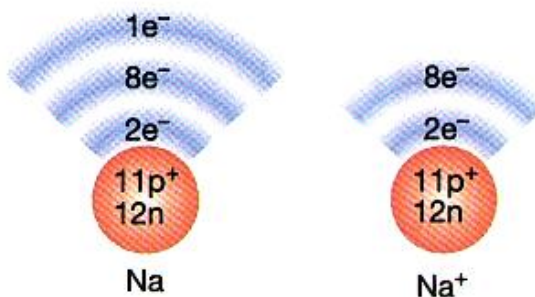


OCTET RULE & IONS

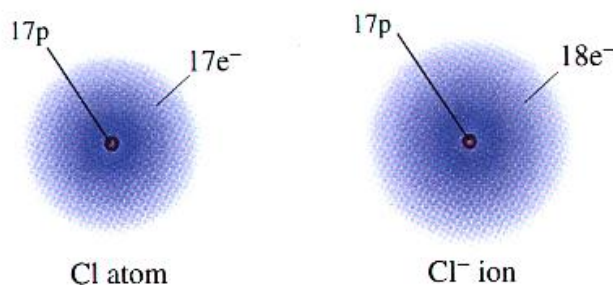
- Most elements, except noble gases, combine to form compounds. Compounds are the result of the formation of chemical bonds between two or more different elements.
- In the formation of a chemical bond, atoms lose, gain or share valence electrons to complete their outer shell and attain a noble gas configuration.
- This tendency of atoms to have eight electrons in their outer shell is known as the *octet rule*.

Formation of Ions:

- An *ion* (charged particle) can be produced when an *atom gains* or *loses* one or more *electrons*.



- A *cation* (+ ion) is formed when a neutral atom *loses an electron*.
- An *anion* (- ion) is formed when a neutral atom *gains an electron*.



IONIC CHARGES

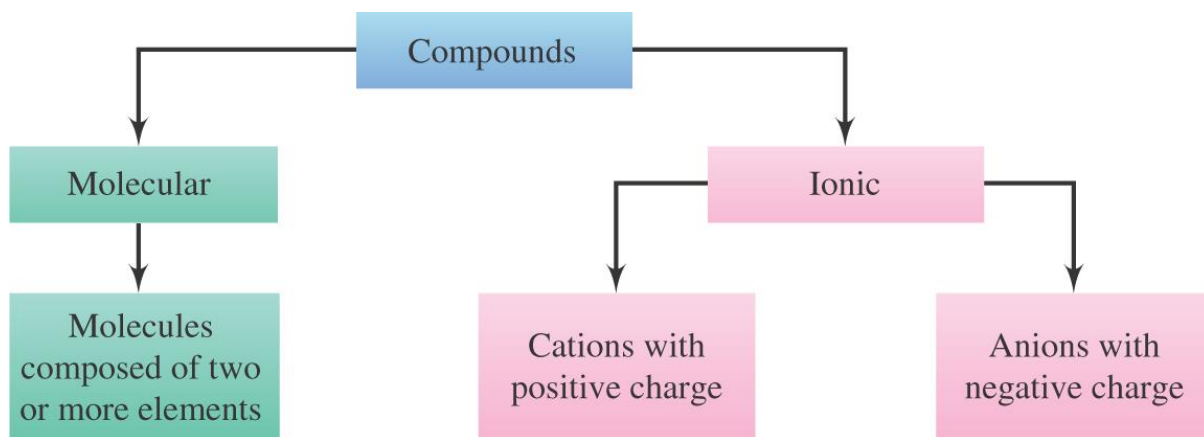
- The ionic charge of an ion is dependent on the number of electrons lost or gained to attain a noble gas configuration.
- For most main group elements, the ionic charges can be determined from their group number, as shown below:

Noble Gases		Metals Lose Valence Electrons			Nonmetals Gain Valence Electrons				Noble Gases
		1A (1)	2A (2)	3A (13)	5A (15)	6A (16)	7A (17)		
He	←	Li ⁺							
Ne	←	Na ⁺	Mg ²⁺	Al ³⁺	N ³⁻	O ²⁻	F ⁻	→	Ne
Ar	←	K ⁺	Ca ²⁺		P ³⁻	S ²⁻	Cl ⁻	→	Ar
Kr	←	Rb ⁺	Sr ²⁺				Br ⁻	→	Kr
Xe	←	Cs ⁺	Ba ²⁺				I ⁻	→	Xe

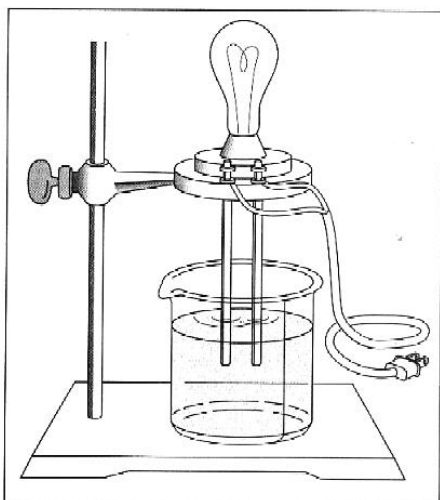
- All other ionic charges need to be memorized and known in order to write correct formulas for the compounds containing them.

