Chemistry: Form TR4-9A Name _____

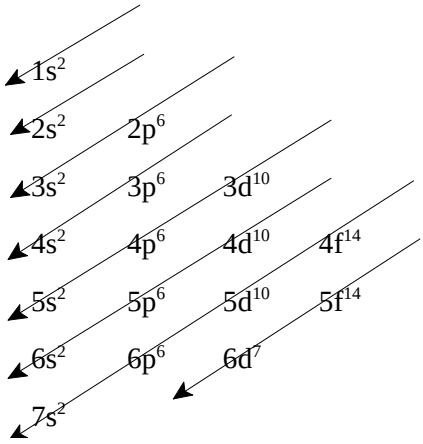
REVIEW Date _____ Period _____

Test Review № 4

Location of electrons. Electrons are in regions of the atom known as orbitals, which are found in subdivisions of the principal energy levels called sublevels. There are up to seven principal energy levels designated by a quantum number, *n*, from 1 to 7. The maximum number of sublevels in a principal energy level is *n*, but none of the existing elements use more than 4 sublevels even in principal energy levels 5–7. Sublevels are designated by the letters s, p, d, and f, in increasing order of energy. Orbitals are regions within a sublevel where electrons of a given energy are likely to be found. There are a maximum of 2 electrons in an orbital. The number of orbitals within a sublevel varies in a predictable pattern. The number of orbitals within a sublevel and the maximum number of electrons is as follows:

Sublevel	s	p	d	f
Number of orbitals	1	3	5	7
Maximum Number of Electrons	2	6	10	14

Rules describing the distribution of electrons. The number of electrons equals the atomic number. Electrons occupy orbitals in sequence beginning with those of the lowest energy. In a given sublevel, a second electron is not added to an orbital until each orbital in the sublevel contains one electron. No more than four orbitals are occupied in the outermost principal energy level



Element	Atomic Number	Sublevel structure	Orbital notation
boron	5	$1s^2 2s^2 2p^1$	↑↓ (1s) ↑↓ (2s) ↑ _ _ (2p)
oxygen	8	$1s^2 2s^2 2p^4$	↑↓ (1s) ↑↓ (2s) ↑↓ ↑ ↑ (2p)
argon	18	$1s^2 2s^2 2p^6 3s^2 3s^6$	↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑↓ (2p) ↑↓ (3s) ↑↓ ↑↓ ↑↓ (3p)

Maximum number of electrons. As a result of the way the rules are applied for determining the maximum number of electrons per orbital, orbitals per sublevel, and sublevels per principal energy level, for any given principal energy level, *n*, the maximum number of orbitals is n^2, and the maximum number of electrons is $2n^2$. An outer energy level, however, never has more than 8 electrons even if it has the room.

Principal Energy Level(*n*)		Number of Orbitals (n^2)	Electrons per Sublevel							Maximum Number of Electrons ($2n^2$)
			s	p	d	f	g	h	i	
			1	3	5	7	9	11	13	
Electrons in each Location	1	1	2	-	-	-	-	-	-	2
	2	4	2	6	-	-	-	-	-	8
	3	9	2	6	10	-	-	-	-	18
	4	16	2	6	10	14	-	-	-	32
	5	25	2	6	10	14	18	-	-	50
	6	36	2	6	10	14	18	22	-	72
	7	49	2	6	10	14	18	22	26	98

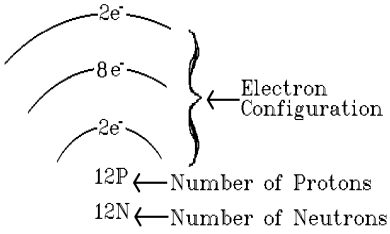
Continue ☞

Subatomic particles.

Type of Particle	Location	Mass	Relative Mass	Charge
Proton	Center	1.67×10^{-27}kg	1 amu	+1
Electron	Outside	9.11×10^{-31}kg	0 amu	-1
Neutron	Center	1.67×10^{-27}kg	1 amu	0

Neutrons. Neutrons were discovered by Sir James Chadwick in 1932. The existence of neutral particle was the only way to explain how atoms of an element could have different masses. Atoms of an element with different masses are called isotopes. The symbols for isotopes are written as follows: ${}^{A}_{Z}\mathrm{X}$, where X = element; A = atomic mass number (mass of isotope); and Z = atomic number (number of protons). The number of neutrons (N) is determined as follows: N = A – Z. The isotopes of hydrogen, for example, all have one proton, but different numbers of neutrons: ${}^{1}_{1}\mathrm{H}$ has no neutrons, ${}^{2}_{1}\mathrm{H}$ has one neutron, and ${}^{3}_{1}\mathrm{H}$ has two neutrons.

