Announcements

- Lon-cap HW 6 – Type 1 due Monday (11/26) and Type 2 due Wednesday (11/28) by 7pm
- “Lab 5: Modern Atomic Theory” write-up due tomorrow in discussion (and text homework)
Announcements

- All previous lectures are posted on course website (“Lectures” link in left menu)
  - Final exam is **cumulative** so you should start reviewing material soon!
- Last textbook homework assignment will be given this Friday (due November 30th)
Atomic Trends

- Note the relation between atomic size and ionization energy!
- Size is related to number of protons and energy levels
- As size *increases*, it becomes easier to remove electron, so ionization energy will *decrease*
When an element becomes an ion, how is size and ionization energy affected?

- When electron is pulled off, there is still the same number of protons; larger positive charge pulls remaining electrons in tighter.
Ionization and atomic size

- When an element becomes an ion, how is size and ionization energy affected?
  - When electron is added, there is still the same number of protons; larger negative charge means electrons are bound less tightly
Trends for isoelectronic species

- Isoelectronic: having the same number of electrons (same electron configuration)

- If two species have the same electron configuration, then how to you determine which is larger or which one has the higher ionization energy?
Trends in Isoelectronic Species

- Which is larger, Na\(^+\) or Ne?
  - Isoelectronic (both contain 10 electrons)
- Have to look at number of protons:
  - Na\(^+\): 11 protons; Ne: 10 protons
- Since Na\(^+\) has 11 protons, it will pull in its electrons tighter

- Therefore Ne will be larger
Trends in Isoelectronic Species

- Which has a higher ionization energy, Na\(^+\) or Ne?
  - Isoelectronic (both contain 10 electrons)
- Already established that Ne is larger
- Since Na\(^+\) is smaller, electrons are pulled in tighter -> harder to remove an electron

Therefore, Na\(^+\) has the higher ionization energy
Which is larger, Cl\(^-\) or Ar?
- both contain 18 electrons

In this case, Cl\(^-\) has an extra electron added so electrons are not bound as tightly; so Cl\(^-\) is larger
Trends in Isoelectronic Species

- Which has a higher ionization energy, Cl\(^-\) or Ar?
  - Isoelectronic (both contain 10 electrons)
- Already established that Cl\(^-\) is larger
- Since Ar is smaller, electrons are pulled in tighter -> harder to remove an electron

- Therefore, Ar has the higher ionization energy
Clicker #1

- Rank the following isoelectronic species in order from **highest** to **lowest** ionization energy.

- \( \text{Ba}^{2+}, \text{Cs}^+, \text{I}^-, \text{Te}^{2-}, \text{Xe} \)

A) \( \text{Xe} > \text{I}^- > \text{Te}^{2-} > \text{Ba}^{2+} > \text{Cs}^+ \)
B) \( \text{Ba}^{2+} > \text{Xe} > \text{Cs}^+ > \text{Te}^{2-} > \text{I}^- \)
C) \( \text{Cs}^+ > \text{Ba}^{2+} > \text{Te}^{2-} > \text{I}^- > \text{Xe} \)
D) \( \text{Te}^{2-} > \text{I}^- > \text{Xe} > \text{Cs}^+ > \text{Ba}^{2+} \)
E) \( \text{Ba}^{2+} > \text{Cs}^+ > \text{Xe} > \text{I}^- > \text{Te}^{2-} \)
Chemical Bonding

- Bonding: “forces” holding molecules together
- What is meant by a “chemical bond”?
- How and why do atoms bond to form molecules?

- Two types of bonding:
  - Intramolecular (Ch. 12) – within molecule
  - Intermolecular (Ch. 14) – between molecules
Chemical Bonding: Ionic

- Contain ionic bonding between a *metal* and *non-metal*

- Electrons are transferred, and ions are electrostatically held together
Notice that valence shells are filled and that there are no unpaired electrons in product.
Example: NaCl

<table>
<thead>
<tr>
<th>Electron configuration</th>
<th>shorthand configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na: 1s(^2)2s(^2)2p(^6)3s(^1)</td>
<td>[Ne]3s(^1)</td>
</tr>
<tr>
<td>To become an ion:</td>
<td></td>
</tr>
<tr>
<td>Na(^+): 1s2s2p(^6)</td>
<td>[Ne]</td>
</tr>
<tr>
<td>Cl: 1s(^2)2s(^2)2p(^6)3s(^2)3p(^5)</td>
<td>[Ne] 3s(^2)3p(^5)</td>
</tr>
<tr>
<td>To become an ion:</td>
<td></td>
</tr>
<tr>
<td>Cl(^-): 1s(^2)2s(^2)2p(^6)3s(^2)3p(^6)</td>
<td>[Ar]</td>
</tr>
</tbody>
</table>

When predicting compounds, ask yourself – is a noble gas configuration formed?
Transferring electrons requires energy (endothermic)
Example: ionization of magnesium

- $\text{Mg} \rightarrow \text{Mg}^+ \rightarrow \text{Mg}^{2+}$
  - First ionization energy: 737 kJ/mol
  - Second ionization energy: 1451 kJ/mol

Takes more energy to remove second electron because atom is now more positively charged (so it holds electrons even tighter)
If it requires so much energy, then how do ionic compounds form?

