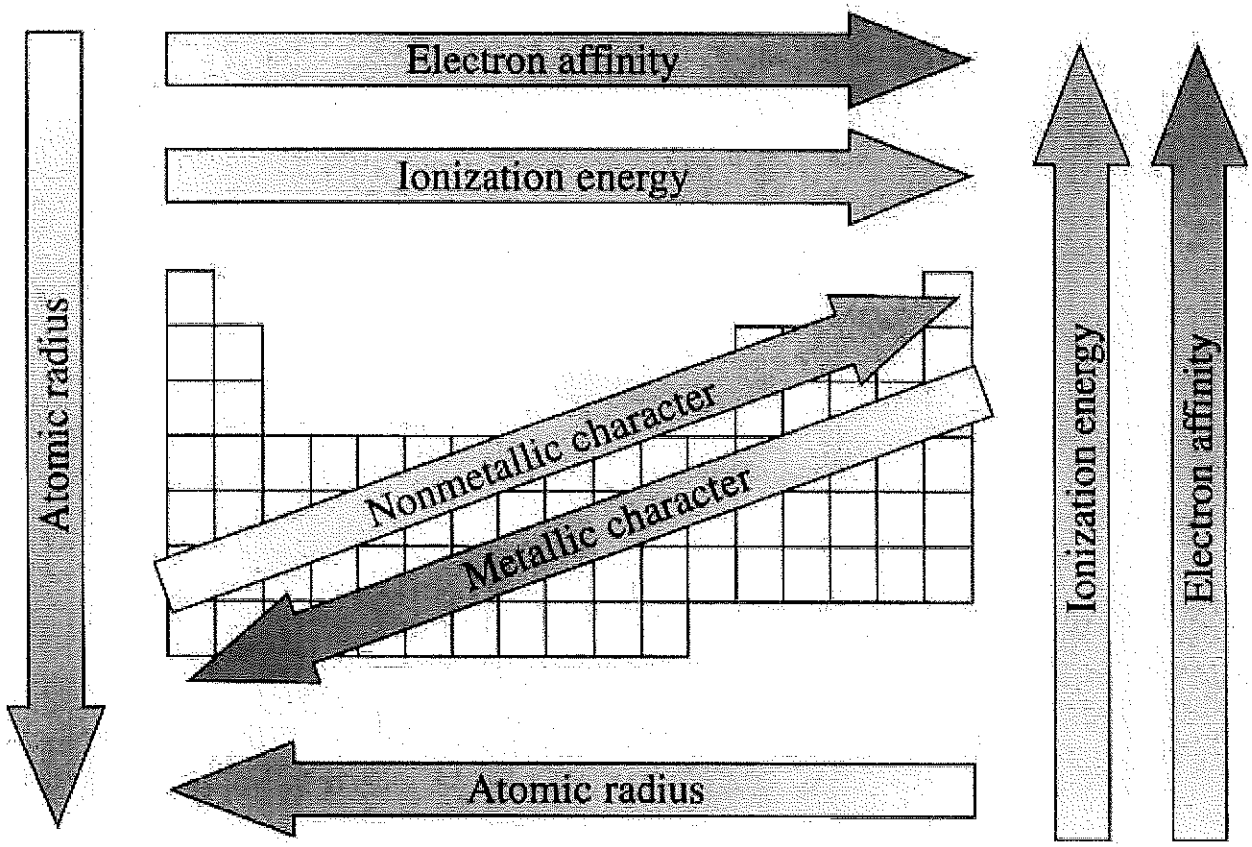
CHAPTER 5:

THE PERIODIC LAW



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Chapter 5: The Periodic Law
Reading Guide

5.1 – History of the Periodic Table *(pgs. 125-129)*

1) What did Dimitri Mendeleev notice when he started to organize the elements in order of increasing atomic mass?

2) What was Henry Moseley's contribution to the development of the periodic table?

3) Define what is meant by **periodic law** and explain how this is related to the periodic table.

5.2 – Electron Configuration and the Periodic Table *(pgs. 130-141)*

4) Explain the difference between the arrangement of **groups** and **periods** as it relates to the periodic table.

