LOCATING MAIN GROUP ELEMENTS ON THE PERIODIC TABLE NOTES

Given the electron configuration or noble gas configuration for an element, it is possible to determine its location on the Periodic Table without actually looking at a Periodic Table.

* To tell which period this element is in... ~ find the H.O.E.L. for this element
You can do this by... ~ _____
The _____ is the number of the period where the element is located.

* To tell which "block" (s, p, d, f) this element is in... ~ find the highest occupied sublevel for this element
You can do this by... ~ _____
The _____ is the "block" where the element is located.

* To tell which group this element is in... ~ find the highest occupied energy level for this element
Then... ~ _____
This gives you the number of valence electrons in the element.
You will then know that 1 valence e- indicates that the element is in Group ____
2 valence e- indicates that the element is in Group ____
3 valence e- indicates that the element is in Group ____
4 valence e- indicates that the element is in Group ____
5 valence e- indicates that the element is in Group ____
6 valence e- indicates that the element is in Group ____
7 valence e- indicates that the element is in Group ____
8 valence e- indicates that the element is in Group ____

Look at the following EXAMPLE: [Ar] $4s^2$ $3d^{10}$ $4p^2$
It is possible to tell the period, group, and "block" where this element is located.
* Period = ____ because _____
* Block = ____ because _____
* Group = ____ because _____
_____

LOCATION OF ELEMENTS PRACTICE

	Noble Gas Config.	Period	Block (s, p, d, f)	Group
1	[Ne] $3s^2$ $3p^2$			
2	[Ar] $4s^2$ $3d^{10}$ $4p^6$			
3	[Xe] $6s^2$			
4	[Kr] $5s^2$ $4d^{10}$ $5p^5$			
5	[Ar] $4s^2$ $3d^{10}$ $4p^1$			
6	[He] $2s^2$ $2p^3$			
7	[Kr] $5s^2$ $4d^{10}$ $5p^4$			
8	[He] $2s^1$			
9	[Xe] $6s^2$ $4f^{14}$ $5d^{10}$ $6p^2$			
10	[Rn] $7s^2$			

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MAIN GROUP (GROUPS 1, 2, AND 13 - 18) ELEMENTS NOTES

GROUP	NAME	ENDING OF CONFIG.	# OF VALENCE e-	ELECTRON DOT DIAGRAM	WANTS TO (lose or gain) e- TO BE LIKE N. GAS?	CHARGE OF ION & OXID-ATION #	ION (smaller or larger) THAN ATOM
1							
2							
13							
14							
15							
16							
17							
18							

HISTORY OF THE PERIODIC TABLE NOTES

I. Mendeleev and Chemical Periodicity
   A. Wanted to organize elements according to their _____
   B. When elements were arranged in order of increasing atomic mass*, similarities in chemical properties appeared at regular intervals (_____)
   C. *Several elements did not quite fit this pattern - Mendeleev put elements with similar _____ in the same column or group rather than trying to keep the trend of increasing mass
   D. 1871 - Mendeleev predicted the existence and properties of several (then undiscovered) elements. These elements were:
   E. Within 15 years, those elements with those properties had been discovered

II. Moseley and the Periodic Law
   A. When elements were arranged in order of increasing _____, there was a distinct regular pattern.

B. _____: The physical and chemical properties of the elements are periodic functions of their atomic numbers.
C. In other words, when elements are arranged in order of increasing atomic number, elements with similar properties appear at regular intervals.
D. Bottom line = elements in the same group have similar properties

III. Modern Periodic Table: arrangement of the elements in order of their atomic numbers so that elements with similar properties fall in the same group

ELECTRON CONFIGURATION & THE PERIODIC TABLE NOTES

I. Stability of Noble Gases
A. Noble gases undergo very few chemical reactions - why?
B. Highest occupied energy level contains _____
C. Electrons in the highest occupied energy level are what determines an element's _____

II. Periods and Blocks of the Periodic Table
A. Horizontal row = _____; 7 on modern Periodic Table
B. Length of period determined by the sublevels being filled in that period
C. Period 1: only _____ sublevel being filled; can hold a maximum of _____ electrons; period contains _____ elements
D. Period 4: ____, ____, and ____ sublevels being filled; s can hold ___ electrons, d can hold ___ electrons, & p can hold ___ electrons; total of _____ electrons; Period 4 contains _____ elements
E. Period can be determined from the element's electron configuration
1. Bromine: [Ar] $4s^2$ $3d^{10}$ $4p^5$
2. Highest number in front of letter is the element's highest occupied _____ - tells which period the element is in
3. For bromine, _____ is highest number, so it is in Period _____

