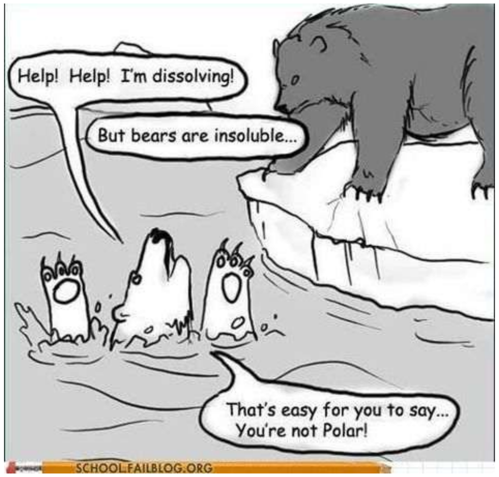
Module 7: Intermolecular Forces – Lecture 1



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"Relax, Ted, it's only a phase!"

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Learning Objective	Openstax 2e Chapter
Intermolecular Forces	10.1

Suggested Practice Problems

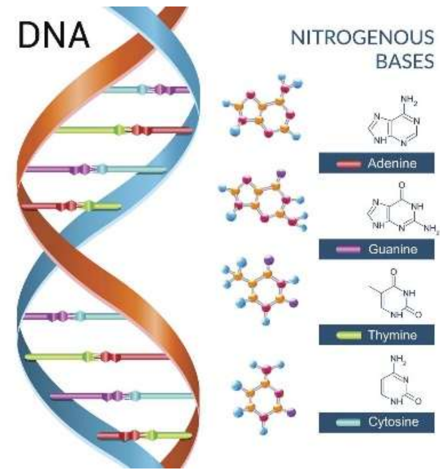
Chapter 10 Exercises – Questions: 7, 9, 11, 13, 15, 17, 21

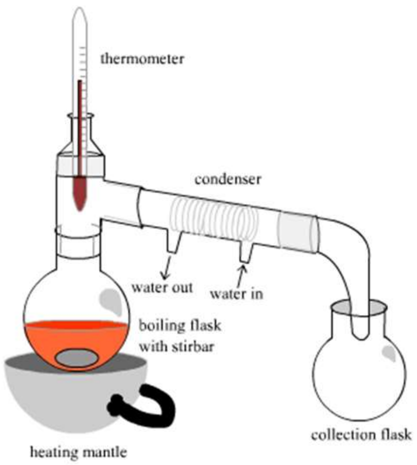
Answers can be found in the Chapter 10 Answer Key

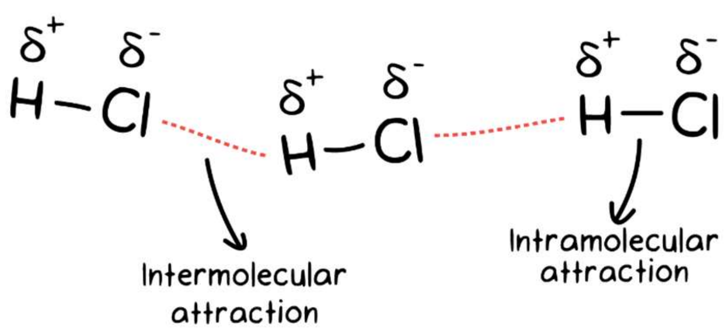
Intermolecular Forces: Bigger Picture Questions

Important properties/applications:

- Superhydrophobic, Hydrophobic, Hydrophilic Surfaces and Coatings
- Chemical Separations (e.g., Gas Chromatography, Azeotropes, Distillation)
- Protein and DNA structure, replication
- Drug/Protein Interactions, Molecular shape

