Chapter 8

Periodic Relationships

Valence Electrons

- **Valence electrons**: outermost s and p electrons
  - N: \(1s^22s^22p^3\) \([\text{He}]2s^22p^3\)
  - Mg: \(1s^22s^22p^63s^2\) \([\text{Ne}]3s^2\)
  - Ge: \(1s^22s^22p^63s^23p^64s^23d^{10}4p^2\) \([\text{Ar}]4s^23d^{10}4p^2\)

- **Core electrons**: inner electrons
  - Full shell (ns\(^2\)np\(^6\)) noble gas electron configuration
Valence Electrons

• Valence electrons easy to figure out from periodic table
  • N: \(1s^22s^22p^3\)
  • Mg: \(1s^22s^22p^63s^2\)
  • Ge: \(1s^22s^22p^63s^23p^64s^23d^{10}4p^2\)

Valence Electrons

• What group of the Periodic Table is each of these elements are found?
  – [Ar]\(4s^23d^{10}4p^4\)
  – [Kr]\(5s^24d^5\)
Representative Elements

• Valence electrons are important for the representative elements
  – Main group elements
  – # Valence electrons = group number

Periodic Relationships

• Properties of representative elements depend on three main factors:

  – Noble gas e\textsuperscript{-} configurations (full outer shell) are very stable

  – Valence electrons do not feel the complete charge of the nucleus because of shielding

  – The n quantum number of the valence electrons increases as you go down a group in the periodic table
Noble Gas Electron Configurations

• An 8-electron outer shell is particularly stable
  – Valence-electron configuration of a noble gas
  – Isoelectronic (same electron-config.) with a noble gas

• A full shell
  – He: $1s^2$ 
  – Ne: $1s^22s^22p^6$ 
  – Ar: $1s^22s^22p^6\ 3s^23p^6$ 
  – Kr: $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$

Ions of Representative Elements

• Ions of representative elements frequently have a full outer shell ($ns^2np^6$)
  – Na: $1s^22s^22p^63s$ $Na^+$ $1s^22s^22p^6$ 
  – O: $1s^22s^22p^4$ $O^-$ $1s^22s^22p^6$
Oxidation Nos. of Represent. Elements

- What about elements in the center?
  - Not always ionic: oxidation numbers important
  - Metals tend to have positive oxidation numbers
  - Nonmetals tend to have negative oxidation numbers
  - Metalloids frequently have both

- Tend to have noble gas $e^-$ configuration: $s^2p^6$
- $s^2$ configuration also common
- Bismuth: +3 and +5
Oxidation Nos. of Represent. Elements

• Predict the most common ions or oxidation numbers of the following elements
  – Ra, Sn, Se, Te

Transition Metal Oxidation Nos.

• Noble gas electron configurations not as important

• Ions of transition elements have ns electrons removed before the (n-1)d
  – Fe \(1s^22s^22p^63s^23p^64s^23d^6\)
  – Fe\(^{2+}\) \(1s^22s^22p^63s^23p^6d^6\)
  – Fe\(^{3+}\) \(1s^22s^22p^63s^23p^6d^5\)

• This is why many transition elements have a +2 ion or oxidation number
  – Two ns electrons removed
Periodic Relationships

- Three main factors influence periodic relationships
  - Noble gas $e^-$ configurations (full outer shell) are very stable
  - Valence electrons do not feel the complete charge of the nucleus because of shielding
  - The $n$ quantum number of the valence electrons increases as you go down a group in the periodic table

Effective Nuclear Charge

- The second factor is due to shielding of valence $e^-$
Effective Nuclear Charges

• Core e\(^-\) shield valence e\(^-\) effectively
  – Core e\(^-\) almost always closer to nucleus than valence

• Valence e\(^-\) do not shield each other effectively
  – Valence e\(^-\) are nearer to nucleus only about half of the time

Effective Nuclear Charge

• Effective nuclear charge = positive charge actually felt by electrons
• Increases across periodic table
  – Adding +1 charge to nucleus each time
  – Adding valence electrons that do not shield effectively

<table>
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<th></th>
<th>Li</th>
<th>Be</th>
<th>B</th>
<th>C</th>
<th>N</th>
<th>O</th>
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<td>6</td>
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<td>8</td>
<td>9</td>
<td>10</td>
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<td>1.9</td>
<td>2.4</td>
<td>3.1</td>
<td>3.8</td>
<td>4.5</td>
<td>5.1</td>
<td>5.8</td>
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</table>
Atomic Radii

- Attraction of nucleus for valence electrons increases as move across table.
- Atoms become smaller as move across table

<table>
<thead>
<tr>
<th>Atomic Radii (pm)</th>
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<tbody>
<tr>
<td>1A</td>
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<tr>
<td>L</td>
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<td>S</td>
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<td>Cl</td>
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<td>Ar</td>
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</table>

Periodic Relationships

- Properties of representative elements depend on three main factors:
  - Noble gas e⁻ configurations (full outer shell) are very stable
  - Valence electrons do not feel the complete charge of the nucleus because of shielding
  - The n quantum number of the valence electrons increases as go down a group in the periodic table
Atomic Radii

• Higher n quantum number means larger atom
• As move down a group:
  – n quantum number increases
    • Li \([\text{He}]2s\)
    • Na \([\text{Ne}]3s\)
    • K \([\text{Ar}]4s\)
    • Etc
  – Atomic radii increase

*The effective nuclear charge also increases down a group. The increase in the n quantum number is more important.

