Chapter 5
The Periodic Law
Pages 132-173

Task Checklist:

Read Section 1
Read Section 2
Read Section 3
Complete all notes
Review chapter PowerPoint

Complete all worksheets
Review chapter packet

"Sorry Ms. Davenport, but budget restraints have forced us to cut down on the periodic table this year."
History of the Periodic Table

At the end of the 1700’s:

- Only about 30 elements were discovered.
- Ex. Cu, Ag, Au, H, O, N & C.

Early 1800’s:

- J.W. Dobereiner organized elements into **triads**.
- Triads - groups of three elements with similar properties. The middle element of the triad was thought to be an approximate average of the properties of the first & third element.

### Dobereiner’s Triads

<table>
<thead>
<tr>
<th>Example</th>
<th>Triad</th>
<th>Atomic Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>chlorine</td>
<td>35.5</td>
</tr>
<tr>
<td></td>
<td>bromine</td>
<td>80</td>
</tr>
<tr>
<td></td>
<td>iodine</td>
<td>127</td>
</tr>
<tr>
<td>2</td>
<td>sulfur</td>
<td>32</td>
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<tr>
<td></td>
<td>selenium</td>
<td>79</td>
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<tr>
<td></td>
<td>tellurium</td>
<td>128</td>
</tr>
<tr>
<td>3</td>
<td>calcium</td>
<td>40</td>
</tr>
<tr>
<td></td>
<td>strontium</td>
<td>88</td>
</tr>
<tr>
<td></td>
<td>barium</td>
<td>137</td>
</tr>
<tr>
<td>4</td>
<td>lithium</td>
<td>7</td>
</tr>
<tr>
<td></td>
<td>sodium</td>
<td>23</td>
</tr>
<tr>
<td></td>
<td>potassium</td>
<td>39</td>
</tr>
</tbody>
</table>

Dobereiner’s Accomplishment:

- First of a long series of chemists who recognized a relation between atomic weights and chemical properties.

Shortcoming of Dobereiner’s Triads:

- There was not a standard method for determining atomic mass.

Many new elements were being discovered as a result of new laboratory techniques. The number of elements doubled & a system of classification was necessary.

  Ex. Spectroscopy
Mid 1800’s

In 1860, a meeting of the First International Congress of Chemists was held in Germany.

- Goal of Meeting: Determine a universal method for determining average atomic mass of elements.
- In 1860, Cannizzaro was credited with developing the universal method of determining standard atomic mass.

J.A.R. Newlands- developed the law of octaves in 1865.

- Law of octaves- stated that if elements were arranged by increasing atomic mass, the properties of the eighth element were similar to properties of the first, the ninth like those of the second, the tenth like those of the third.

Example of the law of octaves:

![Law of Octaves Diagram](image)

Newlands’ Accomplishments:

- Used new method of determining relative atomic masses to organize elements.
- Placed elements in an easy to read table.

Shortcoming of Newlands’ law of octaves:

- He did not always stick to an increase in atomic mass in his organization.
- He sometimes put two elements together because he could not determine which element belonged there.
In 1869, Dmitri Mendeleev published the first periodic table.

- Mendeleev arranged the elements horizontally by increasing atomic mass and placed elements in groups (vertically) based on similar properties.
- Mendeleev’s procedure left many holes throughout the periodic table.
- Mendeleev was able to predict the properties of elements that would eventually fill those holes even though they were not discovered yet.
- Ex. Ekasilicon - Germanium

