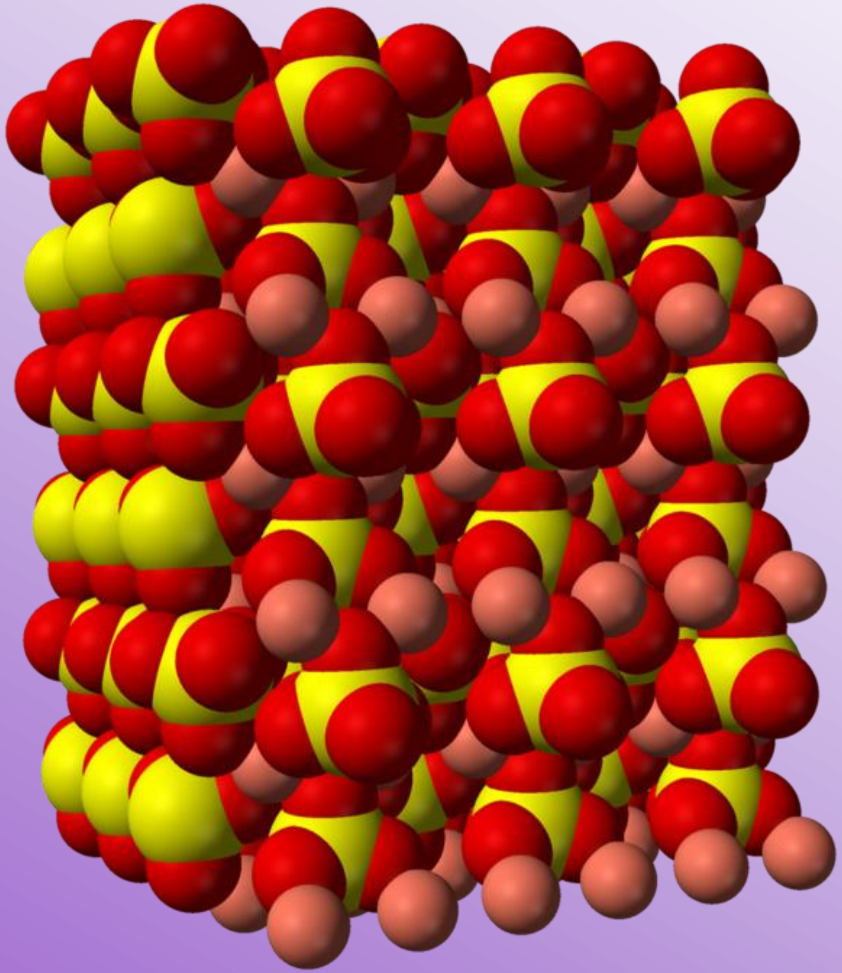
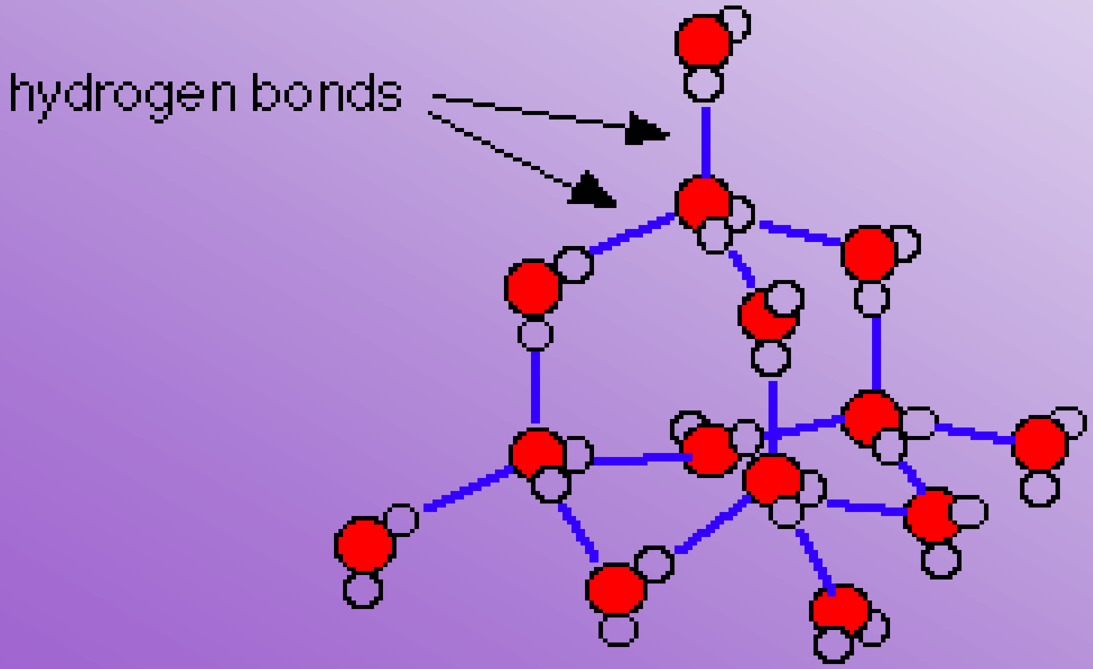


