The Periodic Table: Chapter Problems

Periodic Table
Class Work

1. As you move from left to right across the periodic table, how does atomic number change?
2. What element is located in period 3, group 13?
3. What is atomic number of the element in period 6, group 2?
4. Will the element located at period 6, group 3 have a larger or smaller atomic number than the element in #3?

Homework

5. What is the group # and period # for oxygen (O)?
6. How will the atomic number of the element located directly to the left of oxygen compare to oxygen's atomic number?
7. What is the atomic number of the element in period 5, group 11?
8. What is the atomic number of the element in period 2, group 18?

Special Groups
Class Work

9. To which group on the periodic table does copper belong?
10. To which group on the periodic table does krypton belong?
11. A mystery element is in the same period as gallium. It has a smaller atomic number than gallium and it is highly reactive. What is the mystery element?
12. Two elements are studied. One has atomic number X and one has atomic number X-2. It is known that element X is an alkaline earth metal. To what group does the X-2 element belong?

Homework

13. To which group on the periodic table does magnesium belong?
14. To which group on the periodic table does bromine belong?
15. A mystery metal is in the same period as sulfur. It has a larger atomic number than sulfur and is nonreactive. What is the mystery element?
16. Two elements are studied. One has atomic number X and one has atomic number X+2. It is known that element X is a halogen. To what group does the X+2 element belong?

Periodic Families
Class Work

17. Lithium, sodium and potassium are in the same group of the periodic table. What do they all have in common?
18. The electron configuration ending s^2p^5 belongs in which group of the periodic table?
19. Compare sodium and magnesium. Which is more reactive? Explain your answer using electron configurations as evidence.

Homework

20. Alkaline earth metals are located in group 2 of the periodic table. What do these elements all have in common?
21. The electron configuration s^1 belongs in which group of the periodic table?
22. Compare chlorine and argon. Which is more reactive? Explain your answer using electron configuration as evidence.

Shorthand

Class Work

23. What is the shorthand electron configuration for silicon?
24. What is the shorthand electron configuration for iodine?
25. The electron configuration [Ar] $4s^2$ refers to which element?
26. The electron configuration [Kr] $5s^2$ $4d^2$ belongs to which group of the periodic table?

Homework

27. What is the shorthand electron configuration for iron?
28. What is the shorthand electron configuration for antimony?
29. The electron configuration [Kr] $5s^2$ $4d^8$ refers to which element?
30. The electron configuration [Ne] $3s^1$ belongs to which group of the periodic table?

Stability and Exceptions

Class Work

31. Rank the electron configurations below from most stable to least stable:
    [He] $2s^2$ [Ar] $4s^2$ $3d^5$ [Ne] $3s^2$ $3p^6$
32. Why is a noble gas more stable than an alkaline earth metal?
33. What are the expected and actual electron configurations for molybdenum?
34. What are the expected and actual electron configurations for copper?

Homework

35. Rank the electron configurations below from most stable to least stable:
    [Kr] $5s^2$ $4d^9$ [Ar] $4s^2$ $3d^{10}$ $4p^6$ [Ne] $3s^2$
36. What are the expected and actual electron configurations for chromium?
37. What are the expected and actual electron configurations for silver?
38. Why are the expected and actual electron configurations different for the above elements?

Effective Nuclear Charge

Class Work

39. What is the shielding constant for phosphorus (P)?
40. What is the effective nuclear charge on electrons in the outer most shell of phosphorus?
41. How do the shielding constants for the following elements compare?
    a. Boron and carbon
    b. Neon and sodium
42. The atomic numbers of the elements in 41a and 41b differ from each other by only one. Why is there such a large difference in the answers for 41a and 41b?

Homework

43. What is the shielding constant for calcium (Ca)?
44. What is the effective nuclear charge on electrons in the outer most shell of calcium?
45. Two elements are studied: X and Y. Both elements are in the same period of the periodic table. X has a larger effective nuclear charge than Y. How do the shielding constants of X and Y compare?
46. Two elements are studied: X and Y. Element X has a large effective nuclear charge than Y. Explain what this means.

Coulomb's Law

Class Work

47. What is Z_{eff} for aluminum and silicon?
48. What is the equation that calculates Coulomb's Law for aluminum?
49. What is the equation that calculates Coulomb's Law for silicon?
50. Based on the equations for aluminum and silicon, which element would be larger?

