Test Review № 5

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

Families of Elements. *Alkali metals* (Group 1) are extremely reactive (not found free in nature). They form stable ionic compounds, react with water to form a base, react with air to form oxides, and react with acids to form salts. *Alkaline earth metals* (Group 2) are also reactive (not found free in nature), but not as reactive as group 1 elements. They form stable ionic compounds, react with water to form a base, react with air to form oxides, and react with acids to form salts. The *Nitrogen family* (Group 15) has members that range from typical nonmetals (nitrogen and phosphorus) through metalloids (arsenic and antimony) to metals (bismuth). Nitrogen forms stable diatomic molecules with a triple bond. It is a component of protein and forms some unstable compounds that are used as explosives. Phosphorus is a component of nucleic acids (DNA, RNA). It is more reactive than nitrogen at room temperature. The *Oxygen family* (Group 16) has members that range from typical nonmetals (oxygen and sulfur) through metalloids (selenium and tellurium) to metals (polonium). They are all solids except oxygen. *Halogens* (salt formers - Group 17) are very reactive nonmetals. They have high electronegativities, and are not found free in nature. When they are free, they form diatomic molecules. They react with metals to form salts. They are found in all three phases – solid, liquid and gas. *Noble gases* (Group 18) have complete outer shells. As a result, they are almost inert (not reactive) Krypton, xenon, and radon form compounds with oxygen and fluorine, however. *Transition elements* (Groups 3–12) lose electrons from two outermost energy levels and form colored solutions.

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. Ions that have the electron configuration of the same noble gas are isoelectronic. As nuclear charge of isoelectronic ions increases, radius generally decreases (much as it does for elements in the same period). Cations are smaller than the parent atom due to increased attraction of electrons to the net positive charge. Anions are larger than the parent atom due to increased repulsion of the electrons by the net negative charge. The formation of binary ionic compounds is due to the fact that the aggregated oppositely charged ions in an ionic solid have lower energy than the original elements. The strength of the attraction is indicated by lattice energy.

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.

Bond Type and Polarity. When the electronegativity difference is greater than or equal to 1.7, the atom with the greater electronegativity gains the electron, and an **ionic bond** is formed. Electronegativity differences below 1.7 result in covalent bonds or sharing. If the electronegativity difference is close to zero (<0.4), the atoms share equally and a **nonpolar bond** forms. Higher electronegativity differences (still below 1.7) result in unequal sharing or **polar bonds**.

Electron Dot Diagrams. Electron dot diagrams 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. Electrons will move into the *s* orbital first. Once the *s* is filled, additional electrons will go into each of the three *p* orbitals without pairing until each *p* orbital has one electron. During bonding, however, the outer shell of the atom is composed of four equal orbitals. Electrons do not pair up until each orbital contains an electron.

Lewis Structures. Lewis structures show how valence electrons are arranged among atoms in a compound using dots to represent the valence electrons that are not shared in a covalent bond. To draw Lewis structures showing covalent bonds for elements in periods 1 and 2, do the following: [1] sum the valence electrons; [2] use a pair of electrons to form a bond (bonding electrons); and [3] arrange the remaining electrons to satisfy the duet rule for hydrogen and the octet rule for other elements (lone pairs). Lewis structures show how valence electrons are arranged among atoms in a compound using dots to represent the valence electrons that are not shared in a covalent bond. To draw Lewis structures showing covalent bonds for elements in periods 1 and 2, do the following: [1] sum the valence electrons; [2] use a pair of electrons to form a bond (bonding electrons); and [3] arrange the remaining electrons to satisfy the duet rule for hydrogen and the octet rule for other elements (lone pairs).There may several valid equivalent Lewis structures for some molecules. This is called resonance. When there are several nonequivalent Lewis structures for a molecule, it is possible to choose among them using formal charge. Formal charge is the difference between the number of valence electrons on the free atom and the number of valence electrons assigned to the atom in the molecule. Lone pairs belong entirely to the atom in question. Shared electrons are divided equally among the sharing atoms. Atoms in molecules should have formal charges as close to zero as possible. Negative formal charges reside with the most electronegative element.

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

1. The electron configuration for Cu is (1) $[Ar]4s^23d^9$ (2) $[Kr]4s^23d^9$ (3) $[Ar]4s^13d^{10}$ (4) $[Ar]4s^24d^9$ (5) $[Ar]3s^13d^{10}$

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

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

4. 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

5. 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

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

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

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

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

10. The alkali metals all have the same (1) electronegativity, (2) atomic radius, (3) oxidation number, (4) ionization energy

11. The alkaline earth metals are those elements in Group (1) 1, (2) 2, (3) 11, (4) 12

12. Which Group in the Periodic Table contains the alkali metals? (1) 1 (2) 2 (3) 13 (4) 14

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

14. 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

15. Which is an alkaline earth metal? (1) Na (2) Ca (3) Ga (4) Ta

16. A metallic element whose aqueous ions produce colorless solutions would be found in Period 4 and Group (1) 1 (2) 17 (3) 8 (4) 18