### Mendeleev’s Periodic Table

<table>
<thead>
<tr>
<th>Period</th>
<th>I</th>
<th>II</th>
<th>III</th>
<th>IV</th>
<th>V</th>
<th>VI</th>
<th>VII</th>
<th>VIII</th>
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<tbody>
<tr>
<td>1</td>
<td>H</td>
<td>Be</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
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<td>Na</td>
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<td>Be</td>
<td>Mg</td>
<td>Al</td>
<td>Si</td>
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<td>Cl</td>
<td>Ar</td>
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<tr>
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<td>Ca</td>
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<td>Ti</td>
<td>V</td>
<td>Cr</td>
<td>Mn</td>
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<tr>
<td>5</td>
<td>Rb</td>
<td>Sr</td>
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<td>Se</td>
<td>Br</td>
</tr>
<tr>
<td>6</td>
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<td>Ba</td>
<td>La</td>
<td>Hf</td>
<td>Ta</td>
<td>W</td>
<td>Re</td>
<td>Os</td>
</tr>
<tr>
<td>7</td>
<td>Au</td>
<td>Hg</td>
<td>Tl</td>
<td>Pb</td>
<td>Bi</td>
<td>Po</td>
<td>At</td>
<td>Rn</td>
</tr>
<tr>
<td>8</td>
<td>Th</td>
<td>Tl</td>
<td>Pb</td>
<td>Bi</td>
<td>Po</td>
<td>At</td>
<td>Rn</td>
<td>U</td>
</tr>
</tbody>
</table>

In 1894, John Strutt & William Ramsay added the noble gases to the periodic table.

**Predicted properties**
- Atomic mass: 72 amu
- High Melting Point
- Density: 5.5 g/cm³
- Dark Gray Metal
Early 1900’s

In 1911, Henry Moseley developed the Modern Periodic Table. He arranged periodic table in rows by increasing atomic number and in columns by similar properties.

Moseley’s Accomplishments:

- Did work on metals with Rutherford.
- Discovered that the positive charge of the nucleus increases from one element to the next.
- Led to the modern definition of atomic number.
- Moseley’s periodic table explained some of the contradictions that Mendeleev experienced within his model.

Moseley’s Periodic Table

![Periodic Table](image)

**Periodic Law** - The physical and chemical properties of the elements are a function of their atomic numbers

After 1913:

- Lanthanides/Rare Earth Metals were added to the periodic table.
- Actinides/Radioactive metals were discovered shortly after that.
Periodic Law

Periodic Law- states that the chemical & physical properties of the elements are periodic function of their atomic numbers.

Group/Family- vertical columns of elements on the periodic table with similar properties.

Periods- horizontal rows of elements on the periodic table arranged by increasing atomic number.

Metals, Metalloids, & Nonmetals

Metals
- Make up most of the periodic table.
- Ductile- can be drawn into wire
- Malleable- can be hammered into thin sheets.
- Lustrous- shiny
- Good Conductors of heat & electricity
- Located to the left of the step ladder on the periodic table.

Metallic Character

1. What is the MOST metallic element?
   Cs or Fr

2. What is the MOST nonmetallic element?
   F
Nonmetals

- Brittle-break when hammered.
- Poor conductors of heat & electricity.
- Lack luster
- Located to the right of the step ladder on the periodic table.

Metalloids

- Semimetals.
- Properties of both metals & nonmetals.
- Located along the step ladder on the periodic table.
- Examples: B, Ge, Sb, Te, At, Si, As

Alkali Metals-Group 1

- Most reactive group of metals.
- Usually found in combined form as a salt due to their high reactivity.
- Combine vigorously with nonmetals especially groups 16 & 17.
- React readily with water.
- Soft and silvery appearance.
Alkaline Earth Metals-Group 2
- Found in the earth’s crust but not in the elemental form due to their high reactivity. They are usually found in rock structures.
- 2nd most reactive group of metals.
- More dense than group 1.
- Shiny silvery-white color.

Transition Metals-Groups 3-12
- Compose the d-block.
- Typical metallic properties.
- Good conductors.
- Lustrous.
- Produce colored ions.

Main Block Elements-Groups 13-18
- Compose the p-block.
- Properties of elements vary greatly.
- Contains all of the nonmetals & metalloids as well as some metals.

