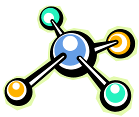
# Chapter 6

Chemical Bonding



# 6.1

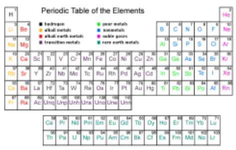
Intro to Chemical Bonding

## Objectives

- **Define** *chemical bond.*
- **Explain** why most atoms form chemical bonds.
- **Describe** ionic and covalent bonding.
- **Explain** why most chemical bonding is neither purely ionic nor purely covalent.
- **Classify** bonding type according to electronegativity differences.

## Quick Review

- Elements in the same family have similar properties. **Why**?
- It is because elements in the same family have the same number of valence electrons.
- These valence electrons are very important!!




## Chemical Bond

- A _________________is the force that holds two atoms together.
  - OR… a mutual electrical attraction between nuclei and valence electrons of different atoms that bind the atoms together
  - OR… Chemical bonds form by the attraction between the positive nucleus of one atom and the negative electrons of another atom.

## How do bonds form?

- Bonds form when atoms, lose, gain, or share electrons
- **Ionic Bonds** form from the attraction between a __________ and an __________
  - _________ and _________ of electrons
  - Typically… _________ and _________
- **Covalent bonds** form when electrons are __________
  - Typically… _________ and _________

## Bonding…

- Bonding between atoms is rarely purely __________ or purely __________
- The bonds usually fall somewhere between these to extremes
- You can use __________ to determine the type of bond

## Bonding and Electronegativity

| Bonding and Electronegativity | |
|---|---|
| **Bond Type** | **Electronegativity Difference** |
| Ionic | |
| Polar-Covalent | |
| Nonpolar-Covalent | |

*These are just guidelines, sometime they are proven incorrect in lab experiments.
**We will assume they are always true

## Bond Types Cont.

- **Nonpolar-covalent bonds** happen when the electron(s) are share __________ by each atom resulting in a __________ distribution of charge
- **Polar-covalent bonds** occur when the atoms __________ the electron(s) equally
- **Polar** means there is an __________ distribution of charge
  - Poles, like on Earth or a magnet

## Practice

- What type of bond will the following atoms form? Which atom is MORE electronegative?
  - Page 153 has electronegativities
  - A. C & H
  - B. C & S
  - C. O & H
  - D. Na & Cl
  - E. Cs & S

## Assignment

- 6.1 Wkst

## 6.2

Covalent Bonding and Molecular Cmpds

## Objectives

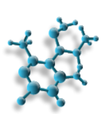
- **Define** *molecule* and *molecular formula.*
- **Explain** the relationships among potential energy, distance between approaching atoms, bond length, and bond energy.
- **State** the octet rule.

## Objectives Continued

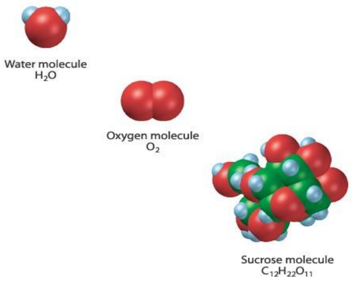
- **List** the six basic steps used in writing Lewis structures.
- **Explain** how to determine Lewis structures for molecules containing single bonds, multiple bonds,
  or both.
- **Explain** why scientists use resonance structures to represent some molecules.

## Molecules

- A __________ is a neutral group of atoms that are held together by covalent bonds.
- A chemical compound whose simplest units are molecules is called a __________ __________.



