1. Which of the following is a <u>physical</u> property of aspirin?
    a. Aspirin can moderate some heart disorders when ingested.
    b. Aspirin does not decompose when tightly sealed in a bottle.
    c. Aspirin yields carbon dioxide and water vapor when burned.
    d. Aspirin can be pressed into tablets when mixed with cornstarch.
    e. Aspirin reacts with water to produce salicylic acid and acetic acid.

2. Which best describes the size and shape of a sample of gas?
    a. Definite volume and definite shape.
    b. Definite volume, but shape is determined by the container.
    c. Volume determined by container, but definite shape.
    d. Volume and shape both determined by the container.
    e. Volume and shape cannot be described.

3. Barium sulfate is described as a white crystalline solid which melts
    at 1580°C and decomposes at 1600°C. At a temperature of 500°C, you
    would expect a sample of barium sulfate to be a
    a. colorless liquid.
    b. white crystalline solid.
    c. yellow liquid.
    d. white cloud of vapor.
    e. form that cannot be determined.

4. Which of the following is a mixture?
    a. cough syrup
    b. iron
    c. helium
    d. sodium hydrogen carbonate
    e. steam

5. Which of the following can be classified as a pure compound?
    a. alcohol in water, C_2H_5OH in H_2O
    b. sugar, $C_{12}H_{22}O_{11}$
    c. carbon, C
    d. iodine, I_2
    e. mercury, Hg

6. What is the chemical symbol for copper?
    a. C   b. Ca   c. Cl   d. Cr   e. Cu

7. What element is represented by the chemical symbol K?
    a. kaolin   b. phosphorus   c. potassium d. silver   e. sodium

8. Which chemical symbol represents a <u>metallic</u> element?
    a. Ar   b. Br   c. Ca   d. H   e. P

9. Which example illustrates a form of <u>potential</u> energy?
    a. charged battery
    b. flying bird
    c. lighted match
    d. moving truck
    e. running water

10. When the prefix *milli* is used in the metric or SI system, a
    fundamental unit of measurement is multiplied by what factor?
    a. 10^{-3} b. 10^{-2} c. 10^{-6} d. 10^3 e. 10^6

11. The measurement most likely to describe the amount of pain reliever in
    a headache tablet is
    a. 1.5 kg   b. 500 mg   c. 1.00 mL   d. 325 mg/mL   e. 0.25 L

12. Which of the following measurements has three significant figures?
    a. 1,207g   b. 4.250g   c. 0.006g   d. 0.0250g   e. 0.03750g

13. What is the numerical value of 1.2 x 1.222? Express your answer using
    the correct number of significant figures.
    a. 1.5   d. 1.4664
    b. 1.47   e. None of the these.
    c. 1.466

14. In scientific notation, the number 185,000,000 is
    a. 185×10^6 d. 185×10^{-8}
    b. 1.85×10^6 e. 1.85×10^{-8}
    c. 1.85×10^8

15. In scientific notation, the number 0.0046 is expressed as
    a. 46×10^{-3} d. 4.6×10^{-1}
    b. 4.6×10^{-3} e. 46×10^{-1}
    c. 4.6×10^{-2}

16. The number 5.320×10^2 in conventional notation is
    a. 532.0   b. 53.20   c. 5.320   d. 0.5320   e. 0.005320

17. How many grams are contained in 1.20 pounds?
    a. 545g   b. 378g   c. 264g   d. 2.2g   e. 1.20g

18. 95.0°F is the same as
    a. 21°C   b. 35°C   c. 85°C   d. 171°C   e. 203°C

19. 68°C is the same as
    a. 341 K   b. 321 K   c. 285 K   d. 205 K   e. 158 K

20. How many calories are released when 500 g of water cools from
    95.0°C to 25.0°C? (The specific heat of water is 1.00 cal/g·°C)
    a. 35.0 cal
    b. 70.0 cal
    c. 1.25×10^4 cal
    d. 3.50×10^4 cal
    e. 4.75×10^4 cal

21. What is the specific heat of a metal if it takes 26.5 calories to raise the temperature of a piece weighing 50.0 g by 5.00 Celsius degrees?
    a. 250 cal/g·°C
    b. 133 cal/g·°C
    c. 6.63 cal/g·°C
    d. 1.89 cal/g·°C
    e. 0.106 cal/g·°C

22. What is the volume of a gold nugget that weighs 2.20 kg? The density of gold is 19 g/cm^3.
    a. 8.60 x 10^3 cm^3
    b. 116 cm^3
    c. 11.6 cm^3
    d. 8.60 cm^3
    e. 0.116 cm^3

23. The smallest amount of an element that retains that element's characteristics is the
    a. atom   b. electron   c. molecule   d. neutron   e. proton

24. Which characteristics correctly describe a proton?
    a. approximate mass 1 amu; charge +1; inside nucleus
    b. approximate mass 5 x 10^{-4} amu; charge –1; outside nucleus
    c. approximate mass 5 x 10^{-4} amu; charge +1; inside nucleus
    d. approximate mass 1 amu; charge 0; inside nucleus
    e. approximate mass 1 amu; charge +1; outside nucleus

25. Atoms of $^{35}_{17}$Cl contain _____ protons and _____ electrons.

    a. 17,17   b. 17,18   c. 18,17   d. 35,18   e. 35,17

26. An atom with a mass number of 58 and with 32 neutrons will have _____ protons.
    a. 16   b. 26   c. 32   d. 58   e. 90

27. The symbol of the element with 23 protons is
    a. Mg
    b. Na
    c. V
    d. Cannot be determined without additional information.
    e. None of the above.

