Parts of Chapter 7 having to do with

(a) Lewis Dot Structures (will provide LD structures on quiz/test)

(b) polar bonds (will not provide on quiz/test)

Lewis Structures of Atoms

- We use dots around the symbol to represent valence electrons. (4 walls – put one electron on each wall until run out of walls then double up electrons on walls with a dot already on it)

one electron wall forms one bond – other atom supplies other electron
two electron wall does not normally form a bond.

1A	2A	3A	4A	5A	6A	7A	8A
Li·	·Be·	·B·	·C·	·N:	·O:	:F:	:Ne:

How many bonds for each of the above main group elements ?

Lewis Structures of Atoms

How many covalent bonds for each of the main group elements ?

usually ionic bond		3 bond	4 bond	3 bond	2 bond	1 bond	no bond
1A	2A	3A	4A	5A	6A	7A	8A
Li·	·Be·	·Ḃ·	·Ċ·	·N̈:	·Ö:	:F̈:	:N̈e:

Procedure for Drawing Electron-Dot Structures (or Lewis Dot Structures)

Step 1: Valence Electrons

- Add up valence electrons for all atoms in the molecule. [ex: H_2O (2*1) + 6 = 8 e]

- Add one electron for each negative charge in an anion, or subtract one electron for each positive charge in a cation.

[ex: SO_4^{-2} 6 + (4*6) + 2 = 32e]

---

valence e for atom = group # for main group elements
(1A to 8A-American system)

Procedure for Drawing Electron-Dot Structures

Step 2: **Connect Atoms**

- Draw lines between all atoms to represent bonds between atoms.

- Hydrogen and halogens usually form only one bond.

- Elements in third row and lower can expand octet (can have more than 8 electrons because have d subshell available even if d subshell is empty)

Procedure for Drawing Electron-Dot Structures

Step 3: Put octets on all atoms

- Complete all atom's octet (bond single line counts as 2 electrons) (except for hydrogen & helium – H, He only gets duet).

(to get octet: use lone pair electrons if not enough bonds)

Procedure for Drawing Electron-Dot Structures

Step 4: **Check # electrons in your Lewis Dot Structure:**

 # electrons in your structure = # valence electrons

done.

Procedure for Drawing Electron-Dot Structures

Step 5: **Put in Multiple Bonds. OR Put in lone pairs.**

If the number of electrons from step 4 does not match up, put in multiple bonds OR add lone pairs to central atom.

(each multiple bond decreases number of e in structure by 2 e) (each lone pair increase e in structure by 2e)

Go back to step 3 & redo iteratively until done. (done: # e in structure = # valence e)

Polar Covalent Bonds: Electronegativity

Electronegativity: atom's ability to attract shared electrons in a covalent bond

partial - charge $\delta -$

partial + charge $\delta +$

$$X:X \qquad \overset{\delta +}{Y}:\overset{\delta -}{X} \qquad M^+ \; X:^-$$

Nonpolar covalent
(electronically symmetrical)

Polar covalent
(partial charges)

Ionic
(full charges)
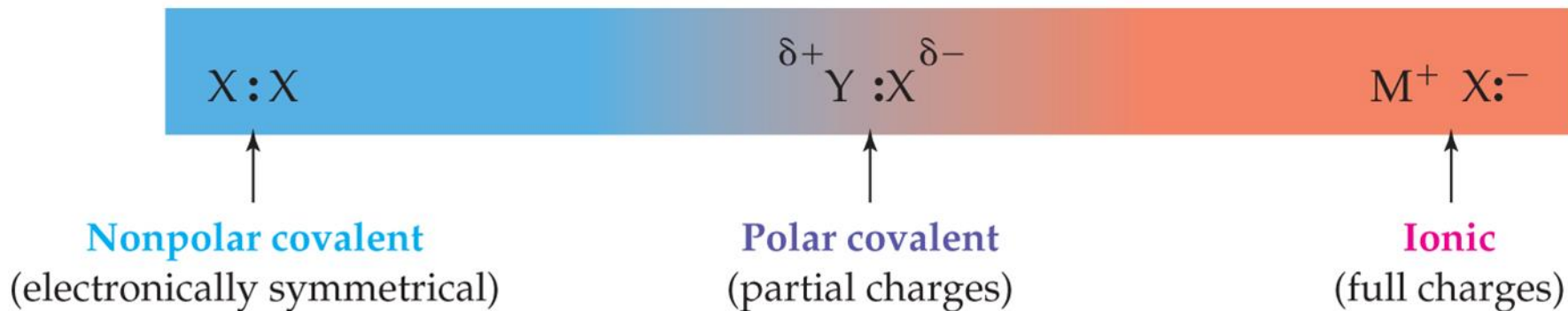
Polar covalent bonds have an unsymmetrical electron distribution in which the bonding electrons, shown as dots, are attracted more strongly by one atom than the other.