Image Credit: wikipedia

Model of the Crystal Lattice Structure in Water



IMF #5: Ion-Dipole Interactions

- Note that even though it takes very high temperatures to separate NaCl Units from one another, Salt and most other ionic compounds easily dissolve in water. In other words, it take tremendous heat to break apart the crystal lattice of table salt and get it to melt, but the crystal lattice can just as easily be broken apart by simply placing the table salt in some water.
- This is because the positive Ion in the compound is attracted to the negative dipole of the water molecule, and the negative ion is attracted to the positive dipole.
- The dissolving process can be modeled using chemical equations:
  - $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$
  - $CaCl_2(s) \rightarrow Ca^{2+(}aq) + 2\ Cl^-(aq)$
  - $Na_3PO_4(s) \rightarrow 3\ Na^+(aq) + PO_4^{3-}(aq)$
- Unlike the first four types of intermolecular forces we discussed, ion-dipole interactions don't exist in pure substances. Instead, **Ion-dipole interactions** are what you get when you mix two substances together. Specifically, it's what you get when you mix a soluble ionic compound and water.
- Ion-Dipole interactions explain why some substances dissolve in water and other don't. We will talk more about this in the next unit.

Video Time !!!

- McGraw Hill Video: "NaCl Dissolving in Water"

Section 5 Additional Resources & Links…

- Tyler Dewitt Video: “What happens when stuff dissolves?”

Section 7: Nomenclature

This part of the Unit is covered on page 218 – 226, 251, & 252 in your textbook

Section 7: Nomenclature / Objectives

Students Will Be Able too…

- …*explain* some of the basic rules for naming ionic compounds
- …*write* the chemical formula for an ionic compound when given it's name
- …*write* the name for a ionic compound when given it's formula.
- …*explain* some of the basic rules for naming molecules
- …*write* the name for a molecule when given a chemical formula.
- …*write* the formula for a molecule when given it's name.
- …*create* a flow chart for naming ionic compounds and molecules.

First Names in Ionic Compounds

- Ionic Compounds have a first name and a last name. The first name is the cation and the second name is the anion.
- If the cation is NOT multivalent meaning it is…
  - From column 1
  - From column 2
  - Ag^+, Al^{3+}, Cd^{2+}, or Zn^{2+}
- Than the first name of the ionic compound is just the element's name.
- Examples:
  - $NaCl$ = "*Sodium* Chloride"
  - $CaCl_2$ = "*Calcium* Chloride"
  - $AlCl_3$ = "*Aluminum* Chloride"
- If the cation is from anywhere else on the periodic table (multivalent), the first name of the ionic compound is the element's name followed by a roman numeral indicating the charge of the cation.
- Examples:
  - $FeCl_2$ = "*Iron(II)* Chloride"
  - $FeCl_3$ = "*Iron(III)* Chloride"
  - $TiCl_2$ = "*Titanium(II)* Chloride"

Last Names in Ionic Compounds

- If the anion is…
  - From Columns 5,6, or 7
- Than the last name of the ionic compound is the elements name but the ending changes to "-ide"
- Examples:
  - Na_3P = "Sodium *Phosphide*"
  - Na_2O = "Sodium *Oxide*"
  - NaBr = "Sodium *Bromide*"
- If the anion is a polyatomic ion than the last name is the name of the polyatomic ion.
- Examples:
  - Na_3PO_4 = "Sodium *Phosphate*"
  - Na_2CO_3 = "Sodium *Carbonate*"
  - $NaNO_3$ = "Sodium *Nitrate*"

Practice Problems: Writing Compound Formulas

Directions: Write the formulas for the following ionic compounds.

1) Lithium Phosphate
2) Osmium(IV) Chlorate
3) Aluminum Cyanide
4) Copper(I) Sulfide
5) Ruthenium(III) Iodide
6) Sodium Bicarbonate
7) Chromium(III) Sulfate
8) Iron(VI) Nitride
9) Magnesium Dichromate
10) Tin(IV) Oxide

Practice Problems: Writing Compound Formulas

Directions: Write the formulas for the following ionic compounds.