## Examples of Molecules




## Chemical Formula

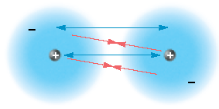
- A __________ **formula** indicates the relative numbers of atoms of each kind in a chemical compound by using atomic symbols and numerical subscripts.
  - Examples: $H_2O$, ____________________
- A __________ **formula** shows the types and numbers of atoms combined in a single molecule of a molecular compound.
  - This is a chemical formula for _________ bonded atoms

## Why do Atoms Bond???

- Atoms gain __________ when they __________ electrons and form __________ bonds.
- __________ __________ states make an atom more __________.
- __________ valence electrons with other atoms results in __________ electron configurations.
  - More stable b/c less _________ NRG

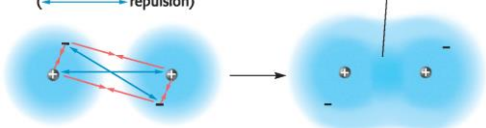
## Example…

- 2 Hydrogen atoms are close enough to each other that their protons and electrons are attracted to each other.
- At the same time, the protons repel each other and the electrons repel each other
- These two forces cancel out to form a covalent bond at a length where the potential energy is at a minimum




For More info go to Page 169

## Formation of a Covalent Bond

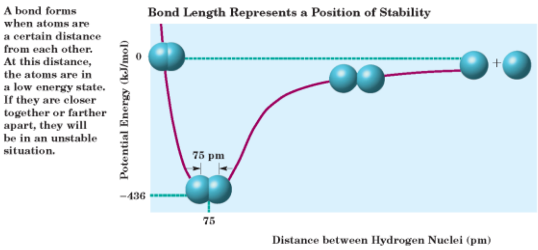



## Characteristics of C Bonds

- **Bond Length** is ________________________
  ______________________________________
- When the bond is formed, _________ is released
  - This is the amount of NRG it would take to break the bonds or _________
  - NRG is reported in _________
    - H-H needs ________ kJ of NRG to break the bond

VC – Bond NRG

## Bond length and Stability




## Bond NRG for Single Bonds

| Bond | Ave Bond Length (pm) | Ave Bond NRG (kJ/mol) |
|---|---|---|
| H-H | | |
| Cl-Cl | | |
| Br-Br | | |
| I-I | | |

Graph Ave Bond Length and Ave Bond NRG.
- What correlation to do you see?

## Electron Dot Notation

- Any Questions?

## Lewis Structure and Structural Formula

- **Lewis Structure**
  - The pair of dots representing a shared pair of electrons in a covalent bond is often replaced by a long dash.

**:F̤̈ F̤̈: OR :F̤̈ - F̤̈:**

- A __________ __________ indicates the kind, number, and arrangement, and bonds but not the __________ of the atoms in a molecule

**F – F**

.

## 8 is GREAT (Octet Rule)

- When atoms bond covalently, they tend to produce __________.
- They are striving to get to _______ (or ______)
  - Nitrogen has 5 v $e^-$ so it needs __________
  - Oxygen has 6 v $e^-$ so it needs __________
  - Fluorine has 7 v $e^-$ so it needs __________
- What about….
  - Cl, C, Si, P, Br, H

## Drawing Lewis Structures

(Page 178)

1. Calculate the total number of valence electrons
2. Divide this number by 2
   1. These are your bonding pairs
      1. Make bonds and electron pairs
3. Make single bonds between elements
4. Subtract 3 from 2
   1. These are the electrons left over for electron pairs and double and triple bonds
5. Finish Lewis Structure

## Drawing Lewis Structures

- Predict the location of certain atoms.
  - Where are Carbon atoms going to be?
  - Where are halogens going to be?
- Determine the number of electrons available for bonding.
  - Electron Dot
- Determine the number of bonding pairs.
  - Every "Single" electron MUST make a bond

## Drawing Lewis Structures

- Place the bonding pairs.
  - Draw all you single bonds first, then if you still need more bonds, try doubles and triples
- Determine whether the central atom satisfies the octet rule.

• Examples to come…

## Lewis Structure (Single Bonds)

• **Single Bond** is a covalent bond where __________

• Ex.
  - F-F
  - N-N
  - $H_2O$

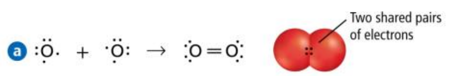
• The $e^-$ that do not bond are called "__________"

## Practice

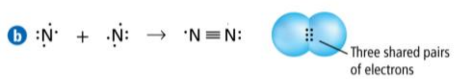
- Page 176
- 1-4

## Multiple Covalent Bonds

- __________ bonds form when __________ pairs of electrons are shared between two atoms.