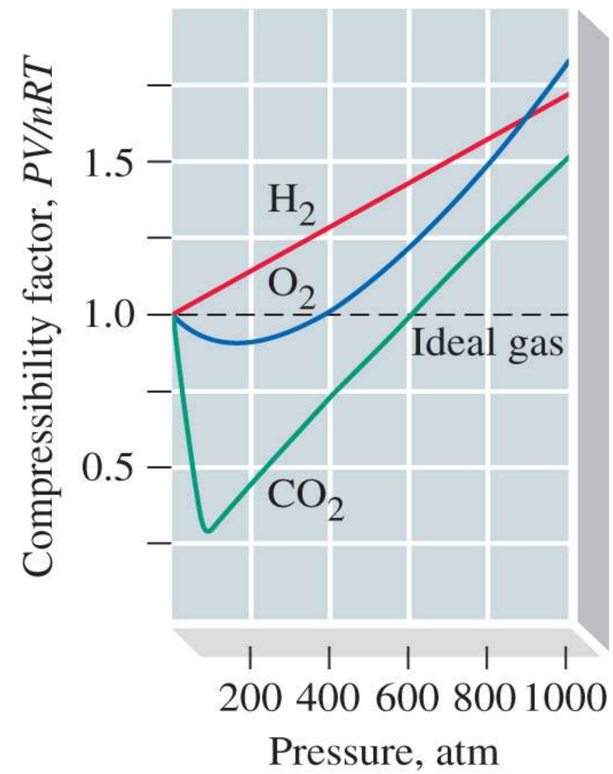






Consequences of Intermolecular Forces

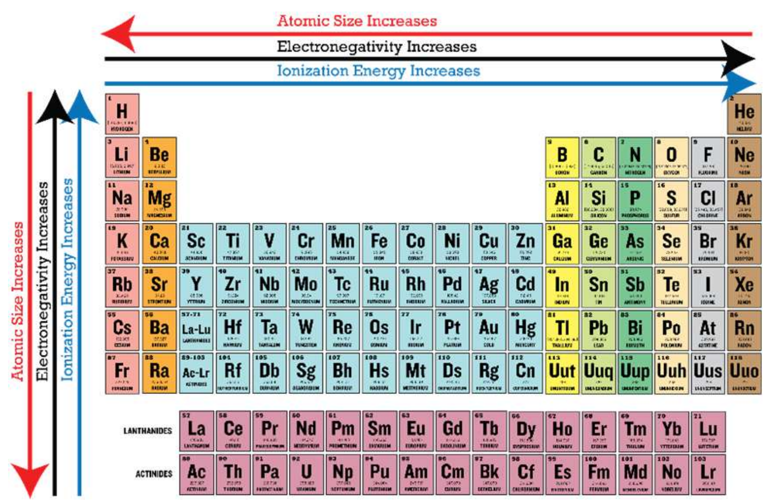
- Earlier we have seen that at *high pressures* and *low temperatures,* **intermolecular forces** cause gas behaviour to depart from ideality
- When these forces are sufficiently strong, a gas condenses to a liquid
- **Intermolecular forces** can keep molecules in such close proximity that a liquid has a certain relatively high density at a given pressure and temperature.

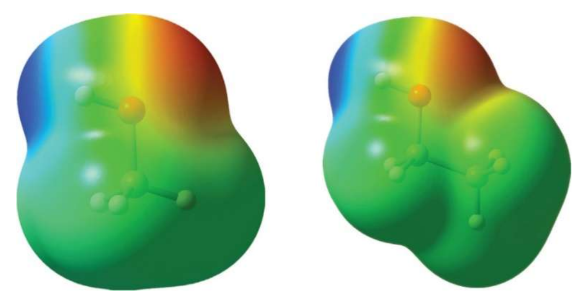


See Mod-3-gases-L4

Van der Waals Forces

- van der Waals forces: All intermolecular forces cause the non-ideal behavior of gases (van der Waals equation from Ch. 9.6)
- Molecular properties such as **dipole moment** and **polarizability** are important physical properties for attractive intermolecular forces.
  - are governed by the distribution of electron density in a molecule, related to the size and electronegativity of atoms in molecules



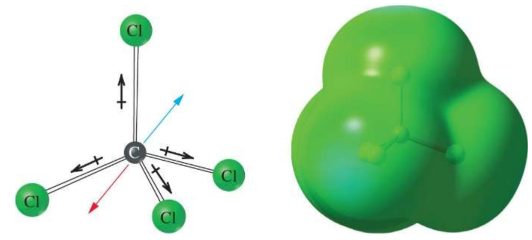


Electrostatic potential maps in methanol and ethanol

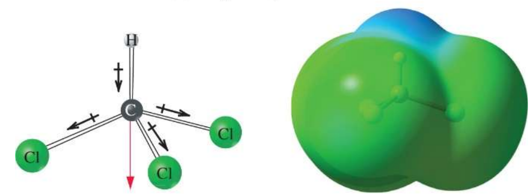
Dipole Moment and Polarizability: for Molecules

- A **dipole moment** (μ) is produced in a molecule when bond dipoles within the molecule do not cancel out.
  - Can examine and sum bond dipoles (magnitude and direction) to determine a net dipole for a molecule.