28. An atom with Z=26 and A=58 contains _____ protons, _____ electrons, and _____ neutrons.
    a. 26; 26; 58
    b. 58; 26; 26
    c. 26; 26; 32
    d. 32; 26; 32
    e. 26; 32; 84

29. An imaginary element X consists of two isotopes having masses of 100.0 amu and 102.0 amu. A sample of X was found to contain 75.0 % of the ^{100}X isotope and 25.0 % of the ^{102}X isotope. Calculate the atomic weight of X.
    a. 100.25amu b. 100.50amu c. 101.00amu d. 101.50amu e. 101.75amu

30. Which elements all belong in the same group?
    a. C, N, O
    b. Fe, Cu, Ni
    c. B, Si, As
    d. F, Cl, Br
    e. Al, Ge, Sb

31. Which element is most likely to have chemical properties similar to those of potassium (atomic number 19)?
    a. Ar   b. Ca   c. Sc   d. Rb   e. Sr

32. Which of the following is an alkali metal?
    a. Al   b. Cl   c. He   d. Na   e. O

33. The maximum number of electrons in any orbital is
    a. 1   b. 2   c. 3   d. 4   e. 5

34. How many electrons can occupy the 4p subshell?
    a. 1   b. 2   c. 6   d. 8   e. 10

35. The shell having n=3 contains _____ subshells, _____ orbitals, and up to _____ electrons.
    a. 2,4,8   b. 3,6,12   c. 3,6,18   d. 3,9,18   e. 3,12,36

36. What is the electron configuration of Mg?
    a. $1s^22s^22p^8$
    b. $1s^22s^22p^63s^2$
    c. $1s^22s^22p^63s^13p^3$
    d. $1s^22s^22p^63s^23p^64s^23d^5$
    e. None of the above.

37. The number of valence electrons in an element with electron configuration $1s^22s^22p^63s^23p^4$ is
    a. 2   b. 4   c. 6   d. 8   e. 16

38. The property that describes the ease with which an atom gives up an electron to form a positive ion is
    a. atomic number   b. electron affinity   c. electronegativity
    d. ionization energy e. None of these

39. The property defined as the energy released on adding an electron to a single atom is
    a. atomic number   b. electron affinity   c. electronegativity
   d. ionization energy e. None of these

40. An element belonging to the halogen family would be expected to have a
    _____ ionization energy and a _____ electron
    affinity.
    a. large, large
    b. large, small
    c. small, small
    d. small, large
    e. None of the above.

41. The term which best describes the crystalline substance that results
    when a large number of metal atoms transfer electrons to a large
    number of non-metal atoms is
    a. covalent compound.
    b. molecule.
    c. ionic solid.
    d. cation.
    e. anion.

42. What is the most likely charge on an ion formed by an element with a
    valence electron configuration of ns^1?
    a. 7^- b. 1^- c. 0   d. 1^+ e. 7^+

43. What fourth period element is represented by the dot structure shown?

     $\ddot{\underset{\displaystyle\cdot}{:X:}}$

    a. K   b. Ca   c. Mn   d. Br   e. Kr

44. The charge on a sulfide ion is
    a. 3^+ b. 2^+ c. 0   d. 2^- e. 3^-

45. What is the formula of the carbonate ion?

46. What is the formula of the nitrate ion?

47. What is the formula of the sulfite ion?

48. Which is the correct formula for the ionic compound containing
    iron(III) ions & oxide ions?
    a. FeO   b. FeO_2 c. Fe_2O_2 d. Fe_2O_3 e. Fe_3O_2

49. The formula for potassium dichromate is
    a. KCr_2O_7 b. $K_2Cr_2O_7$ c. K_2CrO_4 d. PCr_2O_7 e. $PCrO_4$

50. The formula for the compound chromium(II) nitrate is
    a. C_2NO_3 b. Cr_2NO_3 c. $CrNO_3$ d. $Cr(NO_3)_2$ e. $CrNO_2$

51. What is the name of Fe_2S_3?
    a. iron sulfide
    b. iron(II) sulfide
    c. iron(III) sulfide
    d. diiron trisulfide
    e. None of the above.

52. Which of the following formulas represents a compound that is a base?
    a. $CaSO_4$
    b. NH_4Cl
    c. $Mg(OH)_2$
    d. H_2O
    e. None of the above.

53. One definition of an acid is a substance that provides which ion in water solution?
    a. Na^+
    b. H^+
    c. OH^-
    d. NH_4^+
    e. None of the above.

54. One definition of a base is a substance that produces which ion in water solution?
    a. NH_4^+ b. OH^- c. H^+ d. Cl^- e. H^-

55. A chemical bond formed between two identical atoms is a(an) _____ bond.
    a. atomic   b. covalent   c. hydrogen   d. ionic   e. molecular

56. In a covalent compound the bond length can be defined as
    a. the distance between any two pairs of electrons.
    b. the distance between the two largest atoms.
    c. the distance between two nuclei when the repulsion is greatest.
    d. the distance between two nuclei when the attraction is greatest.
    e. the distance between two nuclei when repulsion and attraction are balanced.

57. The element least likely to obey the octet rule in forming chemical bonds is
    a. carbon   b. fluorine   c. nitrogen   d. sodium   e. sulfur

58. In forming covalent bonds where the octet rule is obeyed, sulfur usually forms _____ bonds and chlorine usually forms _____ bonds.
    a. one; one   b. two; two   c. one; two   d. two; one   e. six; seven

59. In a Lewis dot structure the electrons which complete an octet but are not located between two atoms are referred to as
    a. bonding pairs.
    b. delta minus electrons.
    c. excess electrons.
    d. filled shells.
    e. lone pairs.