III. The "s" block elements: Groups 1 and 2
A. Group 1 - Alkali Metals
1. generalized outermost energy level (valence) electron configuration:
2. silvery appearance
3. soft enough to cut with a knife
4. not found in nature as free elements - they're always part of a compound
B. Group 2 - Alkaline Earth Metals
1. generalized valence electron configuration:
2. harder, stronger, more dense than Group 1
3. less reactive than Group 1, but still not found in nature as free elements
C. Exceptions: Hydrogen and Helium
1. Hydrogen (H)
a. electron configuration:
b. properties do not resemble those of any other element on the periodic table
2. Helium (He)
a. electron configuration:
b. in Group 18 because

IV. The "d" block elements: Groups 3 - 12
A. called
B. have typical metallic properties: ductile, malleable, shiny, solid, conduct electricity
C. less reactive than _____ block elements
D. found in nature as free elements
E. usual ending of electron configuration:

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V. The "p" block elements: Grps 13 – 18 ("s" & "p" block elements together referred to as MAIN GROUP elements)
   A. ending electron configurations of _____ through _____
   B. properties vary greatly b/c there are metals, metalloids, and nonmetals
   C. Group 17 - Halogens
      1. most reactive nonmetals
      2. seven electrons in outermost energy level
   D. "p" block metals are harder and more dense than "s" block , but not as hard or dense as the "d" block metals

PERIODIC TRENDS NOTES

Electronegativity (EN)/Electron Affinity (EA):

EN/EA Left to Right across a Period: _____ (not including _____)

Why?

* Elements on the **left** side of the P.T. (_____) want to _____ electrons.
  Elements on the **right** side of the P.T. (_____) want to _____ electrons.
  Trend does not include _____ because these elements do not want to _____ or _____ electrons.

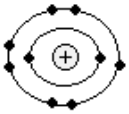
** The more an element wants to gain electrons, the _____ the electronegativity.**

EN/EA Top to Bottom in a Group: _____

Why?

Group 17

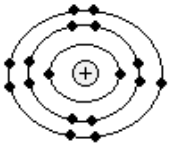
only **2** inner-shell (non-valence) electrons that interfere & lessen the attraction the nucleus has for outside electrons



Explanation: all Group 17 elements want to _____ ____ electron to have a full "s" and "p" sublevel in their HOEL

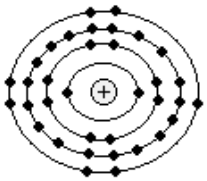
Elements at the top of a group have a _____ attraction for outside electrons (higher EN) because their nucleus (where positive charges are) is physically _____ to the element's HOEL.

more inner-shell electrons = weakened attraction for outside electrons



Additionally, more occupied energy levels not only results in a _____ distance between the nucleus and the HOEL, but there are more electrons in those energy levels.

has **28** inner-shell (non-valence) electrons that interfere & lessen the attraction the nucleus has for outside electrons



These inner-shell electrons interfere (and lessen/decrease) the strength of the attraction that the nucleus has for outside electrons. This interference of inner-shell electrons is called the _____.

The more electrons that shield the nucleus, the harder it is to attract an outside electron. Therefore, the more electrons that shield the nucleus, the _____ the electronegativity.

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Ionization Energy (IE):

IE Left to Right across a Period: _____

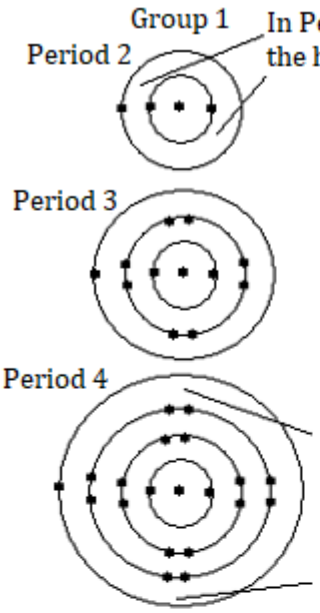
Why?

* Elements on the **left** side of the P.T. (_____) want to _____ electrons. Therefore, it will require a _____ amount of energy to remove an electron.

Elements on the **right** side of the P.T. (_____) want to _____ electrons. Therefore, it will require a _____ amount of energy to remove (take away) an electron.

