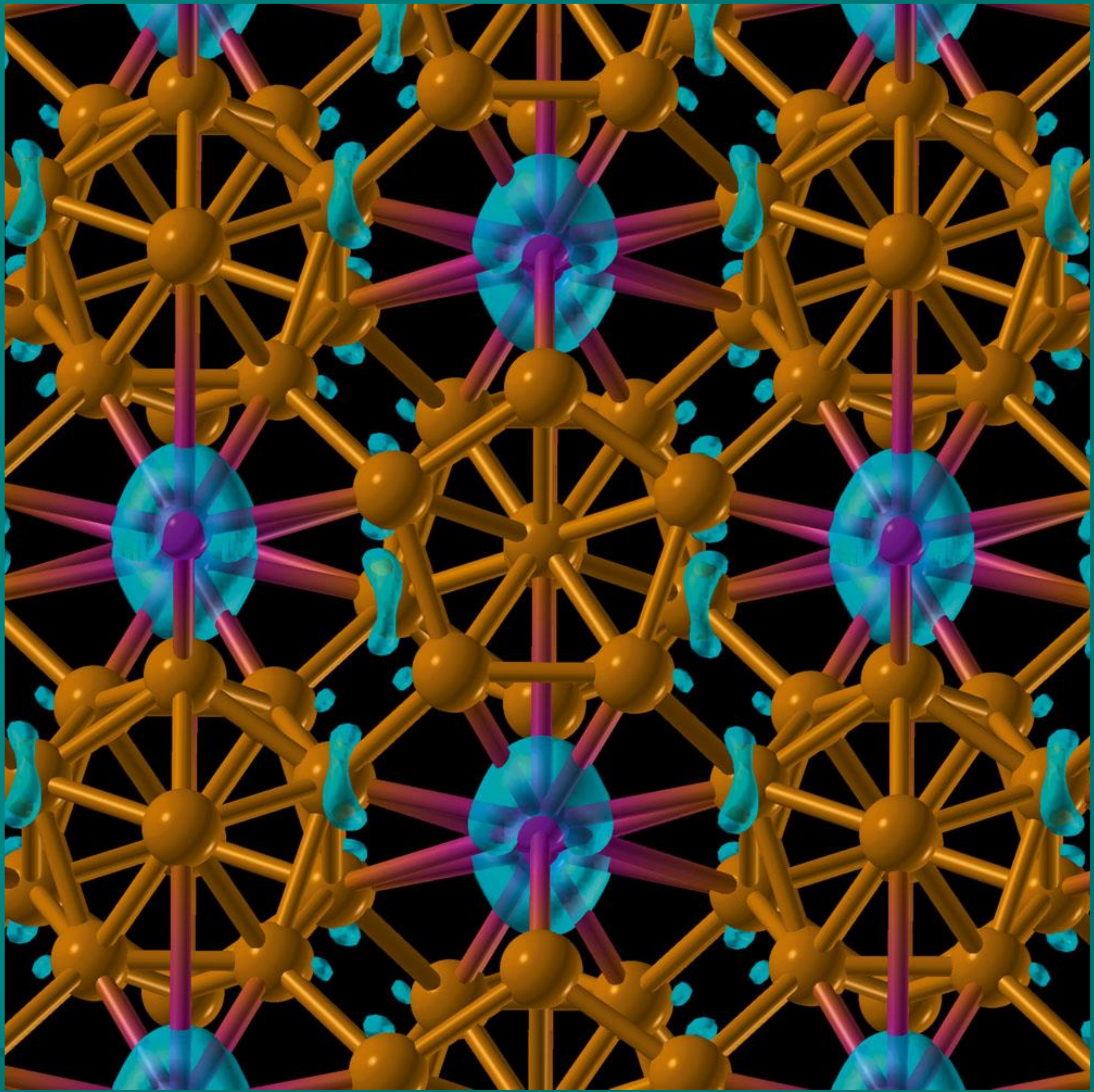
Chapter 12



Chemical Bonding

Types of Chemical Bonds…

Chemical bond: A force that holds groups of two or more atoms/ions together and makes them function as a unit.