60. A molecule in which the central atom has no lone pairs and forms four single bonds is said to have a _____ shape.
   a. bent   b. linear   c. planar   d. pyramidal e. tetrahedral

61. The water molecule has a _____ geometry because its central atom has _____ bonds and _____ lone pairs of electrons.
   a. bent; two; two
   b. linear; two; two
   c. pyramidal; three; one
   d. tetrahedral; four; zero
   e. trigonal; three; one

62. Which element listed is the **most** electronegative?
   a. aluminum   b. bromine   c. chlorine   d. iodine   e. sodium

63. The carbon dioxide molecule is linear. The electronegativities of C and O are 2.5 and 3.5, respectively. Based on these values and on consideration of molecular geometry, the C-O **bond** is _____ and the **molecule** is _____.
   a. polar; polar
   b. non-polar; non-polar
   c. polar; non-polar
   d. non-polar; polar
   e. None of the above.

64. What is the name of ICl_3?
   a. iodine chloride
   b. iodine(III) chloride
   c. triiodine chloride
   d. iodine trichloride
   e. tri(iodine chloride)

65. The formula for phosphorus pentafluoride is
   a. PhF_5 b. PF_5 c. P_5F d. $(PF)_5$ e. P_5F_5

66. If $SiCl_4$ is named as a covalent compound, what would it be called?
   a. silicon chloride
   b. chlorosilicate
   c. silicon tetrachloride
   d. sulfur chloride
   e. sulfur tetrachloride

67. The smallest possible unit of a covalent compound is a(an)
   a. atom   b. cation   c. formula unit   d. molecule   e. polyatomic ion

68. Consider the balanced equation shown, and identify the statement which is **not** true.

   $Na_2SO_4(aq) + BaCl_2(aq) ----> 2\ NaCl(aq) + BaSO_4(s)$
   a. The coefficient of sodium sulfate is one.
   b. Barium sulfate is produced in solid form.
   c. Barium chloride is dissolved in water.
   d. The products are barium sulfate and sodium chloride.
   e. 2 NaCl(aq) could also be correctly written as $Na_2Cl_2(aq)$.

69. Consider the reaction shown and identify the statement that is **not** true.

$$CaCO_3(s) \xrightarrow{825°C} CaO(s) + CO_2(g)$$
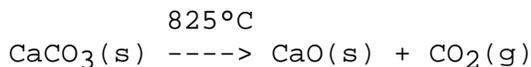
   a. This reaction is balanced as written.
   b. The reactant must be heated for this reaction to occur.
   c. The products are a solid and a gas.
   d. Water must be present for this reaction to occur.
   e. None of the above.

70. Which is the correct equation for the reaction of magnesium with hydrochloric acid to produce hydrogen and magnesium chloride?
   a. $Mg + 2\ HCl \longrightarrow H_2 + MgCl_2$
   b. $Mg + HCl \longrightarrow H + MgCl$
   c. $2\ Mg + 6\ HCl \longrightarrow 3\ H_2 + 2\ MgCl_3$
   d. $Mg + 2\ HCl \longrightarrow 2\ H + MgCl_2$
   e. $Mg + 3\ HCl \longrightarrow 3\ H + MgCl_3$

71. Which of the following equations is **not** balanced?
   a. $2\ Na + 2\ H_2O \longrightarrow 2\ NaOH + H_2$
   b. $C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$
   c. $2\ H_2 + O_2 \longrightarrow 2\ H_2O$
   d. $SO_2 + O_2 \longrightarrow SO_3$
   e. $2\ Al + 6\ HCl \longrightarrow 2\ AlCl_3 + 3\ H_2$

72. When the reaction shown is correctly balanced the coefficients in order are

   Reaction: $HBr + Ca(OH)_2 \longrightarrow CaBr_2 + H_2O$

   a. 2,1,1,1
   b. 1,1,1,2
   c. 2,1,1,2
   d. 2,2,1,1
   e. 2,1,2,2

73. Which sample contains the largest number of **atoms**?
   a. 1.0 mol H_2O
   b. 1.5 mol NaCl
   c. 2.0 mol CH_4
   d. 2.5 mol Au
   e. 3.0 mol Cl_2

74. The formula weight of $Al_2(SO_4)_3$ is _____ grams.
   a. 150.03
   b. 278.02
   c. 315.19
   d. 342.17
   e. 450.14

75. How many molecules are present in 3.25 mol of C_2H_6O?
    a. 3.25
    b. 46.0
    c. 1.85×10^{23}
    d. 6.02×10^{23}
    e. 1.96×10^{24}

76. Which sample contains the **largest** number of ions?
    a. 1.0 mol $(NH_4)_3PO_4$
    b. 1.5 mol $AgNO_3$
    c. 1.0 mol $NaNO_3$
    d. 1.0 mol $Fe(NO_3)_2$
    e. 1.5 mol $SnCl_2$

77. The number of grams in 0.350 mol of Na is
    a. 0.350   b. 8.05   c. 11.0   d. 23.0   e. 65.7

78. $2 \ Al + 6 \ HCl \ ----> \ 2 \ AlCl_3 + 3 \ H_2$
    In the reaction shown, what is the mole ratio that would be used to
    determine the number of moles of H_2 that would be produced when 3.5
    moles of $AlCl_3$ are produced?
    a. $\dfrac{2 \text{ moles Al}}{3 \text{ moles HCl}}$ d. $\dfrac{3 \text{ moles } H_2}{2 \text{ moles } AlCl_3}$

    b. $\dfrac{2 \text{ moles } AlCl_3}{3 \text{ moles } H_2}$ e. $\dfrac{3 \text{ moles } AlCl_3}{2 \text{ moles } H_2}$

    c. $\dfrac{3.5 \text{ moles } AlCl_3}{3 \text{ moles } H_2}$

79. How many grams of C will be consumed when 5.00 grams of Na_2SO_4 react
    according to the balanced reaction shown?
                    $Na_2SO_4 + 2 \ C \ -----> \ Na_2S + 2 \ CO_2$
    a. 0.038 g
    b. 0.211 g
    c. 0.844 g
    d. 1.69 g
    e. 17.1 g

80. The reaction $N_2 + 3 \ H_2 \ ----> \ 2 \ NH_3$ is used to produce ammonia. When
    450. g of hydrogen was reacted with nitrogen, 1575 g of ammonia were
    produced. What is the percent yield of this reaction?
    a. 62.2%
    b. 41.5%
    c. 30.8%
    d. 20.7%
    e. 12.5%

81. The reaction 2 $AgNO_3(aq)$ + $K_2SO_4(aq)$ ----> 2 $KNO_3(aq)$ + $Ag_2SO_4(s)$ is an example of a(an) _____ reaction.
a. acid-base
b. oxidation-reduction
c. precipitation
d. combustion
e. None of the above.