- **Polarizability** (α) of a molecule measures how easily the electron cloud can be perturbed from its "average" shape.
  - Recall for atoms, increases with size.
  - General increase for molecules with larger atoms.
  - *Note: unit of polarizability is same as volume*
  - The perturbation can be caused by the approach of an ion or a molecular dipole, or by the application of an electric field.



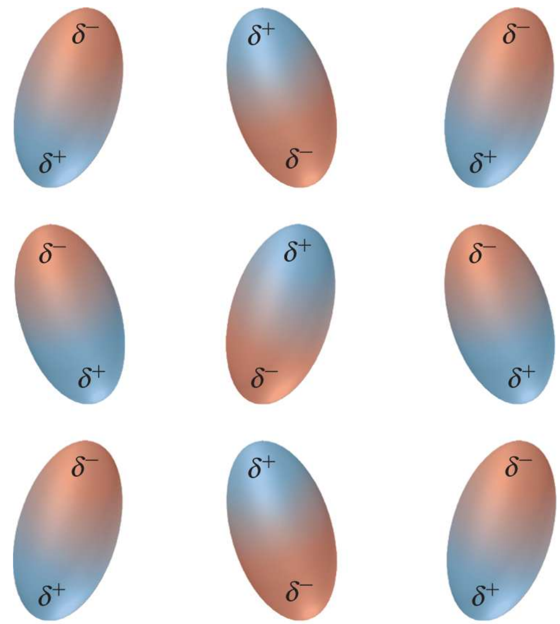
(a) CCl_4: a nonpolar molecule



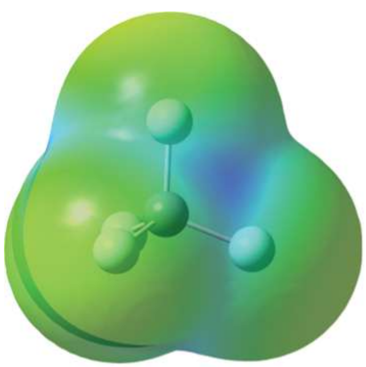
(b) $CHCl_3$: a polar molecule

Dipole-Dipole Interactions

- Interaction between a pair of polar molecules that have a **permanent dipole**
- Molecules line up with a positive end of one molecule aligned with the negative end of another molecule
- Relatively strong

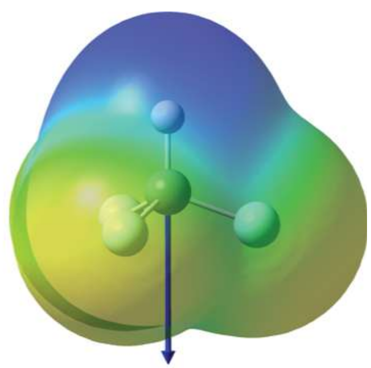


Permanent dipoles attract each other (positive to negative), also called head-to-tail interactions



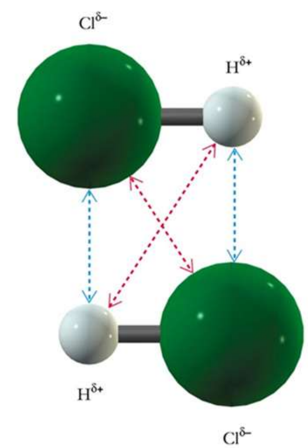
CF_4
$\mu = 0$ (nonpolar)
$\alpha = 3.8 \times 10^{-24}\ \text{cm}^3$