Bond Strength: is the energy required to break a bond.

Valence electrons: Are the electrons involved in bonding. The electrons in the highest occupied energy level. The number of valence electrons can be determined by the group number on the periodic table for a representative element.

Two types of chemical bonds:

1. **Ionic** = A bond formed by the attraction between positive and negative ions. Usually between a metal and a nonmetal. Involves a transfer of electrons from one atom to another.
2. **Covalent** = A bond formed by the attraction of two atoms by a sharing of electrons. Usually between two or more nonmetals.

Ionic Bonds…

Pg. 365-369

Electron dot structures can be helpful in determining how ionic bonds are formed.

A. The element symbol represents the nucleus and the core electrons of an atom.

B. Dots are used to represent the valence electrons.

C. Dots are placed one at a time along the four sides of the element symbol.

D. One dot is placed on each side before the dots are placed two to a side.

Example: sodium Na 1 valence e-

Na

magnesium Mg 2 valence e-

Mg

aluminum Al 3 valence e-

Al

Try: carbon, nitrogen, oxygen, fluorine, neon

Electron configurations formed when Ionic bonds are made. (**Cations and Anions**)

1. Electron configurations for cations

Table 12.2 The Formation of Ions by Metals and Nonmetals

Group	Ion Formation	Electron Configuration: Atom		Electron Configuration: Ion
1	$Na \rightarrow Na^+ + e^-$	$[Ne]3s^1$	$\xrightarrow{e^- \text{ lost}}$	[Ne]
2	$Mg \rightarrow Mg^{2+} + 2e^-$	$[Ne]3s^2$	$\xrightarrow{2e^- \text{ lost}}$	[Ne]
3	$Al \rightarrow Al^{3+} + 3e^-$	$[Ne]3s^23p^1$	$\xrightarrow{3e^- \text{ lost}}$	[Ne]
6	$O + 2e^- \rightarrow O^{2-}$	$[He]2s^22p^4 + 2e^-$	$\rightarrow [He]2s^22p^6 =$	[Ne]
7	$F + e^- \rightarrow F^-$	$[He]2s^22p^5 + e^-$	$\rightarrow [He]2s^22p^6 =$	[Ne]

A. Noble gases are stable and generally unreactive in chemical reactions. (ns^2np^6)

B. The **Octet rule** developed by G.N. Lewis, 1916, is used to predict how atoms achieve stable electron configurations of noble gases.

C. When an element *loses electrons* it becomes a **cation**, positively charged. Metals do this.

D. Cations are smaller than their parent atom.

Example:

Na loses an electron to become a positive ion.

Na $\rightarrow$ **Na^{+} + 1 e-**

$1s^2 2s^2 2p^6 3s^1$ $\qquad$ $1s^2 2s^2 2p^6$

(Ne configuration)

Mg $\rightarrow$ **Mg^{+2} + 2 e-**

$1s^2 2s^2 2p^6 3s^2$ $\qquad$ $1s^2 2s^2 2p^6$

(Ne configuration)

E. Transition metals charges may vary so it is difficult to determine their charges. Some metals do not attain a noble gas configuration.

Example: Silver **Ag**

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 \mathbf{5s^1} 4d^{10}$

Silver loses 1 valence electron and becomes a +1 charged ion.

Ag^+ configuration is:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 4d^{10}$

This is a **pseudo noble gas configuration**.

Cu^+, Ag^+, Au^+, Cd^{+2}, Hg^{+2} and others have these kinds of configurations.