Polar Covalent Bonds: Electronegativity

$^{\delta+}H - Cl^{\delta-}$

$[H :Cl]$
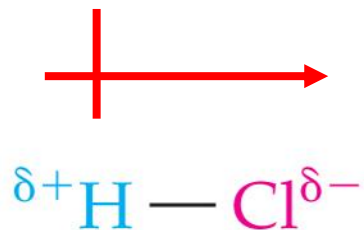
A polar covalent bond. The two bonding electrons (dots) are attracted more strongly by Cl than by H.
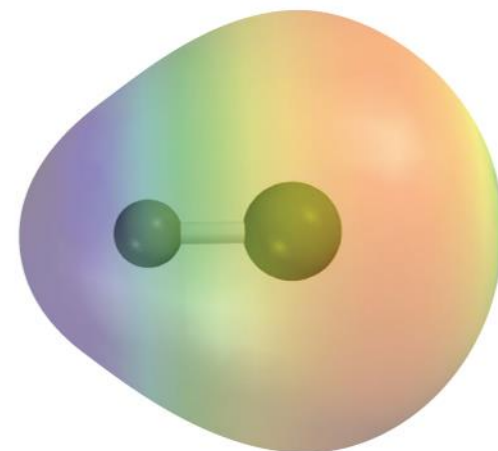
Electron distribution diagram –
yellow no charge, red - charge, blue + charge

Polar Covalent Bonds: Electronegativity

F is most electronegative (memorize this)
(H almost same EN as C)

Parts of Chapter 8 having to do with Intermolecular Forces comes next slides

To do VSEPRT (using chart next slide & handout)

1.   Draw Electron Dot Structure (Lewis Dot Structure).

2.   # charge clouds (# e pairs/e domains) = # bonds + # lone pairs

 # of bonds to central atom (single, double and triple bonds count as one bond for VSEPRT)

(# of lone pairs on the central atom - not # of electrons in lone pairs but # of pairs)

3.   Use geometry e pairs & geometry of molecule for VSEPRT
   (# lone pairs/nonbonding e – these e are invisible for molecular shape)

4.   Use hybridization for Valence Bond

VSEPR & valence bond hybrization chart (e pair/domains = charge cloud)

| # electron pairs(domains)	# lone pair	hybridization	geometry of electron pairs	geometry of molecules	angles
2	0	sp	linear	linear	180
3	0	sp^2	trigonal planar	trigonal planar	120
3	1	sp^2	trigonal planar	bent	
3	2	sp^2	trigonal planar	linear	
4	0	sp^3	tetrahedral	tetrahedral	109.5
4	1	sp^3	tetrahedral	trigonal pyramidal	
4	2	sp^3	tetrahedral	bent	
4	3	sp^3	tetrahedral	linear	
5	0	sp^3d	trigonal bipyramidal	trigonal bipyramidal	120 & 90
5	1	sp^3d	trigonal bipyramidal	see saw	
5	2	sp^3d	trigonal bipyramidal	T shaped	
5	3	sp^3d	trigonal bipyramidal	linear	
6	0	sp^3d^2	octahedral	octahedral	90
6	1	sp^3d^2	octahedral	square pyramidal	
6	2	sp^3d^2	octahedral	square planar	
6	3	sp^3d^2	octahedral	T shaped	
6	4	sp^3d^2	octahedral	bent	

HW 8-1: Do VSEPRT (on central atom)