- __________ bonds form when __________ pairs of electrons are shared between two atoms.




## Multiple Covalent Bonds

- Typically, multiple bonds are __________ than singles bonds.
- Bond Strength

## Lewis Structure (Multiple Bonds)

- See “Drawing Lewis Structures” Wkst
- Ex.
  - $O_2$
  - $CO_2$
  - NOF
  - $CH_3I$
- Many times “Trial and Error” is the best way to do this.

## Exception

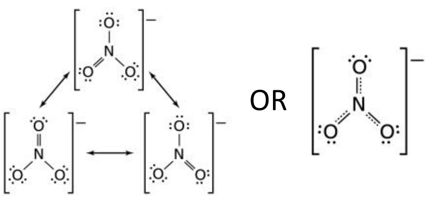
- Main-group elements in Periods 3 and up can form bonds with *expanded valence,* involving *more* than eight electrons
  - Boron tends to bond with 3 elements
    - Don’t worry about this right now

## Practice

- Page 178
- 1-2

## Resonance

- See Video First!
- O-O=O ⟷ O=O-O



## Assignment

- 6.2 Wkst

## 6.3

Ionic Bonding and Ionic Cmpds

## Objectives

- **Compare** a chemical formula for a molecular compounds with one for an ionic compound.
- **Discuss** the arrangements of ions in crystals.
- **Define** *lattice energy* and explain its significance.
- **List** and compare the distinctive properties of ionic and molecular compounds.
- **Write** the Lewis structure for a polyatomic ion given the identity of the atoms combined and other appropriate information.

## Ionic Cmpds

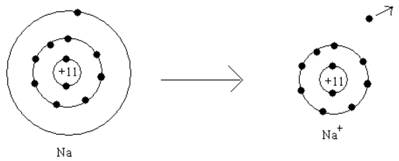
- The electrostatic force that holds oppositely charged particles together in an ionic compound is called an __________
- Compounds that contain __________ bonds are called __________ **Cmpd**.
- __________ ionic compounds contain only **two** different elements—__________
- Ex. NaCl, NaBr, MgO

## Ionic Cmpd

- When 2 or more ions bond, they form a __________ unit
- **Formula Unit** is __________________________
  ________________________________________
- Ex. NaCl, $CaF_2$

## Positive Ion Formation

- A positively charged ion is called a __________.
  - An easy way to remember this is that a "_____" looks like a "______"
- This figure illustrates how sodium loses one valence electron to become a sodium cation.




## Positive Ion Formation

<u>Sodium **ATOM**</u>

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

<u>Sodium **ION**</u>

$1s^2\ 2s^2\ 2p^6\ 3s^0$
So…
$1s^2\ 2s^2\ 2p^6$ with a "+1" charge

There is one more proton than there is electrons

Draw example on board

## Quick ?'s

- How many protons in a Li atom? Electrons?
- How many protons in a Li ion? Electrons?
- How did you figure this out??

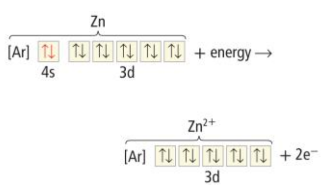
## Positive Ion Formation

- Metals are reactive because they lose valence electrons easily.

| Table 7.2 | Group 1, 2, and 13 Ions | |
|---|---|---|
| Group | Configuration | Charge of Ion Formed |
| 1 | [noble gas] $ns^1$ | 1+ when the $s^1$ electron is lost |
| 2 | [noble gas] $ns^2$ | 2+ when the $s^2$ electrons are lost |
| 13 | [noble gas] $ns^2np^1$ | 3+ when the $s^2p^1$ electrons are lost |