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

18. Which is a transition element? (1) Ag (2) Mg (3) Sb (4) Si

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

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

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

22. 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

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

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

25. Which element has the smallest atomic radius? (1) Mg (2) Ca (3) Sr (4) Ba

26. 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

27. 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

28. The element with the most metallic character in Group 16 is (1) O, (2) S, (3) Se, (4) Te

29. 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

30. 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

31. Which Group 2 element has the greatest tendency to lose electrons? (1) calcium (2) barium (3) strontium (4) magnesium

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

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

34. 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

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

36. 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
(3) metal → nonmetal → noble gas → metalloid
(2) nonmetal →metalloid → metal → noble gas
(4) nonmetal → metal → noble gas → metalloid

37. 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

38. 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

39. 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.

40. 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

41. 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

42. 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

43. 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

44. 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

45. If the members of the halogen family 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

46. 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

47. 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

48. 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

49. The elements that react with water to form strong bases are found in Group (1) 1 (2) 15 (3) 13 (4) 17

50. The alkaline earth metals are those elements in Group (1) 1 (2) 2 (3) 11 (4) 12

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

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

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

54. Elements in which electrons from more than one energy level may be involved in bond formation are called (1) alkali elements (2) transition elements (3) alkaline earth elements (4) halogens

55. Which is a transition element? (1) Rb (2) Au (3) Sb (4) Xe

56. Which type of element frequently forms colored compounds and generally exhibits more than one positive oxidation state? (1) alkaline earths (2) alkali metals (3) transition elements (4) noble gases

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

58. 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

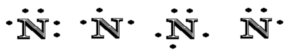
59. A nonmetal which exists in the liquid state at room temperature is (1) aluminum (2) hydrogen (3) mercury (4) bromine

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

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

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

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



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

64. In water, the bond between hydrogen and oxygen is (1) ionic, (2) polar covalent, (3) nonpolar covalent, (4) nonpolar noncovalent.

65. 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.

66. Which of the following is an example of a substance with a nonpolar covalent bond? (1) HCl (2) Cl_2 (3) $HClO_2$ (4) NaCl

67. The electronegativity of sulfur is (1) 16, (2) 239, (3) 2.6, (4) 32.

68. Which of the following elements has the highest electronegativity? (1) fluorine (2) chlorine (3) barium (4) hydrogen

69. The formula for magnesium fluoride is MgF_2. The best explanation for this fact is that when they combine (1) each of two magnesium atoms lose an electron and a fluorine atom gains two, (2) a magnesium atom loses two electrons and each of two fluorine atoms gains one, (3) a magnesium atom shares two electrons with two fluorine atoms, (4) each of two magnesium atoms share an electron with a fluorine atom.

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

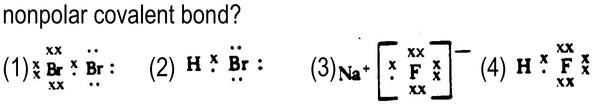
71. Which compound contains a bond with the *least,* ionic character? (1) CO (2) K_2O (3) CaO (4) Li_2O

72. Which type of bond is contained in a water molecule? (1) nonpolar covalent (2) ionic (3) polar covalent (4) electrovalent

73. The bonding in NH_3 most similar to the bonding in (1) H_2O (2) MgO (3) NaCl (4) KF

74. Which is the formula of an ionic compound? (1) SO_2 (2) CH_3OH (3) CO_2 (4) NaOH

75. Which electron dot formula represents a molecule that contains a nonpolar covalent bond?



76. When a reaction occurs between atoms with ground state electron configurations 2–1 and 2–7, the predominant type of bond formed is (1) polar covalent, (2) ionic, (3) nonpolar covalent, (4) metallic.

77. The P—Cl bond in a molecule of PCl_3 is (1) nonpolar covalent, (2) coordinate covalent, (3) polar covalent, (4) electrovalent.

78. 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.

79. In which compound does the bond between the atoms have the least ionic character? (1) HF (2) HCl (3) HBr (4) HI

80. Which substance contains a polar covalent bond? (1) Na_2O (2) Mg_3N_2 (3) CO_2 (4) N_2

81. In which pair do the members have identical electron configurations? (1) S^{2-} and Cl^- (2) S^0 and Ar^0 (3) K^0 and Na^+ (4) Cl^- and K^0

82. When a chlorine atom reacts with a sodium atom to form an ion, the chlorine atom will (1) lose one electron, (2) gain one electron, (3) lose two electrons, (4) gain two electrons.

83. When a calcium atom loses its valence electrons, the ion formed has an electron configuration that is the same as the configuration of an atom of (1) Cl (2) Ar (3) K (4) Sc

84. Which of the following compounds has the most ionic character? (1) KI (2) NO (3) HCl (4) MgS

85. Which atom has the strongest attraction for electrons? (1) Cl (2) F (3) Br (4) I

86. Which compound is ionic? (1) HCl (2) $CaCl_2$ (3) SO_2 (4) H_2O

87. Two atoms of element A unite to form a molecule with the formula A_2. The bond between the atoms in the molecule is (1) electrovalent, (2) nonpolar covalent, (3) ionic, (4) polar covalent.