Homework

51. What is Z_{eff} for lithium and sodium?
52. What is the equation that calculates Coulomb's Law for lithium?
53. What is the equation that calculates Coulomb's Law for sodium?
54. Based on the equations for lithium and sodium, which element would be larger?

Atomic Radii

Class Work

55. As Z_{eff} increases in elements of the same period, how does atomic radius change? Explain why.
56. Put the following elements in order of increasing atomic size:
    Ca, Rb, K, O, Al, As
57. Put the following elements in order of decreasing atomic size:
    Ga, Fr, Br, Si, Na, N
58. Put the following elements in order of decreasing atomic size:
    Po, Sn, Fr, Rb, Cl, Li

Homework

59. As you move down a group, the Z_{eff} remains the same. How does atomic radius change? Explain why.
60. Put the following elements in order of increasing atomic size:
    Ar, Ca, Mg, O, N, At
61. Put the following elements in order of decreasing atomic size:
    B, P, I, Sb, Be, Pb
62. Put the following elements in order of decreasing atomic size:
    N, As, Kr, Fr, S, O

Ionization Energy
Class Work

63. You are working with two unknown elements in the lab: X and Y. You know that X has a lower ionization energy than Y. Which atom has a smaller atomic radius?
64. Put the following elements in order of increasing first ionization energy:
    Ca, Rb, K, O, Al, As
65. Put the following elements in order of increasing first ionization energy:
    Ga, Fr, Br, Si, Na, N
66. Put the following elements in order of increasing first ionization energy:
    Po, Sn, Fr, Rb, Cl, Li

Homework

67. You are working with two unknown elements in the lab: X and Y. You know that X has a larger Z_{eff} than Y. Which element has a higher ionization energy?
68. Put the following elements in order of increasing first ionization energy:
    Ar, Ca, Mg, O, N, At
69. Put the following elements in order of increasing first ionization energy:
    B, P, I, Sb, Be, Pb
70. Put the following elements in order of increasing first ionization energy:
    N, As, Kr, Fr, S, O

Electronegativity
Class Work

71. You are working with two unknown elements in the lab: X and Y. You know that X has a lower ionization energy than Y. Which element has the higher electronegativity?
72. Put the following elements in order of increasing electronegativity:
    Ca, Rb, K, O, Al, As
73. Put the following elements in order of decreasing electronegativity:
    Ga, Fr, Br, Si, Na, N
74. Put the following elements in order of decreasing electronegativity:
    Po, Sn, Fr, Rb, Cl, Li

Homework

75. You are working with two unknown elements in the lab: X and Y. You know that X has a higher electronegativity than Y. Which element has a higher Z_{eff}?
76. Put the following elements in order of increasing electronegativity:
    Ar, Ca, Mg, O, N, At
77. Put the following elements in order of decreasing electronegativity:
    B, P, I, Sb, Be, Pb
78. Put the following elements in order of decreasing electronegativity:
    N, As, Kr, Fr, S, O

Metallic Character
Class Work

79. Put the following elements in order of increasing metallic character:
    P, Cs, Sn, F, Sr, Tl
80. Put the following elements in order of increasing metallic character:
    Ca, Rb, K, O, Al, As
81. Put the following elements in order of decreasing metallic character:
    Ga, Fr, Br, Si, Na, N
82. Put the following elements in order of decreasing metallic character:
    Po, Sn, Fr, Rb, Cl, Li

Homework

83. Put the following elements in order of increasing metallic character:
    Ra, F, Al, Ne, H, He
84. Put the following elements in order of increasing metallic character:
    Ar, Ca, Mg, O, N, At
85. Put the following elements in order of decreasing metallic character:
    B, P, I, Sb, Be, Pb
86. Put the following elements in order of decreasing metallic character:
    N, As, Kr, Fr, S, O

