Chapter 6
“The Periodic Table”

Section 6.1
Organizing the Elements

**OBJECTIVES:**
- Explain how elements are organized in a periodic table.

- Compare early and modern periodic tables.

- Identify three broad classes of elements.

A few elements, such as gold and copper, have been known for thousands of years - since ancient times.
Yet, only about 13 had been identified by the year 1700.
As more were discovered, chemists realized they needed a way to organize the elements.

Chemists used the properties of elements to sort them into groups.
In 1829 J. W. Dobereiner arranged elements into triads – groups of three elements with similar properties.
- One element in each triad had properties intermediate of the other two elements.
Mendeleev’s Periodic Table
- By the mid-1800s, about 70 elements were known to exist
- Dmitri Mendeleev – a Russian chemist and teacher
- Arranged elements in order of increasing atomic mass
- Thus, the first “Periodic Table”

Mendeleev
- He left blanks for yet undiscovered elements
  - When they were discovered, he had made good predictions
- But, there were problems:
  - Such as Co and Ni; Ar and K; Te and I

A better arrangement
- In 1913, Henry Moseley – British physicist, arranged elements according to increasing atomic number
- The arrangement used today
- The symbol, atomic number & mass are basic items included—textbook page 162 and 163

The Periodic Law says:
- When elements are arranged in order of increasing atomic number, there is a periodic repetition of their physical and chemical properties.
- Horizontal rows = periods
  - There are 7 periods
- Vertical column = group (or family)
  - Similar physical & chemical prop.
  - Identified by number & letter (IA, IIA)

Another possibility: Spiral Periodic Table
Three classes of elements are:
1) ** Metals**: electrical conductors, have luster, ductile, malleable
2) ** Nonmetals**: generally brittle and non-lustrous, poor conductors of heat and electricity
3) ** Metalloids**: border the line-2 sides
   - Properties are *intermediate* between metals and nonmetals

Some nonmetals are gases (O, N, Cl); some are brittle solids (S); one is a fuming dark red liquid (Br)
Notice the heavy, stair-step line?

**OBJECTIVES:**
- Describe the information in a periodic table.
- Classify elements based on electron configuration.
- Distinguish representative elements and transition metals.
Groups of elements - family names

- **Group 1A** – alkali metals
  - Forms a “base” (or alkali) when reacting with water (not just dissolved!)

- **Group 2A** – alkaline earth metals
  - Also form bases with water; do not dissolve well, hence “earth metals”

- **Group 7A** – halogens
  - Means “salt-forming”

Electron Configurations in Groups

- Elements can be sorted into 4 different groupings based on their electron configurations:
  1) Noble gases
  2) Representative elements
  3) Transition metals
  4) Inner transition metals

1) **Noble gases** are the elements in Group 8A (also called Group 18 or 0)
   - Previously called “inert gases” because they rarely take part in a reaction; very stable = don’t react
   - Noble gases have an electron configuration that has the outer s and p sublevels completely full

2) **Representative Elements** are in Groups 1A through 7A
   - Display wide range of properties, thus a good “representative”
   - Some are metals, or nonmetals, or metalloids; some are solid, others are gases or liquids
   - Their outer s and p electron configurations are NOT filled

3) **Transition metals** are in the “B” columns of the periodic table
   - Electron configuration has the outer s sublevel full, and is now filling the “d” sublevel
   - A “transition” between the metal area and the nonmetal area
   - Examples are gold, copper, silver
Electron Configurations in Groups

4) **Inner Transition Metals** are located below the main body of the table, in two horizontal rows:
   - Electron configuration has the outer s sublevel full, and is now filling the “f” sublevel.
   - Formerly called “rare-earth” elements, but this is not true because some are very abundant.

**Group 1A** are the alkali metals (but NOT H)
- Elements in the 1A-7A groups are called the representative elements.
- Formerly called “rare-earth” elements, but this is not true because some are very abundant.

**Group 2A** are the alkaline earth metals

**Group 8A** are the noble gases
- Group 7A is called the halogens.

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Do you notice any similarity in these configurations of the alkali metals?
Do you notice any similarity in the configurations of the noble gases?

1s² He
1s²2s²2p⁶ Ne
1s²2s²2p⁶3s²³p⁶ Ar
1s²2s²2p⁶3s²³p⁶4s²3d¹⁰⁴p⁶ Kr
1s²2s²2p⁶3s²³p⁶4s²3d¹⁰⁴p⁶5s²4d¹⁰⁵p⁶ Xe
1s²2s²2p⁶3s²³p⁶4s²3d¹⁰⁴p⁶5s²4d¹⁰⁵p⁶⁰ 5p⁶6s²⁴f²⁴d²⁵p²⁶ Rn

Elements in the s - blocks

- Alkali metals all end in s¹
- Alkaline earth metals all end in s²

- really should include He, but it fits better in a different spot, since He has the properties of the noble gases, and has a full outer level of electrons.