Example: magnesium chloride

- $\text{Mg}^{2+} + 2\text{Cl}^- \rightarrow \text{MgCl}_2$
  - Heat released in forming $\text{MgCl}_2$: 592 kJ/mol

Energy is released when $\text{MgCl}_2$ forms because the electrostatic interaction between $\text{Mg}^{2+}$ and $\text{Cl}^-$ ions causes them to become more stable (lower energy)
While noble gas configurations tend to be stable, magnesium does not “want” to give up a second electron to obtain a noble gas configuration (not the cause of bonding, just a way to remember)

Once you form Mg\(^{2+}\), however, it can electrostatically bond with an anion, which is more stable (lower in energy)

To form a compound as a noble gas configuration will release energy (exothermic) that it took to remove the electrons (endothermic)
What about bonding in H₂? Is it ionic?

Not likely; in order to do that, one H atom would need to lose an electron and one would need to gain one, then electrostatically bond.
Bonding: Covalent

- So why does H₂ form?
- Because we get covalent bonding (sharing of valence electrons)

- Valence electrons spend majority of time in between the two atoms so both hydrogen atoms have a noble gas configuration (not the cause of bonding, just a way to remember)

- This sharing of electrons lowers the energy between the two nuclei
Electron density of $H_2$
Clicker #2

How many electrons total have to be shared in a covalent bond in order for hydrogen to have a noble gas electron configuration?

A) 0
B) 1
C) 2
D) 6
E) 8
Why isn’t He diatomic?

Because it already has a noble gas configuration, there really isn’t a reason for helium to bond to itself.

Thus, covalent bonding has the same idea as ionic bonding, except one atom is not “strong” enough to remove an electron from the other atom.

Instead of transfer of electrons, there is sharing of electrons in covalent molecules.
Bonding: Covalent

- What about when different atoms bond that are not ionic? (non-metal to non-metal)

- In order to predict the type of bonding, we first need to consider how “strong” an atom is (how well can it pull electrons towards itself)
Bonding: Covalent

- Example: carbon monoxide (CO)

- Do these atoms have the exact same attraction for electrons?
  - No, oxygen has more protons, so it has more attraction for electrons than carbon
  - CO molecule will have unequal sharing of electrons, with more electron density around oxygen atom
Electron density of CO

Partial negative charge

Partial positive charge
Electronegativity

- How do we predict whether an atom will have a stronger attraction for electrons?

**Electronegativity**: attraction for an electron by an atom in a molecule

- Based on atomic trends we have already discussed, we can predict the trend for electronegativity of atoms as well.
Atomic Trends

- **Electronegativity**
- Going down a column: atomic radius becomes larger, so the “pull” from the nucleus is not as strong, therefore electronegativity becomes weaker
Atomic Trends

- **Electronegativity**

- Going across a row: electrons are closer to the nucleus, so protons are more apt to attract electrons, resulting in greater “pull”

Electronegativity decreases
Atomic Trends

Remember:
Trends are determined by number of protons and quantum energy level

Atomic size increases
Ionization energy decreases
Electronegativity decreases

Atomic size increases
Ionization energy decreases
Electronegativity decreases

 Atomic size increases
 Ionization energy decreases
 Electronegativity decreases
Which has greater attraction for electrons, Li or F?

A) Li
B) F

Fluorine has greater number of protons, meaning it holds its electrons tighter and the “pull” of the nucleus will be greater.
Types of covalent bonds

- Example: $\text{H}_2$
- Both hydrogen atoms have similar electronegativity, so they “pull” electrons equally
- Refer to this as a **nonpolar covalent bond**
Types of covalent bonds

- Example: HF
- H: very low electronegativity (does not pull electrons well)
- F: very high electronegativity (pulls electrons towards itself very well)
- Refer to this as a polar covalent bond
Think of bonding as a continuum
- (on a scale: not always just one type of bond)

- In general: the closer two atoms are on the periodic table, the more covalent their bonding is

<table>
<thead>
<tr>
<th>Large difference in electronegativity</th>
<th>Type of bonding</th>
<th>No difference in electronegativity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ionic (NaCl) Very polar!</td>
<td>Polar Covalent (HF)</td>
<td>Nonpolar Covalent (H₂)</td>
</tr>
</tbody>
</table>
Clicker #4

- Arrange the following from the most polar to least polar bond.
  I. N-F  II. C-F  III. O-F

A) I > II > III
B) III > II > I
C) II > I > III
D) III > I > II
E) II > III > I
- C-F has the biggest difference in electronegativity so it is the most polar (most unequal sharing)

- O-F has the smallest difference in electronegativity so it is the least polar

- C-F > N-F > O-F
Therefore, from the periodic table we can get:

- Electron configurations
- Charges (most stable for compounds)
- Trends -> size, ionization energy, electronegativity
- Polarity of bonds (ranking)
- Determine formulas of compounds (ionic and covalent)
- Valence electrons (the important ones!)
  - Bonds form so that we have a complete valence shell
Demo: Aluminum Iodide

- \(2\text{Al}(s) + 3\text{I}_2(s) \rightarrow 2\text{AlI}_3(s)\)

- How do we know the product?
- Ionization of Aluminum:
  - \(\text{Al}: 1s^22s^22p^63s^23p^1 \quad \text{Al}^{3+}: 1s^22s^22p^6 = [\text{Ne}]\)
- Ionization of Iodine:
  - \(\text{I}: [\text{Kr}]5s^24d^{10}5p^5 \quad \text{I}^-: [\text{Kr}]5s^24d^{10}5p^6 = [\text{Xe}]\)

- Noble gas configuration is stable (NOT the cause for bonding in ionic compounds, but a good way to remember)