IE Top to Bottom in a Group: _____

Why?



In Period 2, there are only ____ inner-shell electrons "blocking" the hold the nucleus has on the valence electron.

In Period 4, potassium has _____ inner-shell electrons that _____ the attraction that the nucleus has for the valence electron. Therefore, _____ energy is required to remove the valence electron from the Period 4 element than remove the valence electron from the Period 2 element.

Elements in Group 1 want to _____ their ____ valence electron. Elements at the _____ of the Periodic Table lose their valence electron more easily (and with less required _____) than elements at the top of the Periodic Table.

Elements at the _____ of the Periodic Table lose their valence electron easier than elements at the _____ because when there are more inner-shell electrons, the nucleus cannot hold onto the valence electron very well. (And so elements at the _____ of the P.T. have a lower ionization energy than elements at the _____ of the P.T.

This decreased hold (or attraction) is due to the interference of the inner-shell electrons. The more inner-shell electrons, the _____ the shielding effect. A higher shielding effect means that those elements have a _____ ionization energy.

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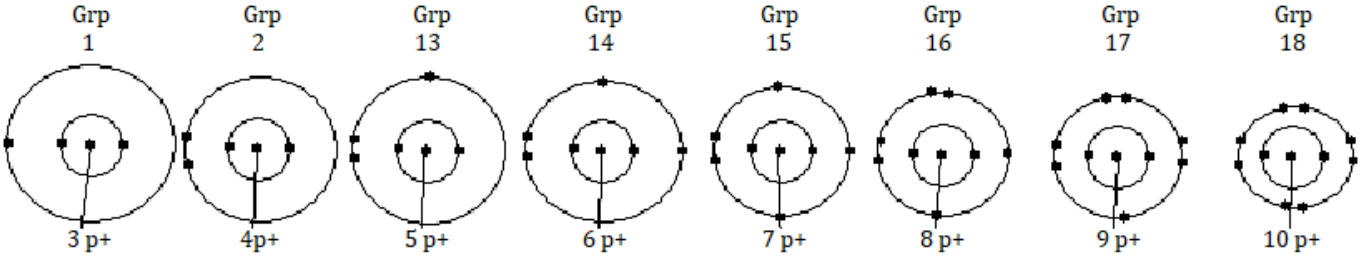
Atomic Radius (AR):

AR Left to Right across a Period: _____

Why?



In all Period 2 elements, there are two (2) _____ _____ _____. As atomic number increases in a period (from left to right), the number of protons in the nucleus of each atom _____.

The more protons in the nucleus (positive charges), the stronger the attraction for the _____ - _____ _____. Even though there are also more _____ in each atom, the electrons are distributed over the same number of _____ _____.

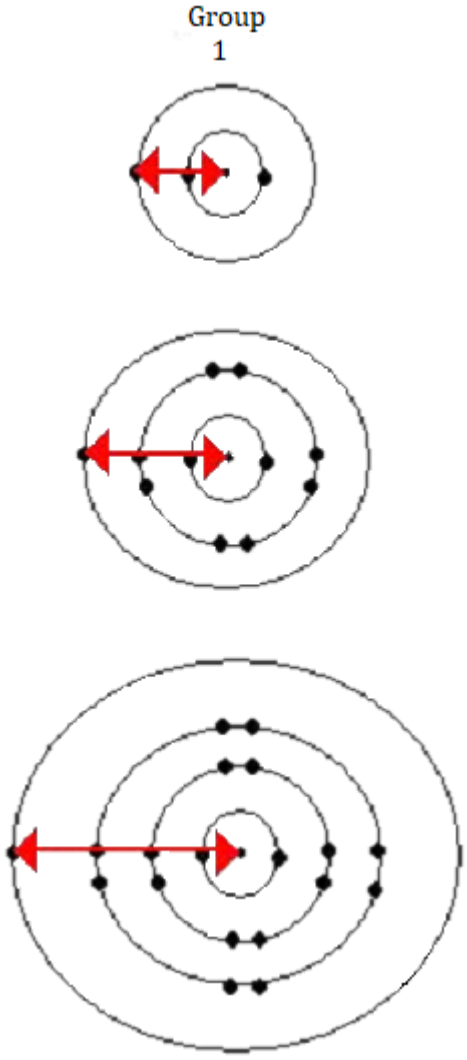
The stronger the attraction (more protons in the nucleus), the _____ the attraction between the positive and negative charges. The stronger attraction pulls the electrons physically closer to the nucleus. The result is a smaller distance between the nucleus and the H.O.E.L.- thus a _____ atomic radius.