1. Draw Lewis Dot Structure. (put in lone pairs to partially complete structures below)

2. # of charge clouds (# e domains, # e pairs) = # bonds (single, double, triple bonds all count as one bond for charge cloud counting) + # lone pairs (not # e in lone pairs) & use chart
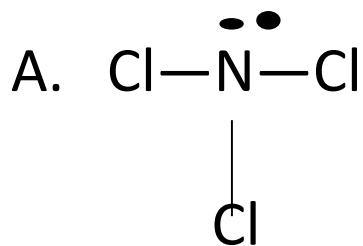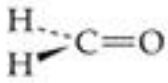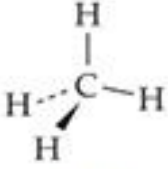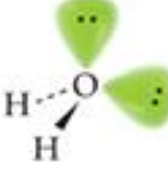
A. Cl—N̈—Cl   N has 4 charge clouds, <u>e geometry</u> - tetrahedral
   |
   Cl   <u>geometry of molecule</u> – trigonal pyramidal
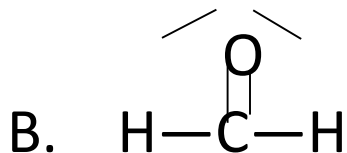
B. H—C—H   C has 3 charge clouds, <u>e geometry</u> –trigonal
   (O on top)   planar, <u>geometry of molecule</u> – trigonal planar

B. PF_5 P has 5 charge clouds, <u>e geometry</u> – trigonal bipyramidal,
             <u>geometry of molecule</u> – trigonal bipyramidal

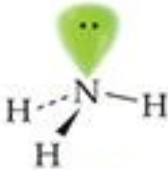**TABLE 8.1** Geometry Around Atoms with 2, 3, 4, 5, and 6 Charge Clouds

Number of Bonds	Number of Lone Pairs	Number of Charge Clouds	Geometry and Shape	Example
2	0	2	Linear	$O=C=O$
3	0	3	Trigonal planar	$H \cdots C = O$ with two H
2	1		Bent	$O \cdots S$
4	0	4	Tetrahedral	$H \cdots C - H$ with H above and H below
3	1		Trigonal pyramidal	$H \cdots N - H$ with H below
2	2		Bent	$H \cdots O$ with H below

$$\begin{bmatrix} 5 & 0 \\ 4 & 1 \\ 3 & 2 \\ 2 & 3 \end{bmatrix}$$

5

Trigonal bipyramidal

Seesaw

T-shaped

Linear

$$\begin{bmatrix} 6 & 0 \\ 5 & 1 \\ 4 & 2 \end{bmatrix}$$

6

Octahedral

Square pyramidal

Square planar

Polar Covalent Bonds and Dipole Moments

Polar Covalent Bonds and Dipole Moments

Cl$^{\delta-}$

C$^{\delta+}$

H····C H

H

Chloromethane, CH$_3$Cl

or

Cl

C

H····C H

H

Chlorine is at the negative end of the bond dipole.

Carbon is at the positive end of the bond dipole.

