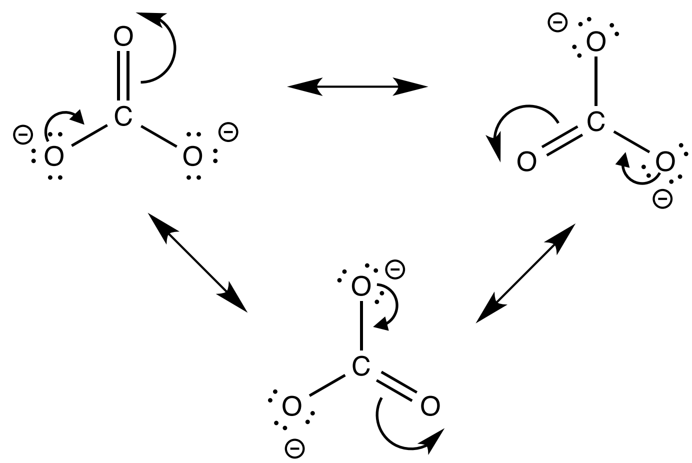
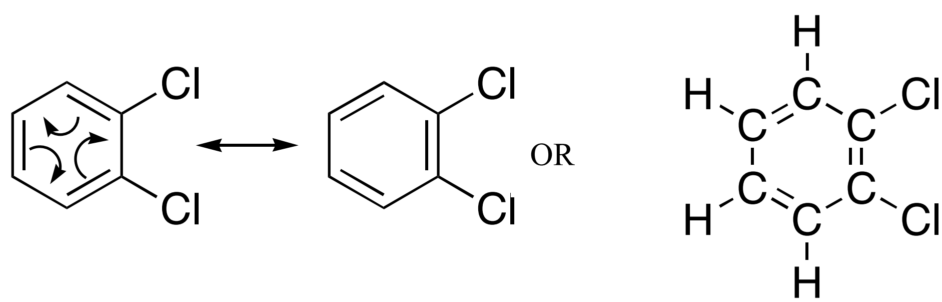
Resonance Structures:

1. Carbonate anion (CO_3^{2-})



The structures above are all equally valid. Only one needs to be drawn.

2. 1,2-Dichlorobenzene



Electronegativity:

- An atom's desire for electrons (negative charge).
- On the periodic table, electronegativity increases as you go from left to right (up to inert gases which are not electronegative) and as you go from down to up
    - i.e. Fluorine is the most electronegative atom (wants to gain the inert gas configuration of Ne)
- It influences acidity of H's attached, as well as the intermolecular forces between molecules.

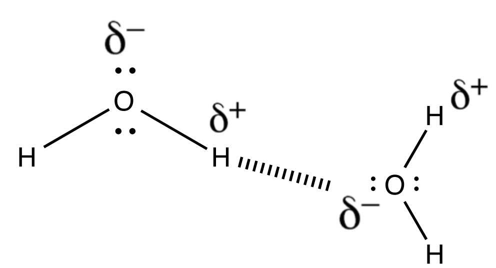
Intermolecular Forces: (forces present between molecules)

- Attractive intermolecular forces:
  - i) Hydrogen bonding – strongest on per atom basis (e.g. base recognition in forming DNA helix)
  - ii) Dipole-dipole interaction
  - iii) London forces (temporary dipole; hydrophobic bonding) – weakest on per atom basis

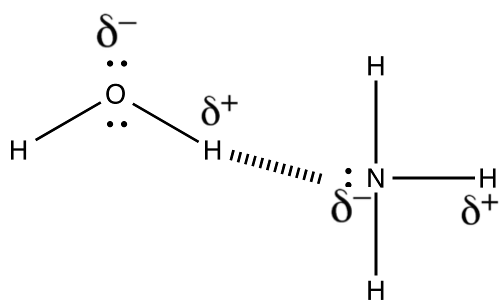
Hydrogen Bonding:

- Donors: H attached to a very electronegative atom (N, O, F, Cl)
- Acceptors: A lone pair of electrons
- Strongest intermolecular attractive force on a per atom basis

e.g. H-O-H (water)



- Oxygen is electronegative and it is sp^3 hybridized
- The partial positive charge on H and the partial negative charge on O lead to their attraction
- Results in high boiling point and high melting point by self-association
- HF, H_2O and NH_3 form hydrogen bonds