The attraction that the positive nucleus has for the negatively-charged electrons is called _____ _____ _____.

AR Top to Bottom in a Group: _____

Why?

* As you go from top to bottom in a group, there are more and more _____ _____ _____.



Metallic Character (MC):

(easier to _____ = _____ metallic = more reactive _____)

Which metal loses its valence electron(s) most easily? _____
Why?

Nonmetallic Character (NMC):

(easier to _____ = _____ nonmetallic = more reactive _____)

Which nonmetal gains electron(s) most easily? _____
Why?

PERIODIC TRENDS (multiple choice w/o using Periodic Table) WORKSHEET

1.) Which element is most metallic? Group 14, Period ___
(A) 2 (B) 3 (C) 4 (D) 5

2.) Which element is most nonmetallic? Group 16, Period ___
(A) 2 (B) 3 (C) 4 (D) 5

3.) Which has the largest atomic radius? Group ___ , Period 2
(A)1 (B) 13 (C) 15 (D) 17

4.) Which has the highest ionization energy? Group 2, Period ___
(A) 3 (B) 4 (C) 5 (D) 6

5.) Which has the most metallic properties? Group ___ , Period 5
(A) 13 (B) 14 (C) 15 (D) 16

6.) Which has the greatest electron affinity?
(A) Grp 16, Pd 4 (B) Grp 16, Pd 5 (C) Grp 17, Pd 5 (D) Grp 17, Pd 4

7.) Which has the smallest atomic radius? Group 15, Period ___
(A) 2 (B) 3 (C) 4 (D) 5

8.) Which has the lowest electron affinity? Group ___, Period 3
(A) 13 (B) 15 (C) 17 (D) 18

9.) Which has the lowest ionization energy? Group 1, Period ___
(A) 2 (B) 3 (C) 4 (D) 5

10.) Which has the most metallic properties?
(A) Grp 15, Pd 5 (B) Grp 16, Pd 5 (C) Grp 15, Pd 6 (D) Grp 16, Pd 6

11.) Which would most easily lose its valence electrons?
(A) Grp 1, Pd 3 (B) Grp 14, Pd 2 (C) Grp 17, Pd 3 (D) Grp 18, Pd 2

12.) Which would most easily gain electrons?
(A) Grp 13, Pd 3 (B) Grp 14, Pd 2 (C) Grp 15, Pd 2 (D) Grp 17, Pd 3

13.) Which has an octet of electrons in its outermost energy level?
(A) Grp 13, Pd 3 (B) Grp 14, Pd 2 (C) Grp 18, Pd 2 (D) Grp 17 , Pd 5

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14.) Which has chemical properties most similar to [Ar] $4s^1$? Group ___, Period 3
(A) 1 (B) 2 (C) 13 (D) 14

15.) Which is most reactive? Group ___, Period 2
(A) 14 (B) 15 (C) 17 (D) 18

16.) Which is most reactive?
(A) Grp 13, Pd 2 (B) Grp 1, Pd 5 (C) Grp 2, Pd 5 (D) Grp 13, Pd 5

17.) Which has chemical properties most similar to [Ne] $3s^2\ 3p^5$?
(A) Grp 16, Pd 3 (B) Grp 18, Pd 3 (C) Grp 17, Pd 4 (D) Grp 18, Pd 2

18.) Which would never be found in the free state?
(A) Grp 1, Pd 4 (B) Grp 13, Pd 3 (C) Grp 15, Pd 3 (D) Grp 14, Pd 4

19.) Which is the least reactive gas? Group ___, Period 2
(A) 16 (B) 15 (C) 17 (D) 18

20.) Which is the most reactive gas? Group ___, Period 2
(A) 16 (B) 15 (C) 17 (D) 18

21.) Which would never be in a compound?
(A) Grp 1, Pd 1 (B) Grp 18, Pd 1 (C) Grp 13, Pd 2 (D) Grp 1, Pd 2

22.) Which would be found in the "d" block of elements?
(A) Grp 1, Pd 3 (B) Grp 11, Pd 4 (C) Grp 17, Pd 5 (D) Grp 14, Pd 2

UNIT 5 REVIEW WORKSHEET

Matching

Alkali metals	Alkaline earth metals	Noble gases
Transition metals	Halogens	

1. The _____ have a single electron in the highest energy level.
2. The _____ achieve the electron configurations of noble gases by losing two electrons.
3. The _____ achieve the electron configuration of noble gases by gaining one electron.
4. The _____ have full s and p orbitals in the highest occupied energy levels.
5. The _____ are stable and unreactive.
6. The _____ are highly reactive nonmetals and readily form compounds with metals.
7. The _____ are metals that are more reactive than the transition elements but less reactive than the alkali metals.

