Final Exam Review – Day 3

Bonding:

Ionic – electrons are transferred, forming cations and anions ($Cl + e^- \rightarrow Cl^-$; $Na \rightarrow Na^+ + e^-$) $\rightarrow Na^+ + Cl^- \rightarrow NaCl$

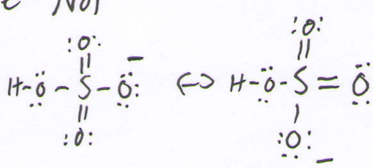
- ΔEN is very large, usually between a metal and non-metal
- Smallest unit is formula unit (e.g. NaCl for sodium chloride)
- high melting and boiling points
- tendency to fracture
- good electrolytes when melted or dissolved in solution

Covalent – e^- shared
- ↗ equally ⇒ non-polar ⇒ ΔEN zero or very small
- ↘ unequally ⇒ polar ⇒ ΔEN large

Smallest unit is a molecule (e.g. H_2O for water)

Lewis structures:
1) Use all and only total # of valence e^- in molecules
2) Least EN atom usually in center (except H)
3) Try to satisfy octet rule on all atoms
4) Minimize formal charges
5) Draw resonance structures
6) Elements with n > 2 can violate octet rule
7) Elements in group 1 usually make one bond (H–H)
   - group 2 " " 2 bonds (:Cl–Be–Cl:)
   - group 3 " " 3 bonds (H–B–H)
   - group 4 " " 1, 2, 3, 4 bonds (C≡O, H–C–H ...)
   - group 5 " " 3 bonds (H–N–H)
   - group 6 " " 2 bonds (H–O:)
   - group 7 " " 1 bond → (H–Cl:)
   - group 8 " " 0 bonds → (Xe, Ne)

Resonance: move e^- NOT nuclei



- * no rotations
- * no new bonds
- * look for resonance any time you see double + single bonds in same molecule, especially if there are any formal charges

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Metallic – e^- delocalized in electron sea; all valence e^- shared by all nuclei

high electrical and thermal conductivity

high malleability and ductility

high melting and boiling points

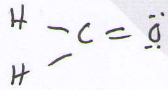
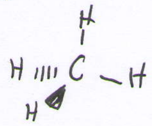
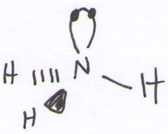
alloys = solutions of two or more elements, the combination of which has metallic properties

VSEPR Theory – geometry around an atom in a molecule puts e^- in bonds / lone pairs as far away from other e^- as possible

lone pairs = more repulsive than e^- in bonds, squeezes bonds closer, smalle

electron group = a lone pair of two e^- OR any bond between two atoms (single / double / triple)

must know these geometries for final exam:

e^- groups	lp	geometry	example
2	0	linear	Ö = C = Ö
3	0	trigonal planar	
4	0	tetrahedral	
4	1	trigonal pyramidal	
4	2	bent	

Polar bond = bond between atoms w/ different electronegativities; $\Delta EN > 0$

IF $EN(X) > EN(Y)$ => $\delta^- X - Y \delta^+$

⟵+ => bond dipole

Polar molecule = molecule w/ overall non-zero molecular dipole

vector sum of bond dipoles

bond dipoles cancel

non-polar molecule ⟸ Ö = C = Ö

molecu dipole

polar molecule ⟸

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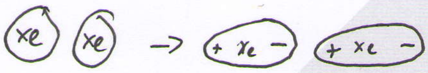
Intermolecular forces: forces between molecules, much weaker than bonds

- Three types:
  1) Dipole-dipole (DD) => in polar molecules; (+) end of one molecule attracted to (−) end of another
     $:\overset{-}{C}\equiv\overset{+}{O}:$ $:\overset{-}{C}\equiv\overset{+}{O}:$
  2) Hydrogen bonding (HB) => X — H --- Y (covalent bond X—H, HB H---Y; X, Y = O, N, F, Cl); very polar bond between small H and very EN atom
     
  3) London Dispersion forces (LDF) => present in all molecules; attraction of instantaneous and induced dipoles
     
- Melting/Boiling involves separating molecules in a sample, overcoming attractive intermolecular forces
- Comparing melting/boiling points: First, identify all IMFs in a sample
  Usually, MP/BP (ionic) > MP/BP (HB) > MP/BP (DD) > MP/BP (LDF)
  If two samples have same IMFs, need to consider other factors

Intermolecular Forces:
Comparing Melting/Boiling Points

Ionic Solids
- * $F = kq_1q_2/r^2$
- * $q_1q_2 < 0$
- * F↓, attraction↑, MP/BP↑
- * Al_2O_3 vs NaCl

Hydrogen Bonding
- * # H-bonds per molecule: LP/H or H/LP
- * if same # H-bonds/molecule, look at strength of HB: EN(F) > EN(O) > EN(N) = EN(Cl)
- * HF vs H_2O
- * HF vs NH_3

Dipole-dipole
- * Compare size of molecular dipoles
- * Not easy

London Dispersion forces
- * Increase w/ increasing # e^- and increasing size

IMPORTANT: though usually ionic > HB > DD > LDF, this is not always true! LDF in large molecules (HBr) can be greater than even DD + HB in smaller molecules (HCl)