1) Iridium(II) Nitride
2) Vanadium(V) Sulfide
3) Sodium Carbonate
4) Niobium(V) Fluoride
5) Barium Phosphate
6) Zirconium(III) Bromide
7) Iron(III) Carbonate
8) Cadmium Bromide
9) Molybendium(III) Cyanide
10) Iridium(IV) Sulfate

Practice Problems: Writing Compound Names

Directions: Give the names for the following ionic compounds.

1) CdF_2

2) Au_2S_3

3) $CuSO_4$

4) $CuCl$

5) $Zn_3(PO_4)_2$

6) $CrCl_2$

7) $CrPO_4$

8) Pd_3P_2

9) $Ti_3(PO_4)_4$

10) $Eu(NO_3)_3$

Practice Problems: Writing Compound Names

Directions: Give the names for the following ionic compounds.

1) VCO_3
2) $V_2(SO_4)_5$
3) $Cr_2(SO_4)_3$
4) $CuCl_2$
5) $Zn(OH)_2$
6) $SnCl_4$
7) Na_3PO_4
8) Ag_3P
9) $Ti(CO_3)_2$
10) $NiNO_3$

Naming Binary Molecular Compounds

- Just like Ionic compounds, molecules have a first and last names.
- Just like Ionic compounds, the first name is the first element in the compound and the second name is a non-metal that changes to "-ide"
- The big difference: In molecular compounds pre-fixes are (sometimes) used to indicate how many atoms there are.
- First name/element rules:
  - If there is only 1 it's simply the element's name.
  - If there is more than 1, you need to use a prefix.
- Second name/element rules:
  - Always gets a prefix; even if there is only 1.
  - Always changed to "-ide" at the end.
- These rules work for molecules that only contain 2 elements. There are many molecules that contain more than 2 elements but we will not worry about those in Chem A.
- These rules don't work for organic compounds either, which are complex carbon compounds. There is an entire system for naming organic compounds but again, we are not going to worry about that in Chem A.

Prefixes Used in Naming Molecules

- Here are the first 10 pre-fixes. You will be given these prefixes on your assessments and do not have to worry about memorizing them:
  - 1 atom = “mono”
  - 2 atoms = “di”
  - 3 atoms = “tri”
  - 4 atoms = “tetra”
  - 5 atoms = “penta”
  - 6 atoms = “hexa”
  - 7 atoms = “hepta”
  - 8 atoms = “octa”
  - 9 atoms = “nona”
  - 10 atoms = “deca”

Molecule Name Examples

- SiF_4 = “Silicon **tetra**flouride”
- CO = “Carbon **mono**xide”
- S_2O = “**Di**sulfur **mono**xide”
- CO_2 = “Carbon **di**oxide”
- PBr_5 = “Phosphorus **penta**bromide”
- H_2O = “**Di**hydrogen **mono**xide”
- P_2Cl_8 = “**Di**phosphorus **octa**chloride”