Atomic radius	Decrease	Noble gases
Electronegativity	Ionization energy	Increase
Metals	Noble gas configuration	Shielding effect
Nonmetals	Metalloid	

8. _____ is the energy required to remove an electron from an atom.
9. The attraction of an atom for an additional electron is called _____.

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10. When they have a(n) _____, ions have a stable, filled outer electron level.
11. Along with the increased distance of the outer electrons from the nucleus, the _____ of the inner electrons causes ionization energy to decrease going down a column of the Periodic Table.
12. A low ionization energy is characteristic of a(n)_____.
13. Ionization energies tend to _____ across periods of the periodic table.
14. An element with an extremely high ionization energy is classified as a(n) _____.
15. The distance from the nucleus to the highest occupied energy level is known as _____.
16. The _____ do not have measured electronegativities since they do not commonly form compounds.

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Predict the oxidation number based on the electron configuration shown.

17. $1s^2\ 2s^2\ 2p^6\ 3s^2$
18. $1s^2\ 2s^2\ 2p^6\ 3s^1$
19. $1s^2\ 2s^2\ 2p^6$
20. $1s^2\ 2s^2\ 2p^5$
21. $1s^2\ 2s^2\ 2p^1$

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Choose the location of the element with the higher ionization energy.

22. Period 2, Group 14 or Period 3, Group 13
23. Period 4, Group 2 or Period 5, Group 2
24. Period 1, Group 18 or Period 2, Group 1
25. Period 3, Group 17 or Period 3, Group 18
26. Period 3, Group 17 or Period 2, Group 17
27. Period 3, Group 17 or Period 3, Group 16

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Arrange the element locations in order of increasing electronegativity.

28. (A) Period 4, Group 13 (B) Period 3, Group 13 (C) Period 5, Group 13
29. (A) Period 4, Group 2 (B) Period 4, Group 16 (C) Period 4, Group 15
30. (A) Period 2, Group 16 (B) Period 2, Group 17 (C) Period 3, Group 16
31. (A) Period 3, Group 15 (B) Period 2, Group 16 (C) Period 4, Group 14

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Choose the location of the element with the larger atomic radius.

32. Period 2, Group 1 or Period 4, Group 1
33. Period 4, Group 2 or Period 4, Group 16
34. Period 4, Group 13 or Period 2, Group 13
35. Period 2, Group 16 or Period 2, Group 14
36. Period 3, Group 17 or Period 4, Group 17
37. Period 2, Group 2 or Period 6, Group 2
38. Period 3, Group 14 or Period 3, Group 16

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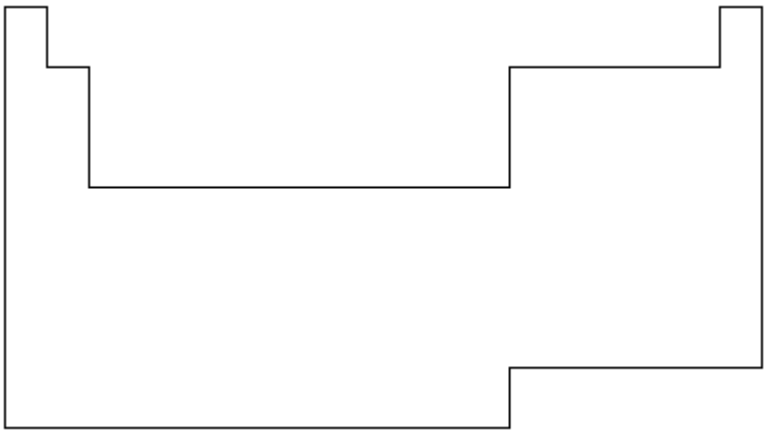
Multiple Choice

39. In any ___, the number of electrons between the nucleus and the outer energy level is the same.
    (A) period (B) group (C) both (D) neither

40. In a ____, electron affinity values decreases as atomic number increases.
    (A) period (B) group (C) both (D) neither

41. The halogens are considered a ____.
    (A) period (B) group (C) both (D) neither



On the Periodic Table, show the location of the element with the…

42. lowest ionization energy
43. most nonmetallic properties
44. smallest atomic radius
45. highest electronegativity
46. largest atomic radius
47. highest ionization energy
48. most metallic properties

Answer the following questions.

