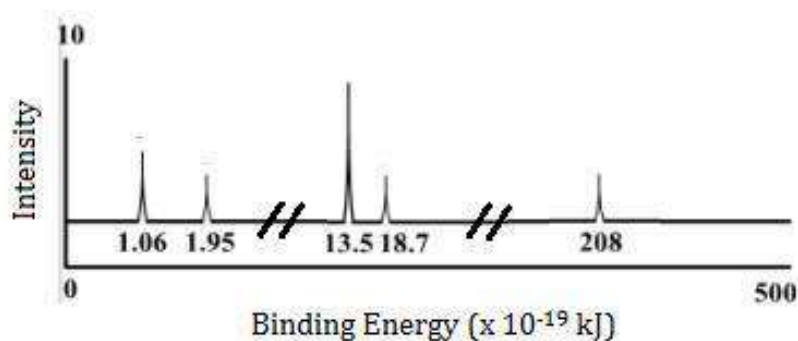


Use the PES spectrum of Phosphorus below to answer questions 1-3.



1. Which peak corresponds to the 1s orbital?
(A) 1.06
(B) 1.95
(C) 13.5
(D) 18.7
(E) 208
2. Which peak corresponds to the valence 3p orbital?
(A) 1.06
(B) 1.95
(C) 13.5
(D) 18.7
(E) 208
3. What would you predict for the intensity and binding energy for the 3p orbital for that of sulfur?
(A) Higher intensity, higher binding energy
(B) Higher intensity, lower binding energy
(C) Lower intensity, higher binding energy
(D) Lower intensity, lower binding energy

4. Which of the following provides the best explanation for the many chemical similarities between sodium and potassium?
- (A) Both have atomic masses between 20 and 40 amu
 - (B) Both have an ns^1 electron configuration for their highest occupied energy level
 - (C) Both are main group elements
 - (D) Both are solids at room temperature and pressure
 - (E) The atomic numbers of the two elements differ by less than ten
5. Electrons from which valence shell are removed first when transition metals ionize?
- (A) s
 - (B) p
 - (C) d
 - (D) f
6. The +3 charge more common than +2 charge for transition metals in the Fe group. What accounts for this trend?
- (A) Decreased ionic radius
 - (B) Increased electronegativity
 - (C) Increased stability of "d" orbitals
 - (D) Decreased stability of "d" orbitals
 - (E) Increased shielding
7. Elements with this valence shell electron configuration will commonly form ions with a -2 charge.
- (A) ns^2
 - (B) ns^2np^1
 - (C) ns^2np^2
 - (D) ns^2np^4
 - (E) ns^2np^6
8. An atom with the electron configuration of $[Ar]4s^23d^3$ would be in the same group as ____ and have a likely charge of ____.
- (A) Cr, +6
 - (B) Ta, +3
 - (C) Nb, +5
 - (D) Y, +2
 - V, +4
9. What would be the expected ionic charges for Sn?

- (A) +1 and +2
- (B) +1 and +3
- (C) +2 and +3
- (D) +3 and +4
- (E) +2 and +4

10. Which gives the correct orbital and order for the electrons lost in the formation of the following cations?

- (A) Ca^{2+} 4p
- (B) Pd^{4+} 5s then 5p
- (C) Co^{3+} 4s then 3d
- (D) Mn^{7+} 3d then 4s
- (E) Zn^{2+} 3d

11. Atoms on the left side of the chart tend to form positive ions because...

- (A) Their principal energy level is almost empty
- (B) Their principal energy level is almost full
- (C) Their atomic number is less than other elements in that period
- (D) Both B and C
- (E) A, B, and C

Atomic and Ionic Radii

12. Atomic radius generally increases as we move

- (A) down a group and from right to left across a period
- (B) up a group and from left to right across a period
- (C) down a group and from left to right across a period
- (D) up a group and from right to left across a period
- (E) down a group; the period position has no effect

13. The effective nuclear charge of an atom is primarily affected by

- (A) inner electrons
- (B) outer electrons
- (C) nuclear charge
- (D) electron distribution
- (E) orbital radial probability