COMPOUNDS

- Compounds are pure substances that contain 2 or more elements combined in a definite proportion by mass.
- Compounds can be classified as one of two types: **Ionic** and **molecular (covalent)**



- **Ionic** compounds are formed by combination of a **metal and a non-metal**. The smallest particles of ionic compounds are ions.
- **Molecular** compounds are formed by combination of 2 or more **non-metals**. The smallest particles of molecular compounds are molecules.
- The nature and type of the **chemical bond** is directly responsible for many physical and chemical **properties** of a substance: (e.g. melting point, conductivity)

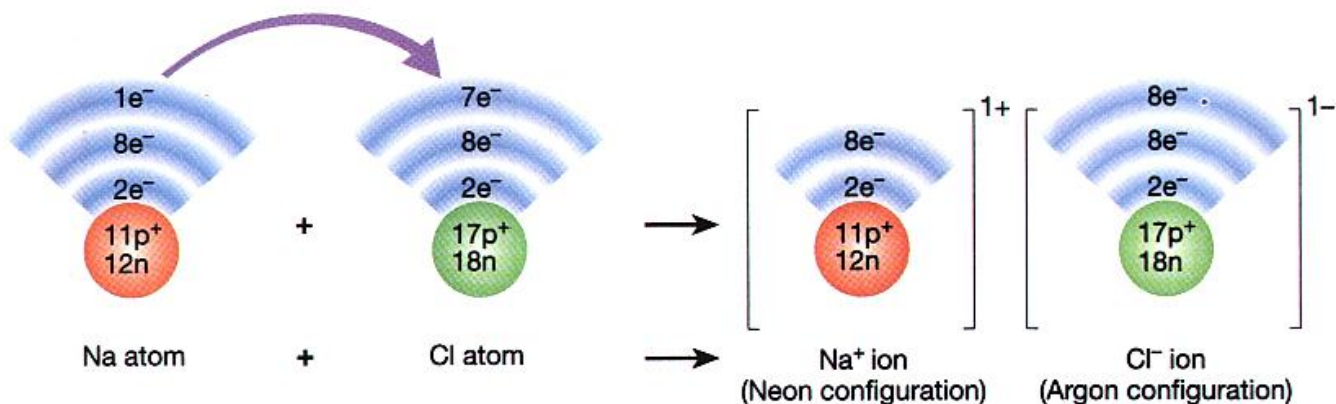


When the **conductivity** apparatus is placed in **salt** solution, the bulb **will light**. But when it is placed in **sugar** solution, the bulb **does not light**.

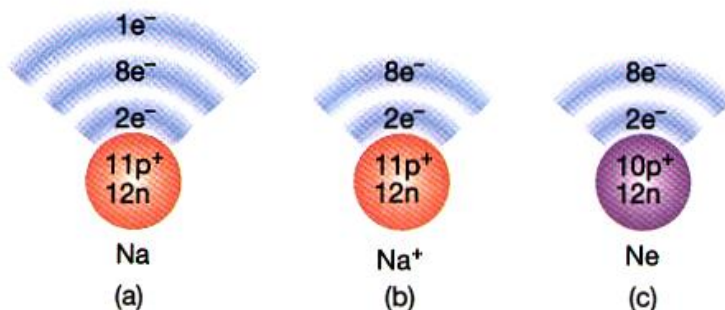
This **difference in conductivity** between salt and sugar is due to the **different types of bonds** between their atoms.

IONIC COMPOUNDS

- **Ionic compounds** contain **ionic bonds**, which occur when electrons are **transferred** between two atoms.



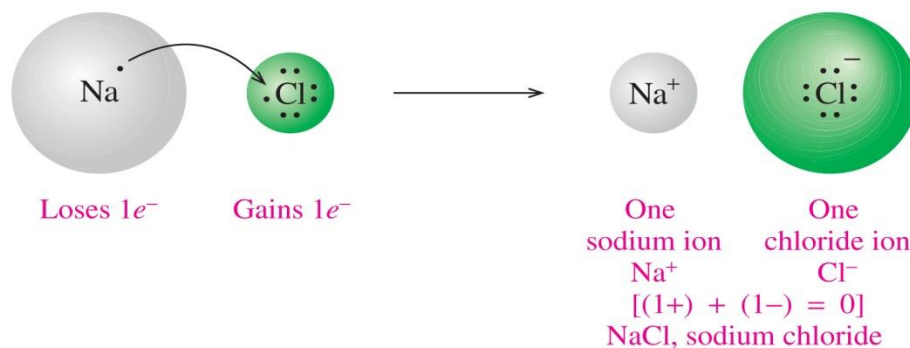
- **Ionic bonds** occur between **metals** and **non-metals**.
- Atoms that lose electrons (**metals**) form positive ions (**cations**).
- Atoms that gain electrons (**non-metals**) form negative ions (**anions**).
- The **smallest** particles of **ionic compounds** are **ions** (not atoms).