Atomic Diagrams. Atomic diagrams show the number of protons and neutrons in the nucleus, and the distribution of electrons around the nucleus in energy levels. Atomic diagrams are extremely useful in predicting the ratios in which elements will combine. The information needed to draw atomic diagrams is found on the periodic table. The periodic table shows the atomic number which equals the number of protons or electrons, the atomic mass, and the electron configuration. It does not show the number of neutrons, but this can be determined by subtracting the atomic number from the atomic mass. This information can be used to draw a diagram.



Electron Dot Diagrams. Electron dot diagrams are a useful way to show the arrangement of outer electrons of an atom. They show valence electrons as dots at 12 o'clock, 3 o'clock, 6 o'clock, and 9 o'clock, and the rest of the atom, known as the kernel, as a symbol. Each of the clock positions represents one of the four outer orbitals. An orbital can hold a maximum of two electrons. The first orbital, represented by any one of the clock positions, is filled with a pair of electrons before putting electrons into the other orbitals. The remaining three orbitals each receive one electron before pairing occurs. Silicon, for example, has four valence electrons. As a result, it will have two electrons in one of the clock positions and one electron in each of two of the remaining three.

Average Atomic Mass

The atomic mass listed on the *Periodic Table* is the average mass of the isotopes. Carbon, for example, has two naturally occurring stable isotopes. The large majority of carbon atoms, 98.89%, are ^{12}C, while only 1.108% are ^{13}C. That is why the average mass is so close to 12.

The average mass is determined by the procedure illustrated in the box to the right. The mass of each isotope is multiplied by its percentage. Then these products are added to find the average.

Average Atomic Mass

$$m_{avg} = p_1m_1 + p_2m_2 + \ldots p_nm_n$$

m_{avg} – average mass; p_1 – percentage of isotope 1;
m_1 – mass of isotope 1; p_2 – percentage of isotope 2;
m_2 – mass of isotope 2; p_n – percentage of isotope n;
m_n – mass of isotope n; n – the number of isotopes

Example

What is the average mass of chlorine if a sample consists of 77.35% Cl–35 and 22.65% Cl–37?

m_{avg} = (0.7735)(35) + (0.2265)(37)
= 27.07 + 8.381
= 35.45

Development of the Periodic Table. Dmitri Mendeleev (1869) prepared a card for each of the known elements listing the symbol, the atomic mass, and the chemical properties. He arranged the cards in order of increasing atomic mass and noticed a pattern: *MENDELEEV'S PERIODIC LAW* – When the elements are arranged in increasing order of atomic mass, the chemical properties repeat themselves periodically. Moseley noticed that when all the elements were arranged in order of mass a few were not in the right family with respect to properties. He used a procedure called X-ray diffraction to determine the atomic number of the elements. When the elements were arranged in increasing order of atomic number, the discrepancies in Mendeleev's table disappeared. *THE PERIODIC LAW* – When the elements are arranged in increasing order of atomic number, the chemical properties repeat themselves periodically. The modern Periodic Table is arranged in order of increasing atomic number.

Organization of the Periodic Table. The modern Periodic Table is arranged in order of increasing atomic number in vertical columns and horizontal rows. The vertical columns are elements with about the same number of outer electrons (valence electrons). They are called groups or families. Elements in the same family have similar properties. Horizontal rows are elements with the same number of shells or energy levels. They are called periods. The major divisions of the Periodic Table are: Alkali metals - Group 1; Alkaline earth metals - Group 2; Halogens - Group 17; Noble gases (Inert gases) - Group 18; Transition metals - Groups 3-12; Lanthanides - Row 6, elements 57 - 71;and Actinides - Row 7, elements 89 - 103.