88. 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.

89. 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.

90. Which compound best illustrates ionic bonding? (1) CCl_4 (2) $MgCl_2$ (3) H_2O (4) CO_2

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

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

93. Which type of bonding is usually exhibited when the electronegativity difference between two atoms is 1.2? (1) ionic (2) metallic (3) network (4) covalent

94. Which element will form an ion with a larger radius than its atom? (1) Na (2) Ba (3) Ca (4) Cl

95. Which element will form an ion whose radius is larger than its atomic radius? (1) F (2) Fr (3) Ca (4) Cs

96. When chlorine reacts with a Group 1 metal, it becomes an ion with a charge of (1) 1-, (2) 2-, (3) 1+, (4) 2+.

97. Which compound contains both covalent and ionic bonds? (1)HCl (2) NH_4Cl (3) $MgCl_2$ (4) CCl_4

98. When a chlorine atom reacts with a sodium atom to form an ion, the chlorine atom will (1) lose one electron, (2) gain one electron, (3) lose two electrons, (4) gain two electrons.

99. When a reaction occurs between atoms with ground state electron configurations $1s^22s^1$ and $1s^22s^22p^5$ the predominant type of bond formed is (1) polar covalent, (2) ionic, (3) nonpolar covalent, (4) metallic.

100. What is the total number of electrons in a Mg^{2+} ion? (1) 10 (2) 2 (3) 12 (4) 24

101. Which compound is ionic? (1) HCl (2) $CaCl_2$ (3) SO_2 (4) N_2O_5

102. What is the electron configuration for Be^{2+} ions? (1) $1s^1$ (2) $1s^2$ (3) $1s^22s^1$ (4) $1s^22s^2$

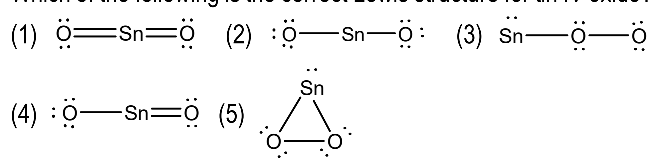
103. As a chemical bond forms between hydrogen and chlorine atoms, the potential energy of the atoms (1) decreases, (2) increases, (3) remains the same.

104. In potassium hydrogen carbonate, $KHCO_3$, the bonds are (1) ionic, only, (2) covalent, only, (3) both ionic and covalent, (4) both covalent and metallic.

105. When potassium and chlorine form a chemical compound, energy is (1) released and ionic bonds are formed, (2) released and covalent bonds are formed, (3) absorbed and ionic bonds are formed, (4) absorbed and covalent bonds are formed.

106. Which compound contains both covalent and ionic bonds? (1)HCl (2) NH_4Cl (3) $MgCl_2$ (4) CCl_4

107. Which of the following is the correct Lewis structure for tin IV oxide?



108. Which of the following molecules has two pairs of nonbonding electrons on the central atom? (1) BH_3 (2) CO_2 (3) H_2O (4) H_2S (5) choices 3 and 4

109. Which type of bonding is usually exhibited when the electronegativity difference between two atoms is 1.2? (1) ionic (2) metallic (3) network (4) covalent

110. How many single bonds are in a molecule of carbon dioxide, CO_2 ? (1) None (2) One (3) Two (4) Three (5) Four

Answers

1. 1 or 3	17. 3	33. 4	49. 1	65. 4	81. 1	97. 2
2. 2	18. 1	34. 2	50. 2	66. 2	82. 2	98. 2
3. 4	19. 2	35. 3	51. 4	67. 3	83. 2	99. 2
4. 1	20. 1	36. 1	52. 4	68. 1	84. 1	100. 1
5. 4	21. 4	37. 2	53. 3	69. 2	85. 2	101. 2
6. 1	22. 1	38. 3	54. 2	70. 1	86. 2	102. 2
7. 3	23. 2	39. 1	55. 2	71. 1	87. 2	103. 1
8. 4	24. 2	40. 3	56. 3	72. 3	88. 4	104. 3
9. 1	25. 1	41. 3	57. 1	73. 1	89. 2	105. 1
10. 3	26. 1	42. 2	58. 4	74. 4	90. 2	106. 2
11. 2	27. 2	43. 1	59. 4	75. 1	91. 1	107. 1
12. 1	28. 4	44. 4	60. 2	76. 2	92. 4	108. 5
13. 4	29. 2	45. 1	61. 4	77. 3	93. 4	109. 4
14. 3	30. 4	46. 1	62. 2	78. 4	94. 4	110. 1
15. 2	31. 2	47. 3	63. 3	79. 4	95. 1	
16. 3	32. 3	48. 1	64. 2	80. 3	96. 1	