polar molecule

Net
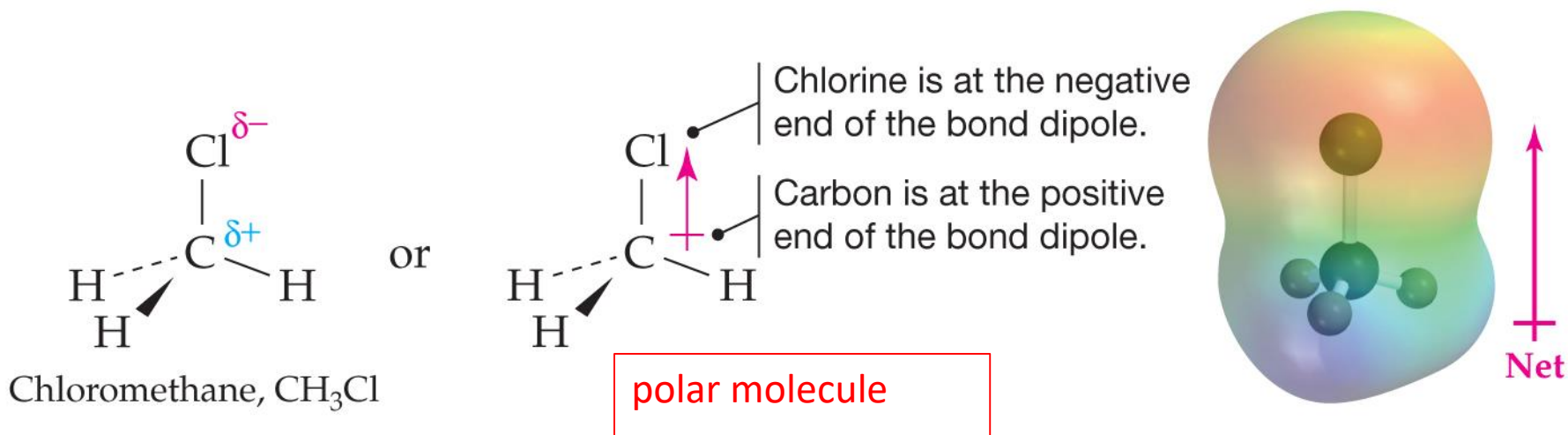
C—Cl bond has a **bond dipole** because of a difference in electronegativities. (Do vector sum of individual dipole moment arrows.)(If vector sum ≠ zero, then molecule is polar.)

Polar Covalent Bonds and Dipole Moments

Ammonia (μ = 1.47 D)
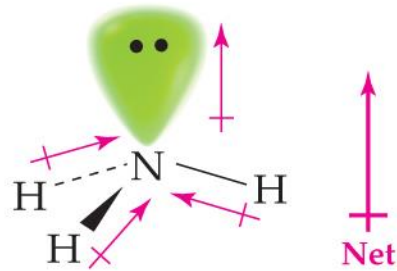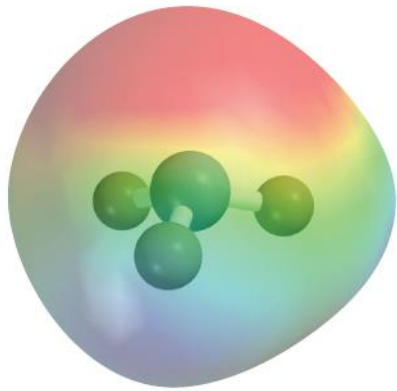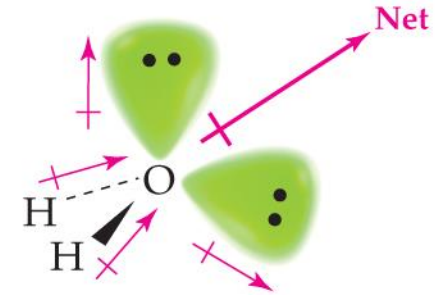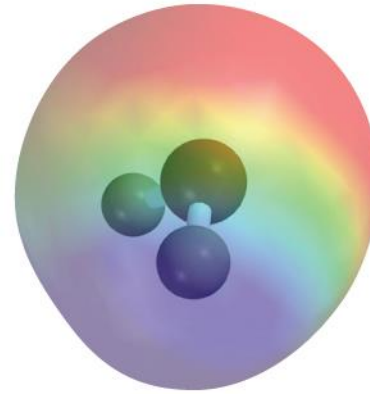
Water (μ = 1.85 D)

The individual bond polarities **do not** cancel. Therefore, the molecule has a dipole moment. In other words, the molecule is **polar**. (vector sum of individual bond dipoles $\neq$ zero, molecule polar)

Polar Covalent Bonds and Dipole Moments

Carbon dioxide ($\mu = 0$)