Atomic Radii

• And so:
Atomic Radii

• In each pair below, which atom would have the largest radius?
  – P or As
  – Sn or Xe

Ionic Radii

• Negative ions are larger than corresponding atom
• Positive ions are smaller than corresponding atom
• For ions of same charge, trends in ionic radii often parallel trends in atomic radii
Ionization Energy

• Minimum energy required to remove electron from gaseous atom in its ground state

• More than one ionization energy
  – 1st IE \[ M(g) \rightarrow M^+(g) + e^- \]
  – 2nd IE \[ M^+(g) \rightarrow M^{2+}(g) + e^- \]

• Variations in IE generally follow trends in atomic size

Ionization Energy

• Small atom: valence electrons close to nucleus
  – Valence electrons held tightly
  – Difficult to remove
  – High IE

• Large atom: valence electrons far from nucleus
  – Valence electrons held loosely
  – Easy to remove
  – Low IE
Ionization Energy

• As move across a period
  – Atoms get smaller
  – In general, the IE increases

• As move down a group
  – Atoms get larger
  – In general, the IE decreases

In each pair, which element will have higher first IE
– Si or S
– Se or Te
Ionization Energy

• What is interesting about these IE?
  – Be = 899 kJ/mole; B = 801 kJ/mole
  – N = 1421 kJ/mole; O = 1314 kJ/mole

Ionization Energy: Irregularities

• Be = 899  B = 801
  – Be: 2s²
  – B: 2s²2p
  – 2p electron in B is shielded by 2s electrons.
  – Outer electron in B is more weakly held than expected
Ionization Energies: Irregularities

• $N = 1421$  $O = 1314$
  
  – $N$: 2p$^3$  
  
  – $O$: 2p$^4$
  
  – Oxygen: Must put fourth p-electron in a pair  
  – Easier to remove than expected

Higher Ionization Energies

• Can remove more than one electron  
  – Valence electrons relatively easy to remove  
  – Core electrons much more difficult to remove (in a full shell)

• B: 1s$^2$2s$^2$2p  
  – 1$^{\text{st}}$ IE: 801 kJ/mole  
    • 1s$^2$2s$^2$2p $\rightarrow$ 1s$^2$2s  
  – 2$^{\text{nd}}$ IE: 2,430 kJ/mole  
    • 1s$^2$2s$^2$ $\rightarrow$ 1s$^2$2s  
  – 3$^{\text{rd}}$ IE: 3,660 kJ/mole  
    • 1s$^2$2s $\rightarrow$ 1s$^2$  
  – 4$^{\text{th}}$ IE: 25,000 kJ.mole  
    • 1s$^2$ $\rightarrow$ 1s (removing a core electron and breaking up a full shell)
Higher Ionization Energies

• In what group is the element with these IE?
  – 1st IE: 899
  – 2nd IE: 1,757
  – 3rd IE: 14,850

Electron Affinity

• You are not responsible for EA on the exam
Metals

• Electron sea model
  – Relatively simple model
    • Explains electrical conductivity
  – More sophisticated models exist
• All atoms in metal give up electrons
  – Form a “sea” of negative charge
• Positive ions are stationary
• Electrons can move

Metals

• In general, elements that give up electrons easily form metals*
  – Larger atoms
  – Left side of a period
  – Bottom of a group

* IE is not the only property involved, but this is a good general rule
Periodic Trends

- Handling somewhat differently than book
- Look for general trends rather than memorizing properties of each group

Group Properties

- Elements in same group generally have similar chemical properties
- Examples:
  - $\text{H}_3\text{PO}_4$ is common acid. Predict formula of an acid of arsenic.
  - What can we predict about the charge of the radium (Ra) ion?
Highly Reactive Elements

Both sides of the periodic table.

Unreactive Elements

Noble gases and center of periodic table
Acidic and Basic Oxides

• Metallic oxides
  – Generally act as bases
    • Remember strong bases: Group IA, Ba hydroxides
  – React with acids

• Nonmetallic oxides
  – Generally act as acids
    • Remember strong acids: HNO₃, H₂SO₄, etc
  – React with bases
Amphoteric Oxides

- Display both acidic and basic properties
  - React with both acids and bases
- Not always metalloids
  - See pp 353-354 in text

Transition Metal Oxides

- Transition metals may have more than one type of oxide
  - If element has multiple oxidation numbers
    - High ON “want” e- more: acidic
    - Low ON “don’t want” e- as much: basic
- Example:
  - CrO: basic
  - Cr₂O₃: amphoteric
  - CrO₃: acidic
Physical State

- Gases tend to be at top of group and on right side
- $O_2$, $N_2$, $F_2$, $Cl_2$, noble gases

Densities

- Most dense elements tend to be in bottom center of table
- Elements on either end are softer and less dense
  - Na and K float on water and can be cut with a knife
  - Several of halogens are gases or liquids
Predictions

• What can we predict about the properties of Rg (Z = 111)?
  – Metal, metalloid, or nonmetal
  – Reactive or unreactive
  – Most likely ions or oxidation state
  – Element(s) it is most likely to react with
  – Acidic or basic oxide

Predictions

• What can we predict about the properties of Iodine?
  – Metal, metalloid, or nonmetal
  – Reactive or unreactive
  – Most likely ions or oxidation states
  – Element(s) it is most likely to react with
  – Acidic or basic oxide
Predictions

• Te is a metalloid and can have both positive and negative oxidation numbers.
  – Predict what the highest positive ON for Te.
  – This ON is found in a binary compound with only one element. What element would that be?

Predictions

• Manganese has several oxidation states
  – Most important: +7, +4, +2
• Predict redox properties of
  – MnO
  – MnO₂
  – Mn₂O₇
• Manganese oxides are insoluble in water
  – Can dissolve with acids and/or bases
  – Would you use HCl, NaOH, or either to dissolve each of the three oxides above?
    • Mn₂O₇ actually decomposes