82.   2 $AgNO_3(aq)$ + $K_2SO_4(aq)$ ----> 2 $KNO_3(aq)$ + $Ag_2SO_4(s)$
The net ionic reaction for the balanced equation shown above is
a. Ag^+ + NO_3^- ----> $AgNO_3$
b.   2 K^+ + SO_4^{2-} ----> K_2SO_4
c. K^+ + NO_3^- ----> KNO_3
d.   2 Ag^+ + SO_4^{2-} ----> Ag_2SO_4
e. H^+ + OH^- ----> H_2O

83. All of the reactions shown are oxidation-reduction reactions **except**
a. $N_2(g)$ + $O_2(g)$ ----> 2 NO(g)
b. 2 $Fe_2O_3(s)$ -----> 4 Fe(s) + 3 $O_2(g)$
c. 2 Zn(s) + 2 HCl(aq) ----> $ZnCl_2(aq)$ + $H_2(g)$
d. 2 NaI(aq) + $Cl_2(g)$ ----> 2 NaCl(aq) + I_2
e. $K_2SO_4(aq)$ + $BaCl_2(aq)$ ----> $BaSO_4(s)$ + 2 KCl(aq)

84. The oxidation number of sulfur in calcium sulfate, $CaSO_4$, is
a. +6   b. +4   c. +2   d. 0   e. -2

85. The element chlorine is very reactive as a(an) _____ agent because it readily _____ electrons to form the chloride ion.
a. oxidizing, loses
b. oxidizing, gains
c. reducing, loses
d. reducing, gains
e. None of the above.

86.   Fe(s) + $CuCl_2(aq)$ ----> Cu(s) + $FeCl_2(aq)$

In the redox reaction shown, _____ is oxidized and becomes _____.
a. Fe, Fe^+ b. Fe, Fe^{2+} c. Cu, Cu^{2+} d. Cu^{2+}, Cu   e. None of these.

87. A process or reaction which takes in heat from the surroundings is said to be
a. conservative
b. endothermic
c. exothermic
d. isothermal
e. endergonic

88. Consider the reaction shown.

    $C_3H_8 + 5\ O_2\ ----> 3\ CO_2 + 4\ H_2O + 488$ kcal

    We can say that this reaction is ___ and that the sign of ΔH is ____.
    a. endothermic; positive
    b. exothermic; positive
    c. endothermic; negative
    d. exothermic; negative
    e. exothermic; neither positive nor negative

89. Entropy can be defined as
    a. the amount of energy required to rearrange chemical bonds.
    b. the amount of energy required to initiate a reaction.
    c. the number of chemical bonds which are changed during a reaction.
    d. the state of equilibrium in a system.
    e. the amount of disorder in a system.

90. All of the statements are true for spontaneous reactions **except**
    a. The value of ΔG is less than zero.
    b. The value of ΔG is unaffected by a catalyst.
    c. They are said to be exergonic.
    d. If the enthalpy change is unfavorable, they occur at a high
       temperature.
    e. The reaction rate is determined by the value of ΔG.

91. $2\ Al_2O_3(s)\ -----> 4\ Al(s) + 3\ O_2(g)$ $\Delta G = +138$ kcal

    Consider the contribution of entropy to the spontaneity of this
    reaction. As written, the reaction is _____, and the entropy
    of the system _____.
    a. spontaneous; increases
    b. spontaneous; decreases
    c. non-spontaneous; increases
    d. non-spontaneous; decreases
    e. non-spontaneous; does not change

92. The number of valence electrons in the nitrite ion (NO_2^-) is
    a. 16   b. 17   c. 18   d. 22   e. 23

93. Using VSEPR, predict the shape of SCl_2.
    a. trigonal planar b. bent   c. linear   d. tetrahedral e. pyramidal

94. The position of the equilibrium for a system where $K = 4.6 \times 10^{-15}$ can
    be described as being favored to _____; the concentration of
    products is relatively _____.
    a. the right; large
    b. the right; small
    c. the left; large
    d. the left; small
    e. neither direction; large

95. $$Fe_2S_3 + 4O_2 \longrightarrow 2FeO + 3SO_2$$
    In the balanced reaction shown, the mole ratio of Fe_2S_3 to O_2 is
    a. 5 to 5   b. 1 to 4   c. 4 to 1   d. 1 to 2   e. 4 to 3

96. Which of the following diagrams represents a slow reaction with a
    small negative enthalpy.

    a. A   b. B   c. C   d. D   e. E
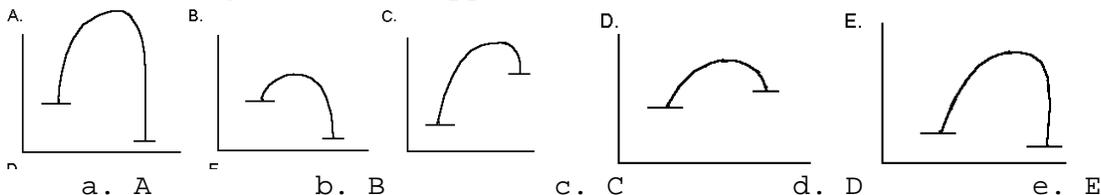
97. A rapid reaction is distinguished by
    a. being unaffected by catalysts.
    b. having a large heat of reaction.
    c. having a small heat of reaction.
    d. having a large value of activation energy.
    e. having a small value of activation energy.