Transition Metals - d block

Note the change in configuration.

The P-block

- Each row (or period) is the energy level for s and p orbitals.

F - block

- Called the “inner transition elements”
The “d” orbitals fill up in levels 1 less than the period number, so the first d is 3d even though it’s in row 4.

The “f” orbitals start filling at 4f, and are 2 less than the period number.

Section 6.3 Periodic Trends

**OBJECTIVES:**
- **Describe trends** among the elements for **atomic size**.

**OBJECTIVES:**
- **Explain** how **ions** form.

**OBJECTIVES:**
- **Describe periodic trends** for first ionization energy, ionic size, and electronegativity.

**First problem: Where do you start measuring from?**
**The electron cloud doesn’t have a definite edge.**
**They get around this by measuring more than 1 atom at a time.**
Atomic Size

Measure the Atomic Radius - this is half the distance between the two nuclei of a diatomic molecule.

ALL Periodic Table Trends

- **Influenced by three factors:**
  1. **Energy Level**
     - Higher energy levels are further away from the nucleus.
  2. **Charge on nucleus (# protons)**
     - More charge pulls electrons in closer. (+ and – attract each other)
  3. **Shielding effect** (blocking effect?)

What do they influence?

- **Energy levels** and **Shielding** have an effect on the GROUP (∟ )
- **Nuclear charge** has an effect on a PERIOD (∟ )

#1. **Atomic Size** - Group trends

- As we increase the atomic number (or go down a group). . .
- each atom has another energy level,
- so the atoms get bigger.

#1. **Atomic Size** - Period Trends

- Going from left to right across a period, the size gets smaller.
- Electrons are in the same energy level.
- But, there is more nuclear charge.
- Outermost electrons are pulled closer.

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Period 2</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>Na</td>
</tr>
<tr>
<td>Na</td>
<td>K</td>
</tr>
<tr>
<td>K</td>
<td>Rb</td>
</tr>
</tbody>
</table>

Atomic Radius (pm)

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Atomic Radius (pm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>3</td>
</tr>
<tr>
<td>Li</td>
<td>10</td>
</tr>
<tr>
<td>Na</td>
<td>3</td>
</tr>
<tr>
<td>K</td>
<td>10</td>
</tr>
<tr>
<td>Rb</td>
<td>3</td>
</tr>
</tbody>
</table>

Period 2
Some compounds are composed of particles called “ions”

- An ion is an atom (or group of atoms) that has a positive or negative charge
- Atoms are neutral because the number of protons equals electrons
- Positive and negative ions are formed when electrons are transferred (lost or gained) between atoms

Metals tend to LOSE electrons, from their outer energy level

- Sodium loses one: there are now more protons (11) than electrons (10), and thus a positively charged particle is formed = “cation”
- The charge is written as a number followed by a plus sign: Na\(^{1+}\)
- Now named a “sodium ion”

Nonmetals tend to GAIN one or more electrons

- Chlorine will gain one electron
- Protons (17) no longer equals the electrons (18), so a charge of -1
- Cl\(^{-}\) is re-named a “chloride ion”
- Negative ions are called “anions”

#2. Trends in Ionization Energy

- Ionization energy is the amount of energy required to completely remove an electron (from a gaseous atom).
- Removing one electron makes a 1+ ion.
- The energy required to remove only the first electron is called the first ionization energy.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>First</th>
<th>Second</th>
<th>Third</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1312</td>
<td></td>
<td></td>
</tr>
<tr>
<td>He</td>
<td>2731</td>
<td>5247</td>
<td>11810</td>
</tr>
<tr>
<td>Li</td>
<td>520</td>
<td>7297</td>
<td>11810</td>
</tr>
<tr>
<td>Be</td>
<td>900</td>
<td>1757</td>
<td>14840</td>
</tr>
<tr>
<td>B</td>
<td>800</td>
<td>2430</td>
<td>3569</td>
</tr>
<tr>
<td>C</td>
<td>1086</td>
<td>2352</td>
<td>4619</td>
</tr>
<tr>
<td>N</td>
<td>1402</td>
<td>2857</td>
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<tr>
<td>O</td>
<td>1314</td>
<td>3391</td>
<td>5301</td>
</tr>
<tr>
<td>F</td>
<td>1681</td>
<td>3375</td>
<td>6045</td>
</tr>
<tr>
<td>Ne</td>
<td>2080</td>
<td>3963</td>
<td>6276</td>
</tr>
</tbody>
</table>

Table 6.1, p. 173
What factors determine IE

- The greater the nuclear charge, the greater IE.
- Greater distance from nucleus decreases IE.
- Filled and half-filled orbitals have lower energy, so achieving them is easier, lower IE.
- Shielding effect.