5) Identify some **key features for elements** that are associated with the following categories:

- **s-block**: (Groups 1 & 2)

  Group 1 - known as _____ (not including H)

  Key features:

  Group 2 - known as _____

  Key features:

- **d-block**: (Groups 3-12)

  Known as the _____ metals

  Key features:

- **p-block:** (Groups 13-18)

  The p-block elements together with the ____-block elements are called the _____ _____ elements.

  The properties of the elements in the p-block _____ greatly.

  Group 17 - known as _____

  Key features:

  Group 18 - known as _____ (see bottom of pg.127)

  Key features:

- **f-block**: (two periods at the bottom of periodic table)

  4f - known as the _____

  Key features:

  5f - known as the _____

  Key features:

5.3 – Electron Configuration and Periodic Properties *(pgs.142-156)*

6) Define **atomic radius (AR)**:

- Period trend: generally _____ across a period.

  Why? -

- Group trend: generally _____ down a group.

  Why? –

7) Define **ionization energy (IE)**:

- Period trend: generally _____ across a period.

  Why? -

- Group trend: generally _____ down a group.

  Why? –

8) Define **electron affinity (EA)**:

How does electron affinity *differ* from ionization energy?

- Period trend: generally _____ across a period.

  Why? -

- Group trend: generally _____ down a group.

  Why? –

9) Define the term **valence electrons**:

- For main-group elements, the valence electrons are the electrons in the **outermost** (highest energy level) _____ & _____ sublevels.

10) Using **Figure 3.10**, complete the following table which identifies the valence electrons in main-groups elements (s & p-block)

Group #	**Group configuration** ("n" = any energy level)	**# of Valence Electrons**
1		
2		
13		
14		
15		
16		
17		
18		

11) Define the term **electronegativity (EN)**:

- Period trend: generally _____ across a period.

  Why? -

- Group trend: generally _____ down a group.

  Why? –

Quick Guide to the Periodic Table

The Periodic Table is a list of known elements. It is organized by increasing **atomic number**. There are two main groups on the periodic table: **metals** and **nonmetals**. The left side of the table contains elements with the greatest metallic properties. As you move from left to right, the elements become less metallic with the far right side of the table consisting of nonmetals. The elements in the middle of the table are called "transition" elements because they are changed from metallic properties to nonmetallic properties. A small group of elements whose members touch the zigzag line (referred to as the "staircase") are called **metalloids** they have both metallic and nonmetallic properties.

The table is also arranged in vertical columns called **"groups"** or **"families"** and horizontal rows called **"periods"**. Each arrangement is significant. The elements in each vertical column or group have similar properties. Group 1 elements all have 1 electron in their outermost shells. This gives them similar chemical properties. Group 2 elements all have two electrons in their outer shells. This also gives them similar properties to one another. Not all groups, however, hold true to this pattern; for example, group 16 begins with a nonmetal (C), includes metalloids (Si & Ge), and includes metals (Sn & Pb). The elements in the first period or row all have electrons within the 1^{st} energy level (thus the atoms only contain one electron shell). The elements in period 2 have electrons in both the 1^{st} and 2^{nd} energy level (thus the atoms contain two electron shells). The elements in period 3 have three electron shells and so on.

There are a number of major groups containing elements with similar properties. These are:

Hydrogen: This element does *not match* the properties of any other group. It is very unique and so it stands alone. It is placed above group one but is not a part of that group. It is a very reactive, colorless, odorless gas at room temperature. (1 valence electron – *only 1 electron)

Group 1: **Alkali Metals** – These metals are extremely reactive and are never found in nature in their pure Form. They are silver colored and shiny. Their density is extremely low so they are soft enough to be cut with a knife. (1 valence electron – ns^1)

Group 2: **Alkaline Earth Metals** – These are slightly less reactive than alkali metals. They are silver colored and denser than alkali metals. (2 valence electrons – ns^2)

Group 3-12: **Transition Metals** – These metals have a moderate range of reactivity and a wide range of properties. In general, they are shiny and good conductors of heat and electricity. They also have higher densities and melting points than groups 1 & 2 (1 or 2 valence electrons depending on the element – ex: Cr= [Ar] $4s^1 3d^5$ or Mn= [Ar] $4s^2 3d^5$)

Lanthanides & Actinides: These are also transition metals that were taken out and placed at the bottom of the periodic table so the table would not be so wide. The elements in each of these two periods share many unique properties, so it is fitting that they are separated on the table. The lanthanides (atomic #s 57-71) are shiny and reactive. The actinides (atomic #s 89-103) are *all* radioactive and are therefore unstable. Elements 95 through 103 do not exist in nature but have been synthetically manufactured in the lab.