Boiling point: −127.8 °C



CHF_3
$\mu = 1.65$ D (polar)
$\alpha = 3.6 \times 10^{-24}\ \text{cm}^3$

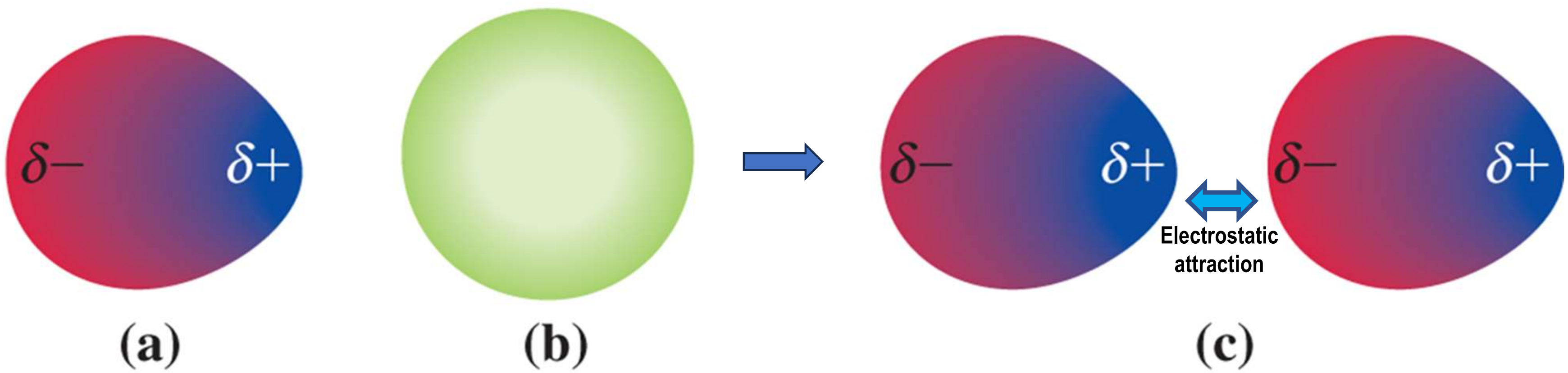
Boiling point: −82.1 °C

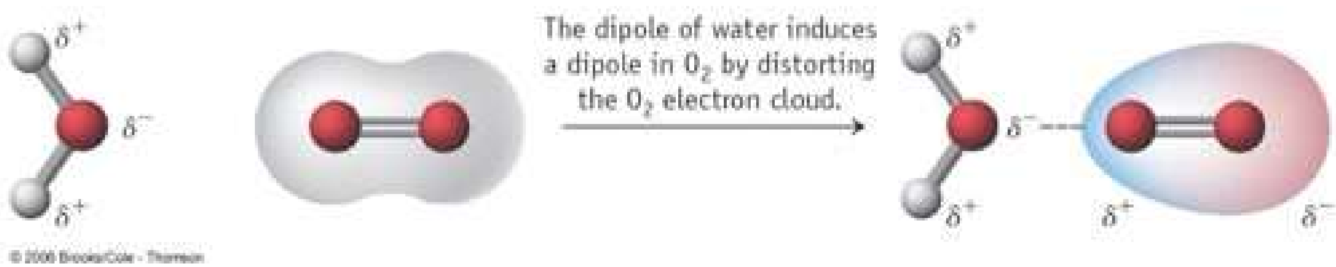


Permanent dipoles can also have side-to-side interactions with dipoles oriented positive-with-negative

Dipole-Induced Dipole Interactions

- Dipole-induced dipole interactions can increase interaction of two molecules as the dipole moment on one molecule can induce a dipole in another non-polar molecule.
- Note: ion can also induce dipole, produce similar ion-induced dipole interaction

