Chapter 22: Bonding in Ionic Compounds

Did you read chapter 22 before coming to class?
A. Yes
B. No

Review: Which color light emitting diode (LED) has the largest band gap?

a) Red  
b) Yellow  
c) Green  
d) Blue

Compare and Contrast: Ionic Compounds vs Metals

- Network Solids
- High melting Ts
- Brittle solids
- Don't conduct heat and electricity in solid
- Often colorless and usually transparent in big chunks (White when powdered)

Metals vs Non-Metals

- **Metals**
  - Large atoms
  - Few valence electrons
  - Low ionization energies
- **Non-metals**
  - Small atoms
  - Many valence electrons
  - High ionization energies

Why do metals and non-metals react? Principles of reactivity:
materials react to lower energy and increase entropy of universe.

How can energy be lowered?

- Metals lose valence electrons
- Non-metals gain valence electrons

Process is downhill energetically

When electrons are moved from one atom to another, ions are produced

- Positively charged
  - Sodium ions
  - Chloride ions
- Negatively charged
  - Chloride ions
  - Sodium ions

Electrons "belong" to individual ions: they are not shared among ions as was the case in metals.
Electron location and mobility is much lower in an ionic substance than in a metal.

Energy levels in an ionic crystal have relatively large spacing (rather than the nearly continuous spacing in metals).

Examples of Ionic Compounds

- **NaCl**: Ions: same charges and similar sizes
- **Al₂O₃**: Ions: different charges and sizes
- **Na₂O**: Ions: similar sizes, but different charges

Describe the structure of each compound: Do ions of one type cluster together? What type of ion immediately surrounds a given ion? How do the answers to these two questions relate to the electric force law? What prediction could you make about the arrangement of ions in any ionic compound?

Formation of a salt crystal

- **2Na + Cl₂ = 2NaCl**

What about entropy change?

- **2Na + Cl₂ = 2NaCl + lots of heat and light**

Heat and light – cause an increase in entropy of the surroundings

How does the model explain properties of salts (ionic compounds)?

- High melting and boiling temperatures?
  - Strong attractions between + and – ions
  - Attractive forces act over fairly large atomic distances
- Britleness?
  - Strong repulsions when ions with like charge come together; material shatters to relieve the stress.
Conductivity

- Don’t conduct as a solid. Why?
- Do conduct when molten or dissolved. Why?

Salts are generally transparent to light

- Why are they transparent?
  - Electron orbitals are localized around individual ions with FEW energy levels

Why are some ionic materials colored?

- Because they contain “transition” metals with more energy levels for electrons
  - Sapphire is a crystalline form of Al₂O₃
  - Chromium substitutions in the lattice allow blue and green light to be absorbed, resulting in a Ruby.
  - Titanium and Iron substitutions allow green and red light absorption, and give the blue color to what we normally think of as Sapphire.

Making a laser

- A ruby laser is possible because of the energy level structure

We can use the periodic table to make predictions of what ions usually form.

- Metals lose their valence electrons.
- Non-metals gain enough valence electrons to become “noble”.

The octet rule

- Atoms will most likely form an ion that has the ns²np⁶ configuration of the closest noble gas atom.
  - Metals take on this configuration by losing electrons
  - Non-metals take on this configuration by gaining electrons
Families

- Chlorine and Fluorine will form the same types of compounds since their valence electrons are the same number and same orbital type.

![](image)

Beryllium (Be) will most likely form an ion with what charge?

- a) -1
- b) -2
- c) +1
- d) +2

What would the chemical formula for magnesium fluoride (a salt of Mg and F) be?

- a) MgF
- b) Mg₂F
- c) MgF₂
- d) MgF₃

Ionic compounds are neutral (no net charge). What are the ionic charges in the following compounds?

- NaCl
- KBr
- MgF₂
- Al₂O₃

- Na⁺¹ and Cl⁻¹
- K⁺¹ and Br⁻¹
- Mg²⁺ and F⁻¹
- Al³⁺ and O²⁻

Naming convention for salts

- The metal comes first with its name unchanged
- The nonmetal comes second, with the suffix “ide” appended

Predicting Formulas for Salts

- Find the number of electrons lost by the metals and gained by the non-metals.
- If they are equal, the atoms combine one to one.
- If they are NOT equal, use the number lost/gained for the other atom’s subscript.

Examples

- Mg and O
- P and Ca
- N and Li

Mg²⁺ & O²⁻ → MgO
P³⁻ & Ca²⁺ → Ca₃P₂
N³⁻ & Li⁺ → LiN
When Potassium (K) and Chlorine (Cl) combine the resulting formula is

A. KCl
B. K₂Cl
C. KCl₂
D. K₂Cl₃
E. K₃Cl₂

When Aluminum (Al) and Chlorine (Cl) combine the resulting formula is

A. AlCl
B. Al₂Cl
C. AlCl₂
D. AlCl₃
E. Al₃Cl₂

When Magnesium (Mg) and Sulfur (S) combine the resulting formula is

A. MgS
B. Mg₂S
C. MgS₂
D. Mg₂S₃
E. Mg₃S₂

When Calcium (Ca) and Phosphor (P) combine the resulting formula is

A. CaP
B. Ca₂P
C. CaP₂
D. Ca₂P₃
E. Ca₃P₂

How about carbon?

- Would carbon like to gain or lose electrons?
- It turns out that it likes to share electrons in covalent bonds, which we’ll talk about on Monday.