2. Configurations for anions

A. The octet rule applies.

B. Nonmetals *gain electrons* becoming negatively charged and are called anions.

C. Halogens form ions called **halide ions**.

D. Anions are *larger* than their parent atom.

Example: Chlorine gains one electron to become a negatively charged ion.

$$Cl \quad + \quad 1\ e^- \quad \rightarrow \quad Cl^-$$

$1s^2 2s^2 2p^6 3s^2 3p^5$ $\qquad$ $1s^2 2s^2 2p^6 3s^2 3p^6$

$O + 2e^- \rightarrow O^{2-}$

$1s^2 2s^2 2p^4$ → $1s^2 2s^2 2p^6$

(Ne configuration)

Table 12.3 Common Ions with Noble Gas Configurations in Ionic Compounds

Group 1	Group 2	Group 3	Group 6	Group 7	Electron Configuration
Li^+	Be^{2+}				[He]
Na^+	Mg^{2+}	Al^{3+}	O^{2-}	F^-	[Ne]
K^+	Ca^{2+}		S^{2-}	Cl^-	[Ar]
Rb^+	Sr^{2+}		Se^{2-}	Br^-	[Kr]
Cs^+	Ba^{2+}		Te^{2-}	I^-	[Xe]

3. Ionic compounds form when these types of changes take place and the oppositely charged ions attract each other. Very strong attraction called an electrostatic force.

A. Ionic compounds are neutral even though they are made of charged particles.

Positive charge = Negative charge

Examples: sodium chloride (NaCl)

$$Na \quad + \quad Cl \quad \rightarrow \quad Na^{+} \quad Cl^{-}$$

$1s^22s^22p^63s^1$ $\quad$ $1s^22s^22p^63s^23p^5$ $\quad$ $1s^22s^22p^6$ $\quad$ $1s^22s^22p^63s^23p^6$

NaCl (1:1)

Alumiunum bromide

Al + Br → Al^{+3} Br^{-}

Br Br^{-}

Br Br^{-}

(aluminum: $1s^2 2s^2 2p^6 \mathbf{3s^2 3p^1}$)

(bromine: $1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{4s^2} 3d^{10} \mathbf{4p^5}$)

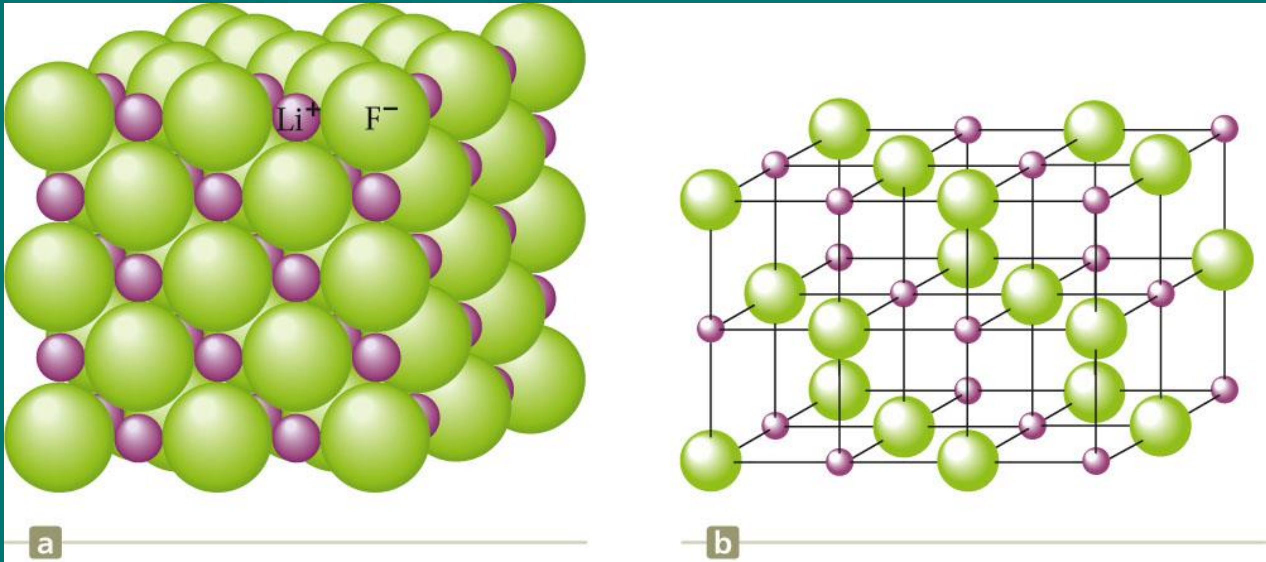
(aluminum ion: $1s^2 2s^2 2p^6$ same as Ne)

(bromide ion: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$ same as Kr)

$\mathbf{AlBr_3}$ (1:3)

B. The compound formulas NaCl and $AlBr_3$ are called formula units and they represent the lowest whole number ratio of ions in the compound.