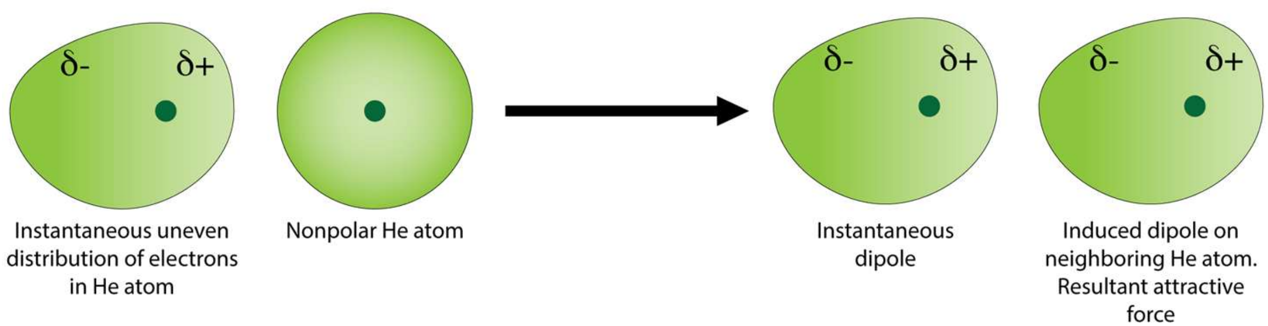




London Dispersion Forces

- London dispersion forces are attractive forces, even between non-polar molecules
- Begins with a temporary dipole (also called an instantaneous dipole or spontaneous dipole) caused by random shift of electron density to one side of a non-polar molecule or atom, depends on polarizability
- Instantaneous dipole can induce a dipole in another non-polar molecule or atom. This is called an instantaneous dipole-induced dipole interaction
- London dispersion forces are relatively weak
- For non-polar molecules (and pure elements), more polarizable molecules generally have higher boiling points

B.P.
He: -268.9 °C
Ne: -246.05 °C
Ar: -185.8 °C



Some Properties of Selected Nonpolar Compounds

weak IMF

strong IMF

Melting and Boiling Points of Halogens				
Molecule	Total Number of Electrons	Melting Point (°C)	Boiling Point (°C)	Physical State at Room Temperature
F_2	18	−220	−188	gas
Cl_2	34	−102	−34	gas
Br_2	70	−7	59	liquid
I_2	106	114	184	solid

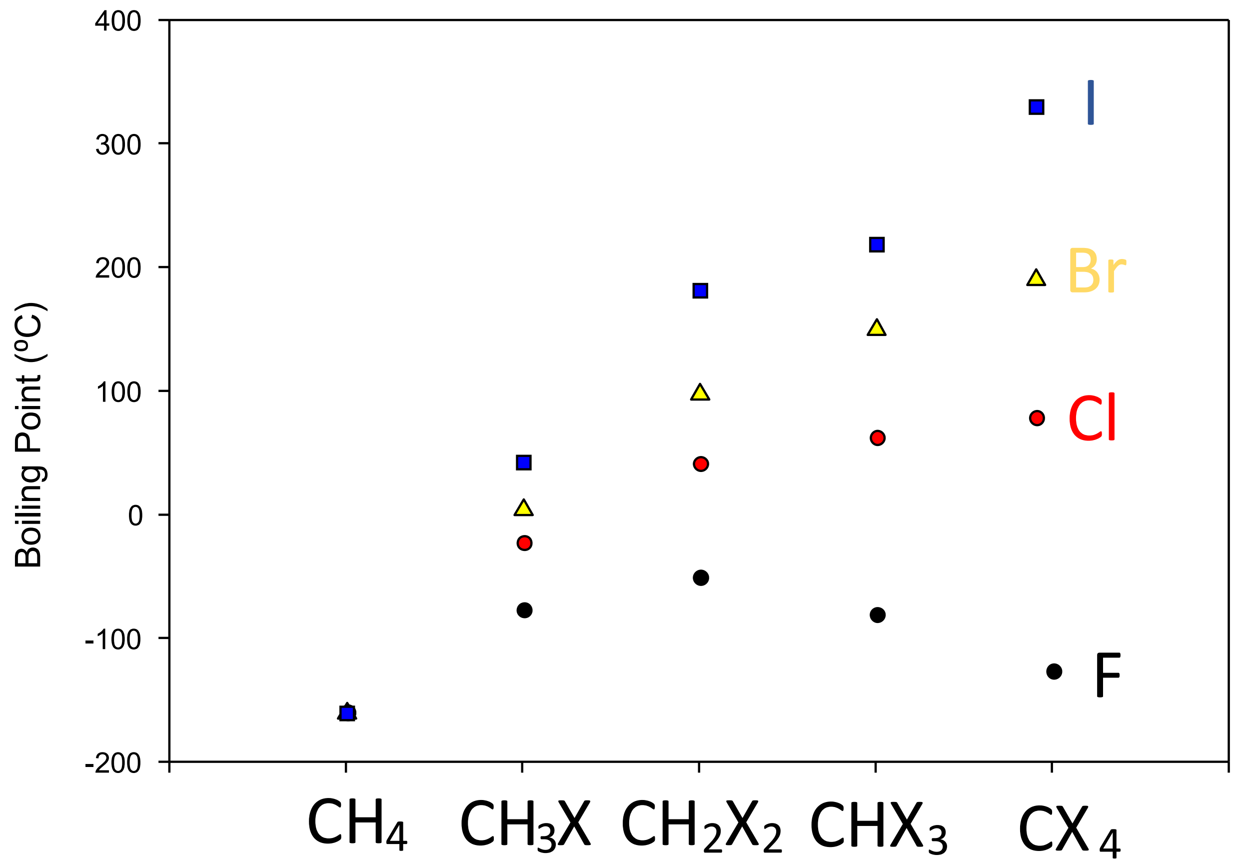
TABLE 12.1 Some Properties of Selected Nonpolar Compounds