14. Electrons in the 1s subshell are much closer to the nucleus in Ar than in He due to the larger _____ in Ar.

- (A) nuclear charge
- (B) paramagnetism
- (C) diamagnetism
- (D) Hund's rule
- (E) azimuthal quantum number

15. Of the following, which gives the correct order for atomic radius for Mg, Na, P, Si and Ar?

- (A) $\text{Mg} > \text{Na} > \text{P} > \text{Si} > \text{Ar}$
- (B) $\text{Ar} > \text{Si} > \text{P} > \text{Na} > \text{Mg}$
- (C) $\text{Si} > \text{P} > \text{Ar} > \text{Na} > \text{Mg}$
- (D) $\text{Ar} > \text{P} > \text{Si} > \text{Mg} > \text{Na}$
- (E) $\text{Na} > \text{Mg} > \text{Si} > \text{P} > \text{Ar}$

16. Which of the following is expected to have the largest radius?

- (A) P^{3-}
- (B) S^{2-}
- (C) Cl^-
- (D) Ar
- (E) K^+

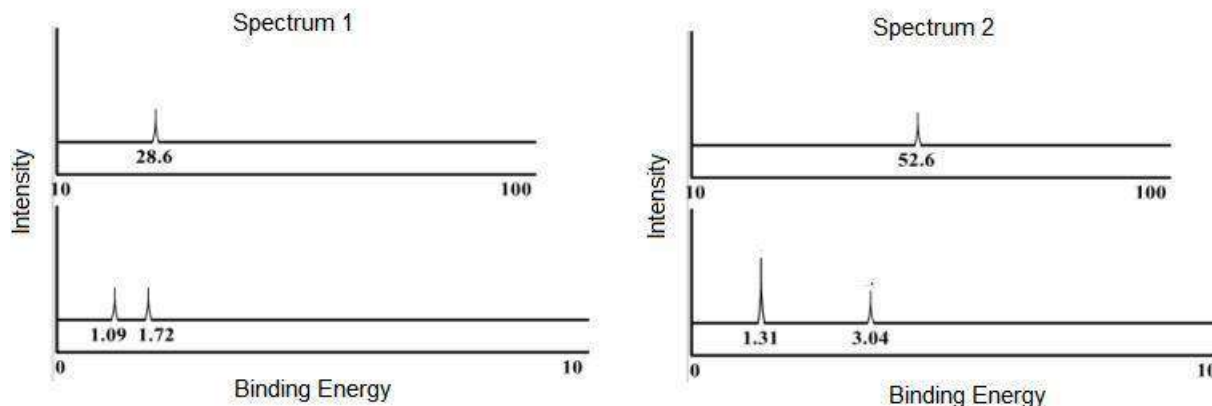
17. Which of the following is expected to have the smallest radius?

- (A) S^{2-}
- (B) Ca^{2+}
- (C) Cl^-
- (D) K^+
- (E) P^{3-}

18. Which of the following pairs correctly shows the proper relationship between the two atoms/ions in terms of atomic/ionic radii?

- (A) $\text{Na} < \text{Na}^+$
- (B) $\text{Cl} > \text{Cl}^-$
- (C) $\text{Ti} < \text{Zn}$
- (D) $\text{N}^{3-} > \text{N}$
- (E) $\text{O} > \text{S}$

Use the PES spectra below to answer questions 19-21. The spectra represent elements found in the same period on the periodic table.



19. Which of these two elements would have the smaller atomic radii?

- (A) Spectrum 1
- (B) Spectrum 2
- (C) They are both the same
- (D) There is not enough information

20. What are the identities of the elements represented by each spectrum respectively?

- (A) Fluorine and nitrogen
- (B) Carbon and beryllium
- (C) Carbon and oxygen
- (D) Sodium and carbon
- (E) Fluorine and carbon

21. Which of the following is **not** true regarding the spectra above?

- (A) The 2s peak at 1.72 in spectrum 1 has a lower binding energy than the 2s peak at 3.04 in spectrum 2 due to a decreased effective nuclear charge.
- (B) There are 4 valence electrons represented in spectrum 1
- (C) The valence electrons in spectrum 1 experience more shielding than those represented in spectrum 2.
- (D) The peak at 52.6 in spectrum 2 is held closer to the nucleus than the peak at 28.6 in spectrum 1.
- (E) The 1s peak in the PES spectrum of Li would have a lower binding energy than both 1s peaks from spectrum 1 and spectrum 2.