98. In the reaction A + B ----> AB, which of the following will **not**
    increase the rate?
    a. adding A.
    b. adding B.
    c. increasing the temperature.
    d. decreasing the temperature.
    e. adding a catalyst.

99. $$CH_4(g) + NH_3(g) \longrightarrow HCN(g) + 3\,H_2(g)$$
    The equilibrium expression for this reaction is

    a. $\dfrac{[CH_4][NH_3]}{[HCN][H_2]^3}$ b. $\dfrac{[HCN][H_2]^3}{[CH_4][NH_3]}$

    c. $\dfrac{[HCN]^3[H_2]}{[CH_4][NH_3]}$ d. $\dfrac{[CH_4][NH_3]}{[HCN][H_2]}$

    e. $\dfrac{[CH_4][NH_3]}{[HCN]^3[H_2]}$

100. Which change to this reaction system would cause the equilibrium to
     shift to the right?

    $$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g) + heat$$
    a. Addition of a catalyst.
    b. Addition of $NH_3(g)$.
    c. Removal of $H_2(g)$.
    d. Raising the temperature.
    e. Lowering the temperature.

101. Which description best fits a solid?
    a. Definite shape and volume; strong intermolecular attractions
    b. Definite volume; shape of container; moderate intermolecular
       attractions
    c. Definite volume; shape of container; no intermolecular attractions
    d. Volume and shape of container; no intermolecular attractions
    e. Volume and shape of container; strong intermolecular attractions

102. Which transformation is evaporation?
    a. liquid ---> solid   d. solid ---> gas
    b. liquid ---> gas   e. gas ----> solid
    c. solid ---> liquid

103. Which process is exothermic?
    a. gas ---> liquid   d. solid ---> gas
    b. liquid ---> gas   e. None of the above.
    c. solid ---> liquid

104. Which of the assumptions of the kinetic-molecular theory best explains
     the observation that a balloon collapses when exposed to liquid
     nitrogen (which is much colder than a cold winter day!!)?
    a. Gas molecules move at random with no attractive forces between them.
    b. The velocity of gas molecules is proportional to their Kelvin
       temperature.
    c. The amount of space occupied by a gas is much greater than the space
       occupied by the actual gas molecules.
    d. In collisions with the walls of the container or with other
       molecules, energy is conserved.
    e. Collisions with the walls of the container or with other molecules
       are elastic.

105. What would be the new pressure if a 400 mL gas sample at 380 mm Hg is
     expanded to 800 mL with no change in temperature?
    a. 190 mmHg   b. 380 mmHg   c. 570 mmHg   d. 760 mmHg   e. 950 mmHg

106. What will be the new volume when 128 mL of gas at 20.0°C is heated to
     40.0°C while pressure remains unchanged?
    a. 64.0 mL   b. 120. mL   c. 128 mL   d. 137 mL   e. 256 mL

107. How many moles of gas are present in a 10.0 liter sample at STP?
    a. 224 mol   b. 22.4 mol   c. 10.0 mol   d. 2.24 mol   e. 0.446 mol

108. A sample of helium has a volume of 480 mL at 47.0°C and 740 mm Hg. The
     temperature is lowered to 22.0°C and the pressure to 625 mm Hg. What
     is the new volume?
    a. 266 mL
    b. 373 mL
    c. 524 mL
    d. 616 mL
    e. 1214 mL

109. How many grams of O_2 are in a 25.0 L sample at 5.2 atm and 28.0°C?
    a. 1810 g
    b. 168 g
    c. 84.2 g
    d. 5.26 g
    e. 0.164 g

110. Consider a sample of helium and a sample of neon, both at 30.0 degrees Celsius and 1.5 atm. Both samples have a volume of 5.0 liters. Which statement concerning these samples is **not** true?
    a. Each sample contains the same number of atoms of gas.
    b. Each sample weighs the same amount.
    c. Each sample contains the same number of moles of gas.
    d. The density of the neon is greater than the density of the helium.
    e. None of the above.

111. At a high altitude water boils at 95°C instead of 100°C as at sea level because
    a. the atmospheric pressure is greater.
    b. the atmospheric pressure is less.
    c. the climate is cooler.
    d. the vapor pressure of water is greater.
    e. the vapor pressure of water is less.

112. The amount of energy associated with changing a liquid into a gas is called the
    a. heat of vaporization   d. joule
    b. heat of fusion   e. calorie
    c. heat of combustion

113. All of the following statements describing solutions are true **except**
    a. Making a solution involves a physical change.
    b. Solutions are homogeneous.
    c. The particles in a solution are atomic or molecular in size.
    d. Solutions are colorless.
    e. Solutions are transparent.

114. A substance represented by a formula written as $M_XLO_Y•zH_2O$ is called a
    a. colloid   b. solid hydrate   c. solute   d. solvent   e. suspension

115. When a solid dissolves, each molecule is removed from the crystal by interaction with the solvent. This process of surrounding each ion with solvent molecules is called
    a. dilution   b. solvation   c. electrolysis   d. hemolysis   e. crenation

116. The solubility of nitrogen in water exposed to the atmosphere, where the partial pressure of nitrogen is 593 mm Hg, is 5.3×10^{-4} M. At the same temperature, what would be the solubility of pure nitrogen, at a pressure of 760 mm Hg?
    a. 4.1×10^{-4} M
    b. 6.8×10^{-4} M
    c. 1500 M
    d. 2400 M
    e. None of the above.