Trends in the Periodic Table. Going across the table from left to right within a row or period the number of protons increases, so the pull on the electrons increases. As a result the covalent atomic radius decreases and metallic properties decrease (except in the transition elements). In addition the number of valence electrons increases and the number of shells remains the same. Going down the table within a group or family the number of protons also increases, but the number of shells increases too. As a result, the atomic radius increases, the pull on the electrons decreases, and metallic properties increase. In a family the number of valence electrons remains the same. This results in the following organization of the Periodic Table:

	1							18
1	1 **M** 1	2	13	14	15	16	17	2 **N** 2 **O**
2	3 **E** 2-1	4 2-2	5 2-3	6 2-4	7 2-5	8 2-6	9 **N** 2-7 **O**	10 **B** 2-8 **L**
3	11 **T** 2-8-1	12 2-8-2	13 2-8-3	14 2-8-4	15 2-8-5	16 2-8-6	17 **N** 2-8-7 **M**	18 **E** 2-8-8
4	19 **A** 2-8-8-1	20 2-8-8-2					**E** **T**	**G** **A**
5	**L**						**A** **L**	**S** **E**
6	**S**						**S**	**S**
7								

Bonding. The electrons of one atom are attracted to the protons of another. When atoms combine, there is a tug of war over the valence electrons. The combining atoms either lose, gain, or share electrons in such a way that they complete their outer shells. Whether atoms gain, lose, or share electrons depends how tightly they hold onto their own electrons and how strongly they pull on the electrons of another atom.

Ionic Bonds. Ionic bonds are caused by the attraction between oppositely charged ions. Ions form as follows: The electrons of one atom are attracted to the protons of another. Metals hold onto electrons loosely while nonmetals hold onto electrons tightly. As a result, metals lose electrons and nonmetals gain electrons in such a way that they complete their outer shells. Atoms that gain or lose electrons become electrically charged. Metals become positively charged ions by losing electrons. Nonmetals become negatively charged ions by gaining electrons. Metal cations and nonmetal anions become ionically bonded because they are oppositely charged. Atoms gain or lose electrons in such a way that they complete their outer shells. This gives them the same electron configuration as a noble gas. For example, potassium, with an electron configuration of 2-8-8-1 loses an electron to become K^+ with an electron configuration of 2-8-8, the same as argon. Chlorine, with an electron configuration of 2-8-7, gains an electron to become Cl^-, with an electron configuration also of 2-8-8.

Covalent Bonds. Covalent bonds are bonds formed by sharing electrons. The electrons of one atom are attracted to the protons of another, but neither atom pulls strongly enough to remove an electron from the other. Covalent bonds form when the electronegativity difference between the elements is less than 1.7 (see the Electronegativity table on the back of the Periodic Table) or when hydrogen behaves like a metal. When a covalent bond forms, no valence electrons are transferred, rather, they are shared. During covalent bonding, unpaired electrons pair up in such a way that the atoms complete their outer shells. This can be illustrated with electron dot diagrams.

Answer the questions below by circling the number of the correct response

1. What is the maximum number of orbitals in the second principal energy level? (1) 1 (2) 2 (3) 3 (4) 4

2. The sublevel of lowest energy is (1) 2s (2) 3s (3) 2p (4) 3d

3. The electron configuration of an atom is $1s^2 2s^2 2p^6 3s^2 3p^3$. The atomic number of the atom is (1) 15 (2) 6 (3) 3 (4) 5

4. The total number of protons in the nucleus of this element $1s^2 2s^2 2p^6 3s^2 3p^2$ is (1) 7 (2) 8 (3) 14 (4) 28

5. What is the total number of protons in the nucleus of this atom $1s^2 2s^2 2p^6 3s^2 3p^4$? (1) 5 (2) 11 (3) 16 (4) 27

6. Which electron configuration represents an atom in an excited state? (1) $1s^2 2s^2 2p^6 4s^1$ (2) $1s^2 2s^2 2p^6$ (3) $1s^2 2s^2$ (4) $1s^2 2s^2 2p^6 3s^2 3p^1$

7. Which is the electron configuration for a neutral atom with an Atomic Number of 18? (1) $1s^2 2s^2 2p^6 3s^1 3p^7$ (2) $1s^2 2s^2 2p^6 3s^2 3p^6$ (3) $1s^2 2s^2 2p^6 3s^7 3p^1$ (4) $1s^2 2s^2 2p^8 3s^2 3p^4$