Ionization Energy

22. The first five ionization energies, in kJ/mol, for a particular element are shown below.

I ₁	I ₂	I ₃	I ₄	I ₅
786	1577	3232	4356	16,091

The element is likely to form an ionic compound in which the charge is:

- (A) 1+
- (B) 2+
- (C) 3+
- (D) 4+
- (E) 5+

23. Of the choices below, which gives the order for first ionization energies?

- (A) Kr > Se > Br > Ga > Ge
- (B) Kr > Br > Se > Ge > Ga
- (C) Ga > Br > Ge > Kr > Se
- (D) Ga > Ge > Se > Br > Kr
- (E) Br > Se > Ga > Kr > Ge

24. An element having which of the following electronic configurations would have the greatest ionization energy?

- (A) [He] 2s² 2p³
- (B) [He] 2s² 2p⁵
- (C) [Ne] 3s² 3p³
- (D) [Ne] 3s² 3p⁵

25. The electron configuration belonging to the atom with the highest second ionization energy is

- (A) 1s²2s²2p⁶3s¹
- (B) 1s²2s²2p⁶3s²
- (C) 1s²2s²2p⁶3s²3p¹
- (D) 1s²2s²2p⁶3s²3p⁴
- (E) 1s²2s²2p⁶3s²3p⁵

26. Sodium is much more apt to exist as a cation than is chlorine. This is because _____.

- (A) chlorine is a gas and sodium is a solid
- (B) chlorine has a greater ionic radius than sodium does

- (C) chlorine has a greater atomic radius than sodium
- (D) chlorine has a greater ionization energy than sodium does
- (E) chlorine is more metallic than sodium

Use the following responses to answer questions 27-29 below.

- (A) $\text{Cl} \rightarrow \text{Cl}^+ + \text{e}^-$
- (B) $\text{Cl}^+ + \text{e}^- \rightarrow \text{Cl}$
- (C) $\text{Cl}^+ \rightarrow \text{Cl}^{2+} + \text{e}^-$
- (D) $\text{Cl}^{2+} + \text{e}^- \rightarrow \text{Cl}^+$

27. ____ The process representing the first ionization energy of chlorine
28. ____ The process representing the second ionization energy of chlorine
29. ____ The process requiring the most energy

Use the following responses to answer questions 30-32. Use an answer once, more than once, or not at all.

- (A) Mg & Al
- (B) As & Se
- (C) Cl & F
- (D) Cr & Mo
- (E) Cu & Ag

30. __ __ Show a reversal in the trend for first ionization energy because of electron-electron repulsions.
31. ____ Show a reversal in the trend for first ionization energy because of shielding by full orbitals.
32. ____ Typically form ions with charges of +3 and +6

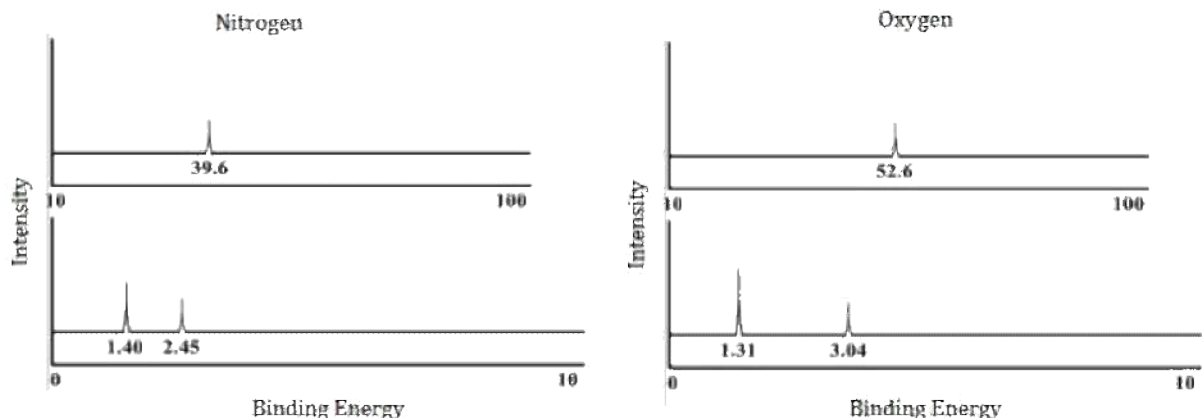
Use the following responses to answer questions 33-37. Use an answer once, more than once, or not at all.