## Positive Ion Formation

- Transition metals commonly form 2+ or 3+ ions, but can form greater than 3+ ions.
- Other relatively stable electron arrangements are referred to as pseudo-noble gas configurations.
- Zn has a full outer s, p, and d orbital




## Positive Ion Formation

- Transition metals can form many different ions, with different charges
- The is because sometimes they do lose their “d” orbital electrons
- You will not be expected to determine what charged ion a transition metal would produce
- We will run into transition metals with more than one ion
- Ex. $Cu^{+2}$ $Cu^{+3}$

## A positive ion is also called a…

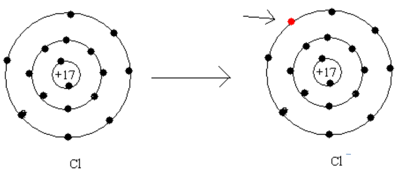
1. Cathode
2. Cation
3. Kation
4. None of the above

## Which of the following will form a cation?

1. Oxygen
2. Calcium
3. Copper
4. Bromine
5. More than one of the above
6. All of the above
7. None of the above

## Negative Ion Formation

- An __________ is a negatively charged ion.
  - An easy way to remember this one is that it is the opposite of a cation. ☺
- The figure shown here illustrates chlorine gaining an electron to become a chlorine ion.



## Negative Ion Formation

- <u>Chlorine **ATOM**</u>  <u>Chlorine **ION**</u>

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

So…

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ with a "-1" charge

There is one more electron than there is protons

Draw example on board

## Practice

- Write the electron config for …
  - Na Ion
  - O ion
  - Ca ion

## Quick ?'s

- How many protons in a Cl atom? Electrons?
- How many protons in a Cl ion? Electrons?

- How did you figure this out??

## How many protons in an S ion?

1. Numeric

## How many valence electrons in an S atom?

1. Numeric

## Negative Ion Formation

- Nonmetal ions gain the number of electrons required to fill an octet.
- Some nonmetals can gain or lose electrons to complete an octet.

**Table 7.3 Group 15–17 Ions**

| Group | Configuration | Charge of Ion Formed |
|---|---|---|
| 15 | [noble gas] $ns^2np^3$ | 3– when three electrons are gained |
| 16 | [noble gas] $ns^2np^4$ | 2– when two electrons are gained |
| 17 | [noble gas] $ns^2np^5$ | 1– when one electron is gained |

## A cation is…

1. Negatively charged
2. Positively charged
3. Neutral
4. None of the above

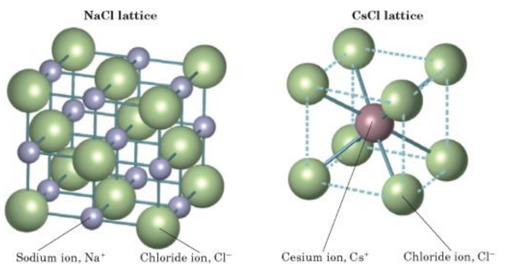
## Which of the following would produce a negatively charged ion?

1. Ca
2. S
3. Cs
4. All of the above
5. None of the above

## Formation of Ionic Cmpds

- In an ionic crystal, ions minimize their potential energy by combining in an orderly arrangement known as a __________.
  - Attractive forces exist between oppositely charged ions within the lattice.
  - Repulsive forces exist between like-charged ions within the lattice.
- The combined attractive and repulsive forces within a crystal lattice determine:
  - ________________________________
  - ________________________________

## NaCl and CsCl Crystal Lattices




## Lattice NRG

- The _________ required to separate _________ of ions in an _________ compound is referred to as the **lattice NRG.**
- Lattice energy is _________ related to the size of the ions that are bonded.
- _________ ions form compounds with more closely spaced ionic charges, and require _________ energy to separate.
- Electrostatic force of attraction is _________ related to the distance between the opposite charges.
- The _________ the ion, the _________ the attraction

## Energy and the Ionic Bond

- The value of lattice energy is also affected by the charge of the ion.

Which compound has the strongest bonds?