8. An atom with the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ has an incomplete (1) 2nd principal energy level (2) 2s sublevel (3) 3rd principal energy level (4) 3s sublevel

9. The electron configuration of a phosphorous atom is (1) $1s^2 2s^2 2p^3$ (2) $1s^2 2s^2 2p^6 3s^2 3p^3$ (3) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ (4) $1s^2 2s^2 2p^1$

10. How many orbitals are half filled in an atom: $1s^2 2s^2 2p^6 3s^2 3p^4$ of this element in the ground state? (1) 1 (2) 2 (3) 3 (4) 6

11. What is the number of orbitals in a 4d sublevel? (1) 1 (2) 5 (3) 3 (4) 7

12. A neutral atom in the ground state has an atomic number of 8. How many electrons are in the 2p sublevel? (1) 1 (2) 2 (3) 3 (4) 4

13. What is the electron configuration for a neutral atom of the radioisotope $_{15}P^{32}$ in its ground state? (1) $1s^2 2s^2 2p^6 3s^2 3p^2$ (2) $1s^2 2s^2 2p^6 3s^1 3p^4$ (3) $1s^2 2s^2 2p^6 3s^2 3p^3$ (4) $1s^2 2s^2 2p^6 3s^2 3p^6$

14. What is the number of filled orbitals in a neutral atom of sulfur-32 in the ground state? (1) 1 (2) 6 (3) 7 (4) 9

15. Which represents the outermost electron configuration of an Na atom in the ground state? (1) $1s^1$ (2) $2s^1$ (3) $3s^1$ (4) $4s^1$

16. What is the maximum number of orbitals in a d sublevel? (1) 1 (2) 5 (3) 3 (4) 7

17. Which orbital in an atom of calcium would contain electrons with the highest energy? (1) 3s (2) 3p (3) 2p (4) 4s

18. What is a possible electronic configuration of a nitrogen atom? (1) $1s^1 2s^3 2p^3$ (2) $1s^2 2s^2 2p^3$ (3) $1s^2 2s^3 2p^2$ (4) $1s^3 2s^3 2p^1$

19. A completely filled principal energy level contains 32 electrons. The principal quantum number (*n*) of this level is (1) 5 (2) 2 (3) 3 (4) 4

20. What is the total number of unpaired electrons in an atom with the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^4$? (1) 6 (2) 2 (3) 3 (4) 4

21. What is the maximum number of electrons in the third principal energy level? (1) 6 (2) 2 (3) 10 (4) 18

22. In the third principal energy level, the sublevel of highest energy is (1) s (2) p (3) f (4) d

23. How many orbitals are half filled in an atom: $1s^2 2s^2 2p^6 3s^2 3p^4$ of this element in the ground state? (1) 1 (2) 2 (3) 3 (4) 6

24. Which of the following particles is negatively charged? (1) electron (2) proton (3) neutron (4) cation

25. How many neutrons does $^{35}_{17}\mathrm{Cl}$ have? (1) 35 (2) 17 (3) 52 (4) 18

26. Isotopes are atoms which have different (1) atomic masses, (2) atomic radii, (3) atomic numbers, (4) electron configurations

27. An atom that contains 35 protons, 45 neutrons, and 35 electrons has an atomic number of (1) 35, (2) 80, (3) 45, (4) 115

28. Two isotopes of the same element will have the same number of (1) neutrons and electrons, (2) neutrons and nucleons, (3) protons and nucleons, (4) protons and electrons

29. An atomic mass unit is defined as exactly (1) $^1/_{12}$ the mass of a ^{12}C atom, (2) $^1/_{14}$ the mass of a ^{14}N atom, (3) $^1/_{16}$ the mass of a ^{16}O atom, (4) $^1/_{19}$ the mass of a ^{19}F atom

30. Which correctly represents all atom of neon containing 11 neutrons? (1) $^{11}_{10}\mathrm{Ne}$ (2) $^{21}_{10}\mathrm{Ne}$ (3) $^{20}_{11}\mathrm{Ne}$ (4) $^{21}_{11}\mathrm{Ne}$