- (A) Increased shielding of valence electrons
- (B) Effective nuclear charge
- (C) $\frac{1}{2}$ filled orbitals
- (D) Fully filled orbitals

33. ____ Why gold is a better conductor of electricity than copper.

34. ____ Why the atomic radius decreases as the atomic number increases from 11-17 across the periodic table.
35. ____ Why atomic radius increases down a group on the periodic table.
36. ____ Why gallium has an unexpected smaller radius than aluminum?
37. ____ Why magnesium has an unexpectedly higher ionization energy than aluminum.

Use the PES spectrum of Nitrogen and Oxygen below to answer questions 38-40.



38. Which of the following best explains why the 1s peak in the nitrogen spectrum has a lower binding energy compared to the 1s peak in the oxygen spectrum?

- (A) Nitrogen atoms have a 3 electrons in their p subshell.
- (B) Oxygen experience more electron-electron repulsions.
- (C) Nitrogen has a larger atomic radius.
- (D) Electrons in the p subshell of oxygen atoms provide more shielding than those in nitrogen atoms.
- (E) Nitrogen atoms have a smaller nuclear charge and coulombic attraction than oxygen atoms.

39. Given the fact that both elements are in the same period, what information from the PES spectra indicates that oxygen would have a smaller atomic radius?

- (A) The number of peaks
- (B) The intensity of the peaks
- (C) Binding energy of the peaks
- (D) A and C
- (E) B and C

40. What information from the PES spectra above best accounts for nitrogen having smaller 1^{st} ionization energy than oxygen?

- (A) The number of peaks
- (B) The intensity of the core electron peaks
- (C) The intensity of the valence electron
- (D) Binding energy of the valence electron peaks
- (E) A and C

41. What trend would you predict for the binding energies of valence electrons moving down a group on the periodic table?
- (A) Increasing binding energy due to increased shielding.
 - (B) Increasing binding energy due to decreased shielding.
 - (C) Decreasing binding energy due to increased shielding.
 - (D) Decreasing binding energy due to decreased shielding.
 - (E) Decreasing binding energy due to decreased nuclear charge.
42. Which of the following elements would require the shortest wavelength of light to ionize?
- (A) Ga
 - (B) B
 - (C) Al
 - (D) C
 - (E) N
43. What would be the necessary wavelength (nm) required to remove one valence electron from sodium's outermost "s" orbital? The first ionization energy for sodium is 495.8 kJ/mol.
- (A) 350 nm
 - (B) 242 nm
 - (C) 124 nm
 - (D) 824 nm
 - (E) 496 nm

Metallic Character

44. Between which two elements is the difference in metallic character the greatest?
- (A) Rb and O
 - (B) O and I
 - (C) Rb and I
 - (D) Li and O
 - (E) Li and Rb
45. The list that correctly indicates the order of metallic character is _____.
- (A) $B > N > C$
 - (B) $F > Cl > S$
 - (C) $Si > P > S$
 - (D) $P > S > Se$
 - (E) $Na > K > Rb$

46. Which of the following best describes why tin is considered a metal and carbon a nonmetal despite being in the same group?

- (A) Carbon has smaller atomic radius
- (B) The atomic mass of tin is nearly 10 times that of carbon.
- (C) Tin has less shielding of its valence electrons and a higher ionization energy.
- (D) Tin has more shielding of valence electrons and a lower ionization energy.
- (E) Tin has more shielding of its valence electrons and a higher ionization energy.