Tetrachloromethane ($\mu = 0$)
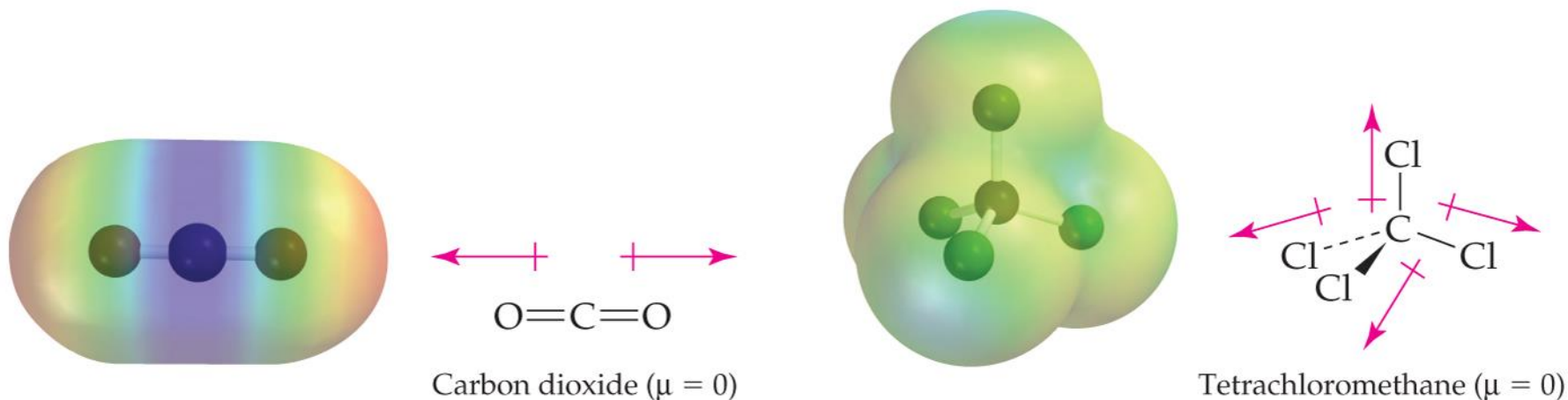
The individual bond polarities cancel. (vector sum of individual dipoles = zero) Therefore, the molecule does not have a dipole moment. In other words, the molecule is **nonpolar**.

Intermolecular Forces

Intermolecular Forces: Attractions between "molecules" that hold them together. These forces are electrical in origin and result from the mutual attraction of unlike charges or the mutual repulsion of like charges. (like interstate highways)

Types of Intermolecular Forces (higher Intermolecular Force, higher MP/BP)
- Ion–dipole forces (variable strength)
- Van der Waals forces
  - hydrogen bonds (strongest intermolecular)
  - dipole–dipole forces
  - London dispersion forces (weakest intermolecular)

Intermolecular Forces

Dipole–Dipole Forces: The result of electrical interactions among permanent dipoles on neighboring molecules
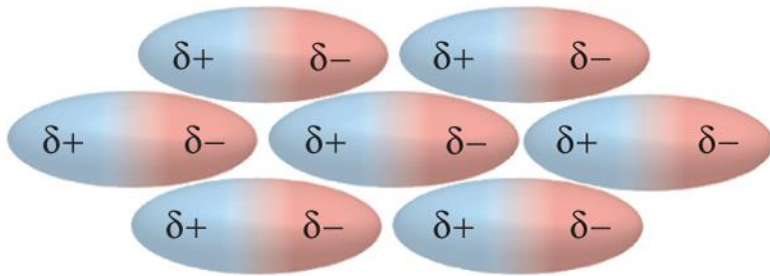
Polar molecules **attract** one another when they orient with unlike charges close together, but …
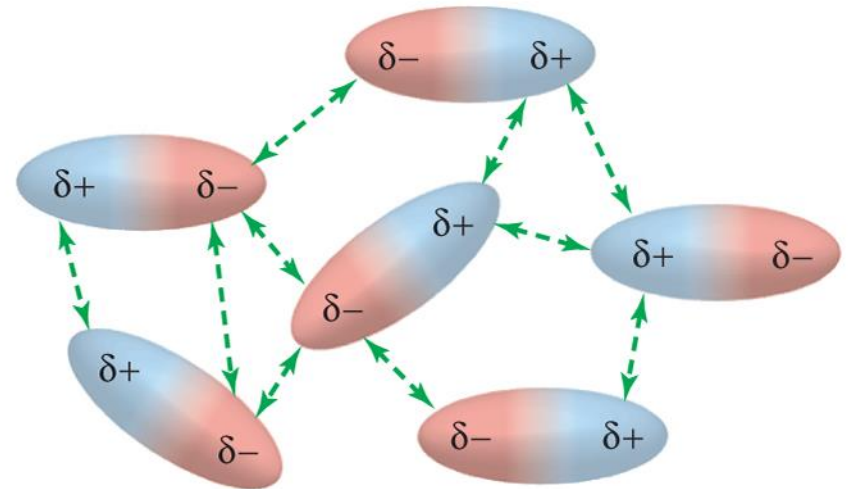
… they **repel** one another when they orient with like charges together.