31. How many electrons are in a neutral atom of $^7_3\mathrm{Li}$? (1) 7 (2) 10 (3) 3 (4) 4

32. The nucleus of a fluorine atom has a charge of (1) 1^+, (2) 19^+, (3) 9^+, (4) 0

33. What is the total number of neutrons in an atom of $^{39}_{19}\mathrm{K}$?
(1) 19 (2) 20 (3) 39 (4) 58

34. Name the subatomic particles contained in the nucleus of the atom.

35. State the charge associated with each type of subatomic particle contained in the nucleus of the atom.

36. What is the net charge of the nucleus?

37. Which of the following particles is negatively charged? (1) electron (2) proton (3) neutron (4) cation

38. How many neutrons does $^{35}_{17}Cl$ have? (1) 35 (2) 17 (3) 52 (4) 18

39. What is the atomic number of helium? (1) 1 (2) 2 (3) 3 (4) 4

40. Hydrogen–3 differs from hydrogen–1 in that hydrogen–3 has (1) 1 more proton, (2) 2 more protons, (3) 1 more neutron, (4) 2 more neutrons.

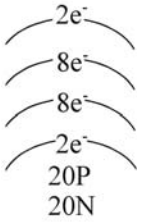
41. What is the mass number of carbon–14? (1) 12 amu (2) 14 amu (3) 6 amu (4) 8 amu

42. The property of all carbon atoms that is the same is (1) the mass, (2) the number of neutrons, (3) the number of protons, (4) the number of nucleons [particles in the nucleus]

43. The number of neutrons in a typical atom with an electron configuration of 2–8–7 is (1) 17, (2) 18, (3) 35, (4) 7.

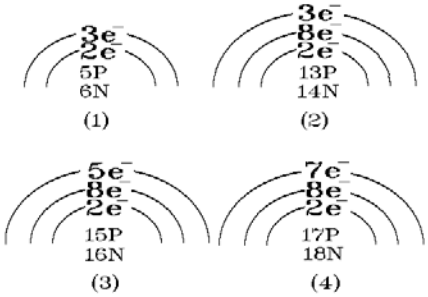
44. Frischium comes in three isotopes with the following abundances: 90.000 percent Fs–500; 8.0000 percent Fs–501; and 2.0000 percent Fs–503. The average mass is (1) 598.2 amu (2) 501.33 amu (3) 499.85 amu (4) 500.14 amu

45. Below is a Bohr-Rutherford diagram of an element.



Which element could be represented by this diagram? (1) calcium (2) cadmium (3) chlorine (4) no known element

46. Which of the following is a correct diagram of aluminum [Al]?



47. Which of the following is the correct electron dot diagram for helium?

(1) He· (2) He (3) He (4) :He:

48. In the Periodic Table, the elements are arranged in order of increasing (1) atomic size, (2) atomic number, (3) atomic mass, (4) ionization energy

49. The chemical properties of the elements are periodic functions of their atomic (1) spin, (2) isotopes, (3) mass, (4) number.

50. Which pair contains elements which have the most similar chemical properties? (1) Mg and Ca (2) N and S (3) H and Li (4) Na and Cl

51. The element with an atomic number of 34 is most similar in its chemical behavior to the element with an atomic number of (1) 19 (2) 31 (3) 36 (4) 16

52. Silicon is most similar in chemical activity to (1) carbon, (2) lead, (3) sulfur, (4) nitrogen

53. The element 2–8–6 belongs in Period (1) 6, (2) 2, (3) 3, (4) 4

54. Most of the elements in the Periodic Table are classified as (1) metalloids, (2) nonmetals, (3) noble gases, (4) metals

55. Phosphorus is best classified as a (1) nonmetal, (2) metalloid, (3) metal, (4) transition element

56. The Group 1 metals all have the same (1) electronegativity, (2) atomic radius, (3) oxidation state, (4) ionization energy

57. Which Group in the Periodic Table contains the most active metals? (1) 1 (2) 2 (3) 13 (4) 14

58. In which Group of the Periodic Table would this element, 2–5, most likely be found? (1) 1 (2) 2 (3) 13 (4) 15

59. As the elements in Period 3 are considered in order of increasing atomic number, the number of principal energy levels in each successive element (1) decreases (2) increases (3) remains the same