Group 13: **Boron Family/Group** – Contains 1 metalloid and 4 metals. Reactive. Aluminum is in this group. It is the most abundant metal in the earth's crust. (3 valence electrons – ns^2p^1)

Group 14: **Carbon Family/Group** – Contains 1 nonmetal, 2 metalloids, and 2 metals. Varied reactivity (4 valence electrons - ns^2p^2)

Group 15: **Nitrogen Family/Group** – Contains 2 nonmetals, 2 metalloids, and 1 metal. Varied reactivity. (5 valence electrons - ns^2p^3)

Group 16: **Chalcogens** – Contains 3 nonmetals, 1 metalloid, and 1 metal. Reactive group. (6 valence electrons - ns^2p^4)

Group 17: **Halogens** – Contains 4 nonmetals and 1 metalloid. Very reactive. Poor conductors of heat and electricity. Tend to form salts with metals, NaCl= sodium chloride (7 valence electrons- ns^2p^5)

Group 18: **Noble Gases** – Unreactive nonmetals. All are colorless, odorless, gases at room temperature. All are found in earth's atmosphere in small quantities (8 valence electrons - ns^2p^6)

Periodic Table Questions:

1) The vertical columns on the periodic table are called _____.
2) The horizontal rows on the periodic table are called _____.
3) Most of the elements in the periodic table are classified as _____.
4) The elements along the zigzag line or the staircase are classified as _____.
5) The elements in the upper far right corner are classified as _____.
6) Elements in the first group have _____ outer shell electron and are very reactive. They are called the _____.
7) Elements in the second group have _____ outer shell electrons and are also very reactive. They are called the _____.
8) Elements in groups 3-12 have many useful properties and are called the _____.
9) Elements in group 16 have _____ outer shell electrons and are known as the _____.
10) Elements in group 17 are known as "salt formers" and have _____ outer shell electrons. They are called _____.
11) Elements in group 18 are stable and chemically unreactive. They are said to be "inert" and have _____ outer shell electrons. We call these _____.
12) The elements at the bottom of the periodic table were pulled out to keep the table from becoming too wide. The first period along the bottom of the table is called the _____.
13) The second period along the bottom of the table which consists of radioactive, mostly synthetic, elements is called the _____.

Color Coding the Periodic Table

Directions: Color code the attached periodic table by following and checking off each instruction.

States of Matter:

_____ draw a BLUE box around the elements that are **liquids at room temperature**

_____ draw a RED box around the elements that are **gases at room temperature**

_____ leave all the elements that are **solids at room temperature** blank

Types of Elements:

_____ trace the **zigzag line** that separates the metals and nonmetals in BLACK

_____ draw a BLACK box around the elements that are **metalloids-**

(B, Si, Ge, As, Sb, Te, At)

Periods:

_____ number the **periods** in the left margin (#1-7)

Families/Groups:

_____ number the **families/groups** at the top of each column (#1-18)

_____ color the **alkali metals** BLUE – *except for hydrogen* (leave H blank!)

_____ color the **alkali earth metals** GREEN

_____ color the **transition metals** YELLOW

_____ color the **lanthanides** YELLOW and draw a small **"L"** in the corner of each box

_____ color the **actinides** YELLOW and draw a small **"A"** in the corner of each box

_____ color the **boron family** members PURPLE

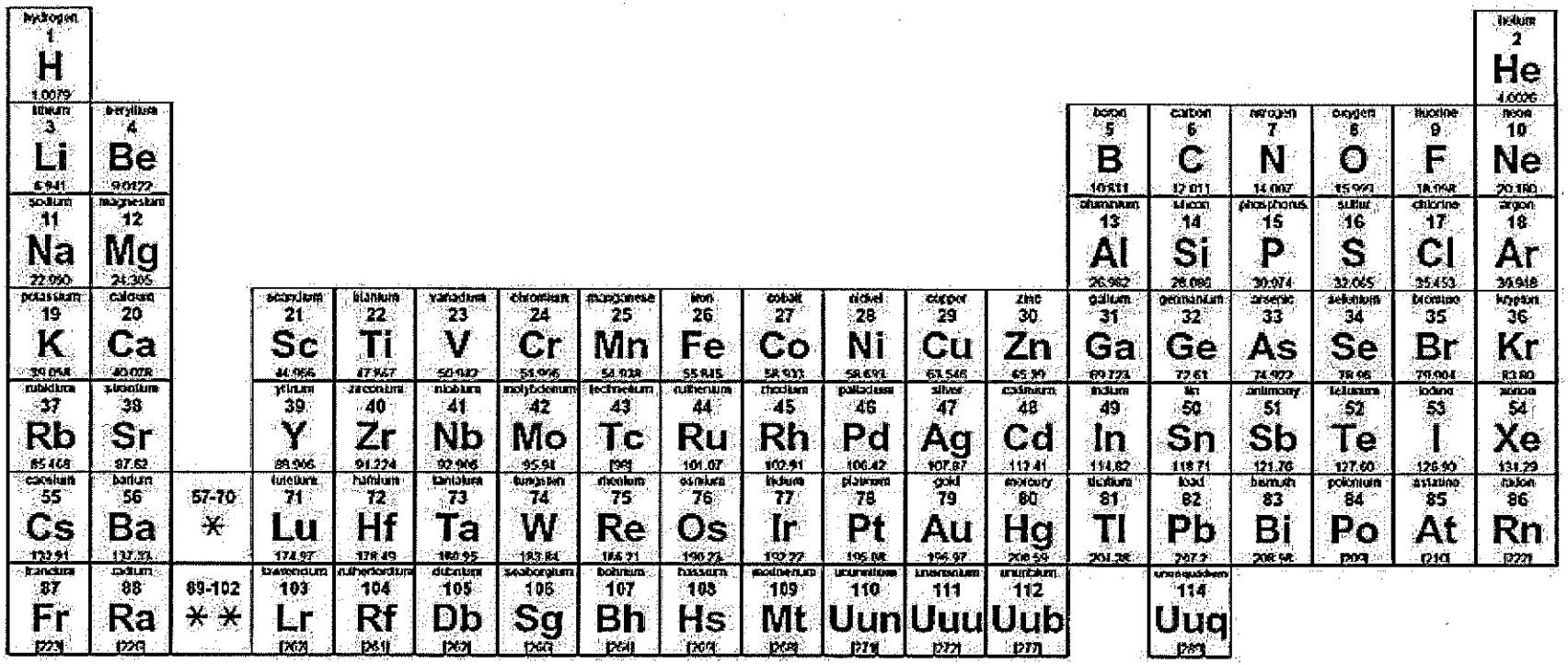
_____ color the **carbon family** members GREY

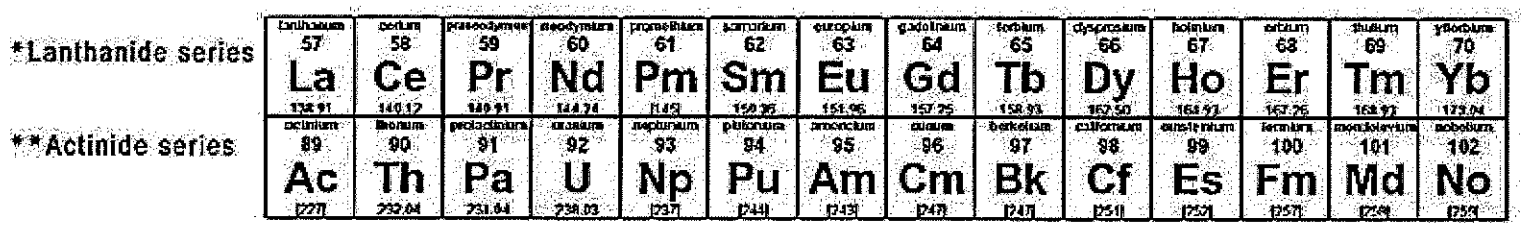
_____ color the **nitrogen family** members PINK

_____ color the **chalcogens (oxygen family)** ORANGE

_____ color the **halogens** BROWN

_____ color the **noble gases** RED





KEY:

- ☐ Metals
- ☐ Nonmetals
- ☐ Metalloids
- ☐ Solid
- ☐ Liquid
- ☐ Gas
- ☐ Alkali Metals
- ☐ Alkali Earth Metals
- ☐ Transition Metals
- ☐ Lanthanides
- ☐ Actinides
- ☐ Boron Family
- ☐ Carbon Family
- ☐ Nitrogen Family
- ☐ Oxygen Family
- ☐ Halogens
- ☐ Noble Gases

In-Class Notes for Chapter 5

1) Define **valence electrons**:

2) Indicate the number of **valence electrons** for each group on the periodic table.