Intermolecular Forces

Dipole–Dipole Forces

TABLE 8.4	Comparison of Molecular Weights, Dipole Moments, and Boiling Points		
Substance	Mol. Wt.	Dipole Moment (D)	bp (K)
$CH_3CH_2CH_3$ London	44.10	0.08	231
CH_3OCH_3 dipole-dipole	46.07	1.30	248
CH_3CN dipole-dipole	41.05	3.93	355

As the *dipole moment increases*, the *intermolecular forces increase*.

As the *intermolecular forces increase*, the *boiling point increases*.

Intermolecular Forces

London Dispersion Forces: The result of the motion of electrons that gives the molecule a short-lived dipole moment. This induces temporary dipoles in neighboring molecules. (other texts call this Van der Waals forces)
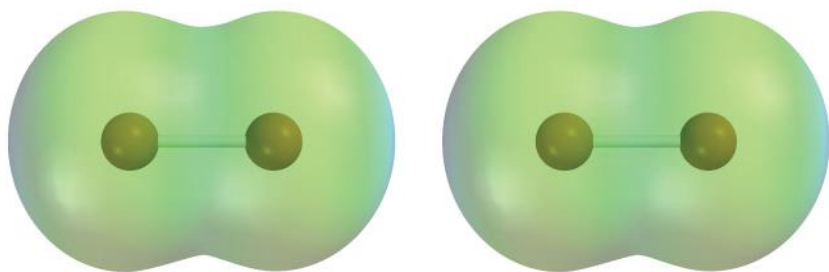
Averaged over time, the electron distribution in a Br_2 molecule is **symmetrical**.
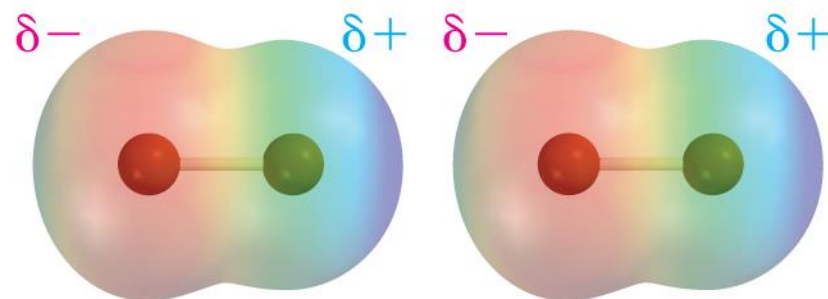
At any given instant, the electron distribution in a molecule may be **unsymmetrical**, resulting in a temporary dipole and inducing a complementary attractive dipole in neighboring molecules.

Intermolecular Forces

London Dispersion Forces

TABLE 8.5	Melting Points and Boiling Points of the Halogens	
Halogen	**mp (K)**	**bp (K)**
F_2	53.5	85.0
Cl_2	171.6	239.1
Br_2	265.9	331.9
I_2	386.8	457.5

As the *dispersion forces* increase, the *intermolecular forces* increase. As the *intermolecular forces* increase, the *boiling point* increases.   (larger molecules have higher London force, higher MP/BP)