**Table 7.6 Lattice Energies of Some Ionic Compounds**

| Compound | Lattice Energy (kJ/mol) | Compound | Lattice Energy (kJ/mol) |
|---|---|---|---|
| KI | 632 | KF | 808 |
| KBr | 671 | AgCl | 910 |
| RbF | 774 | NaF | 910 |
| NaI | 682 | LiF | 1030 |
| NaBr | 732 | $SrCl_2$ | 2142 |
| NaCl | 769 | MgO | 3795 |

## Ionic vs Covalent Bonds/Cmpds

- The force that holds ions together in an ionic compound is a _____________ electrostatic attraction.
- In contrast, the forces of attraction _____________ molecules of a covalent compound are much weaker.
- This difference in the strength of attraction between the basic units of molecular and ionic compounds gives rise to different properties between the two types of compounds.

## Ionic vs Covalent Bonds/Cmpds

- Molecular (covalent) cmps
  1. ________________________
  2. ________________________
- Ionic cmpds
  1. __________________________
  2. __________________________

## Comparing Melting and Boiling pts

| Compound name | Formula | Type of compound | Melting point °C | Melting point K | Boiling point °C | Boiling point K |
|---|---|---|---|---|---|---|
| Magnesium fluoride | $MgF_2$ | ionic | 1261 | 1534 | 2239 | 2512 |
| Sodium chloride | NaCl | ionic | 801 | 1074 | 1413 | 1686 |
| Calcium iodide | $CaI_2$ | ionic | 784 | 1057 | 1100 | 1373 |
| Iodine monochloride | ICl | covalent | 27 | 300 | 97 | 370 |
| Carbon tetrachloride | $CCl_4$ | covalent | –23 | 250 | 77 | 350 |
| Hydrogen fluoride | HF | covalent | –83 | 190 | 20 | 293 |
| Hydrogen sulfide | $H_2S$ | covalent | –86 | 187 | –61 | 212 |
| Methane | $CH_4$ | covalent | –182 | 91 | –164 | 109 |

## How to identify the cmpd as ionic or covalent

- In a laboratory your may…
  - Lq or gas = _____________
  - If solid…
    - Tap gently and if it breaks, it should not turn into a powder (it should fracture)
    - Heat it up (Ionic have _____________ melting points) and if it melts
      - Check for conductivity (_____________ conduct electricity)
    - _____________ in water and check for conductivity
      - _____________ conduct electricity

## Polyatomic Ions

- _____________ are covalently bonded atoms that have a charge
- Ex.
  - $NH4^{+1}$ is ammonium
    - The +1 is the charge for the entire molecule
  - There are a lot of polyatomic ions
    - See "Polyatomic Ions Sheet"

## Assignment

- 6.3 Wkst

# 6.4

Metallic Bonding

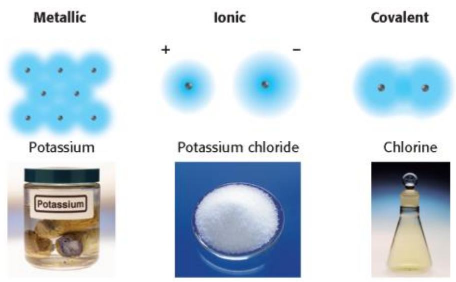
## Objectives

- **Describe** the electron-sea model of metallic bonding, and explain why metals are good electrical conductors.
- **Explain** why metal surfaces are shiny.
- **Explain** why metals are malleable and ductile but ionic-crystalline compound are not.

## Metallic Bonding

- Chemical bonding in metals is different than in ionic, molecular, or covalent-network cmpds
- They have very unique properties:
  1. _____________
  2. _____________
  3. _____________
  4. _____________
  5. _____________

## Comparisons




## Metallic Bond

- Within the crowded lattice, the outer energy levels of metal atoms overlap.
- The _____________ proposes that all metal atoms in a metallic solid contribute their valence electrons to form a "sea" of electrons.
- The electrons are free to move around and are referred to as _____________, forming a metallic _____________.
  - They don't "_____________" to any one atom

## Metallic Bond

- The chemical bonding that results from the attraction between _____________ atoms and the surrounding _____________ of electrons is called _____________ **bonding.**

## Enthalpy of Vaporization

- Enthalpy of Vaporization is the amount of NRG absorbed as heat when a specified amount of substance vaporizes at a constant pressure.