Group 1 (Alkali metals): _____ Group 15 (Nitrogen family): _____

Group 2 (Alkaline-earth metals: _____ Group 16 (Chalcogens): _____

Group 13 (Boron family): _____ Group 17 (Halogens): _____

Group 14 (Carbon family): _____ Group 18 (Noble gases): _____

3) ***Bohr diagrams*** show how many ***total electrons*** are in an atom and to ***which energy levels*** those electrons belong. Draw the example Bohr diagrams from the board in the space below.

Generic Bohr Diagram	*Helium:* _____
Oxygen: _____	*Argon:* _____

4) **Lewis structures** show only the **valence electrons** for an atom. Draw the example Lewis structures from the board in the space below.

Bohr Diagrams & Lewis Dot Structures

Directions:

Bohr Diagrams: Draw a Bohr diagram for each element in the upper right hand corner of each box. Rather than drawing individual protons and neutrons in the nucleus, you may simply write how many of each there are in the nucleus (ex: 2p, 2n). Then draw the individual electrons on the appropriate energy levels (keep in mind the maximum number of electrons allowed on each level).

Lewis Dot Structure: For each element, draw the number of *valence electrons* (represented as dots) around the symbol of the element.

1) Beryllium p^+ _____ n^0 _____ e^- _____ Group Name: _____ Group #: _____ Period #: _____ **Lewis Dot Structure:**	2) Sodium p^+ _____ n^0 _____ e^- _____ Group Name: _____ Group #: _____ Period #: _____ **Lewis Dot Structure:**
3) Sulfur p^+ _____ n^0 _____ e^- _____ Group Name: _____ Group #: _____ Period #: _____ **Lewis Dot Structure:**	4) Fluorine p^+ _____ n^0 _____ e^- _____ Group Name: _____ Group #: _____ Period #: _____ **Lewis Dot Structure:**

5) Calcium p^+ _____ n^0 _____ e^- _____ Group Name: _____ Group #: _____ Period #: _____ **Lewis Dot Structure:**	6) Argon p^+ _____ n^0 _____ e^- _____ Group Name: _____ Group #: _____ Period #: _____ **Lewis Dot Structure:**
7) Carbon p^+ _____ n^0 _____ e^- _____ Group Name: _____ Group #: _____ Period #: _____ **Lewis Dot Structure:**	8) Nitrogen p^+ _____ n^0 _____ e^- _____ Group Name: _____ Group #: _____ Period #: _____ **Lewis Dot Structure:**

Questions:

9) What information does the group # tell you? _____

10) What information does the period # tell you? _____

11) If two elements are in the same group, what does this mean?

_____

12) If two elements are in the same period, what does this mean?

_____

13) What do you notice about the arrangement of electrons for the elements in group 18?