Halogens-Group 17
- Most reactive groups of non-metals.
- React vigorously with metals (especially groups 1 & 2) to produce salts.
- Fluorine is a poisonous pale yellow gas, chlorine is a poisonous pale green gas, bromine is a toxic and caustic brown volatile liquid, and iodine is a shiny black solid which easily sublimes to form a violet vapor on heating.
- Found in nature in the combined state.
Noble Gases-Group 18

- Least reactive of all elements.
- Often called inert gases.
- All are gases.
- The noble gases are all found in minute quantities in the atmosphere, and are isolated by fractional distillation of liquid air. Helium can be obtained from natural gas wells where it has accumulated as a result of radioactive decay.

Inner Transition Metals- Periods 6 & 7

- Compose the f-block.
- Fill in Between Groups 3 & 4 on the Periodic Table.

Lanthanides (Period 6)- Rare Earth Metals

- Shiny reactive metals

Actinides (Period 7)

- Unstable & radioactive metals.
- Most are laboratory made.

Representative Elements

- Consist of the s & p block elements.
Reactivity of Metals Trend

Period Trend- Metals increase in activity from right to left on the periodic table.
  ▪ The alkali metals are the most reactive group of metals.

Group Trend- Metals increase in reactivity from top to bottom with a group.
  ▪ Ra is the most reactive alkaline earth metal.

Reactivity of Nonmetals Trend

Period Trend- Nonmetals increase in activity from left to right on the periodic table with the exception of the noble gases.
  ▪ The halogens are the most reactive group of nonmetals.

Group Trend- Nonmetals increase in reactivity from bottom to top with a group.
  ▪ F is the most reactive halogen.

1. What is the most reactive metal on the periodic table? Explain.

   Fr, furthest left and lowest on periodic table.

2. Circle the most reactive nonmetal in each row:

   a. Te  Po  S
   b. Br  I  Cl
Atomic Radii Trend

Atomic Radius - \( \frac{1}{2} \) the distance between the nuclei of two identical atoms that are bonded together.

- Measure from nucleus to nucleus and divide by two.
- The larger the atomic radius the larger the atom.

General Group Trend (representative elements)- Atomic radii increases as you travel down a group in the periodic table.

Reason: The number of energy levels increases as you travel down a group because the number of electrons increases. Each energy level added is farther from the nucleus, thus increasing the size of the atom.

General Period Trend (representative elements)- Atomic radii increases as you travel left across a period on the periodic table.

Reason: As you go right across a period, electrons are added to the same energy level. Simultaneously protons are being added to the nucleus of the atom. The more protons in the nucleus the higher the atom’s nuclear charge. This in turn creates a stronger force of attraction, which pulls the electrons closer to the nucleus resulting in a smaller atomic radius as your travel right. As you travel left across a period, the nuclear charge is less, thus the radius is larger.

Atomic Radii increases as you travel down & left on the periodic table!
1. Circle the atom with the **largest** atomic radii.
   
   a. I   F   Cl

   b. O   B   N

   c. S   Sb   Sr

2. Put the following elements in order of increasing atomic radii.
   
   a. Na, P, Rb

      P, Na, Rb

3. Which element has the largest atomic radii?

   Cs or F
Ionic Radii Trend

Ionic Radius- the radius of an ion.

Parent atom- atom from which an ion is formed.

Example: O is the parent atom of O\textsuperscript{2-}.

Ion- atom with a positive or negative charge due to the loss or gain of electrons.

Cation- a positive ion formed from the loss of electrons. Ex. Ca\textsuperscript{2+}

- Metals typically form cations.
- Cations are smaller than their parent atoms.

Reason: The electrons in the outermost energy levels are lost first. AS the highest energy levels are removed, the electron cloud gets smaller thus decreasing the radius.

Anion- a negative ion formed from the gain of electrons. Ex. Cl\textsuperscript{-}

- Nonmetals typically form anions.
- Anions are larger than their parent atoms.

Reason: As electrons are added to the electron cloud, electron repulsions forces cause them to spread out further thus increasing the ionic radii.