C. There are no separate molecular units, only a continuous array of ions alternately arranged in a repeating 3-D pattern.



This structure represents the ions as packed spheres.

This structure shows the positions (centers) of the ions. The spherical ions are packed in the way that maximizes the ionic attractions.

D. In ionic compounds containing polyatomic ions, the polyatomic ion acts as one ion. However, the atoms in the polyatomic ion are covalently bonded together.

4. Properties of ionic compounds.

A. crystalline solids at room temperature.

B. Very stable due to strong attractive forces of the ionic bonds.

C. High melting and boiling points.

Ex. $NaCl$ mp = 800°C (1472°F)

D. Conduct electricity when melted and when dissolved in water.

Covalent Bonds…
Pg. 370-391

1. Covalent bonds involve the sharing of electrons so that atoms acquired stable electron configurations.

They can be:

Single bonds - one pair of e- shared

Double bonds – two pairs of e- shared

Triple bonds – three pairs of e- shared

2. Structural formulas (**Lewis Structures**)

pg. 370-374

shows the arrangement of valence e- in the molecule

- Lines are used to represent shared pairs of e-.
- Unshared pairs are represented by dots.

F + F → F-F

B. Steps for Lewis Structures:

1. draw a skeleton structure
2. determine the number of valence e-
3. subtract 2 e- for each line in the skeleton structure
4. distribute unshared pairs, one atom at a time starting with the outside atoms, to satisfy the octet rule

Ex. 1. **N-N**

2. N is in group 5, 5 valence electrons. Two nitrogens ; 2 x 5 = 10 (total valence e- for the compound)

3. 10-2 = 8 (these are the electrons that need to be distributed)

4. **N-N** (not correct, why?)

use another pair of electrons for another bond

5. **N=N** (not correct)

6. **N≡N** Done!

Ex. **CH_4**

```
      H
      |
H  –  C  –  H
      |
      H
```

2) 4 + (4 x 1) = 8

3) 8 – 8 = 0

Done!

Try **HBr**

Why are there no dots around hydrogen?

C. **Polyatomic ions**; same rules but…. for positive ions, subtract charge, for negative ions add charge.

Ex. SO_4^{2-}

1)

```
      O        2-
      |
  O – S – O
      |
      O
```

2) S = 6

O = 6 x 4 = 24

+ 2 e-

(6 + 24 + 2 = 32)

3) 32 – 8 = 24 dots

4. $$\left(\begin{array}{ccc} & O & \\ & | & \\ O - & S & - O \\ & | & \\ & O & \end{array}\right)^{2-}$$

Try **NH_4^+**

D. Some covalent bonds are formed when one atom donates both electrons. This is called a **coordinate covalent bond**.

Ex. **NH_4^+**

These bonds are just like any other covalent bond, only both electrons were donated from one atom.

E. Not all molecules obey the octet rule. Can't draw Lewis Structures for everything.

3. **Resonance** pg. 376-378

This is when two or more valid Lewis Structures can be written for a molecule.