117. What is the molarity of a solution prepared by dissolving 3.50 mol
NaCl in enough water to make 1.50 L of solution?
a. 0.429 M   b. 2.33 M   c. 5.25 M   d. 87.8 M   e. 137 M

118. How many moles of HCl are present in 75.0 mL of a 0.200 M solution?
a. 15.0 mol
b. 2.67 mol
c. 0.375 mol
d. 0.275 mol
e. 0.0150 mol

119. How many grams of $FeSO_4$ are present in a 20.0 mL sample of a 0.500 M
solution?
a. 65.8 g
b. 6.08 g
c. 3.04 g
d. 1.52 g
e. 0.760 g

120. How many grams of NaOH are needed to make 750 mL of a 2.5% (w/v)
solution?
a. 3.9 g
b. 7.5 g
c. 19 g
d. 20. g
e. 50. g

121. What is the % (w/v) concentration of a solution containing 12 grams of
solute in 400 mL of solution?
a. 1.2 %
b. 3.0 %
c. 4.0 %
d. 6.0 %
e. 12 %

122. How many mL of 0.105 M $AgNO_3$ are needed for an experiment that
requires 0.00510 mol of $AgNO_3$?
a. 0.536 mL
b. 17.8 mL
c. 18.70 mL
d. 20.6 mL
e. 48.6 mL

123. What is the final concentration if 100 mL of water is added to 25.0 mL
of 6.0 M NaCl?
a. 1.2 M
b. 1.5 M
c. 6.0 M
d. 24 M
e. 30 M

124. Which substance is **not** an electrolyte?
a. NaCl   b. HCl   c. NH_4NO_3 d. KOH   e. CH_4

125. Considering 0.1 M solutions of each substance, which contains the **smallest** concentration of ions?
    a. $(NH_4)_3PO_4$
    b. $Ca(NO_3)_2$
    c. $FeSO_4$
    d. K_2CO_3
    e. Na_2SO_4

126. How many grams are in 10.0 mEq of Ca^{2+}?
    a. 0.201 g
    b. 40.1 g
    c. 20.1 g
    d. 0.802 g
    e. 0.401 g

127. If a normal sample contains 4.5 mEq/L of calcium ion, how many mg of calcium are contained in a 25.0 mL blood sample?
    a. 9.0 mg
    b. 5.6 mg
    c. 2.8 mg
    d. 2.3 mg
    e. 1.4 mg

128. Which has the **highest** boiling point?
    a. 0.1 M Na_2SO_4
    b. 0.1 M glucose, $C_6H_{12}O_6$
    c. 0.1 M $MgCl_2$
    d. 0.1 M $Al(NO_3)_3$
    e. pure water

129. The passage of a solvent across a semipermeable membrane because of concentration differences is called
    a. dialysis   b. hemolysis c. hydration d. osmosis   e. solvation

130. If "A" contains 2 % NaCl and is separated by a semipermeable membrane from "B" which contains 10 % NaCl, which event will occur?
    a. NaCl will flow from "A" to "B".
    b. NaCl will flow from "B" to "A".
    c. Water will flow from "A" to "B".
    d. Water will flow from "B" to "A".
    e. None of the above.

131. What is the osmolarity of a 0.20 M solution of KCl?
    a. 0.10 Osmol
    b. 0.20 Osmol
    c. 0.30 Osmol
    d. 0.40 Osmol
    e. 0.80 Osmol

132. Red blood cells are placed in a solution and neither hemolysis nor crenation occurs. Therefore the solution is
    a. hypertonic
    b. hypotonic
    c. isotonic
    d. isotopic
    e. concentrated

133. Which statement concerning Arrhenius acid-base theory is **not** correct?
    a. An Arrhenius acid produces hydrogen ions in water solution.
    b. An Arrhenius base produces hydroxide ions in water solution.
    c. A neutralization reaction produces water plus a salt.
    d. Acid-base reactions must take place in aqueous solution.
    e. None of the above.

134. Which of the following compounds is a salt?
    a. HBr   b. KNO_3 c. H_2SO_4 d. NaOH   e. $C_6H_{12}O_6$

135. A Bronsted-Lowry base is a substance which
    a. produces hydrogen ions in aqueous solution.
    b. produces hydroxide ions in aqueous solution.
    c. donates protons to other substances.
    d. accepts protons from other substances.
    e. accepts hydronium ions from other substances.

136. $C_5H_5N + H_2CO_3 -----> C_5H_6N^+ + HCO_3^-$
    In the reaction shown, the conjugate acid of C_5H_5N is

    a. C_5H_5N
    b. H_2CO_3
    c. $C_5H_6N^+$
    d. HCO_3^-
    e. H_3O^+

137. Which of the following is a diprotic acid?
    a. acetic acid   d. sulfuric acid
    b. hydrochloric acid   e. nitric acid
    c. phosphoric acid

138. Aqueous ammonia reacts with acids because
    a. it contains the hydroxide group.
    b. it is neutral.
    c. it is itself an acid.
    d. it is a salt.
    e. it produces hydroxide ions when placed in water.

139. What is the conjugate acid of HSO_4^-?
    a. SO_4^{2-}
    b. H_2SO_4
    c. H_3O^+
    d. OH^-
    e. H_2SO_3

140. Water and HSO_4^- can either accept protons or donate protons. Such substances are said to be
    a. amphoteric
    b. conjugate
    c. diprotic
    d. monoprotic
    e. triprotic

141. Which net ionic equation correctly represents the neutralization of a solution of barium hydroxide by a solution of nitric acid?
    a. $Ba^{2+} + 2\ NO_3^- ----> Ba(NO_3)_2$
    b. $H^+ + NO_3^- ----> HNO_3$
    c. $Ba^{2+} + 2\ OH^- ----> Ba(OH)_2$
    d. $H^+ + OH^- ----> H_2O$
    e. $Ba(NO_3)_2 + H_2O ----> Ba^{2+} + 2\ NO_3$