## Assignment

- 6.4 Wkst

## 6.5

Molecular Geometry

## Objectives

- **Explain** VSEPR theory.
- **Predict** the shapes of molecules or polyatomic ions using VSEPR theory.
- **Explain** how the shapes of molecules are accounted for by hybridization theory.

## Objectives (cont.)

- **Describe** dipole-dipole forces, hydrogen bonding, induced dipoles, and London dispersion forces and their effects on properties such as boiling and melting points.
- **Explain** the shapes of molecules or polyatomic ions using VSEPR theory.

## Molecular Geometry

- The properties of molecules depend not only on the bonding of atoms but also on ______________ *:* the three-dimensional arrangement of a molecule's atoms.
- The polarity of each bond, along with the geometry of the molecule, determines ______________, or the uneven distribution of molecular shape.
- Molecular polarity strongly influences the forces that act ______________ molecules in liquids and solids.

## VSEPR

- Valence Shell Electron Pair Repulsion (VSEPR)
- Electron pairs ______________ each other and cause molecules to be in ______________ positions relative to each other.
- ______________ electron pairs also determine the ______________ of a molecule.
- Electron pairs are located in a molecule as ______________ as they can be.

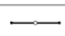
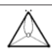
## VSEPR

- Bond Angles
- Think of tying balloons together.
  - The point where they are connected represents the central atom
  - The balloons represent electron-dense regions
- What is it going to look like when you tie 2 balloons together?
- 3 balloons, 4 balloons…

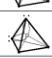
## VSEPR

- When Drawing them…
  - All bonds are represented with a single line
    - Single, Double, and triple bonds = 1 line
- Go Over Drawing Lewis Structure HO
- Practice
  - $CBr_4$
  - $BCl_3$
  - $H_2O$
  - ONF

## VSEPR and Molecular Geometry

| | Molecular shape | Atoms bonded to central atom | Lone pairs of electrons | Formula example | Lewis structure |
|---|---|---|---|---|---|
| Linear |  | 2 | 0 | $BeF_2$ | :F–Be–F: |
| Bent |  | 2 | 1 | $SnCl_2$ | Sn, Cl, Cl |
| Trigonal-planar |  | 3 | 0 | $BF_3$ | B, F, F, F |
| Tetrahedral |  | 4 | 0 | $CH_4$ | H–C–H, H, H |

## VSEPR and Molecular Geometry

| | Molecular shape | Atoms bonded to central atom | Lone pairs of electrons | Formula example | Lewis structure |
|---|---|---|---|---|---|
| Trigonal-pyramidal |  | 3 | 1 | $NH_3$ | N, H H H |
| Bent |  | 2 | 2 | $H_2O$ | O, H H |

## VSEPR and Molecular Geometry

- What are their molecular shapes? (Name and draw)
  - $CBr_4$
  - $BCl_3$
  - $H_2O$
  - ONF
  - $CO_2$

## Hybridization

- For this purpose, we use the model of _____________, which is the mixing of two or more atomic orbitals of similar NRG on the same atom to produce new hybrid atomic orbitals of equal NRG
- **Hybrid orbitals** are orbitals of _____________ produced by the combination of two or more orbitals on the same atom.
- Hybridization explains the _____________ and _____________ of many molecules.