Compound	Molar Mass, u	Polarizability,* 10^{-25} cm^3	Boiling Point, K
H_2	2.016	8.04	20.35
O_2	32.00	15.7	90.19
N_2	28.01	17.4	77.35
CH_4	16.04	25.9	109.15
CH_3CH_3	30.07	44.7	184.55
Cl_2	70.90	46.1	238.25
$CH_3CH_2CH_3$	44.10	62.9	231.05
CCl_4	153.81	112	349.95

*Sometimes polarizability is referred to as polarizability volume. Note that the units of polarizability given above have the units of volume. Thus, polarizability provides a measure of the atomic or molecular volume. Polarizability values are from the *CRC Handbook of Chemistry and Physics*, 93rd edition.

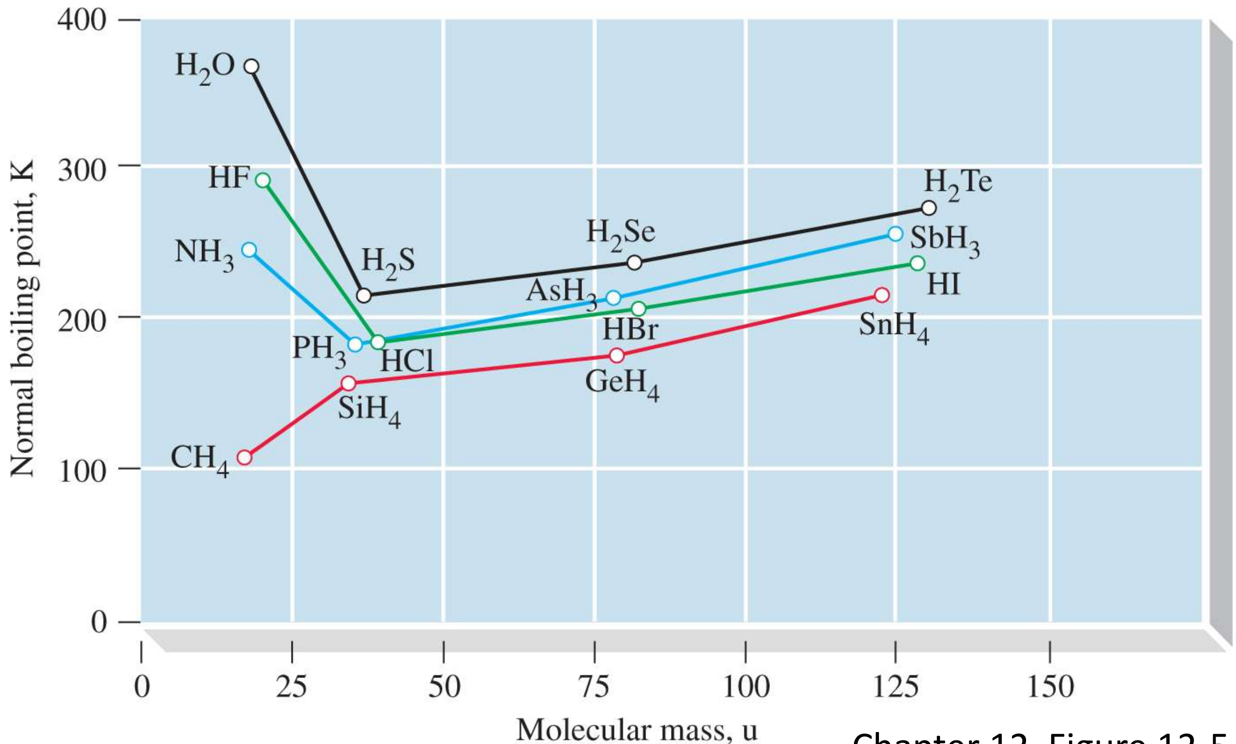
Another trend: More electrons (correlates with larger atoms) results in greater polarizability, stronger London dispersion forces