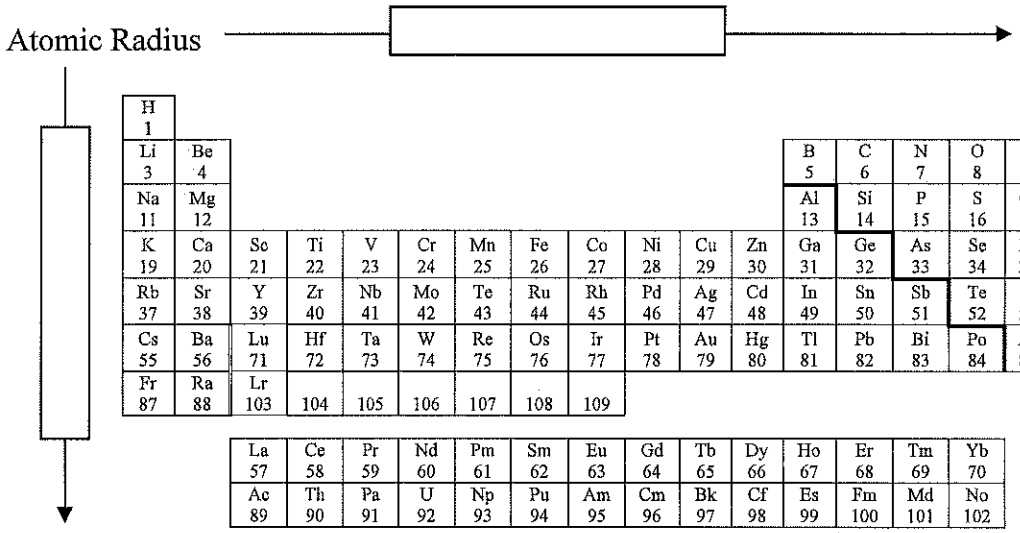
_____

Periodic Trends: Atomic Radii

1. Define atomic radius:

2. What element has the largest atomic radius on the periodic table? _____

3. Using your hint element above, fill in the boxes below with either *decreases* or *increases* to appropriately indicate the trends observed for atomic radii across a period and down a group.



H 1																	He 2
Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Te 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
Cs 55	Ba 56	Lu 71	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
Fr 87	Ra 88	Lr 103	104	105	106	107	108	109									118

La 57	Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70
Ac 89	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102

4. Explain what causes the trend observed for the size of the atom *across a period* in terms of nuclear charge, the location of the electrons, and electrostatic attraction.

_____

_____

_____

5. Explain what is meant by *shielding effect* and use this term to explain what causes the trend observed for size of the atom *down a group.* _____

_____

_____

_____

6. For each of the following pairs of atoms, circle the atom that has the larger radius:

   a. Ca or Sr
   b. Na or P
   c. Li or Cs
   d. K or Ca

7. List the following atoms in order of increasing atomic radii (smallest to largest):

   Li F Na Mg

Periodic Trends: Ionization Energy

1. Define ionization energy:

2. What element has the highest ionization energy on the periodic table? _____

3. Using your hint element above, fill in the boxes below with either *decreases* or *increases* to appropriately indicate the trends observed in the ionization energies across a period or down a group.

Ionization Energy → [ ]

↓

H 1																	He 2
Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Te 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
Cs 55	Ba 56	Lu 71	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
Fr 87	Ra 88	Lr 103	104	105	106	107	108	109									118

La 57	Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70
Ac 89	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102

4. Explain what causes the trend observed for the ionization energies *across a period* in terms of the relative size of the atom and electrostatic attraction. _____

_____

_____

5. Explain what causes the trend observed for the ionization energies *down a group* in terms of nuclear charge and the location of the electrons. _____

_____

_____

6. For each of the following, circle the atom that has the larger ionization energy:
   a. Ca or Sr
   b. Na or Cl
   c. Li or Cs
   d. K or Ca

7. Noting the positions of metals versus nonmetals on the periodic table, would the metals or the nonmetals generally have the lower ionization energy? _____

8. List the following atoms in order of increasing ionization energy (smallest to largest):

   Li F Na Mg

Atomic Radii and Ionization Energy Practice Worksheet

1) Which element on the periodic table has the highest ionization energy? _____

2) Ionization energy _____ across a period and _____ down a column.

3) Rank each of the following sets of atoms from lowest to highest ionization energy.

a) Mg, Si, S

b) Mg, Ba, Ca

c) Ne, Cu, Ba

d) Si, P, N

4) In each of the following sets, circle the element with the *highest* ionization energy:

a) Li, K, Cs

b) S, Cl, Ar

c) Br, I, Te

d) Cl , Na, Al

5) Which element on the periodic table has the largest atomic radius? _____

6) Atomic radius _____ across a period and _____ down a column.

