SP 08 Chem 110 Exam I
Mon. 2/11
Time 6:30-7:45 PM

Location:
posted on web; know your section!!

Bring:
#2 pencils
PSU photo ID
calculator - may not be text programmable

Know:
Chem. 110 section number
Your PSU student ID number

Do NOT bring:
Cell phones, PDA's, or any other electronic devices

Do the sample practice exams on the web; select handouts
Do the practice exams on Angel!
Do all homework problems!
Go to the review sessions!
Go to SI and office hours!

What Should You Know?

Energy Level Diagrams
H atom vs. many-electron atoms; which orbitals are degenerate? Be able to place the orbitals in order of increasing energy. Explain why the two are different.

Calculate the orbital energy for a given principal quantum number n
- For an H atom
- For a many-electron atom

Effective Nuclear Charge
- Know the definition, explain the effect of core electrons
- Compare Z_{eff} for orbitals; be able to tell which orbital experiences the greatest Z_{eff}
- Compare Z_{eff} for atoms; be able to tell which atom experiences the greatest Z_{eff}

Electron Spin
- What is the spin quantum number, and what are the choices of values?
- What is the Pauli Exclusion Principal, and how does it govern the arrangement of electrons in orbitals?
Chapter 8: What Should You Know?

Electron Configurations

- What is Hund’s rule and why does it determine the most stable configuration?
- Write the electron configuration of any element
- Write electron configurations using condensed notation (noble gas in brackets)
- Know the difference between core and valence electrons
- Recognize the special stability of a half-filled or full shell of electrons. Know which 2 families of elements have unexpected s and d orbital configurations because they prefer to have half-filled or full shells.
- Know how to use the periodic table to determine electron configurations.

Lecture 8: Ionic bonding

Read: BLB 2.7–2.8; 8.1–8.2
HW: BLB 2:49,51,53d&e,55,59,61; 8:16,22,25
Sup 2:4; 8:1

Know:
Ionic Bonding
- Lewis Symbols
- Nomenclature
- Lattice energy

Bonus deadline for Skill Check Tests 3 & 4 is Mon., 2/4.

Exam 1: Monday, Feb 11 6:30;
Conflict Sign-up due 2/6
Data sheet p 49-50 will be on exam
Bring:
- pencils,
- student ID
- and a calculator—NO text-programmable calculators or wireless devices

Supplemental Instruction (SI) times on web
3 Types of Bonding

- **IONIC**
electrons traded to form separate ions

  both have noble gas __________

  \[ \text{Na}^+ \quad \text{Cl}^- \]

- **COVALENT**
electrons shared between a few nuclei distinct molecules

  _____ e^- __ __ e^- 

- **METALLIC**
electrons shared among all nuclei

  “sea of electrons”

For ionic and covalent bonding:

**Keep track of electrons with Lewis Dot Symbols**

1. __________ electrons are involved in bonding; represent each electron with one dot

   C \[ \cdot \cdot \cdot \]

   4 valence electrons

   S [Ne] 3s^2 3p^4 \[ \cdot \cdot \cdot \]

   ____ valence electrons

2. **Octet rule:** elements tend to gain, lose or share electrons so as to gain an __________ configuration of 8 electrons.

   (Duet for H and He instead of octet)

   \[ \Rightarrow \text{Inert gases have 8 valence electrons} \]
Knowing the Number of Valence Electrons, You Should be Able to Draw These:

### TABLE 8.1 Lewis Symbols

<table>
<thead>
<tr>
<th>Element</th>
<th>Electron Configuration</th>
<th>Lewis Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>[He]2s(^1)</td>
<td>Li(^+)</td>
</tr>
<tr>
<td>Be</td>
<td>[He]2s(^2)</td>
<td>Be(^+)</td>
</tr>
<tr>
<td>B</td>
<td>[He]2s(^2)2p(^1)</td>
<td>B(^+)</td>
</tr>
<tr>
<td>C</td>
<td>[He]2s(^2)2p(^2)</td>
<td>C(^-)</td>
</tr>
<tr>
<td>N</td>
<td>[He]2s(^2)2p(^3)</td>
<td>N(^-)</td>
</tr>
<tr>
<td>O</td>
<td>[He]2s(^2)2p(^4)</td>
<td>O(^-)</td>
</tr>
<tr>
<td>F</td>
<td>[He]2s(^2)2p(^5)</td>
<td>F(^+)</td>
</tr>
<tr>
<td>Ne</td>
<td>[Ne]3s(^2)</td>
<td>Ne(^+)</td>
</tr>
<tr>
<td>Na</td>
<td>[Ne]3s(^1)</td>
<td>Na(^+)</td>
</tr>
<tr>
<td>Mg</td>
<td>[Ne]3s(^2)</td>
<td>Mg(^+)</td>
</tr>
<tr>
<td>Al</td>
<td>[Ne]3s(^2)3p(^1)</td>
<td>Al(^+)</td>
</tr>
<tr>
<td>Si</td>
<td>[Ne]3s(^2)3p(^2)</td>
<td>Si(^+)</td>
</tr>
<tr>
<td>P</td>
<td>[Ne]3s(^2)3p(^3)</td>
<td>P(^+)</td>
</tr>
<tr>
<td>S</td>
<td>[Ne]3s(^2)3p(^4)</td>
<td>S(^-)</td>
</tr>
<tr>
<td>Cl</td>
<td>[Ne]3s(^2)3p(^5)</td>
<td>Cl(^-)</td>
</tr>
<tr>
<td>Ar</td>
<td>[Ne]3s(^2)3p(^6)</td>
<td>Ar(^+)</td>
</tr>
</tbody>
</table>

 Ionic Bonding

Involves Transfer of _____________ to form charged _____________
IONIC BONDING

Ionic bond is due to electrostatic attraction.

Coulomb’s Law:

\[ Q_1 - \text{charge on 1}^{\text{st}} \text{ atom} \]
\[ Q_2 - \text{charge on 2}^{\text{nd}} \text{ atom} \]
\[ d - \text{distance between charges} \]

Electrons are exchanged to from separate ions with complete _____________.

Na\(^+\) Cl\(^-\)

1s\(^2\) 2s\(^2\) 2p\(^6\) 1s\(^2\) 2s\(^2\) 2p\(^6\) 3s\(^2\) 3p\(^6\)

[Ne] [Ar] complete octets

IONIC COMPOUNDS

- distinct +/− ions
- ∞ crystal lattice
- strong omnidirectional electrostatic bonds

Characteristics

1. low electrical conductivity as solids (high as liquids)
2. very high melting, boiling points
3. hard but brittle
4. soluble only in polar solvents (water)

Ionic compound: a compound that consists of positive and negative ions.

Metal + nonmetal

Examples: NaCl
K\(_2\)SO\(_4\)
Mg(NO\(_3\))\(_2\)
NH\(_4\)Cl

Ionic compounds have ________ melting points and are brittle (many are ________).

Molecular compounds: compounds consisting of individual molecules

All nonmetals or nonmetals plus metalloids.

Examples: HCl
H\(_2\)O
CH\(_3\)COOH
NH\(_3\)

Molecular compounds usually have ________ melting points (many are ________ or ____________).
Naming Inorganic Compounds and Ions

See Section 2.8

<table>
<thead>
<tr>
<th>Charge</th>
<th>Symbol</th>
<th>Common Cations*</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1+</td>
<td>H⁺</td>
<td>Hydrogen ion</td>
<td>NH₄⁺</td>
<td>Ammonium ion</td>
</tr>
<tr>
<td></td>
<td>Li⁺</td>
<td>Lithium ion</td>
<td>Cu⁺</td>
<td>Copper(I) or cuprous ion</td>
</tr>
<tr>
<td></td>
<td>Na⁺</td>
<td>Sodium ion</td>
<td>K⁺</td>
<td>Potassium ion</td>
</tr>
<tr>
<td></td>
<td>Ca²⁺</td>
<td>Calcium ion</td>
<td>Mg²⁺</td>
<td>Magnesium ion</td>
</tr>
<tr>
<td></td>
<td>Sr²⁺</td>
<td>Strontium ion</td>
<td>Ba²⁺</td>
<td>Barium ion</td>
</tr>
<tr>
<td></td>
<td>Zn²⁺</td>
<td>Zinc ion</td>
<td>Cd²⁺</td>
<td>Cadmium ion</td>
</tr>
<tr>
<td></td>
<td>Ag⁺</td>
<td>Silver ion</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Charge</th>
<th>Symbol</th>
<th>Common Anions*</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1⁻</td>
<td>H⁻</td>
<td>Hydride ion</td>
<td>H₂O⁻</td>
<td>Hydroxide ion</td>
</tr>
<tr>
<td></td>
<td>F⁻</td>
<td>Fluoride ion</td>
<td>Cl⁻</td>
<td>Chloride ion</td>
</tr>
<tr>
<td></td>
<td>Br⁻</td>
<td>Bromide ion</td>
<td>I⁻</td>
<td>Iodide ion</td>
</tr>
<tr>
<td></td>
<td>CN⁻</td>
<td>Cyanide ion</td>
<td>OH⁻</td>
<td>Hydroxide ion</td>
</tr>
<tr>
<td></td>
<td>O₂⁻</td>
<td>Oxide ion</td>
<td>CO₃²⁻</td>
<td>Carbonate ion</td>
</tr>
<tr>
<td></td>
<td>O₃²⁻</td>
<td>Peroxide ion</td>
<td>S₂⁻</td>
<td>Thiosulfide ion</td>
</tr>
</tbody>
</table>