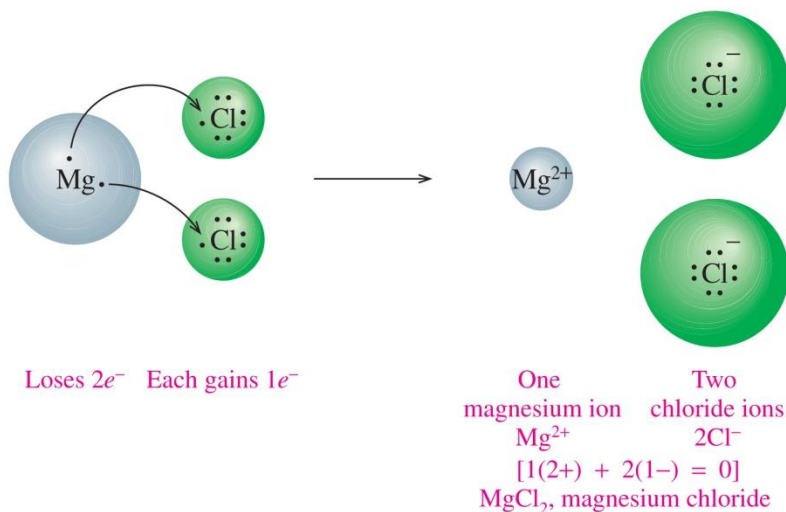
Comparison between sodium atom (a), sodium ion (b) and neon atom (c)

IONIC CHARGES & FORMULAS

- The formula of an ionic compound indicates the number and kinds of ions that make up the ionic compound.
- The sum of the ionic charges in the formula is always zero, which indicates that the total number of positive charges is equal to the total number of negative charges.
- For example, the +1 charge on the sodium ion is cancelled by the -1 charge on the chloride ion, to form a net zero charge.



- When charges between the two ions do not balance, subscripts are used to balance the charges.
- For example, since each magnesium loses 2 electrons, and each chloride gains one electron, 2 chlorides are needed to balance the charge of the magnesium ion. Therefore, magnesium chloride is written as MgCl_2 .



NAMING & WRITING IONIC FORMULAS

- When writing ionic formula, knowing the charge of the ions are important since the net charge on the compound must be zero.
- Some elements produce only one ion (Type I) while others produce two or more ions (Type II).
- Differentiating between type I and II ions is important, since the naming system is different for each. Shown below are the common ions of each type:

1 Group 1A		2 Group 2A											13 Group 3A					14 Group 4A	15 Group 5A	16 Group 6A	17 Group 7A	18 Group 8A	
				3	4	5	6	7	8	9	10	11	12										
				3B	4B	5B	6B	7B	8B		1B	2B											
H ⁺																							
Li ⁺														Al ³⁺									
Na ⁺	Mg ²⁺																						
K ⁺	Ca ²⁺						Cr ²⁺	Mn ²⁺	Fe ²⁺	Co ²⁺	Ni ²⁺	Cu ⁺	Zn ²⁺										
							Cr ³⁺	Mn ³⁺	Fe ³⁺	Co ³⁺	Ni ³⁺	Cu ²⁺											
Rb ⁺	Sr ²⁺											Ag ⁺	Cd ²⁺										
														Sn ²⁺									
														Sn ⁴⁺									
Cs ⁺	Ba ²⁺											Au ⁺	Hg ₂ ²⁺										
												Au ³⁺	Hg ²⁺	Pb ²⁺									
														Pb ⁴⁺									

- Note that most main-group elements are type I (except Sn and Pb), and most transition elements are type II (except Ag, Cd and Zn).

NAMING & WRITING IONIC FORMULAS
Binary Ionic Compounds (Type I):

- *Binary* compounds contain only *two elements*.
- *Type I* are those *cations* that form *only one ion*.
- In these compounds, *charges of the cations must equal the charges of the anions* since the *net charge is zero*.
- *Subscripts* are used to *balance* the charges between cations and anions.

Name	Sodium bromide	Potassium sulfide
Ions	Na⁺, Br⁻	K⁺, S⁻²
Formula	NaBr	K₂S

When naming ionic compounds:

- Name the *cation* first, the *anion* last.
- The *cation* name is the same as the name of the *metal* it forms from.
- The *anion* name takes the *root of non-metal* and the ending “*-ide*”.

MgCl₂ **magnesium chloride**

NaI **sodium iodide**

AlF₃ **aluminum fluoride**

Examples:

1. Write formulas for the following ionic compounds:

calcium chloride: _____

sodium sulfide: _____

2. Name the following ionic compounds:

Na₃P : _____

BaCl₂: _____

NAMING & WRITING IONIC FORMULAS
Binary Ionic Compounds (Type II):

- *Type II* ions are those cations that *form more than one ion*.
- When naming compounds formed from these ions, include the *ionic charge as Roman numeral*, in parentheses, after the metal's name.

? -1 FeCl₂	+2 -1 FeCl₂	Iron(II) chloride
? -1 FeCl₃	+3 -1 FeCl₃	Iron(III) chloride
? -2 Cu₂O	+1 -2 Cu₂O	Copper(I) oxide
? -2 CuO	+2 -2 CuO	Copper(II) oxide

Examples:

1. Name each of the following Type II ionic compounds:

SnCl₂: _____

Cu₂S: _____

Fe₂O₃ _____

2. Write formulas for each of the following compounds:

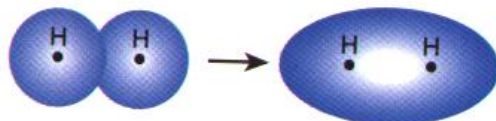
Tin(II) bromide: _____

Lead (IV) oxide: _____

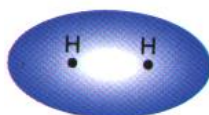
Nickel (III) sulfide: _____

COVALENT COMPOUNDS

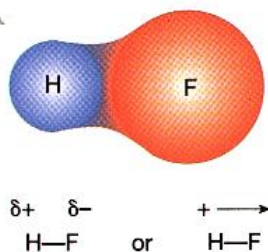
- **Covalent** bonds occur when electrons are *shared* between two atoms.