1. Circle the larger ion.
   a. Cl\textsuperscript{-1} Cl
   b. Na Na\textsuperscript{+1}
   c. N N-3
   d. Fe\textsuperscript{+2} Fe\textsuperscript{+3}
   e. C\textsuperscript{+4} C\textsuperscript{-4}
Electronegativity Trend

Electronegativity- measure of the ability of an atom in a chemical compound to attract electrons.

- The higher the electronegativity the greater the attraction for electrons.
- The electronegativity scale was determined by Linus Pauling.
- The electronegativity scale ranges from 0 to 4.
- Since some noble gases do not form compounds they have electronegativity values of 0.
- Used to determine the polarity of a bond.

General Group Trend (representative elements)- Electronegativity increases as you travel up a group in the periodic table.

General Period Trend (representative elements)- Electronegativity increases as you travel right across a period on the periodic table.

Electronegativity increases as you travel up & right on the periodic table!

1. Circle the atom with the highest electronegativity.
   
   a. F S Sn

   b. Li N O
Ionization Energy Trend

Ionization Energy- the energy required to remove an electron from an atom. An electron can be removed from any atom if enough energy is applied.

REMEMBER: Full or ½ orbits lend more stability to the atom.

- Nonmetals have higher ionization energies than metals.
- The lower the ionization energy of an atom the easier it is to remove an electron.

Ionization Energy is usually measured in KJ/mol

Ex. Li → Li⁺ + 1e⁻

More than one electron can be removed from an atom.

- Each electron removed requires more energy.
- 1st ionization energy- first electron removed
- 2nd ionization energy- second electron removed

General Group Trend (representative elements)- Ionization energy increases as you travel up a group in the periodic table.

Reason: As you travel up a group there are less principle energy levels, therefore, electrons are located closer to the nucleus. It is harder to remove electrons that are closer to the nucleus to due the nuclear charge attraction.

General Period Trend (representative elements)- Ionization energy increases as you travel right across a period on the periodic table.

Reason: As you go right across a period, the atomic number or the number of protons increases therefore so does the positive nuclear charge. Thus, the negatively charged electrons are more attracted to the nucleus making them harder to remove.

Ionization energy increases as you travel up & right on the periodic table!
1. Circle the atom with the **highest** ionization energy.
   a. F  O  B
   b. As  N  P
   c. F  C  Li

2. Why is there a jump in ionization energy between N and O?

   Electronic configuration of Oxygen is 1s² 2s² 2p⁴
   Electronic configuration of Nitrogen is 1s² 2s² 2p³

   An atom containing half filled or fully filled orbitals are very stable. Nitrogen has a half filled 2p orbital. So when it is already stable, why would it want to lose an electron upon supplying energy and become unstable? This is why one requires large amounts of energy to ionized the nitrogen atom.
   On the other hand, since Oxygen is already unstable relative to Nitrogen, by losing one electron it attains a stable half filled 2p orbital. So oxygen undergoes ionization at a relatively lower energy. This is why nitrogen has a higher

3. Why does Na have a large 2nd ionization energy but not a large 1st ionization energy?

   Na wants to lose 1 electron but does not want to lose a second electron so it takes a lot of energy to remove it.

4. An unknown element has the following successive ionization energies:

   1st ionization energy: 589 kJ/mol
   2nd ionization energy: 1144 kJ/mol
   3rd ionization energy: 4095 kJ/mol
   4th ionization energy: 6465 kJ/mol

   Predict this element’s family.

   Alkaline Earth Metals want to lose first two electrons but not 3rd. +2 charge.
Electron Affinity Trend

Electron Affinity- energy change that occurs when an electron is acquired by a neutral atom.

- Nonmetals have a more negative (LOWER) electron affinity than metals.
- Metals typically have a more positive (HIGHER) electron affinity than nonmetals.
- Noble gases have an electron affinity between metals & nonmetals.