Free Response

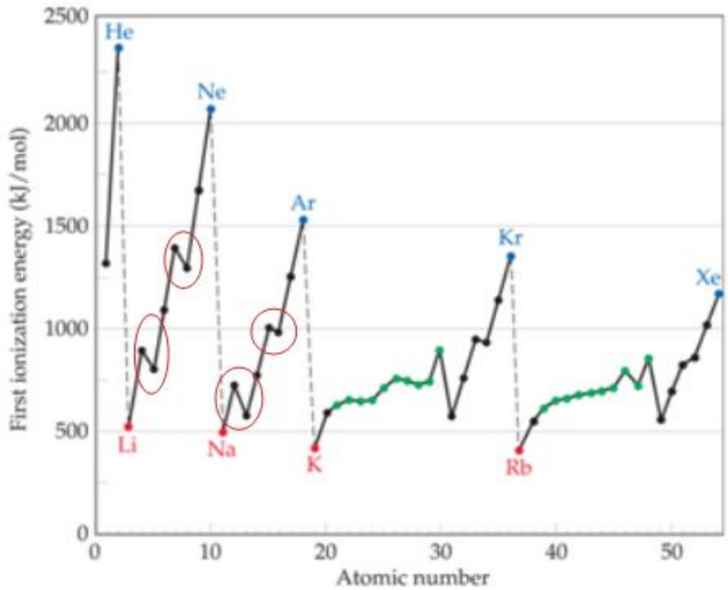
Key Terms

Effective Nuclear Charge	*Atomic Radius*	*Energy Level*
Shielding	*Principal Quantum Number*	*Electron configuration*
Coulomb's Law		*Valence electrons*

1. Suppose that a stable element with atomic number 119, symbol X, has been discovered.

    a. Write the shorthand electron configuration for X.
    b. Would X be a metal or a nonmetal? Explain in terms of electron configuration.
    c. On the basis of periodic trends, would X have the largest atomic radius in its group or would it have the smallest? Explain in terms of electronic structure.

2. The figure below shows trends in first ionization energy across a period.



a. What is the general trend in first ionization energy across a period?
b. Although boron is to the right of beryllium in period 2, it has a lower first ionization energy than beryllium. Explain why using key terms to justify your answer.
c. Why does oxygen have a lower first ionization energy than nitrogen?

3. Elements X, Y and Z are on the third row of the Periodic Table. The first four ionization energies for each are given below.

	IE_1	IE_2	IE_3	IE_4
X	498.8 kJ/mol	4562 kJ/mol	6910.3 kJ/mol	9543 kJ/mol
Y	577.5 kJ/mol	1816.7 kJ/mol	2744.8 kJ/mol	11577 kJ/mol
Z	737.7 kJ/mol	1450.7 kJ/mol	7732.7 kJ/mol	10542.5 kJ/mol

Which of the following: F, Ne, Al, Na, Mg could be element Z? Justify your answer using one or more of the key terms.

4. Some transition metals have actual electron configurations that differ from the expected electron configurations.
    a. Choose two examples of elements that have exceptional electron configurations and provide their actual shorthand electron configurations.

    b. Explain why their electron configurations are exceptions to the normal filling order using the key terms above.

5. 99% of the human body is made up of just 11 elements. Nitrogen makes up about 3.2% of the body and phosphorous around 1%.
    a. Write the shorthand electron configurations for nitrogen and phosphorous.
    b. Compare the atomic radii of phosphorous and nitrogen using one or more of the key terms in your response.
    c. Compare the electronegativity of nitrogen and phosphorous using one or more the key terms in your response.
    d. Which has a greater first ionization energy, nitrogen or phosphorous? Use the key terms to justify your response.

6. Sodium and chloride react to form a very common compound – table salt.
    a. Write the shorthand electron configurations for sodium and chloride.
    b. Compare the atomic radii of sodium and chloride using one or more of the key terms in your response.
    c. Compare the electronegativity of sodium and chloride using one or more of the key terms in your response.
    d. Which has a greater first ionization energy, sodium or chloride? Use the key terms to justify your response.

<u>Answers</u>

1. Atomic number increases from left to right.
2. Aluminum
3. 56
4. Since it is to the right of barium, it will have a higher atomic number.
5. Period 2, group 16
6. Since it is to the left of oxygen, it will have a lower atomic number than oxygen.
7. 47
8. 10
9. Transition metals
10. Noble gas
11. Potassium (K)
12. Noble gas
13. Alkaline earth metal
14. Halogen
15. Argon (Ar)
16. Alkaline earth metals
17. They are all alkali metals with the ending electron configuration of s^1.
18. Halogens
19. Sodium has an ending electron configuration of s^1 while magnesium has an electron configuration of s^2. Sodium is much more reactive than magnesium because sodium needs just one more electron to fill its outer shell while magnesium already has a full outer shell.