Intermolecular Forces

London Dispersion Forces

(higher with higher surface area of contact)
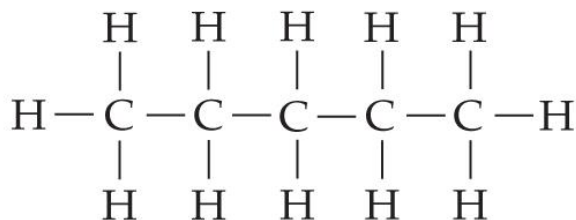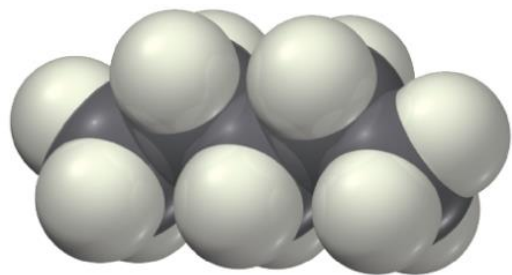
Pentane (bp = 309.2 K)
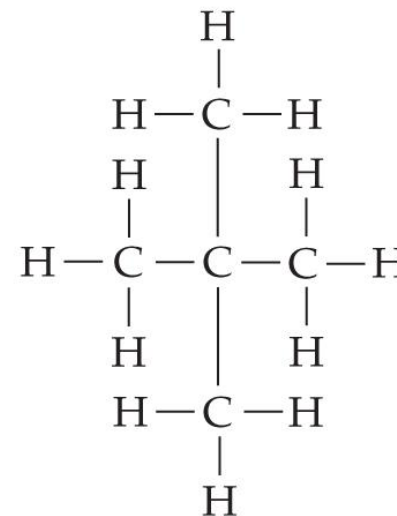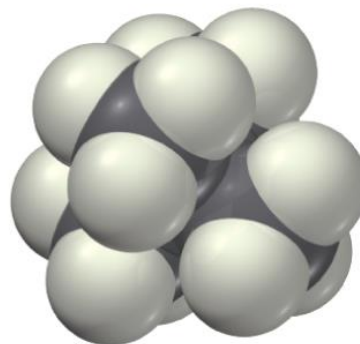
2,2-Dimethylpropane (bp = 282.6 K)

Longer, **less compact molecules** like pentane feel stronger dispersion forces and consequently have higher boiling points.

higher London, higher MP/BP

More compact molecules like 2,2-dimethylpropane feel weaker dispersion forces and have lower boiling points.

lower London, lower MP/BP

Intermolecular Forces

Hydrogen Bond: hydrogen atom directly bonded to a very electronegative atom (F, O, N) (almost FUN) and an unshared electron pair on another electronegative atom
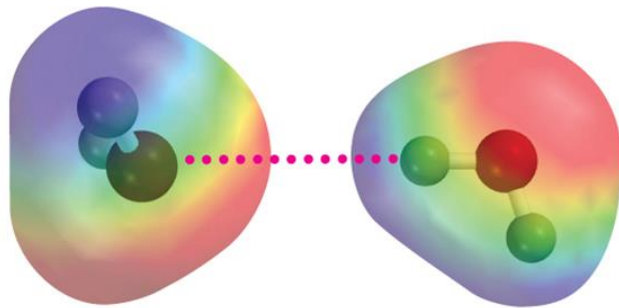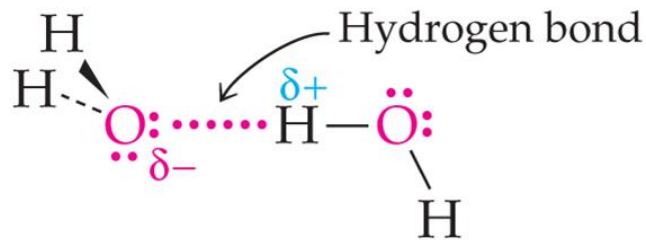
Water
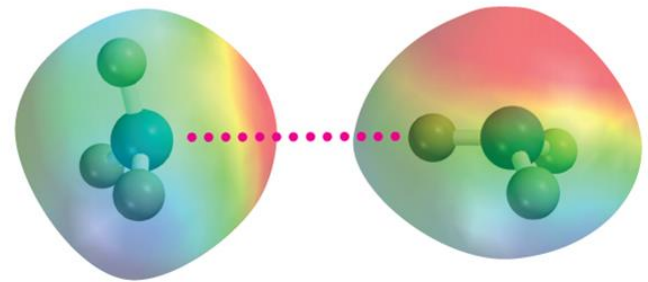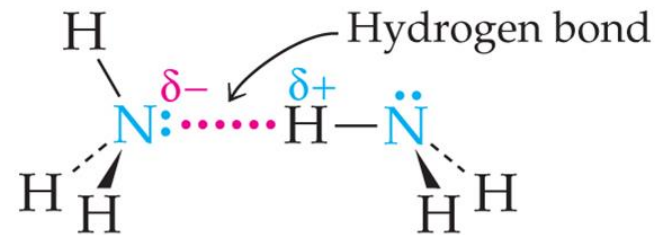
Ammonia