Ex. Ozone O_3

O–O=O or **O=O–O**

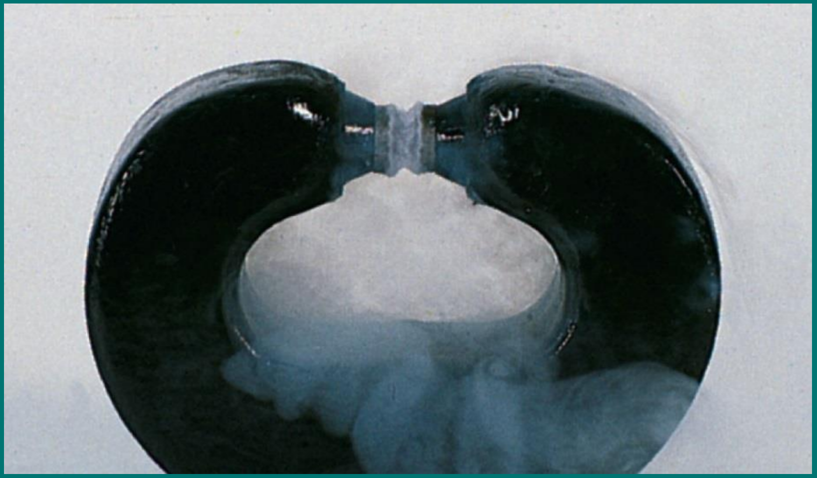
Resonance occurs when the multiple bond can be drawn in different locations.

Try: NO_3^-

4. Exceptions to the Octet Rule pg. 379-380

A. When the total number of valence e- is an odd number:

Paramagnatism occurs: This is caused by <u>one or more unpaired</u> electrons. (The compound shows a strong attraction to an external magnetic field)



Ex. Oxygen O_2 can be O=O, but really is

O–O known by experimentation

B. Expanded octets, more than 8 electrons around the central atom.

PCl_5 SF_6

Expanded octets can occur on central atoms that are nonmetals in period 3 or beyond. (Using *d* sublevels to share e^-)

5. VSEPR Model

A. **"Valence Shell Electron Pair Repulsion"**

This theory helps determine the shape of the molecule. The structure around a given atom is determined by minimizing the repulsions between electron pairs.

B. Use Lewis Structures to determine shape.

C. Unshared pairs around the central atom "push" down on bonds and decrease bond angles.

Ex. CH_4

```
      H
      |
  H — C — H
      |
      H
```

tetrahedral (109.5°)

NH_3

```
  H — N — H
      |
      H
```

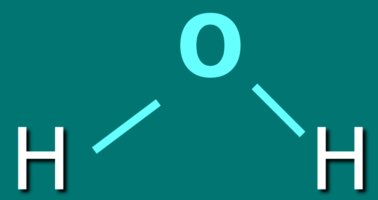
unshared pairs push the bonds away

```
      N
   /  |  \
  H   H   H
```

pyramidal

H_2O

H — **O** — H



bent

CO_2

O ═ **C** ═ O

e- repel equally, carbon has no unshared pairs to cause it to be **linear**

Shape Name

The name of the shape of a molecule is based on the position of the <u>atoms</u> in the molecule.

Refer to Table 12.4, page 388.

(The last column shows the position of the atoms. The name is based on the geometry of the <u>atom</u> location.)

Electron Arrangement

Electron arrangement is the arrangement of the electrons around an atom. We are *usually* only concerned with the central atom.

Refer to Table 12.4, page 388.

(The third and fourth column represent the electron arrangement. This is based on the position of the *electron pairs*.)

Sometimes the electron arrangement and shape name are the same:

CCl_4

Shape name: tetrahedral

E⁻ arrangement: tetrahedral

Sometimes the electron arrangement and shape name are different:

NH_3

Shape name: trigonal pyramidal

E⁻ arrangement: tetrahedral

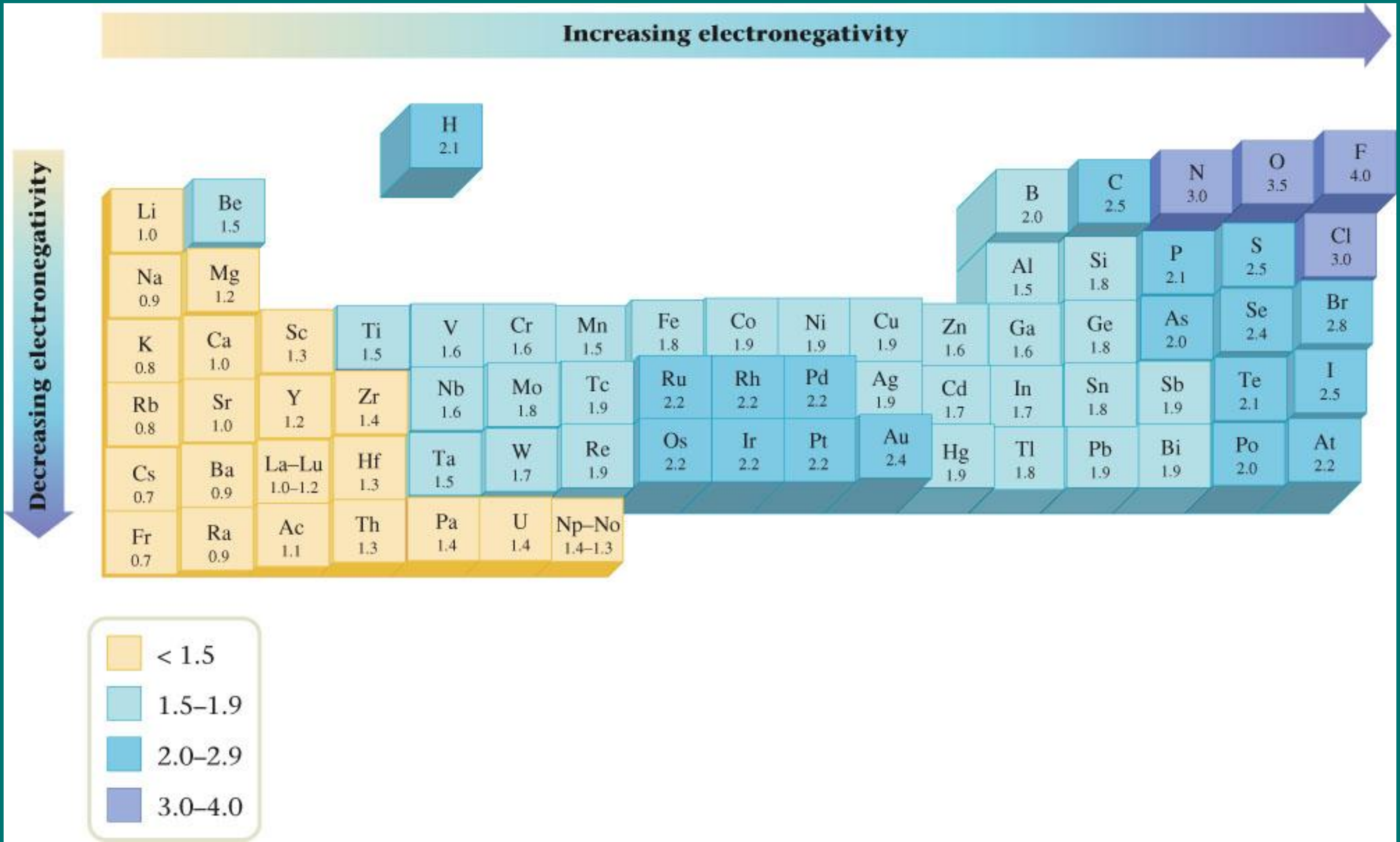
Polar Bonds…
pg. 361-363

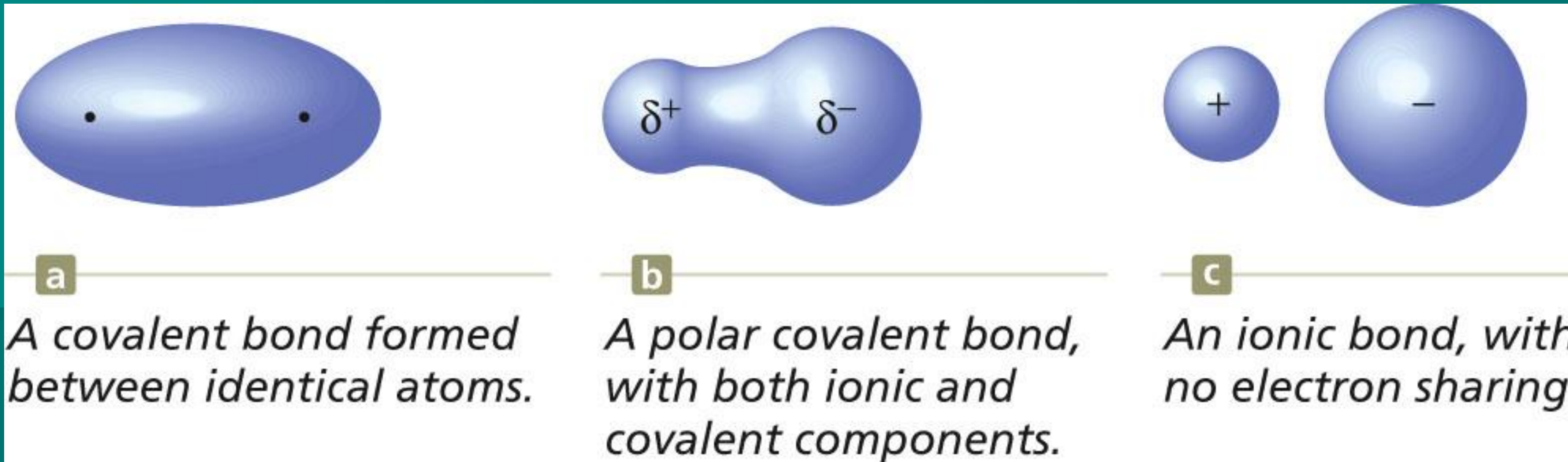
A. When electrons are shared, they are not always shared equally.

1. **Nonpolar covalent bonds** – atoms in the bond are the same element; *e- are shared equally.*

2. **Polar covalent bonds** – atoms in the bond are different; *e- are shared unequally.*

3. This is based on electronegativities of the elements. **Electronegativity** *is the ability of an atom in a molecule to attract shared electrons to itself.*





B. Each atom can be partially positive or negative.

The symbols: δ+ and δ- represent this.

	δ+ δ-		
Ex.	**H — Cl**	or	**H — Cl**
	2.1 3.0		

C. Differences in the electronegativities determines the bond type. Use this table to tell whether the bond is:

1) non-polar **covalent** 0.0 – 0.4

2) **polar covalent**;

greater than 0.4 – less than 2.0

3) **ionic** ≥ 2.0

**very few compounds are purely covalent or purely ionic.*

Ex. **What type of bond is between H and Cl?**

H(2.1) **Cl**(3.0) 3.0 – 2.1 = 0.9

polar covalent

7. Polar Molecules

- If polar bonds are present a molecule may be polar, sometimes not. It depends on the shape of the molecule.