Shielding

- The electron on the outermost energy level has to look through all the other energy levels to see the nucleus.
- Second electron has the same shielding, if it is in the same period.

Ionization Energy - Group trends

- As you go down a group, the first IE decreases because...
  - The electron is further away from the attraction of the nucleus, and
  - There is more shielding.

Ionization Energy - Period trends

- All the atoms in the same period have the same energy level.
- Same shielding.
- But, increasing nuclear charge
- So IE generally increases from left to right.
- Exceptions at full and 1/2 full orbitals.

He has a greater IE than H.
- Both elements have the same shielding since electrons are only in the first level
- But He has a greater nuclear charge.
- Li has lower IE than H
- more shielding
- further away
- These outweigh the greater nuclear charge

- Be has higher IE than Li
- same shielding
- greater nuclear charge

- B has lower IE than Be
- same shielding
- greater nuclear charge
- By removing an electron we make s orbital half-filled

- Oxygen breaks the pattern, because removing an electron leaves it with a 1/2 filled p orbital
First Ionization Energy

- Ne has a lower IE than He
- Both are full,
- Ne has more shielding
- Greater distance

Driving Forces

- **Full Energy Levels** require lots of energy to remove their electrons.
  - Noble Gases have full orbitals.
  - Atoms behave in ways to try and achieve a noble gas configuration.

2nd Ionization Energy

- For elements that reach a filled or half-filled orbital by removing 2 electrons, 2nd IE is lower than expected.
- True for s²
- Alkaline earth metals form 2+ ions.
3rd IE
- Using the same logic $s^2p^1$ atoms have an low 3rd IE.
- Atoms in the aluminum family form 3+ ions.
- 2nd IE and 3rd IE are always higher than 1st IE!!

Trends in Ionic Size: Cations
- Cations form by losing electrons.
- Cations are smaller than the atom they came from – not only do they lose electrons, they lose an entire energy level.
- Metals form cations.
- Cations of representative elements have the noble gas configuration before them.

Ionic size: Anions
- Anions form by gaining electrons.
- Anions are bigger than the atom they came from – have the same energy level, but a greater area the nuclear charge needs to cover.
- Nonmetals form anions.
- Anions of representative elements have the noble gas configuration after them.

Configuration of Ions
- Ions always have noble gas configurations ($= a$ full outer level)
- Na atom is: $1s^22s^22p^63s^1$
- Forms a 1+ sodium ion: $1s^22s^22p^6$
- Same configuration as neon.
- Metals form ions with the configuration of the noble gas before them - they lose electrons.

Configuration of Ions
- Non-metals form ions by gaining electrons to achieve noble gas configuration.
- They end up with the configuration of the noble gas after them.

Ion Group trends
- Each step down a group is adding an energy level.
- Ions therefore get bigger as you go down, because of the additional energy level.

Li$^{1+}$  Na$^{1+}$  K$^{1+}$  Rb$^{1+}$  Cs$^{1+}$
Ion Period Trends
- Across the period from left to right, the nuclear charge increases - so they get smaller.
- Notice the energy level changes between anions and cations.

Size of Isoelectronic ions
- Iso- means “the same”
- Isoelectronic ions have the same # of electrons
  - Al\(^{3+}\), Mg\(^{2+}\), Na\(^{+}\), Ne\(^{-}\), F\(^{-}\), O\(^{2-}\), and N\(^{3-}\)
    - all have 10 electrons
  - all have the same configuration: 1s\(^2\)2s\(^2\)2p\(^6\) (which is the noble gas: neon)

Size of Isoelectronic ions?
- Positive ions that have more protons would be smaller (more protons would pull the same # of electrons in closer)

#3. Trends in Electronegativity
- Electronegativity is the tendency for an atom to attract electrons to itself when it is chemically combined with another element.
- They share the electron, but how equally do they share it?
- An element with a big electronegativity means it pulls the electron towards itself strongly!

Electronegativity Group Trend
- The further down a group, the farther the electron is away from the nucleus, plus the more electrons an atom has.
- Thus, more willing to share.
- Low electronegativity.

Electronegativity Period Trend
- Metals are at the left of the table.
- They let their electrons go easily
- Thus, low electronegativity
- At the right end are the nonmetals.
- They want more electrons.
- Try to take them away from others
- High electronegativity.
The arrows indicate the trend: Ionization energy and Electronegativity increase in these directions.

Atomic size and Ionic size increase in these directions:

Summary Chart of the trends:
Figure 6.22, p.178

End of Chapter 6