Electron affinity is usually measured in KJ/mol
Ex. $S + 2e^- \rightarrow S^{2-}$

$A + e^- \rightarrow A^- \text{ (an anion)} + \text{energy}$

- Most atoms release energy when they acquire an electron.
- If energy is released, then the electron affinity value is negative (-kJ/mol). [more stable]

$A + e^- + \text{energy} \rightarrow A^- \text{ (an anion)}$

- Some atoms must be “forced” to take an electron by adding energy.
- If energy is absorbed, then the electron affinity value is positive (+kJ/mol). [less stable]

* The “-” or “+” only indicated energy released or gained. Therefore -300kJ is greater than -200kJ.

1. Circle the atom with the highest electron affinity.
   a. N   S   Ar
   b. O   Ne   Ca
   c. Cu   Ar   Kr
History of the Periodic Table

Dobereiner

1. How did he group elements?

2. Why did he choose this method for grouping elements?

3. Explain the law of triads.

4. What did he notice about the atomic mass of the middle element in a triad?

Newlands

5. How did he group elements?

6. Explain the law of octaves.

7. Why did some scientists dispute his work as coincidence?
Mendeleev

8. He proposed arranging elements in order of increasing ________________.

9. In what year did he arrange the first periodic table?

10. Why did Mendeleev leave “gaps” in his periodic table?

Moseley

11. He arranged elements in order of increasing ________________.

12. What is periodic law?
ARRANGEMENT OF THE PERIODIC TABLE

1. What group on the periodic table contains the most active nonmetals? ________________
2. What group on the periodic table contains the most active metals? ________________
3. As you go from left to right across a period, atomic size ( decreases / increases ).
4. As you travel down a group, atomic size ( decreases / increases ).
5. A anion is ( larger / smaller ) than its parent atom.
6. A cation is ( larger / smaller ) than its parent atom.
7. As you travel down a group, the first ionization energy ( decreases / increases ).
8. Group 1 elements are called ________________________________.
9. Group 2 elements are called ________________________________.
10. Group 3-12 elements are called ________________________________.
11. Most of the periodic table is made up of ( metals / nonmetals ).
12. Nonmetals are located on the ( right / left ) of the periodic table.
13. Group 17 elements are called ________________________________.
14. Group 18 elements are called ________________________________.
15. The most reactive metal on the periodic table is ________________.
16. The most reactive nonmetal on the periodic table is ________________.
17. A colored ion usually indicates a ________________________________.
18. Elements are arranged in periods according to their ________________________________.
19. Elements are arranged in groups based upon similar ________________________________.
20. An element with both metallic and nonmetallic properties is called a(n) ________________.
ATOMIC & IONIC RADII TRENDS

1. Circle the larger atom.
   a. K, Ga          b. Rb, Si          c. Mg, Ba          d. P, Ra

2. Circle the larger of the pair.

3. Explain why atomic radii increases as you travel down a group on the periodic table.

4. Explain why atomic radii increases as you travel from right to left across a period on the periodic table.

5. Explain why anions are larger than their respective parent atoms.
**Ionization Energy**

**Introduction:** Ionization energy is the amount of energy required to remove an electron from an atom. A higher ionization energy (IE) for an electron means the electron is harder to remove from the atom than an electron with a lower IE. The energy required to remove the first electron from an atom is denoted as $IE_1$; the energy required to remove the second electron is denoted as $IE_2$, and so on.

**Directions:** Below is a table of the ionization energies for the first 18 elements. Each row contains data for a different element; the columns indicate the energy required to remove the first, second, third, fourth, and fifth electrons of each atom. For example, for helium, 2,372 kJ/mol are required to remove its first electron; once that electron has been removed, an additional 5,250 kJ/mol are required to remove the second electron.

As you can see, there are several numbers missing from the table. You have two tasks: 1) examine any trends you see in the data, and 2) predict the values of the missing numbers. Be ready to defend your answers!