- **Covalent bonds** occur between *two non-metals*.
- The *smallest* particle of a **covalent** compound is a *molecule*.
- Two types of **covalent** bonds exist: *polar and nonpolar*.
- **Nonpolar** covalent bonds occur between *similar atoms*. In these bonds the *electron pair* is shared *equally* between the two protons.



- **Polar** covalent bonds occur between *different* atoms. In these bonds the *electron pair* is shared *unequally* between the two protons. As a result there is a *charge separation* in the molecule, and *partial charges* on each atom.



Examples:

Classify each of the bonds below as ionic, polar covalent or non-polar covalent:

1. PCl_3
2. MgF_2
3. O_2
4. SO_2

NAMING & WRITING COVALENT FORMULAS
Binary Covalent Compounds:

- These compounds are named similar to ionic compounds, with the second element named based on its *root* and suffix “*-ide*”.
- *Greek* prefixes are used to indicate the *number of atoms* in these compounds:

<i>Number</i>	<i>Prefix</i>	<i>Number</i>	<i>Prefix</i>
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

CS_2 carbon *di*sulfide

PCl_5 phosphorus *pent*achloride

N_2O_4 *di*nitrogen *tetr*oxide

P_4O_{10} *tetra*phosphorous *dec*oxide

- The *first atom* uses a prefix only when *more than one* atom is present.
- The *second atom always* uses a prefix.

Examples:

1. Name the following molecular compounds:

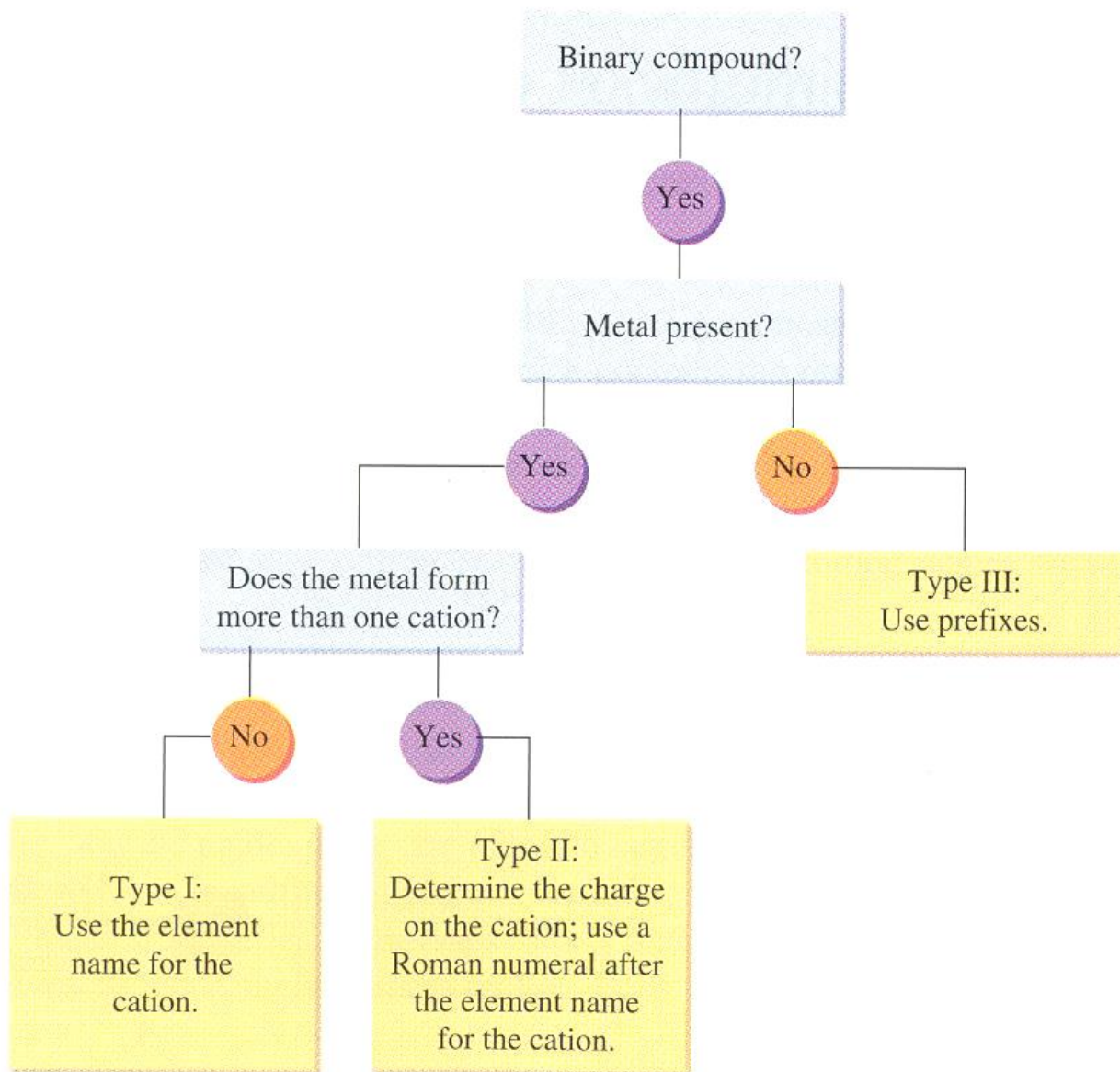
P_2O_5 : _____

IF_7 : _____

2. Write formulas for the following molecular compounds:

carbon tetrachloride: _____

dichlorine monoxide: _____

**SUMMARY OF NAMING
BINARY COMPOUNDS**

POLYATOMIC IONS

- Some ionic compounds contain *polyatomic ions*, an ion composed of *several atoms bound together*.
- Some common polyatomic ions are:

NH ₄ ⁺	ammonium	OH ⁻	hydroxide
NO ₃ ⁻	nitrate	CN ⁻	cyanide
SO ₄ ²⁻	sulfate	C ₂ H ₃ O ₂ ⁻	acetate
PO ₄ ³⁻	phosphate	HCO ₃ ⁻	bicarbonate
CO ₃ ²⁻	carbonate		

- When writing formulas for compounds containing polyatomic ions, treat the polyatomic ion as one group.

potassium nitrate	$1+ \quad 1-$ $\text{K}_2 (\text{NO}_3)_2$	KNO₃
calcium hydroxide	$2+ \quad 1-$ $\text{Ca}_2 (\text{OH})_2$	Ca(OH)₂
ammonium acetate	$1+ \quad 1-$ $(\text{NH}_4)_2 (\text{C}_2\text{H}_3\text{O}_2)_2$	NH₄C₂H₃O₂
sodium sulfate	$1+ \quad 2-$ $\text{Na}_2 (\text{SO}_4)_2$	Na₂SO₄
copper(II) nitrate	$2+ \quad 1-$ $\text{Cu}_2 (\text{NO}_3)_2$	Cu(NO₃)₂

Examples:

- Write formulas for the following polyatomic compounds:

sodium carbonate: _____

ammonium sulfide: _____

magnesium bicarbonate: _____

POLYATOMIC IONS

- *Polyatomic* ionic compounds are named by naming the cation first, followed by the polyatomic ion.

**sodium phosphate****ammonium bromide****copper (I) nitrate or cuprous nitrate****lead (IV) carbonate or plumbic carbonate****Examples:**

2. Name the following polyatomic compounds:

 $\text{Mg}(\text{OH})_2$: _____ NaCN : _____ $\text{Fe}_2(\text{SO}_4)_3$: _____

NAMING ACIDS

- Acids are molecular compounds that form ions when dissolved in water.

Naming Binary Acids:

- Formulas are written similar to binary ionic compounds, assigning a +1 charge to hydrogen.
- When naming the acids, use **hydro-** prefix, followed by the name of the non-metal with an **-ic** ending, followed with the word **acid**.

HCl hydrochloric acid

H₂S hydrosulfuric acid

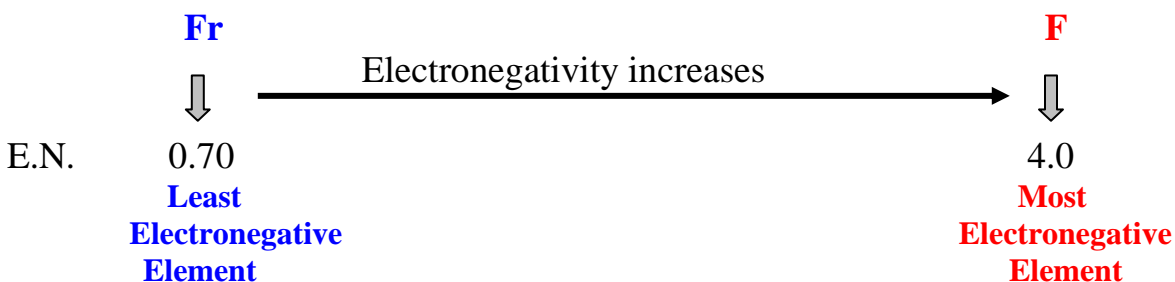
Naming Polyatomic Acids:

- Several polyatomic acids are important in the study of chemistry, and their names must be learned. These acids and the polyatomic ions that form from their ionization are listed below:

Name	Formula	Polyatomic formed from ionization of acid
Nitric acid	HNO ₃	NO ₃ ⁻ (nitrate)
Sulfuric acid	H ₂ SO ₄	SO ₄ ²⁻ (sulfate)
Phosphoric acid	H ₃ PO ₄	PO ₄ ³⁻ (phosphate)
Carbonic acid	H ₂ CO ₃	CO ₃ ²⁻ (carbonate) HCO ₃ ⁻ (bicarbonate)
Acetic acid	HC ₂ H ₃ O ₂	C ₂ H ₃ O ₂ ⁻ (acetate)
Hydrocyanic acid	HCN	CN ⁻ (cyanide)

ELECTRONEGATIVITY

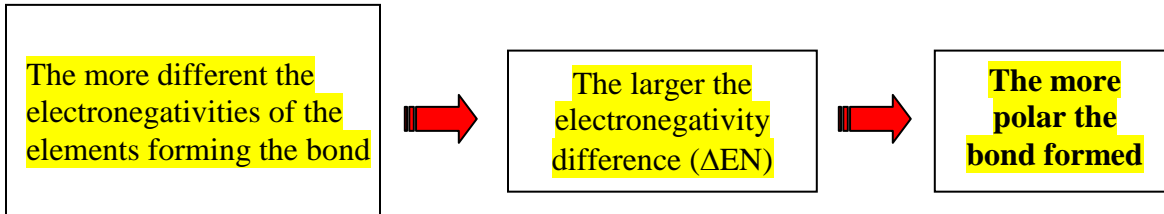
- **Electronegativity (E.N.)** is the ability of an atom involved in a covalent bond to attract the bonding electrons to itself.
- **Linus Pauling** derived a relative Electronegativity Scale based on Bond Energies:



Electronegativities of the Elements

H																	He
2.20																	n.a.
Li	Be											B	C	N	O	F	Ne
0.98	1.57											2.04	2.55	3.04	3.44	3.98	n.a.
Na	Mg											Al	Si	P	S	Cl	Ar
0.93	1.31											1.61	1.90	2.19	2.58	3.16	n.a.
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
0.82	1.00	1.36	1.54	1.63	1.66	1.55	1.83	1.88	1.91	1.90	1.65	1.81	2.01	2.18	2.55	2.96	3.00
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
0.82	0.95	1.22	1.33	1.60	2.16	1.90	2.20	2.28	2.20	1.93	1.69	1.78	1.96	2.05	2.10	2.66	2.60
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
0.79	0.89	1.10	1.30	1.50	2.36	1.90	2.20	2.20	2.28	2.54	2.00	1.62	2.33	2.02	2.00	2.20	n.a.
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	—	Uuq	—	—	—	—
0.70	0.89	1.10	n.a.	n.a.	n.a.	n.a.	n.a.	n.a.	n.a.	n.a.	n.a.		n.a.				

BOND POLARITY & ELECTRONEGATIVITY



CONCLUSION: *Polarity* is a measure of the *inequality* in the sharing of bonding electrons.

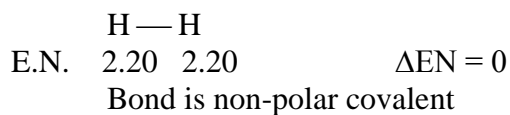
Electronegativity Difference Between the Bonding Atoms	Bond Type
Zero	Covalent
↓	↓
Intermediate	Polar covalent
↓	↓
Large	Ionic

Δ EN	Bond Type
0	NON-POLAR COVALENT
$0 < \Delta$ EN < 1.7	POLAR COVALENT
$1.7 < \Delta$ EN	IONIC

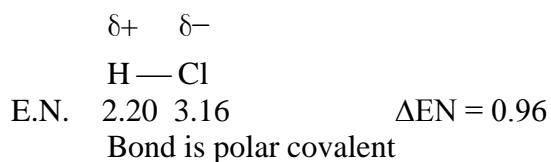
Examples:

Classify each of the bonds below as ionic, polar covalent or non-polar covalent:

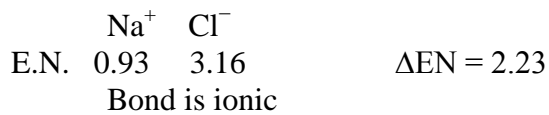
1. Hydrogen molecule



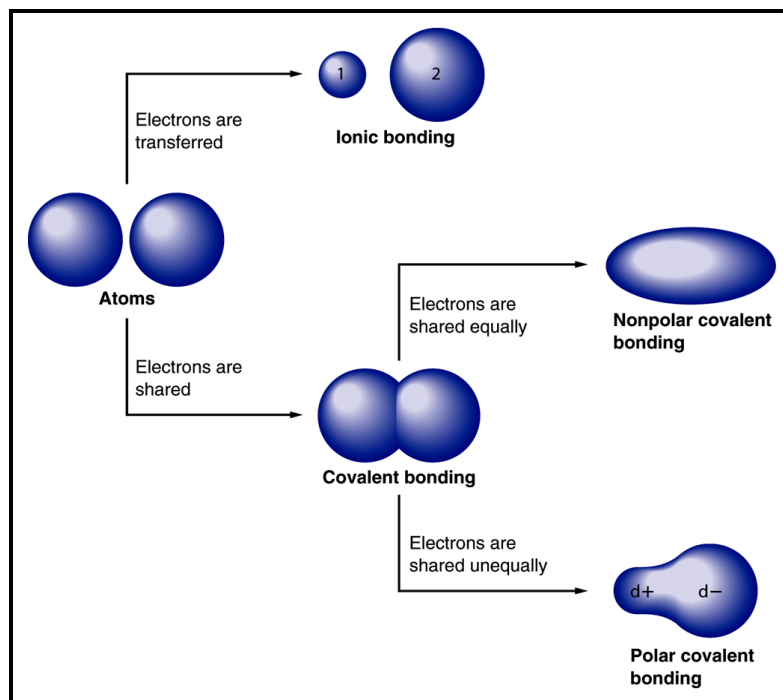
2. Hydrogen chloride



3. Sodium chloride



SUMMARY OF BONDING



COMPARISON OF PROPERTIES OF IONIC & COVALENT COMPOUNDS

	<i>Ionic</i>	<i>Covalent</i>
Structural Unit	Ions	Atoms or Molecules
Attractive Force	Strong	Moderate to Strong
Melting point	High	Generally low
Boiling point	High	Generally low
Solubility in Water	High	Low or None
Hardness	Hard & brittle	Soft to very hard
Electrical Conductivity	Low (solid) High (sol'n)	None
Examples	AgBr NaCl	H ₂ H ₂ O

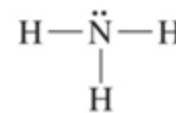
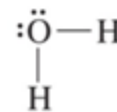
LEWIS STRUCTURES

- *Lewis structures* use Lewis symbols to show *valence electrons* in *molecules* and *ions* of compounds.