Intermolecular Forces

Hydrogen Bond

Liquid water contains a vast three-dimensional network of hydrogen bonds resulting from the attraction between positively polarized hydrogens and electron pairs on negatively polarized **oxygens**.

An **oxygen** can form two hydrogen bonds, represented by dotted lines.

TABLE 8.7 A Comparison of Intermolecular Forces

Force	Strength	Characteristics
Ion–dipole	Highly variable (10–70 kJ/mol)	Occurs between ions and polar molecules
Dipole–dipole	Weak (3–4 kJ/mol)	Occurs between polar molecules
London dispersion	Weak (1–10 kJ/mol)	Occurs between all molecules; strength depends on size, polarizability
Hydrogen bond	Moderate (10–40 kJ/mol)	Occurs between molecules with O—H, N—H, and F—H bonds

ion-ion > ion-dipole > H bonding > dipole- dipole > London dispersion
Largest Intermolecular Force   Smallest Intermolecular Force

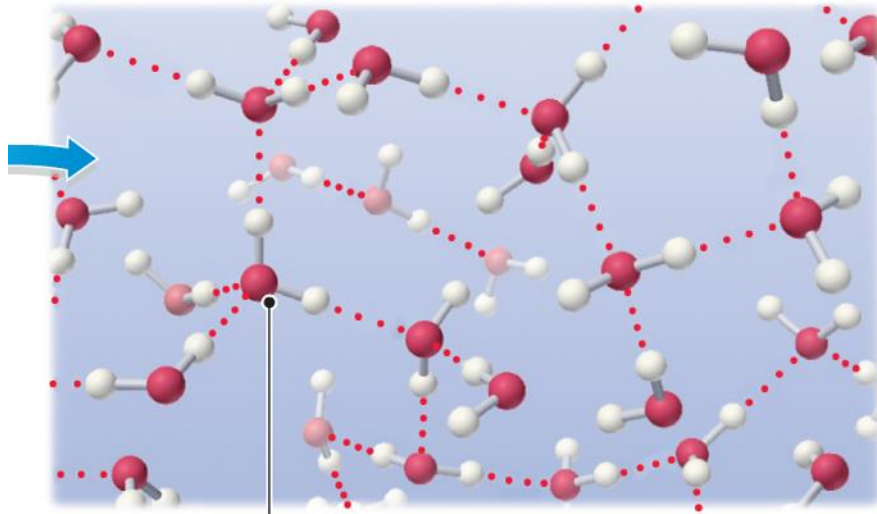Everything has London.   If also has one of the other intermolecular forces (ion-dipole, dipole-dipole or H bonding) than list that as the predominant Intermolecular Force.

Compare 2 molecules of comparable size for the above (largest to smallest intermolecular force.) If one molecule has a really large formula mass, dispersion may be a larger intermolecular force than dipole or H bonding.

HW 8-4: Intermolecular Forces

**What is the predominant intermolecular force
in the following ? Explain. (Lewis Dot, VSEPRT, Vector sum of dipoles) (H bonding,
dipole-dipole, London)**

a) CCl_4 London dispersion (tetrahedral, zero dipole moment)

b) $CHCl_3$ dipole-dipole   (tetrahedral, non zero dipole moment)

c) NF_3 dipole-dipole   (trigonal pyramidal, non-zero dipole moment)

d) NF_2H hydrogen bonding   (trigonal pyramidal, non zero dipole
moment but also H directly attached to F,O,N – so H bond)

e) H_3C O H   hydrogen bonding   (Carbon has tetrahedral, carbon has non
zero dipole moment, but H directly attached to F,O,N – so H bond)