49. Explain the relationship between the relative size of an ion to its atom and the charge on the ion.

50. Explain why noble gases are inert and do not form ions.

51. Why do elements in the same family generally have similar properties?

52. If element X is a very reactive nonmetal, then the element with atomic number X + 1 should have what properties?

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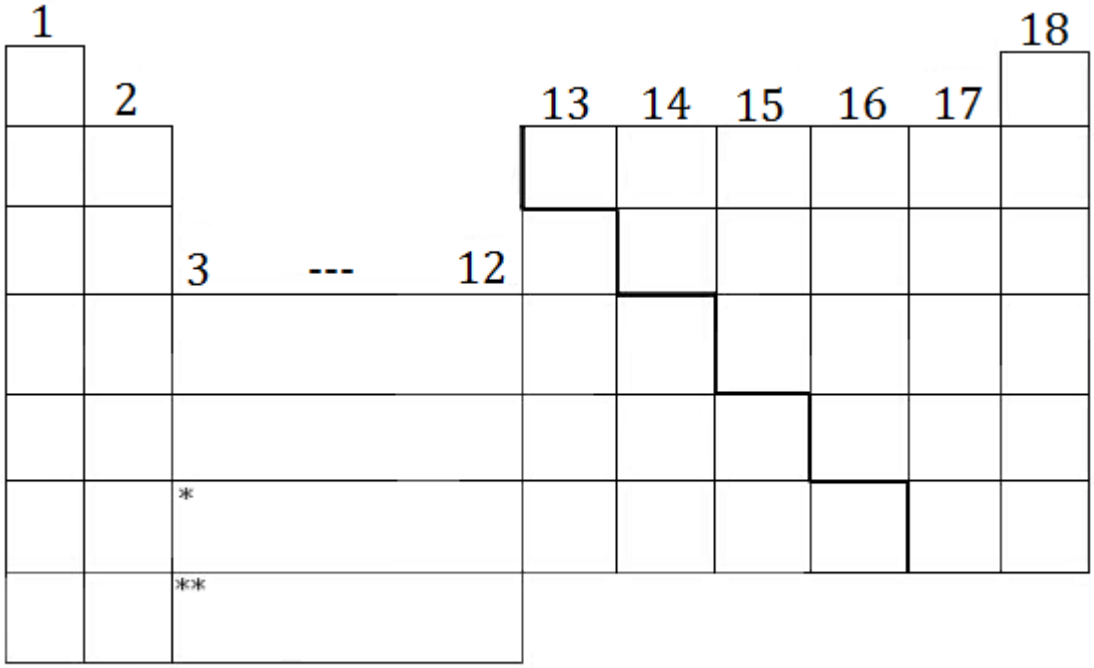
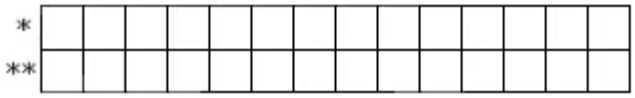
UNIT 5 REVIEW/SUMMARY WORKSHEET





1.) Color the "s" block area red.
2.) Color the "p" block area blue.
3.) Color the "d" block area green.
4.) Color the "f" block area orange.
5.) Draw an X in the boxes that represent the unreactive elements
6.) Draw a diagonal line (from upper left to lower right) in the area that represents the very reactive nonmetals.
7.) Draw a diagonal line (from upper right to lower left) in the area that represents the very reactive metals.
8.) Draw a purple capital letter "R" with a circle around it at the location that represents the element with the largest
atomic radius.
9.) Draw a blue capital letter "I" with a diamond around it at the location that represents the element with the highest ionization energy.
10.) Draw a dark green capital letter "E" with a triangle around it at the location that represents the element with
the highest electronegativity/electron affinity.
11.) Outline in black the boxes where metalloids with more nonmetallic properties are located.
12.) Outline in red the boxes where metalloids with more metallic properties are located.
13.) Draw a star in the location that represents the most metallic element (or most reactive metal).
14.) Draw a heart in the location that represents the most nonmetallic element (or most reactive nonmetal).

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