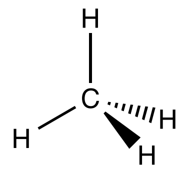
- Water is a liquid at RT while ammonia is a gas
- Oxygen is more e-neg than nitrogen, so the protons on water have a higher positive partial charge than the protons on ammonia
- In an ammonia solution, water would be the hydrogen bond donor and ammonia would be the acceptor
- Water dissolves ammonia very well – up to 18M

Dipole-Dipole Interactions:

Dipole drawing convention:

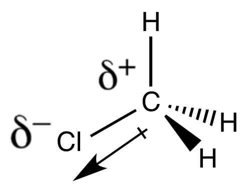
Partial positive charge is the "plus" end, partial negative charge is the arrow head

1. Methane; CH_4



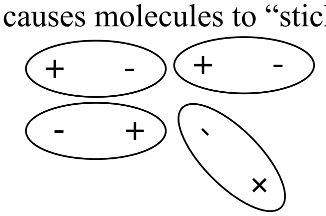
- C and H have ~same electronegativity
- Non-polar (net-zero ~dipole)
- Gaseous
- Low BP   -164 ^oC
- Low MP -182 ^oC

2. Chloromethane, methyl chloride; CH_3Cl

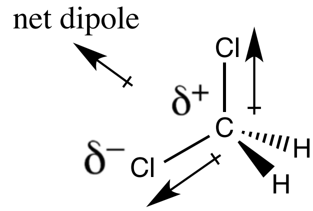


- H and C have similar electronegativity values (non-polar bond)
- Cl is very electronegative due to the fact that it only needs one electron to get inert gas configuration.
- Electron density is pulled toward the chlorine atom, creating a net dipole toward chlorine atom. A net dipole is the vector sum of individual bond dipoles.

Dipoles in different molecules tend to line-up temporarily with each other (partial positive / negative charge on the molecule) – causes molecules to "stick" to each other

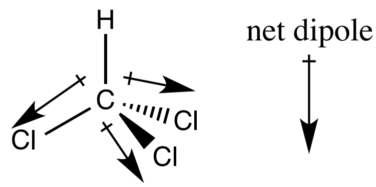


3. Dichloromethane, methylene chloride; CH_2Cl_2



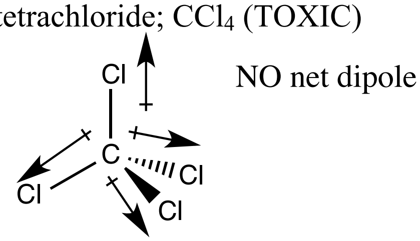
- Liquid at room temperature BP 40 ^{o}C   MP – 95 ^{o}C
- More polar than chloromethane
- Not miscible with water

4. Trichloromethane, chloroform; $CHCl_3$



- More polar than methylene chloride   BP 61 ^{o}C   MP – 64 ^{o}C

5. Tetrachloromethane, carbon tetrachloride; CCl_4 (TOXIC)

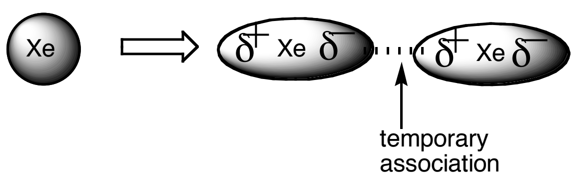


- Non-polar molecule (net-zero dipole)
- Has temporary dipoles since chlorine is polarizable (see below), BP ~77
- Historically used as a dry cleaning fluid

London Forces (Temporary Dipoles):

- Also know as London dispersion forces
- Distortion of filled outer shell electrons
- Principal effect in hydrophobic interactions

Atoms	Boiling Point	
He	-269 ^oC	Small atom/ Low polarizability
Ne	-246 ^oC	↓
Ar	-186 ^oC	
Kr	-153 ^oC	
Xe	-108 ^oC	Large atom/ High polarizability



- The larger the atom (expanded electron density), the easier the formation of temporary dipoles

CH_4 associates with CH_4 due to London forces

C_9H_{20} hydrophobic bonding