*The most common ions are in boldface.

Figure 2.26

Figure 2.27

Figure 2.28
Be able to fill in information as in the table below:

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>potassium ion</td>
<td>CO₃²⁻</td>
</tr>
<tr>
<td>ammonium ion</td>
<td></td>
</tr>
<tr>
<td>sodium bicarbonate</td>
<td>H₂SO₄</td>
</tr>
<tr>
<td>Combine Ca^{2+} and PO₄^{3−}</td>
<td></td>
</tr>
<tr>
<td>Cu(II) and sulfate ion</td>
<td></td>
</tr>
</tbody>
</table>

Formation of ionic salts from the elements is very exothermic.
Na⁺(g) + Cl⁻(g) → NaCl(s)
ΔH = -788 kJ/mol

BUT
- Removing an electron from a metal is ___________________________.
  Na(g) → Na⁺(g) + e⁻
  ΔH = +496 kJ/mol

- And adding an electron to a non-metal is moderately ________________.
  Cl(g) + e⁻ → Cl⁻(g)
  ΔH = -349 kJ/mol

Overall the process seems like it would be endothermic.
Why is formation of an ionic salt so favorable?

The added stability is due to the attraction of the ions for each other. The attraction brings the ions close, releases energy, and forms a ________________.
Strength of Ionic Bond  
(Lattice Energy)

Lattice Energy measures the strength of the ionic interaction.

Definition:

\[ \text{NaCl}(s) \rightarrow \text{Na}^+(g) + \text{Cl}^-(g) \]
\[ \Delta H = +786 \text{ kJ/mol} = \text{Lattice Energy} \]

- LE cannot be determined experimentally
- Use Hess’ Law (and known atomic properties) to get the value of the lattice energy.

Born-Haber Cycle
- By Hess’s Law, the enthalpy of formation is equal to the sum of the energies of several individual steps.
- Reasons for the stability of ionic compounds is made clear.

Lattice Energy

<table>
<thead>
<tr>
<th>Compound</th>
<th>E (kJ)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl</td>
<td>786</td>
</tr>
<tr>
<td>MgO</td>
<td>3934</td>
</tr>
</tbody>
</table>

Why the big difference???

\[ E \propto \frac{Q_1 Q_2}{d} \]

Different salts have different Q’s and d’s

<table>
<thead>
<tr>
<th>Compound</th>
<th>NaCl</th>
<th>MgO</th>
</tr>
</thead>
<tbody>
<tr>
<td>Radii</td>
<td>Na⁺ 0.97Å</td>
<td>Mg⁺² 0.66Å</td>
</tr>
<tr>
<td></td>
<td>Cl⁻ 1.81Å</td>
<td>O⁻² 1.40Å</td>
</tr>
<tr>
<td></td>
<td>2.78Å</td>
<td>2.06Å</td>
</tr>
</tbody>
</table>

**SO:** lattice energy for MgO is more than 4x greater
Salts and Ionic Lattices

ION SIZES:
Ion sizes are important in ionic bonds

<table>
<thead>
<tr>
<th>Ion</th>
<th>Melting Point (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaF</td>
<td>993</td>
</tr>
<tr>
<td>NaCl</td>
<td>801</td>
</tr>
<tr>
<td>NaBr</td>
<td>747</td>
</tr>
<tr>
<td>NaI</td>
<td>661</td>
</tr>
<tr>
<td>MgO</td>
<td>2800</td>
</tr>
</tbody>
</table>

Melting point increases as lattice energy
______________.

Both ______________ with increasing charge and size.

Which one of the following expresses the correct relationship between the melting points of the 3 ionic solids NaCl, MgS, and BaS?

1. NaCl > MgS > BaS
2. NaCl > BaS > MgS
3. MgS > BaS > NaCl
4. MgS > NaCl > BaS
5. BaS > MgS > NaCl