<table>
<thead>
<tr>
<th>Atomic #</th>
<th>Element</th>
<th>$IE_1$</th>
<th>$IE_2$</th>
<th>$IE_3$</th>
<th>$IE_4$</th>
<th>$IE_5$</th>
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</thead>
<tbody>
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<td>1,817</td>
<td>2,745</td>
<td>11,578</td>
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<tr>
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<td>3,931</td>
<td>5,771</td>
<td>7,238</td>
</tr>
</tbody>
</table>
ALL PERIODIC TRENDS

1. Circle the MOST reactive of the pair.
   a. K, Ga      b. Rb, Si      c. Mg, Ba      d. F, Br
   e. S, Ar      f. N, F      g. Ne, Br

2. Circle the larger atom.
   a. Na, Al     b. Ti, Zn     c. Ca, Ba     d. Si, Ra
   e. K, Ga     f. Rb, Si    g. Mg, Ba     h. P, Ra

3. Circle the largest of the pair.
   a. Li, Li\(^+\)  b. Fe\(^{+2}\), Fe\(^{+3}\)  c. O, O\(^{-2}\)  d. P, P\(^{-3}\)
   e. B, B\(^{+3}\) f. F, F

4. Circle the element with the largest electronegativity in each pair.
   a. Mg, Ba     b. Ba, C     c. F, Ne     d. Cl, P
   e. K, Se     f. N, As   g. Fe, Ne     h. Se, Ne

5. Circle the element with the largest 1st ionization energy in each pair.
   a. Na, Al     b. Ti, Zn     c. Ca, Ba     d. Si, Ra
   e. Rb, I     f. N, Sb    g. N, O

6. Circle the element with the largest electron affinity in each pair. (Remember look at the absolute value.)
   a. Cl, Ca     b. Na, O     c. Kr, Li     d. Ne, F
   e. C, F     f. C, Ne
Honors Chapter 5 Review

Part 1: Multiple Choice

Directions: Choose the letter that best completes or answers the following statements.

1. Who published the first periodic table? (Do you also know the year?)
   A. Dobereiner    B. Newlands    C. Mendelev    D. Moseley    E. Einstein

2. Mendelev's periodic table was arranged according to ______ while the modern periodic table is arranged according to ______.
   A. decreasing atomic number, decreasing atomic mass
   B. decreasing atomic mass, decreasing atomic number
   C. increasing atomic number, increasing atomic mass
   D. increasing atomic mass, increasing atomic number

3. Which of the following is a metallic list of elements?
   A. Li, Na, K, Rb    B. H, He, Li, Be    C. F, Cl, Br, I    D. B, Al, Ga, In
   A. Na    B. Mg    C. Al    D. Cl    E. Ar

5. The group name of the group containing calcium is
   A. alkali metals    B. noble gases    C. alkaline earth metals    D. halogens    E. representative elements

6. What are known as the electrons in the highest energy level?
   A. the octet    B. the core electrons    C. electronegativity    D. the valence electrons

7. Dobereiner classified elements into
   A. octets    B. a periodic table    C. triads    D. periods    E. groups of ten

8. Who rearranged the periodic table according to increasing atomic number?
   A. Dobereiner    B. Newlands    C. Mendelev    D. Moseley    E. Einstein

9. Which one of the following elements has the greater electron affinity value?
   A. K    B. Br

10. Which of the following would have the largest value change in ionization energy between the 4th ionization energy and 5th ionization energy (the 5th IE would be extremely large)?
    A. Na    B. Mg    C. Al    D. Si    E. P

11. The amount of energy required to remove one electron from a neutral atom is called the
    A. atomic radii    B. electronegativity    C. ionic radii    D. ionization energy    E. valence shell

12. The properties of the elements repeat periodically when the elements are arranged in increasing order by their atomic numbers is known as
    A. the law of octaves    B. the octet rule    C. the triad rule    D. the periodic law    E. the Mendeleev law

13. All of the alkali metals have which of the following in common?
    A. They all are relatively unreactive
    B. They all form compounds readily with oxygen
    C. They all have different numbers of valence electrons
    D. They all form diatomic molecules

14. Which group of elements do you expect to form a -1 charge?
    A. alkali metals    B. alkaline earth metals    C. halogens    D. noble gases

15. What is known as the ability of an atom to attract electrons in a chemical bond?
    A. atomic radii    B. ionization energy    C. electronegativity    D. electron affinity
16. What is known as the energy change that occurs when an atom takes an electron?
   A. atomic radii   B. ionization energy   C. electronegativity   D. electron affinity.