Boiling Point Trends for Halocarbons



What causes the differences in BPs and their trends for different halocarbons?

Boiling points of some hydrides of the elements of groups 14, 15, 16, and 17

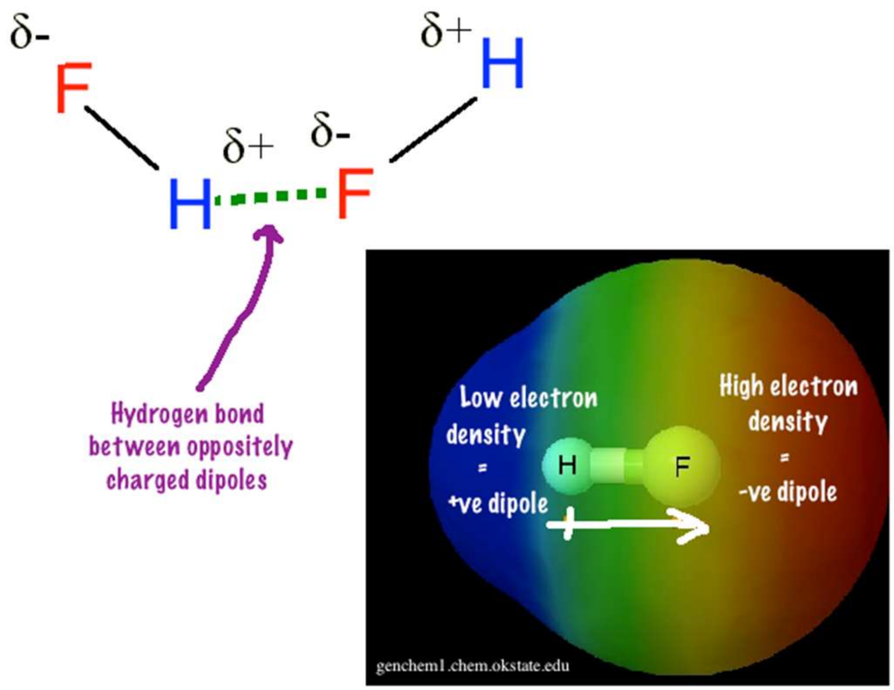


Chapter 12, Figure 12-5

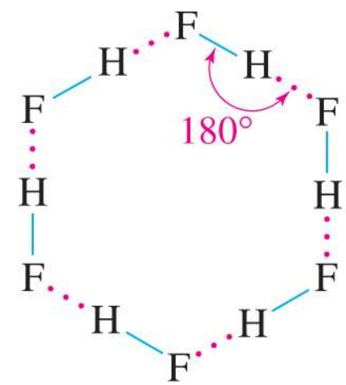
The values for NH_3, H_2O, and HF are unusually high compared with those of other members of their groups. Why?

Hydrogen Bonding

In a hydrogen bond, an H atom is covalently bound to a highly electronegative atom (e.g., **F**, **O, N,** Cl, Br)



Hydrogen Fluoride (HF)



In gaseous HF, many of the molecules are arranged into cyclic $(HF)_6$ structures

Hydrogen Bonding

- H atom is covalently bonded to a highly electronegative atom (typically F, O, or N, sometimes Cl or Br)

- In these examples of hydrogen bonding, we have the ***bond between two separate molecules***, which is termed ***intermolecular Hydrogen Bonding.***