## Intermolecular Forces

- Intermolecular forces are the attraction forces between _____________
  - Boiling points are a good measure of these forces
    - Higher BP, _____________ intermolecular forces
  - These forces tend to be weaker than bonds within a molecule or formula unit
    - ___________

## ??

- Which has stronger I Forces? Ionic Bonds or Covalent Bonds?
- Explain

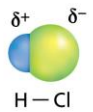
| Type of substance | Common use | State at room temperature | Melting point (°C) | Boiling point (°C) |
|---|---|---|---|---|
| **Ionic substances** | | | | |
| Potassium chloride, $KCl$ | salt substitute | solid | 770 | sublimes at 1500 |
| Sodium chloride, $NaCl$ | table salt | solid | 801 | 1413 |
| Calcium fluoride, $CaF_2$ | water fluoridation | solid | 1423 | 2500 |
| **Covalent substances** | | | | |
| Methane, $CH_4$ | natural gas | gas | –182 | –164 |
| Ethyl acetate, $CH_3COOCH_2CH_3$ | fingernail polish | liquid | –84 | 77 |
| Water, $H_2O$ | (many) | liquid | 0 | 100 |
| Heptadecane, $C_{17}H_{36}$ | wax candles | solid | 22 | 302 |

## Polar Bonds

- These are the strongest intermolecular forces
- They have an ______________ charge distribution
  - This leads to dipoles
  - ____________ is made when an electrons spend more time around one atom than another

ond

Electronegativity Cl = 3.16
Electronegativity H = 2.20
Difference = 0.96




## Polar

- A dipole is represented by an arrow with its head pointing toward the negative pole and a crossed tail at the positive pole. The dipole created by a hydrogen chloride molecule is indicated as follows:

H−Cl

## Dipole - Dipole

- The ______________ attraction is very similar to an ionic attraction
- The oppositely charge ends of a ______________ covalent molecule are attracted to each other
- The ______________ the polar difference, the ______________ the attraction

## Dipole-Dipole

- Dipole-Dipole bonds form when the dipole from one molecules is attracted ______________
- Page 195 – Fig. 5.9

## Polar Covalent Bonds

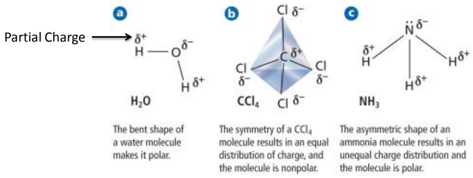
- Covalently bonded molecules are either ______________ or ______________.
- ______________ molecules are not attracted by an electric field.
- ______________ molecules align with an electric field.

## Polar Covalent Bonds

- Compare water and $CCl_4$.
- Both bonds are polar, but only water is a polar molecule because of the shape of the molecule.
- Quick Demo – water and balloon

## Polar Covalent Bonds

- The electric charge on a $CCl_4$ molecule measured at any distance from the center of the molecule is identical to the charge measured at the same distance on the opposite side.




## H Bonds

- Some hydrogen-containing compounds have unusually high boiling points. This is explained by a particularly strong type of dipole-dipole force.
- In compounds containing H–F, H–O, or H–N bonds, the large electronegativity differences between hydrogen atoms and the atoms they are bonded to make their bonds highly polar.

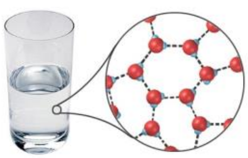
## H Bonds

- The small size of the hydrogen atom allows the atom to come very close to an unshared pair of electrons in an adjacent molecule.
- The intermolecular force in which a hydrogen atom that is bonded to a highly electronegative atom is attracted to an unshared pair of electrons of an electronegative atom in a nearby molecule is known as ______________.

## H Bonds

- Hydrogen bonds are usually represented by dotted lines connecting the hydrogen-bonded hydrogen to the unshared electron pair of the electronegative atom to which it is attracted.
- An excellent example of hydrogen bonding is that which occurs between water molecules. The strong hydrogen bonding between water molecules accounts for many of water's characteristic properties

## H Bonds



## London Dispersion Forces

- In all atoms and molecules the electrons are moving
- This can result in the electron distribution being ______________ at any moment in time
  - This can produce a _____________ dipole
- This temporary dipole can induce a dipole-dipole attraction called ______________
- Page 197 – Fig. 5.13

## Assignment

- 6.5 Wkst