Practice Problems: Writing Binary Molecular Formulas

Directions: Give the formula for the following molecules

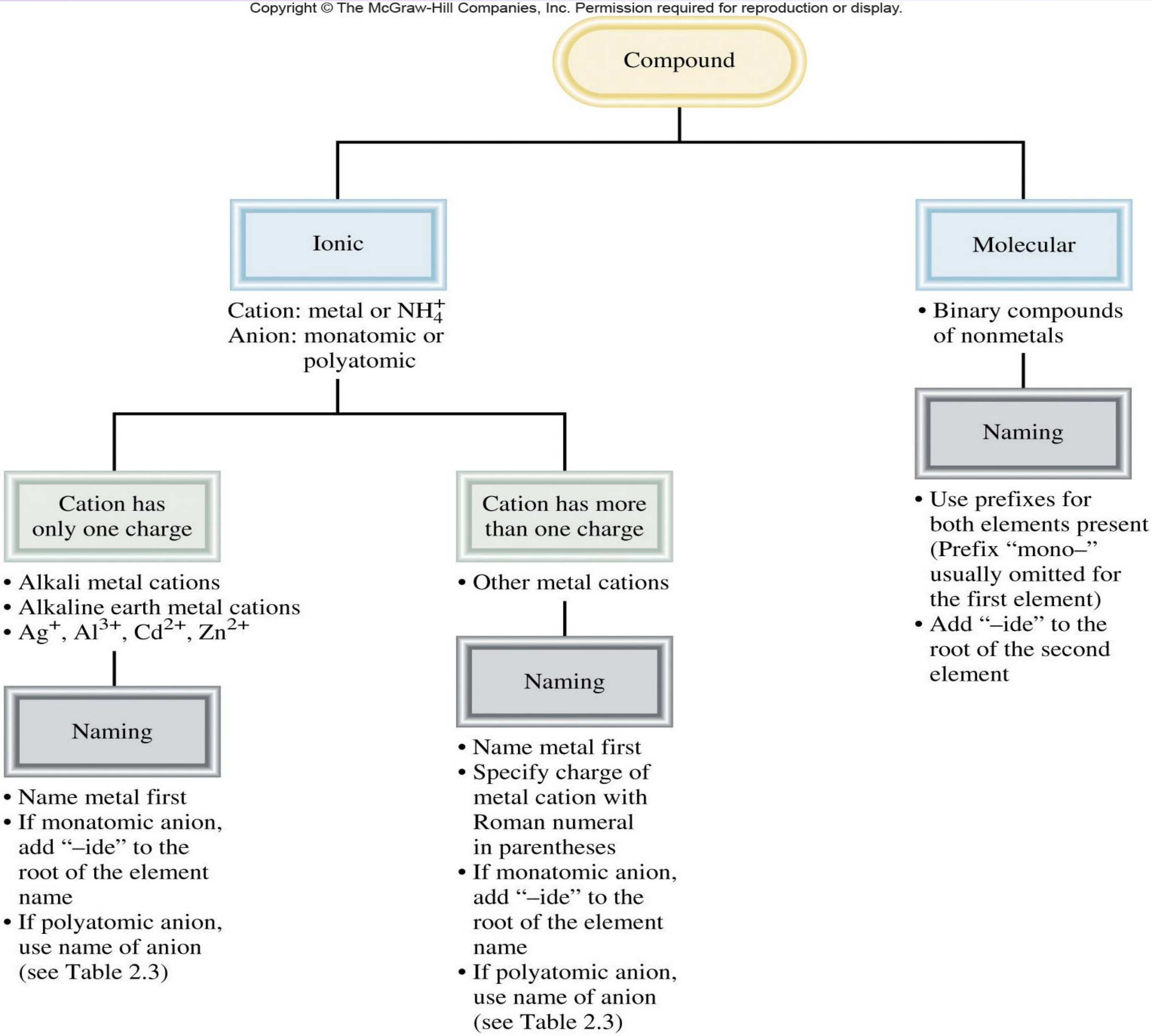
1) Diphosphorus Pentoxide
2) Silicon Dioxide
3) Carbon Tetrabromide
4) Dinitrogen Monoxide
5) Sulfur Dioxide
6) Phosphorus Pentabromide
7) Iodine Trichloride
8) Nitrogen Triiodide
9) Dinitrogen Trioxide

Practice Problems: Naming Binary Molecular Compounds

Directions: Give the name for the following molecules

1) CS_2
2) OF_2
3) PCl_3
4) NF_3
5) N_2O_5
6) SCl_6
7) SO_3
8) NO
9) P_2F_6

Nomenclature Flow Chart #1



Compound

Ionic

Cation: metal or NH_4^+
Anion: monatomic or polyatomic

Cation has only one charge

- Alkali metal cations
- Alkaline earth metal cations
- Ag^+, Al^{3+}, Cd^{2+}, Zn^{2+}

Naming

- Name metal first
- If monatomic anion, add "–ide" to the root of the element name
- If polyatomic anion, use name of anion (see Table 2.3)

Cation has more than one charge

- Other metal cations

Naming

- Name metal first
- Specify charge of metal cation with Roman numeral in parentheses
- If monatomic anion, add "–ide" to the root of the element name
- If polyatomic anion, use name of anion (see Table 2.3)

