Lecture 3

Thursday, April 03, 2008
7:55 PM

References Used:

1. G Barrow, **Physical Chemistry, 5th Edition**, *McGraw Hill,* 1988.

Course Notes:

- **Question:** when are the two mid term exams?
  Answer:
  - 1st exam will be on Monday April 21st
  - 2nd Exam will be on Monday May 19th

- There were NO LABs this week. The first labs will be next week.
- The "Calendar and Homework" link on the website shows the course schedule/ homework assignments/ and Lab Schedule

Review of Last Lecture:

- We talked about amorphous materials and nanocrystalline materials
- We talked about how big the proton, electron, and atom are
- We talked about what an amu is and about the periodic table
- We talked about how everything wants to be in the lowest possible energy state
- We talked about how all the elements want to have a noble gas electron configuration
- We talked about electron orbitals and electron configurations
- We talked about the Leonard Jones Potential Diagram describing molecular bonding energy and position

What does that mean that the atoms are bonded together? How do they bond together?

Primary (Chemical) Bonds:

Ionic Bonding (Ceramics)

- Combination of metallic and non-metallic elements
- Metals easily give up their electrons
- Non-metals will take electrons
- Everything acquires inert gas electron configurations
- The ions are then attracted to one another by coulombic forces (electrical charge)
- Non-directional bonding
- Most ceramics have ionic bonding
- Bond energies 600- 1500 kJ/mol
- Hard and Brittle
- Electrically insulative

Bond Type	Substance	Bonding Energy (kJ/mol)	Melting Temp (°C)
Ionic	MgO	1000	2800
Covalent	C(diamond)	713	>3550
Metallic	Hg	68	-39
	Fe	406	1538
	W	849	3410
van der Waals	Ar	7.7	-189
Hydrogen	H_2O	51	0

Covalent Bonding (Plastics)

- Inert gas configurations are created by "sharing" electrons
- Sharing orbitals -- overlapping atomic orbitals:
- Actually we are solving the wave function for two nuclei and two electrons
- End up with bonding and antibonding functions
  - Molecular hydrogen: H-H : overlaping s orbitals
  - Molcular Cl-Cl: over lapping p orbitals

- Carbon (critical to all organic chemistry, plastics, and carbon energy sources) can form 4 bonds
  - Methane example is shown in the book -- book is wrong!
  - Why?:
    - Does this make any sense?
    - Carbon has 6 electrons
    - What is the atomic structure of carbon: $1s^2 2s^2 2p^2$
    - How many valence electrons does carbon have? Ans.: 2
    - How do we get 4 bonds then?
    - Professor Linus Pauling from California Institute of Technology won the Nobel Prize in Chemistry in 1954 for answer
    - Theory of hybridization
    - 2s and 2p orbitals can be combined to form new orbitals -- hybridization
    - The most suitable solutions can be found by forming wave functions which project farthest from the central atom -- concentrated along tetrahedral directions
    - Create sp^3 hybrid orbitals which are tetrahedrally oriented and describe the bonds in CH_4

	Atomic Number	Atomic Weight (amu)	Atomic Radius (nm)	Melting Point (°C)	
Hydrogen	1	1.008	--	-259	
Carbon	6	12.011	0.071	3367	sublimes
Oxygen	8	16.00	--	-218.4	
Aluminum	13	26.98	0.143	660.4	
Iron	26	55.85	0.124	1538	
Tungsten	74	183.84	0.137	3410	
Lead	82	207.2	0.175	327	

Metallic Bonding (Metals)

- Metallic elements (d orbitals) -- have many valence electrons 1, 2, 3
- Valence electrons form a "cloud" of electrons around ion cores
- Electrons charges shield the cores from each other and prevent repulsion
- Non-directional bonding
- Free electrons act as "glue"
- Bonding energies vary
- Good conductors -- why?
- Ductile behavior -- we'll discuss a in a lot more detail through course

Secondary (Physical) Bonds

Van Der Waals Bonding

- Dipole bonding -- either molecular or atomic
- Fluctuating induced dipoles
- Polar molecule induced dipoles
- Permanent dipole bonds -- hydrogen bonding

End of Chapter 2