7) Rank each of the following sets of atoms from lowest to highest atomic radius.

a) N, C, Li

b) Ne, O, C

c) O, P, Si

d) Mg, K, P

8) In each of the following sets, circle the element with the *largest* atomic radius:

a) Sr, Rb, Ba

b) O, Se, Po

c) Cu, Ni, C

d) P, N, As

Periodic Trends: Electron Affinity and Electronegativity

1. Define electron affinity:

2. What is the difference between electron affinity and ionization energy? _____

3. Define electronegativity:

4. Based on the definitions above and your knowledge of valence electrons, which family has a value of zero on Pauling's electronegativity scale and the lowest electron affinity? Why do you think this is? _____

5. Which element has the highest electron affinity and electronegativity on the periodic table? _____

6. Using your hint element above, fill in the boxes below with either *decreases* or *increases* to appropriately indicate the overall trends observed in the electronegativity values across a period or down a group.

Electronegativity → [ ]

1 H 2.20																	2 He
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98	10 Ne
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16	18 Ar
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96	36 Kr
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Te 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66	54 Xe
55 Cs 0.79	56 Ba 0.89	71 Lu	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2	86 Rn
87 Fr 0.70	88 Ra 0.90	103 Lr	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt									118

7. Explain what causes the trend observed for electronegativity *across a period* in terms of nuclear charge and the location of the electrons. _____

8. Explain what causes the trend observed for electronegativity *down a group* in terms of nuclear charge and the location of the electrons. _____

_____

_____

9. For each of the following, circle the atom that has the higher electronegativity:
   a. Ca or Sr
   b. Na or P
   c. Li or Cs
   d. Ge or O

10. List the following atoms in order of decreasing electronegativity (largest to smallest):

Li F Na Mg

Periodic Trends: Practice Worksheet

1) Rank the following elements by **increasing atomic radius:**

   ➔ carbon, aluminum, oxygen, potassium

2) Rank the following elements by **decreasing electron affinity:**

   ➔ barium, nitrogen, zinc, silicon

3) Rank the following elements by **decreasing ionization energy:**

   ➔ strontium, phosphorous, chlorine, cesium

4) Rank the following elements by **increasing electronegativity:**

   ➔ sulfur, oxygen, neon, aluminum

5) What is the difference between electron affinity and ionization energy?

6) <u>Why</u> does fluorine have a higher ionization energy than iodine? Be specific about the reason – don't just say "because it's higher up in the group".

7) Why do elements in the same family generally have similar chemical properties?

Chapter 5 Review: Answer each of the following questions.

1. Who published the first classification of the elements and how was this table organized?

2. Who reorganized the periodic table into the format we still use today and what property did he use to do this?

3. What is a period? How many are there in the periodic table?

4. What is a group (also called a family)? How many are there in the periodic table?

5. State the number of valence electrons in an atom of:

   a) sulfur _____ b) calcium _____ c) chlorine _____ d) arsenic _____

6. List, by number, both the period and group of each of these elements:

	Symbol	Period	Group
a) beryllium			
b) iron			
c) lead			

7. Which of the following pairs of elements below belong to the same period?

   a) Na and Cl b) Na and Li c) Na and Cu d) Na and Ne

8. Which of the following pairs of elements below belong to the same group?

   a) H and He b) Li and Be c) C and Pb d) Ga and Ge

9. In what type of orbitals (s, p, d or f) are the actinide and lanthanide electrons found?

10. Would you expect strontium to be, chemically, more similar to calcium or rubidium and WHY?

11. In going from top to bottom of any group, each element has _____ more occupied energy level(s) than the element above it.

12. What are the Group 1 elements called? _____

13. What are the Group 2 elements called? _____

14. What are the Group 16 elements called? _____

15. What are the Group 17 elements called? _____

16. What are the Group 18 elements called? _____

17. What is the family name given to the group of elements that have the following valence shell electron configurations?

a) s^2 _____ b) s^2p^6 _____

c) s^1 _____ d) s^2p^5 _____

18. Which alkali metal belongs to the sixth period? _____

19. Which halogen belongs to the fourth period? _____

20. What element is in the fifth period and the eleventh group? _____

21. The members of a family or group have similar properties because they all have the same number of _____ electrons.

22. What major factor causes atomic radius to decrease across a period?

23. What major factor causes atomic radius to increase down a column?

24. List the following elements in order of **increasing atomic radius**: Mg, Cl, Al, Ar

25. List the following elements in order of **decreasing ionization energy**: Sn, Ge, C, Si

26. List the following elements in order of **decreasing electronegativity**: Cl, K, Cu, F