60. Which Group contains elements which are metalloids? (1) 1 (2) 11 (3) 14 (4) 4

61. The elements with the least chemical reactivity are in Group (1) 1, (2) 18, (3) 3 (4) 16

62. Which element is a metalloid? (1) arsenic (2) neon (3) potassium (4) bromine

63. Which Group of elements exhibits all three phases of matter at room temperature? (1) 2 (2) 14 (3) 15 (4) 17

64. What are two properties of most nonmetals?
    (1) high ionization energy and poor electrical conductivity
    (2) high ionization energy and good electrical conductivity
    (3) low ionization energy and poor electrical conductivity
    (4) low ionization energy and good electrical conductivity

65. Which element is classified as a noble gas at STP? (1) hydrogen (2) neon (3) oxygen (4) nitrogen

66. In which shell are the valence electrons of the elements in Period 2 found? (1) 1 (2) 2 (3) 3 (4) 4

67. Of the following, which element has the smallest atomic radius? (1) Mg (2) Ca (3) Sr (4) Ba

68. As one proceeds from lithium to fluorine in the Periodic Table, the tendency for the elements to lose electrons (1) decreases, (2) increases, (3) remains the same

69. As the elements in Period 3 are considered from left to right, the ability of each successive element to gain electrons (1) decreases, (2) increases, (3) remains the same

70. Of the following, which is the element with the most metallic character in Group 16 is (1) O, (2) S, (3) Se, (4) Te

71. As the elements in Group 14 are considered in order of increasing atomic number, the metallic properties of successive elements (1) decreases, (2) increases, (3) remains the same

72. In Period 3 of the Periodic Table, the element with the smallest atomic radius is in Group (1) 1 (2) 2 (3) 15 (4) 17

73. Of the following, which Group 2 element has the greatest tendency to lose electrons? (1) calcium (2) barium (3) strontium (4) magnesium

74. Which Group in the Periodic Table contains atoms that have –2 oxidation states? (1) 1 (2) 2 (3) 16 (4) 17

75. The elements in Group 2 have similar chemical properties primarily because they have the same (1) ionization energies, (2) reduction potentials, (3) number of principal energy levels, (4) number of electrons in the outermost shell

76. As one proceeds from left to right across Period 2 of the Periodic Table, the decrease in atomic radius is primarily due to an increase in the number of (1) orbitals, (2) protons, (3) neutrons, (4) principal energy levels

77. The most active metal in Period 4 of the Periodic Table is (1) Fe, (2) Sc, (3) K, (4) Ca.

78. In Period 3, as the atomic numbers increase, the pattern according to which the properties of the elements change is
(1) metal → metalloid → nonmetal → noble gas
(2) metal → nonmetal → noble gas → metalloid
(3) nonmetal →metalloid → metal → noble gas
(4) nonmetal → metal → noble gas → metalloid

79. In going down the Group 15 elements on the Periodic Table, the metallic properties of the elements (1) decrease, (2) increase, (3) remain the same

80. As one proceeds from left to right across Period 3 of the Periodic Table, there is a decrease in (1) ionization energy (2) electronegativity (3) metallic characteristics (4) valence electrons

81. As one proceeds from fluorine to astatine in Group 17, the electronegativity (1) decreases and the atomic radius increases, (2) decreases and the atomic radius decreases, (3) increases and the atomic radius decreases, (4) increases and the atomic radius increases.

82. As the elements in Period 3 are considered in order of increasing atomic number, the number of principal energy levels in each successive element (1) decreases, (2) increases, (3) remains the same

83. If the elements are considered from top to bottom in Group 17 the number of electrons in the outermost shell will (1) decrease, (2) increase, (3) remain the same

84. Which represents the correct order of activity for the Group 17 elements [> means greater than]
(1) bromine > iodine > fluorine > chlorine
(2) fluorine > chlorine > bromine > iodine
(3) iodine > bromine > chlorine > fluorine
(4) fluorine > bromine > chlorine > iodine

85. Which is most characteristic of metals with very low ionization energies? (1) they are very reactive (2) they have a small atomic radius (3) they form covalent bonds (4) they have a high electronegativity

86. Metallic elements usually possess
(1) low electronegativities and high ionization energies
(2) high electronegativities and low ionization energies
(3) high electronegativities and high ionization energies
(4) low electronegativities and low ionization energies

87. If the members of Group 17 are arranged in order of increasing electronegativity, they are also arranged in order of increasing (1) ionization energy, (2) atomic radius, (3) atomic mass, (4) nuclear charge

88. As the elements are considered from top to bottom in Group15 of the Periodic Table, the ionization energy (1) decreases, (2) increases, (3) remains the same