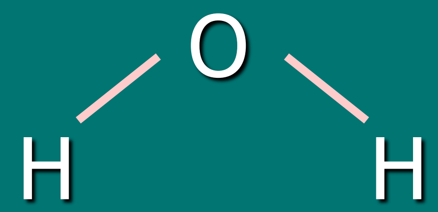
Ex. **HBr** = **H — Br**

This is a polar molecule because there are different atoms in the molecule and they have different *electronegativities*.

Ex. **CO_2** $O = C = O$

This is a **nonpolar** molecule because the polar bonds are directly opposite each other and cancel each other out. It is a linear molecule with like atoms bonded on either side of the central atom.

Ex. **H_2O**



This is polar because the oxygen pulls e- more than the hydrogen and there is nothing to cancel the pull of the oxygen.

All bent molecules are polar.

Summary of polarity of molecule:

Look for symmetrical shape.

- -tetrahedral, trigonal planar, linear

If symmetrical shape, then look at atoms bonded to central atom.

- -if like atoms bonded to central atom, molecule is non-polar.
- -if unlike atoms bonded to central atom, molecule is polar.

ALL non-symmetrical shapes are POLAR.

- -bent and trigonal pyramidal.

8. Properties of molecular substances

A. low melting/boiling points

B. usually gases at room temperature

C. nonelectrolytes (do not conduct electricity)

D. usually *not* soluble in water

E. **UNUSUAL**: diamonds, sand (silicon dioxide) are very stable and strong covalent substances. They are called **network solids**.