142. Hydrogen cyanide, HCN, is a weak acid. Which equation **best** represents its aqueous chemistry?
    a. $HCN + H_2O ----> CN^- + H_3O^+$
    b. $HCN + H_2O ----> H_2CN^+ + OH^-$
    c. $HCN ----> H^+ + CN^-$
    d. $HCN ----> H^- + CN^+$
    e. $H_2O ----> H^+ + OH^-$

143. At 25°C, the value of K_w is
    a. 1.00 b. 1.00×10^{-7} c. 1.00×10^{-14} d. 1.00×10^7 e. 1.00×10^{14}

144. If the concentration of H_3O^+ is 3.5×10^{-3} M, the concentration of OH^- is _____ M.
    a. 2.9×10^{-12}
    b. 1.0×10^{-12}
    c. 1.0×10^{-7}
    d. 3.5×10^{-11}
    e. 10.5×10^{-3}

145. In an aqueous solution that is basic, $[H_3O^+]$ is _____ than 1.0×10^{-7} and _____ than $[OH^-]$.
    a. greater; less
    b. less; greater
    c. greater; greater
    d. less; less
    e. None of the above.

146. If the $[H^+]$ of a water sample is 1×10^{-4} M, the pH of the sample is _____, and the sample is _____.
    a. -4; acidic   d. 10; basic
    b. 4; acidic   e. -10; basic
    c. 4; basic

147. Calculate the hydronium ion concentration in a solution with
pH = 6.35.
a. 7.65 M
b. 6.35 M
c. 4.5 x 10^{-7} M
d. 0.80 M
e. 2.2 x 10^{-8} M

148. To prepare a buffer using sodium dihydrogen phosphate, which of the
following would also be needed?
a. hydrochloric acid
b. ammonium hydroxide
c. ammonium phosphate
d. phosphoric acid
e. sodium hydroxide

149. How many mL of 0.100 M NaOH are needed to neutralize 24.0 mL of
0.150 M HCl?
a. 12.0 mL
b. 18.0 mL
c. 24.0 mL
d. 36.0 mL
e. 48.0 mL

150. What is the normality of a solution containing 49 g of H_2SO_4 in enough
water to make 400 mL of solution?
a. 2.5 N
b. 5.0 N
c. 1.0 N
d. 10 N
e. 0.20 N

151. How many mL of 0.241 M H_2SO_4 will be needed to neutralize a 50.0 mL
sample of 0.191 M KOH?
a. 19.8 mL
b. 31.5 mL
c. 39.6 mL
d. 79.3 mL
e. 126 mL

152. The term nucleon refers to
a. electrons belonging to an atom that undergoes nuclear decay.
b. electrons that are emitted from a nucleus in a nuclear reaction.
c. the nucleus of a specific isotope.
d. both protons and neutrons.
e. None of these.

153. Which is the best description of an alpha particle?
a. charge +2; mass of 4 amu; high penetrating power
b. charge +2; mass of 4 amu; low penetrating power
c. charge -1; mass of 0 amu; medium penetrating power
d. charge -1; mass of 0 amu; high penetrating power
e. charge 0; mass of 0 amu; high penetrating power

154. The emission of a particle from an unstable nucleus is called
    a. mutation
    b. nuclear decay
    c. fission
    d. fusion
    e. translocation

155. Which product is formed by alpha emission from uranium-235? The atomic number of uranium is 92.
    a. $^{231}_{90}Th$ b. $^{233}_{90}Th$ c. $^{235}_{93}Np$ d. $^{235}_{92}U$ e. $^{236}_{92}U$

156. Which product is formed by beta emission from phosphorus-32? The atomic number of phosphorus is 15.
    a. $^{28}_{13}Al$ b. $^{30}_{13}Al$ c. $^{32}_{16}S$ d. $^{32}_{15}P$ e. $^{33}_{16}S$

157. Which nuclear reaction is an example of alpha emission?
    a. $^{235}_{92}U \rightarrow {}^{4}_{2}He + {}^{231}_{90}Th$

    b. $^{75}_{34}Se \rightarrow {}^{0}_{-1}e + {}^{75}_{35}Br$

    c. $^{123}_{53}I \rightarrow {}^{123}_{53}I + energy$

    d. $^{235}_{92}U + {}^{1}_{0}n \rightarrow {}^{139}_{56}Ba + {}^{94}_{36}Kr + 3\,{}^{1}_{0}n$

    e. $^{14}_{7}N + {}^{4}_{2}He \rightarrow {}^{17}_{8}O + {}^{1}_{1}H$

158. Which nuclear reaction is **not** balanced?
    a. $^{10}_{5}B + {}^{4}_{2}He \rightarrow {}^{13}_{7}N + {}^{1}_{0}n$

    b. $^{238}_{92}U + {}^{4}_{2}He \rightarrow {}^{241}_{95}Am + {}^{1}_{0}n$

    c. $^{40}_{18}Ar + {}^{1}_{1}H \rightarrow {}^{40}_{19}K + {}^{1}_{0}n$

    d. $^{14}_{7}N + {}^{4}_{2}He \rightarrow {}^{17}_{8}O + {}^{1}_{1}H$

    e.   None of the above.

159. Rubidium-87 results from the beta emission of what radioisotpe? The atomic number of rubidium is 37.
    a. $^{83}_{34}Se$ b. $^{87}_{35}Br$ c. $^{87}_{36}Kr$ d. $^{87}_{37}Rb$ e. $^{91}_{38}Sr$

160. If the half-life of vanadium-48 is 16 days, it is true that
    a. vanadium-48 is a beta emitter.
    b. the decay rate would be different if the chemical environment of vanadium-48 is changed.
    c. after 32 days a sample of vanadium-48 would have completely decayed.
    d. after 16 days 50 % of a sample of vanadium-48 would have decayed.
    e. vanadium-48 would decay faster in its first half-life than in later half-lives.