89. An element that has both a high ionization energy and a high electronegativity is most likely a (1) metal (2) metalloid (3) nonmetal (4) noble gas

90. The element with the lowest first ionization energy in any given Period will always belong to Group (1) 1 (2) 2 (3) 17 (4) 18

91. An element that exhibits the largest variety of oxidation states is (1) Li (2) O (3) C (4) N

92. Which Group in the Periodic Table contains both metals and nonmetals? (1) 11 (2) 2 (3) 18 (4) 14

93. This element assumes only a +3 oxidation state in chemical combination (1) Na (2) Si (3) Al (4) Cl

94. Which Period contains elements that are all gases at STP ? (1) 1 (2) 2 (3) 3 (4) 4

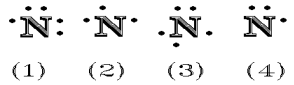
95. Which Group 18 element in the ground state has a maximum of 2 completely filled principal energy levels? (1) Kr (2) Xe (3) He (4) Ne

96. A nonmetal which exists in the liquid state at room temperature is (1) aluminum (2) hydrogen (3) mercury (4) bromine

97. The only metal which is a liquid at STP is in Period (1) 5 (2) 6 (3) 3 (4) 4

98. Which Group contains an element that is a liquid at room temperature? (1) 18 (2) 2 (3) 16 (4) 17

99. Which of the following is the correct electron dot diagram for nitrogen?



(1) (2) (3) (4)

100. Barium combines by (1) gaining two electrons, (2) losing two electrons, (3) sharing two electrons, (4) sharing 3 electrons.

101. Which of the following occurs during covalent bonding? (1) Electrons are lost. (2) Electrons are gained. (3) Valence electrons fall from the excited state to the ground state. (4) Unpaired electrons form pairs.

102. When calcium combines, it usually (1) loses two electrons, (2) gains six electrons, (3) shares two electrons, (4) shares six electrons.

103. A Ca^{2+} ion differs from a Ca atom in that the Ca^{2+} ion has (1) more protons, (2) more electrons, (3) fewer protons, (4) fewer electrons.

104. When an ionic bond is formed, the atom that transfers its valence electron is the atom that has the (1) higher electronegativity value, (2) lower atomic number. (3) higher atomic mass, (4) lower ionization energy.

105. When an ionic bond is formed, the atom that transfers its valence electron becomes an ion with (1) positive charge and more protons, (2) positive charge and no change in the number of protons, (3) negative charge and more protons, (4) negative charge and no change in the number of protons.

106. An atom that loses or gains one or more electrons becomes (1) an ion, (2) an isotope, (3) a molecule, (4) an electrolyte

107. Which kind of bond is formed when two atoms share electrons to form a molecule? (1) ionic (2) metallic (3) electrovalent (4) covalent

Answers

1. 4	23. 2	44. 4	66. 2	88. 1			
2. 1	24. 1	45. 1	67. 1	89. 3			
3. 1	25. 4	46. 2	68. 1	90. 1			
4. 3	26. 1	47. 2	69. 2	91. 4			
5. 3	27. 1	48. 2	70. 4	92. 4			
6. 1	28. 4	49. 4	71. 2	93. 3			
7. 2	29. 1	50. 1	72. 4	94. 1			
8. 3	30. 2	51. 4	73. 2	95. 4			
9. 2	31. 3	52. 1	74. 3	96. 4			
10. 2	32. 3	53. 3	75. 4	97. 2			
11. 2	33. 2	54. 4	76. 2	98. 4			
12. 4	34. protons,	55. 1	77. 3	99. 3			
13. 3	neutrons	56. 3	78. 1	100. 2			
14. 3	35. $p = +1$; $n = 0$	57. 1	79. 2	101. 4			
15. 3	36. positive	58. 4	80. 3	102. 1			
16. 2	37. 1	59. 3	81. 1	103. 4			
17. 4	38. 4	60. 3	82. 3	104. 4			
18. 2	39. 2	61. 2	83. 3	105. 2			
19. 4	40. 4	62. 1	84. 2	106. 1			
20. 2	41. 2	63. 4	85. 1	107. 4			
21. 4	42. 3	64. 1	86. 4				
22. 4	43. 2	65. 2	87. 1				