1A						8A	
H·						He:	
	2A	3A	4A	5A	6A	7A	
Li·	·Be·	·B·	·C·	·N·	:O·	·F·	·Ne·
Na·	·Mg·	·Al·	·Si·	·P·	:S·	·Cl·	·Ar·

**Lewis symbols for the first 3 periods
of Representative Elements**

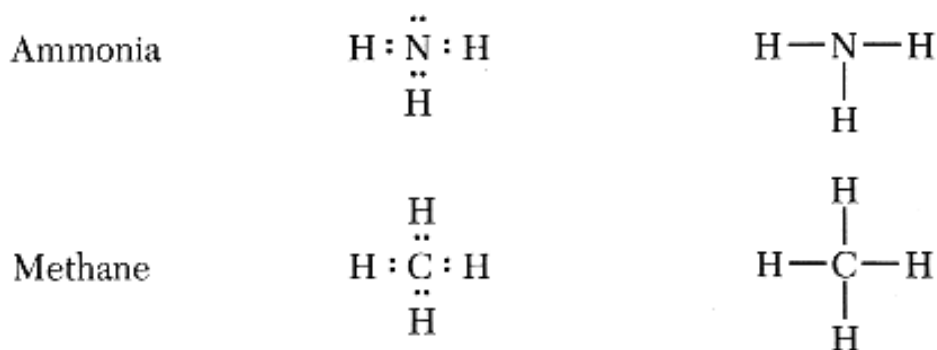
- In a Lewis structure, a *shared electron pair* is indicated by *two dots* between the atoms, or by a *dash* connecting them.
- *Unshared* pairs of valence electrons (called *lone pairs*) are shown as belonging to individual atoms or ions.
- Writing correct Lewis structures for covalent compounds requires an understanding of the *number of bonds* normally *formed by common nonmetals*.



4A	5A	6A	7A	8A
·X·	·X·	:X·	:X·	:X:
4 bonds	3 bonds	2 bonds	1 bond	0 bonds
$\begin{array}{c} \\ \text{—C—} \\ \end{array}$	$\begin{array}{c} \text{—N—} \\ \end{array}$	$\begin{array}{c} \text{:O—} \\ \end{array}$	$\begin{array}{c} \text{:F—} \end{array}$:Ne:

LEWIS STRUCTURES

- Covalent structures are best represented with *electron-dot symbols* or *Lewis structures*. Structures must satisfy *octet rule* (except hydrogen).



- More complex Lewis structures can be drawn according to the following steps:
 - Count the total number of valence electrons to be used in the structure. For ions, add one electron for each negative charge, and subtract one electron for each positive charge.
 - Write a skeleton structure, arranging the atoms in the most symmetrical pattern. Remember the number of bonds commonly formed by atoms.
 - Connect each atom by a dashed line representing 2 electrons each.
 - Determine the number of electrons left by subtracting number of electrons used from the total determined in step 1.
 - Distribute the remaining electrons around the atoms in pairs in order to achieve eight electrons around each (*octet rule*). Hydrogen is an exception (doublet).
 - If there are not enough electrons to achieve octet rule for each atom, form double or triple bonds by sharing more than one pair of electrons.

LEWIS STRUCTURES

Examples:1. water (H_2O)

Total number of electrons: _____

2. carbon dioxide (CO_2)

Total number of electrons: _____

3. carbonate ion (CO_3^{2-})

Total number of electrons: _____

4. ammonia (NH_3)

Total number of electrons: _____

5. chlorate ion (ClO_3^-)

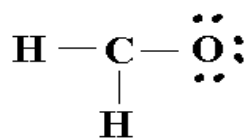
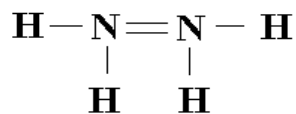
Total number of electrons: _____

EVALUATING LEWIS STRUCTURES

- When evaluating Lewis structures for correctness, two points must be considered:
 1. Are the correct number of electrons present in the structure?
 2. Is octet rule satisfied for all elements? (Hydrogen is an exception)

Examples:

Determine if each of the following Lewis structures are correct or incorrect. If incorrect, rewrite the correct structure.

1) H₂CO2) N₂H₄

MOLECULAR SHAPES

- The three-dimensional shape of the molecules is an important feature in understanding their properties and interactions.
- All *binary* molecules have a *linear* shape since they only contain two atoms.
- More complex molecules can have various shapes (linear, bent, etc.) and need to be predicted based on their Lewis structures.
- A very simple model, **VSEPR** (Valence Shell Electron Pair Repulsion) Theory, has been developed by chemists to predict the shape of large molecules based on their *Lewis structures*.
- Based on VSEPR, the *electron pair groups* in a molecule will repel one another and seek to minimize their repulsion by arranging themselves around the central atom as far apart as possible.
- Electron pair groups can be defined as any one of the following: bonding pairs, non-bonding pairs and multiple bonds.
- A summary of VSEPR predictions are listed below:

Number of electron pair groups around central atom		Molecular Shape	Bond Angle	Examples	Notation
Bonding	Non-bonding				
2	0	Linear	180	CO ₂	AX ₂
3	0	Trigonal planar	120	BF ₃	AX ₃
2	1	Bent	120	SO ₂	AX ₂ E
4	0	Tetrahedral	109.5	CH ₄	AX ₄
3	1	Pyramidal	109.5	NH ₃	AX ₃ E
2	2	Bent	109.5	H ₂ O	AX ₂ E ₂

A = central atom

X=bonding electron pairs

E=non-bonding electron pairs

MOLECULAR SHAPES

- Examples of various shapes are shown below.