47. Which of the following orbitals of scandium would require the lowest frequency of light to ionize?

- (A) 2p
- (B) 3s
- (C) 3p
- (D) 4s
- (E) 3d

Electronegativity

48. Which of the following best describes the variation of electronegativity with respect to an element's position on the periodic table?

- (A) Increases across a period; increases down a group.
- (B) Increases across a period; decreases down a group.
- (C) Decreases across a period; increases down a group.
- (D) Decreases across a period; decreases down a group.

49. Which of the following groups of elements is arranged correctly in order of increasing electronegativity?

- (A) B < O < Al < F
- (B) Al < B < O < F
- (C) B < O < F < Al
- (D) F < O < B < Al

50. The ability of an atom in a molecule to attract electrons is best quantified by the _____.

- (A) paramagnetism

- (B) diamagnetism
- (C) electronegativity
- (D) electron change-to-mass ratio
- (E) first ionization potential

Use the following responses to answer for questions 51 - 53.

- (A) O
- (B) Fe
- (C) Rb
- (D) Mg
- (E) N

51. ____What is the most electronegative element of the above?
52. ____Which element readily forms cations with multiple oxidation states?
53. ____Which of the elements above has the smallest ionic radius for its most commonly found ion?
54. An element with a lower electronegativity value will likely have
- (A) Valence shell PES peaks with low binding energies
 - (B) A lower first ionization energy
 - (C) A low nuclear charge and a high amount of shielding
 - (D) A and B
 - (E) All of the above
55. Which of the following BEST explains why silicon has a lower electronegativity than chlorine?
- (A) Si has a lower nuclear charge and similar shielding than Cl
 - (B) Si has a higher nuclear charge and similar shielding of Cl
 - (C) Si has the equivalent nuclear charge and less shielding than Cl
 - (D) Si has the equivalent nuclear charge and more shielding than Cl
 - (E) None of these

Group Trends

56. Which is not true of nonmetals?

- (A) Most of their oxides are acidic
- (B) They are poor conductors of heat
- (C) They are poor conductors of electricity
- (D) Many are gases at room temperature
- (E) Most tend to lose electrons readily

57. All of the following statements concerning the characteristics of the halogens are true EXCEPT:

- (A) The first ionization energies (potentials) decrease as the atomic numbers of the halogens increase.
- (B) They only require one more electron to have full "s" and "p" orbitals.
- (C) Fluorine atoms have the smallest radii.
- (D) They readily form cations.
- (E) Fluorine is the most electronegative of the halogens.

58. Element M reacts with chlorine to form a compound with the formula MCl_2 . Element M is more reactive than magnesium and has a smaller radius than barium. This element is ___?

- (A) Sr
- (B) K
- (C) Na
- (D) Ra
- (E) Be

59. The only noble gas that does not have the $ns^2 np^6$ valence electron configuration is _____.

- (A) radon
- (B) neon
- (C) helium
- (D) krypton

(E) All noble gases have the $ns^2 np^6$ valence electron configuration.

60. An alkaline earth metal forms a compound with oxygen with the formula _____. (The symbol M represents any one of the alkaline earth metals.)

- (A) MO
- (B) M_2O
- (C) MO_2
- (D) M_2O_2
- (E) MO_3

61. Alkali metals tend to be more reactive than alkaline earth metals because _____.

- (A) alkali metals have lower densities
- (B) alkali metals have lower melting points
- (C) alkali metals have greater atomic radii
- (D) alkali metals are not more reactive than alkaline earth metals
- (E) alkali metals have lower ionization energies

62. The element in the periodic table that looks like a metal, is a poor thermal conductor, and acts as an electrical semiconductor is _____.

- (A) Sn
- (B) B
- (C) As
- (D) Si
- (E) Ge

63. Of the following metals, _____ exhibits multiple oxidation states.

- (A) Al
- (B) Cs
- (C) V
- (D) Ca
- (E) Na