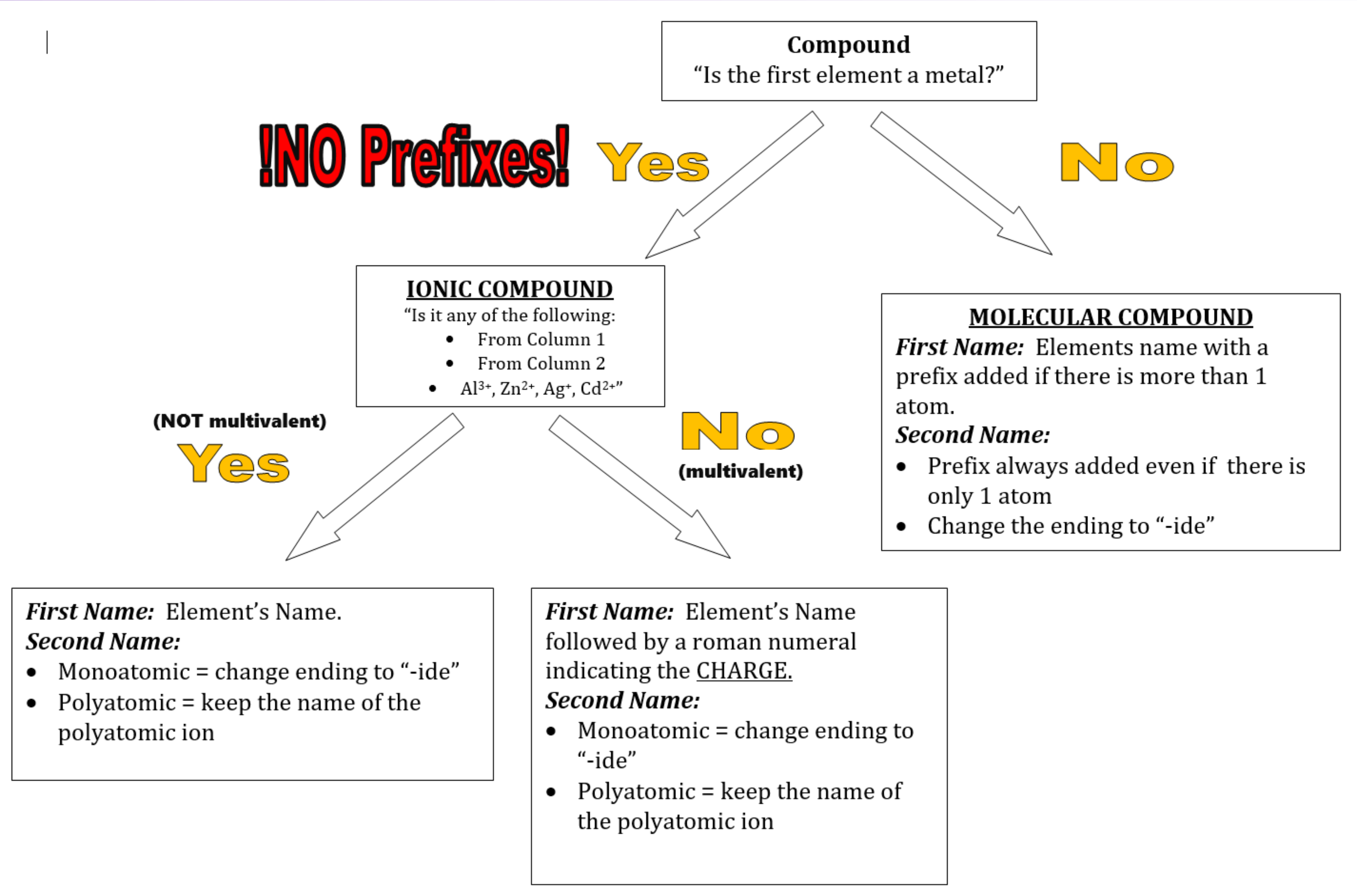
Molecular

- Binary compounds of nonmetals

Naming

- Use prefixes for both elements present (Prefix "mono–" usually omitted for the first element)
- Add "–ide" to the root of the second element

Nomenclature Flow Chart #2



Compound
"Is the first element a metal?"

!NO Prefixes!

Yes

No

IONIC COMPOUND
"Is it any of the following:
- From Column 1
- From Column 2
- Al^{3+}, Zn^{2+}, Ag^{+}, Cd^{2+}"

(NOT multivalent)
Yes

No
(multivalent)

MOLECULAR COMPOUND

First Name: Elements name with a prefix added if there is more than 1 atom.

Second Name:
- Prefix always added even if there is only 1 atom
- Change the ending to "-ide"

First Name: Element's Name.

Second Name:
- Monoatomic = change ending to "-ide"
- Polyatomic = keep the name of the polyatomic ion

First Name: Element's Name followed by a roman numeral indicating the CHARGE.

Second Name:
- Monoatomic = change ending to "-ide"
- Polyatomic = keep the name of the polyatomic ion

Practice Problems: Nomenclature Mix

Directions: Give the name for the following substances. You will first need to determine if the are ionic compounds or molecules.

1) P_4S_8

2) $BaSO_3$

3) N_2O_4

4) $ZnCl_2$

5) Ni_2O

6) $MoCl_3$

7) S_3O_5

8) YBr_3

9) S_3O_6

Video Time !!!

- CCC #11 Video: “How to Speak Chemistrian”
  Note: This video will discuss naming of Acids which is not in our objectives.

Section 7 Additional Resources & Links…

- Tyler Dewitt's 6 Video Playlist on Naming Ionic Compounds
- Tyler Dewitt's Video: "Naming Covalent Molecular Compounds"