Linear (AX₂):

Bond angle: _____

Polarity: _____

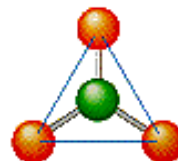
Exceptions:

**Trigonal Planar (AX₃):**

Bond angle: _____

Polarity: _____

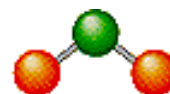
Exceptions:

**Bent (AX₂E):**

Bond angle: _____

Polarity: _____

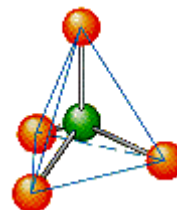
Exceptions:

**Tetrahedral (AX₄):**

Bond angle: _____

Polarity: _____

Exceptions:

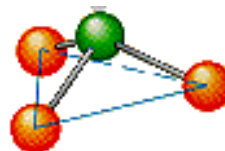


MOLECULAR SHAPES
Pyramidal (AX₃E):

Bond angle: _____

Polarity: _____

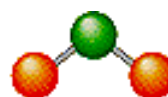
Exceptions:


Bent (AX₂E₂):

Bond angle: _____

Polarity: _____

Exceptions:


Summary of Shapes:

- Symmetrical shapes that have polar bonds, but are usually non-polar:

Linear

Trigonal planar

Tetrahedral

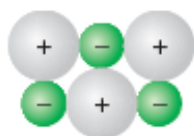
- Unsymmetrical shapes that have polar bonds, but are always polar:

Bent

Pyramidal

ATTRACTIVE FORCES IN MOLECULES

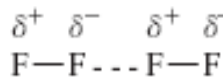
- Many differences in the properties of solids and liquids are associated with the attractive forces that keep their molecules together. These forces range from very strong to very weak and are responsible for the wide range of properties observed in these substances.
- Ionic solids have high melting points that are due to the strong nature of the ionic bond and the forces of attraction between the cation and the anion.



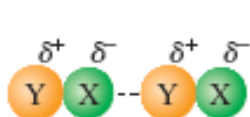
- Covalent molecules also contain attractive forces, but of a weaker nature than the ionic compounds. The forces between covalent molecules can be divided into three types:

1. Dispersion Forces:

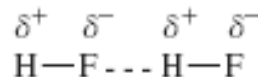
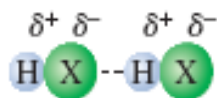
These forces exist between non-polar molecules, and are the weakest of the 3 types of covalent forces. These forces are caused by temporary shifts in distribution of electrons in a non-polar molecule that cause a temporary dipole.


2. Dipole-Dipole Attractions:

These forces exist in polar molecules and are moderate strength. The negative end of a polar molecule is attracted to the positive end of another such molecule by these forces.

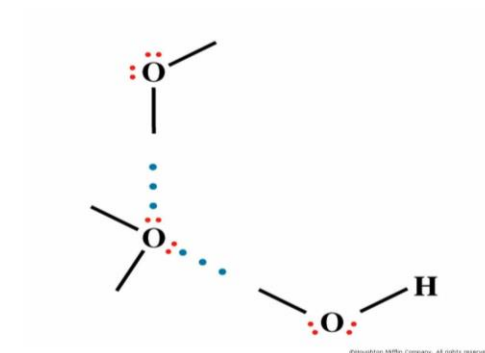
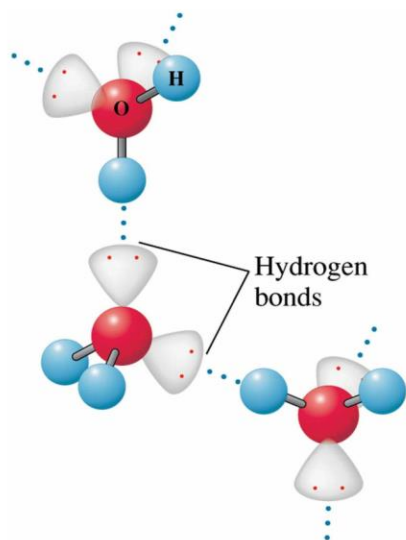

3. Hydrogen Bonds:

These are the strongest of the 3 types of forces and are formed between molecules that have a hydrogen attached to the very electronegative elements (F, O and N).



ATTRACTIVE FORCES IN MOLECULES

- Hydrogen bonding in water is responsible for its many unique properties and behavior in the body. Two representations of this bonding is shown below:



ATTRACTIVE FORCES IN MOLECULES**Examples:**

1. Indicate the major type of molecular forces expected of each of the following:
 - a) HF
 - b) I₂
 - c) PCl₃

2. Identify the strongest attractive forces between molecules of each of the following:
 - a) CH₃OH
 - b) CCl₄
 - c) HCl
 - d) CH₃CH₃

3. Identify the substance in each pair that would have the higher boiling point:
 - a) HF or HBr
 - b) CH₄ or CH₃OH
 - c) MgBr₂ or PBr₃