20. Alkaline earth metals all have the ending electron configuration of s^2.
21. Alkali metals
22. Chlorine has an ending electron configuration of s^2p^5 while argon has an ending electron configuration of s^2p^6. Argon is nonreactive because it has a full outer shell. Argon, however, needs just one more electron to fill its outer shell and so is highly reactive.
23. [Ne] $3s^2$ $3p^2$
24. [Kr] $5s^2$ $4d^{10}$ $5p^5$
25. Calcium
26. Transition metals (zirconium)
27. [Ar] $4s^2$ $3d^6$
28. [Kr] $5s^2$ $4d^{10}$ $5p^3$
29. Palladium
30. Alkali metals (sodium)
31. [Ne] $3s^2$ $3p^6$ [He] $2s^2$ [Ar] $4s^2$ $3d^5$
32. A noble gas has a full energy level while an alkali metal has a full subshell.
33. Expected: [Kr] $5s^2$ $4d^4$
    Actual: [Kr] $5s^1$ $4d^5$
34. Expected: [Ar] $4s^2$ $3d^9$
    Actual: [Ar] $4s^1$ $3d^{10}$
35. [Ar] $4s^2$ $3d^{10}$ $4p^6$ [Ne] $3s^2$ [Kr] $5s^2$ $4d^9$
36. Expected: [Ar] $4s^2$ $3d^4$
    Actual: [Ar] $4s^1$ $3d^5$
37. Expected: [Kr] $5s^2$ $4d^9$
    Actual: [Kr] $5s^1$ $4d^{10}$
38. The s and d orbitals are very close to each other. By moving one electron from an s orbital to a d orbital, these elements can have the stability of a half full subshell.
39. 10
40. 5
41.
    a. Boron: 2; Carbon: 2
    b. Neon: 2; Sodium: 10
42. The outer most electrons in boron and carbon are located in the same energy level, so they have the same number of shielding electrons. The outer most electrons of sodium are located in a higher energy level than neon, so sodium has a higher shielding constant.
43. 18
44. 2
45. If X and Y are in the same period, then they are in the same energy level. This means that they both have the same number of shielding electrons. Their shielding constants will be the same.
46. If X has a larger effective nuclear charge, the force between the protons and valence electrons is larger than in element Y.
47. Aluminum: +3
    Silicon: +4
48. $F = k(3e)^2 / r^2$

49. $F = k(4e)^2 / r^2$
50. Aluminum and silicon have the same initial radius because their valence electrons are in the same energy shell. The Z_{eff} on silicon is larger. This makes the force larger. The nucleus pulls tighter on silicon's valence electrons than on aluminum's valence electrons. Aluminum has a larger radius than silicon.
51. Lithium: +1
    Sodium: +1
52. $F = ke^2 / r^2$
53. $F = ke^2 / r^2$
54. Lithium and sodium have the same Zeff. However, sodium has electrons in a higher energy level than lithium. This makes the radius of sodium larger than lithium. The resulting force on sodium is smaller than the force on lithium. Sodium has a larger radius than lithium.
55. As Z_{eff} increases, the force between the nucleus and the valence electrons increases. The electrons are pulled closer to the nucleus and atomic radius decreases.
56. O, Al, As, Ca, K, Rb
57. Fr, Na, Ga, Si, Br, N
58. Fr, Rb, Po, Sn, Li, Cl
59. Moving down a group, Z_{eff} remains the same but the energy level increases. Valence electrons are located farther from the nucleus. The increased distance decreases the force between the nucleus and valence electrons and atomic radius increases.
60. O, N, Ar, Mg, Ca, At
61. Pb, Sb, I, Be, P, B
62. Fr, As, Kr, S, N, O
63. As atoms get smaller, force increases causing ionization energy to increase. If Y has a higher ionization energy then it has a smaller radius than X.
64. Rb, K, Ca, As, Al, O
65. Fr, Na, Ga, Si, N, Br
66. Fr, Rb, Li, Po, Sn, Cl
67. As Z_{eff} increases, force increases causing ionization energy to increase. If X has a higher Z_{eff}, then it also has a higher ionization energy.
68. Ca, Mg, At, O, N, Ar
69. Pb, Be, Sb, B, P, I
70. Fr, As, S, O, N, Kr
71. Ionization energy increases as electronegativity increases. If Y has a higher ionization energy than X, then it also has a higher electronegativity.
72. Rb, K, Ca, Al, As, O
73. N, Br, Si, Ga, Na, Fr
74. Cl, Po, Sn, Li, Rb, Fr
75. As Z_{eff} increases, the force increases causing electronegativity to increase. If X has a higher electronegativity, then it also has a higher Z_{eff}.
76. Ca, Mg, Ar, At, N, O
77. I, P, B, Sb, Pb, Be
78. O, N, S, As, Kr, Fr
79. F, P, Sn, Tl, Sr, Cs
80. O, As, Al, Ca, K, Rb