17. A certain element belongs to the p-block. It exhibits great conductivity, is a solid at room temperature, and is malleable and ductile. Further, it contains 4s² electrons, but not 5s² electrons. It is also known to have chemical properties similar to those of carbon. Identify the element.
   A. Sn   B. Sb   C. Ge   D. Pb   E. O

18. Which of the following metals would be the most reactive?
   A. Rb   B. Ca   C. Mg   D. Al

19. Which of the following nonmetals would be the most reactive?
   A. He   B. F   C. S   D. Se

20. You have just discovered three new elements. Element 1 is a solid at room temperature and lustrous. Element 2 is a gas at room temperature and a poor conductor of heat and electricity. Element 3 is a solid at room temperature, a fair conductor of heat and electricity, and not ductile. Classify each element as either a metal, nonmetal, or semimetal.

21. What is the main reason why metals decrease in size when they ionize? Ex: Na > Na⁺

22. What is the main reason why nonmetals increase in size when they ionize? Ex: Te²⁻ > Te⁻

23. How do scientists determine the radius of an atom?

**Ionization Energy Trends:** For each problem below, choose the particle with the LARGEST 1st ionization energy.

24. A. N   B. As   C. Bi

25. A. F   B. Ba   C. Si

26. A. Ne   B. Na   C. Rb

**Radii Trends:** For each problem below, choose the LARGEST particle.

27. A. N   B. As   C. Bi

28. A. F   B. Ba   C. Si

29. A. Mg   B. Mg²⁺

30. A. N   B. N³⁻

31. A. Br   B. I⁻

**Electronegativity:** For each problem below, choose the one with the LARGEST electronegativity value.

32. A. N   B. As   C. Bi

33. A. O   B. Li   C. Ne
Part 2: Periodic Table Groups

Directions: Label the following on the periodic table:
- halogens
- transition metals
- noble gases
- alkaline earth metals
- alkali metals
- actinides
- lanthanides
- main block elements
- metalloids

![Periodic Table Image]

Part 3: History of the Periodic Table

Directions: Name the scientist.

1. _______________________ Developed the law of octaves.
2. _______________________ Introduced triads.
3. _______________________ Added noble gases to the periodic table.
4. _______________________ Arranged the periodic table by atomic mass.
5. _______________________ Arranged the periodic table by atomic number.
6. _______________________ Developed method for determining average atomic mass.

Part 4: Period Trends

Directions: Name the periodic trend.

1. _______________________ Determines the size of an ion.
2. _______________________ Energy required to remove an electron from an atom.
3. _______________________ Determines the size of an atom.
4. _______________________ Energy change that occurs when an atom acquires an electron.
5. _______________________ Ability of an atom to attract electrons in a chemical bond.
Part 6: Trends

Atomic Radii Trends
Directions: Circle the atom with the LARGEST atomic radii.

1. Be  Ra  Ba
2. F  Te  Po

Ionic Radii Trends
Directions: Circle the atom with the LARGEST ionic radii.

3. Mg$^{2+}$  Mg
4. S  S$^{2-}$

Ionization Energy Trends
Directions: Circle the atom with the LARGEST ionization energy.

5. Mg  Si  Cl
6. Pb  N  P

Electron Affinity Trends
Directions: Circle the atom with the LARGEST electron affinity.

7. S  O  Kr
8. Na  Ar  P

Electronegativity Trends
Directions: Circle the atom with the LARGEST electronegativity.

9. Li  Rb  Cs
10. B  C  N
The Periodic Table of the Elements (with Ionization Energies)

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<th>Symbol</th>
<th>Atomic #</th>
<th>First ionization energy (kJ/mol)</th>
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*lanthanides
**actinides

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