81. Fr, Na, Ga, Si, Br, N
82. Fr, Rb, Li, Po, Sn, Cl
83. He, Ne, F, H, Al, Ra
84. Ar, O, N, At, Mg, Ca
85. Br, Pb, Sb, B, P, I
86. Fr, As, S, N, O, Kr

Free Response

1.
    a. $8s^1$
    b. It would be a metal (specifically an alkali metal) because it ends with the configuration s^1 and has a loosely held outer electron.
    c. It would have the largest atomic radius in its group because its valence electron is in a higher energy level.

2.
    a. In general, first ionization energy increases across a period.
    b. Boron's valence electron configuration is ns^2p^1 and beryllium's is ns^2. Boron has lower first ionization than beryllium because the electrons in the outer s orbital create a shielding effect that reduces the Coulombic attraction between its nucleus and the one outer electron in the p orbital, making that outer electron easier to remove.
    c. Nitrogen has a valence electron configuration of ns^2p^3 and oxygen has a valence electron configuration of ns^2p^4. Since each nitrogen's outer p orbitals has one electron, it is more stable, and, therefore, requires more energy to remove an electron than oxygen.

3. There is a big jump in energy between IE_2 and IE_3, which indicates a transition from removing valence electrons that are more loosely held due to shielding of inner electrons to removing inner electrons. Magnesium is the element in period 3 that has the valence electron configuration ns^2, so it must be element Z.

4.
    a. There are two main exceptions to electron configuration: chromium and copper.
    Chromium expected: $[Ar]4s^23d^4$
    Chromium actual: $[Ar]4s^13d^5$
    Copper expected: $[Ar]4s^23d^9$
    Copper actual: $[Ar]4s^13d^{10}$

    b. In these cases, a completely full or half full d sub-level is more stable than a partially filled d sub-level, so an electron from the 4s orbital is excited and rises to a 3d orbital.

5.

a. Nitrogen: [He]$3s^2 3p^3$ and Phosphorus: [Ne]$4s^2 4p^3$
b. Nitrogen has a smaller atomic radius than phosphorous because it has a lower principal quantum number than phosphorus. Phosphorous has a higher principal quantum number than so the electrons in the outer energy level are farther away from the nucleus.
c. Nitrogen is more electronegative than phosphorous. Although it has the same number of valence electrons as phosphorus, it has a smaller atomic radius. According to Coulomb's Law, electric force is inversely proportional to r^2 so the smaller the radius, the larger the electric force of attraction between the outer electrons and the nucleus.
d. Nitrogen has a greater first ionization energy then phosphorus, since the Coulombic force of attraction between the outer electrons and nucleus is greater due to a smaller radius, it takes more energy to remove an electron from nitrogen than it does from phosphorus.

6.

a. Sodium: [Ne]$3s^1$; Chlorine: [Ne]$3s^2 3p^5$
b. Sodium has a larger atomic radius than chlorine because it has a lower effective nuclear charge than chlorine. Although both elements have the same principal quantum number, sodium only has one outer electron and the force of attraction between the nucleus and outer electrons is less than the Coulombic attraction between chlorine's outer electrons and its nucleus.
c. Chlorine is much more electronegative than sodium because it has a higher effective nuclear charge.
d. Chlorine has a much greater first ionization energy than sodium because chlorine's nucleus has a higher positive charge and it has more outer electrons; it requires much more energy to remove an electron from chlorine than it does to remove the one outer electron from sodium.