161. Approximately how old is a fossil that contains 3.13 % of its original carbon-14? The half-life of carbon-14 is 5730 years.
    a. 2870 years
    b. 5730 years
    c. 11,500 years
    d. 22,900 years
    e. 28,700 years

162. Which of the following is **not** an example of ionizing radiation?
    a. X-rays
    b. gamma rays
    c. beta particles
    d. alpha particles
    e. ultraviolet rays

163. What is the missing reactant in the reaction shown?

$$^{27}_{13}Al \ + \ \underline{\hspace{1cm}} \ \rightarrow \ ^{30}_{15}P \ + \ ^{1}_{0}n$$

    a. $^{4}_{2}He$ b. $^{1}_{1}H$ c. $^{2}_{1}H$ d. $^{1}_{0}n$ e. $^{0}_{-1}e$

164. When a nucleus is bombarded with particles and breaks into two similarly sized nuclei plus one or more small particles, the process is called
    a. fission
    b. fusion
    c. spontaneous decay
    d. induced decay
    e. mutation

165. What is the boiling point of an aqueous solution prepared by adding 150.g of NaCl to 1 kg of water. The boiling point elevation constant for water is 0.51 °C·kg/mol
    a. 1.3 °C
    b. 2.6 °C
    c. 100.3 °C
    d. 101.3 °C
    e. 102.6 °C

166. A vessel contains 0.10mol N_2, 0.20 mol O_2, and 0.30 mol Ne at a total pressure of 2.0 atm. What is the partial pressure of O_2?
    a. 0.067 atm
    b. 0.25 atm
    c. 0.67 atm
    d. 1.25 atm
    e. 3.0 atm

167. What is the pH of a buffer system that contains 1.0 M acetic acid ($HC_2H_3O_2$) and 1.5 M sodium acetate ($NaC_2H_3O_2$)? The Ka for acetic acid is 1.8×10^{-5}.
    a. 1.20
    b. -4.57
    c. 4.57
    d. -4.92
    e. 4.92

Unit 1 `

1. d
2. d
3. b
4. a
5. b
6. e
7. c
8. c
9. a
10. a
11. b
12. d
13. a
14. c
15. b
16. a
17. a
18. b
19. a
20. d
21. e
22. b
23. a
24. a
25. a
26. b
27. c
28. c
29. b
30. d
31. d
32. d
33. b
34. c
35. d
36. b
37. c
38. d
39. b
40. a
41. c
42. d
43. d
44. d
45. CO_3^{2-}
46. NO_3^{-}
47. SO_3^{2-}
48. d
49. b
50. d
51. c
52. c
53. b
54. b

Unit 2 `

55. b
56. e
57. e
58. d
59. e
60. e
61. a
62. c
63. c
64. d
65. b
66. c
67. d
68. e
69. d
70. a
71. d
72. c
73. c
74. d
75. e
76. e
77. b
78. d
79. c
80. a
81. c
82. d
83. e
84. a
85. b
86. b
87. b
88. d
89. e
90. e
91. c
92. c
93. b
94. d
95. b
96. e
97. e
98. d
99. b
100. e

Unit 3

101. a
102. b
103. a
104. b
105. a
106. d
107. e
108. c
109. b
110. b
111. b
112. a
113. d
114. b
115. b
116. b
117. b
118. e
119. d
120. c
121. b
122. e
123. a
124. e
125. c
126. a
127. d
128. d
129. d
130. c
131. d
132. c
133. e
134. b
135. d
136. c
137. d
138. e
139. b
140. a
141. d
142. a
143. c
144. a
145. d
146. b
147. c
148. d
149. d
150. a
151. a
152. d
153. b
154. b
155. a
156. c
157. a
158. b
159. c
160. d
161. e
162. e
163. a
164. a
165. e
166. c
167. e

Formulas, Constants, and Conversion Factors

$N_A = 6.022 \times 10^{23}$
1 lb = 454 g
1 kg = 2.205 lb
1 in = 2.54 cm
39.37 in = 1 m
1.057 qt = 1 L
1 gal = 3.785 L
1 atm = 760 mm Hg = 760 torr = 14.7 psi

Combined Gas Law
$\dfrac{P_1 V_1}{n_1 T_1} = \dfrac{P_2 V_2}{n_2 T_2}$

Expressed as a

Dalton's Law
$P_T = P_1 + P_2 + P_3$ $(P_1 = \%_1 \cdot P_T)$

Henry's Law
$\dfrac{C_1}{P_1} = \dfrac{C_2}{P_2}$

Ideal Gas Law
$PV = nRT$ $R = 0.0821$ L · atm / K · mol

Molar volume of a gas = 22.414 L at STP

$^\circ C = (^\circ F - 32)\, 5/9$
$^\circ F = (^\circ C \times 9/5) + 32$
$K = {}^\circ C + 273$

Heat of a temp. change: $\Delta H_{TC} = m \cdot c \cdot \Delta T$
Heat of a phase change: $\Delta H_{PC} = m \cdot H_{PC}$

$\Delta G = \Delta H - T\Delta S$

density = $\dfrac{\text{mass}}{\text{volume}}$

Dilutions
$M_1 V_1 = M_2 V_2$

% yield = $\dfrac{\text{actual yield}}{\text{theoretical yield}}$ x 100

Titrations
$N_a V_a = N_b V_b$

Molarity (M) = $\dfrac{\text{moles of solute}}{\text{liters of solution}}$

Molality (m) = $\dfrac{\text{moles of solute}}{\text{kg of solvent}}$

Normality (N) = $\dfrac{\text{equivalents of solute}}{\text{liters of solution}}$

Osmolarity = $\dfrac{\text{total moles of solute particles}}{\text{liters of solution}}$
= (M) x (# particles)

Equivalent wt of an ion = $\dfrac{\text{molar mass of ion (g)}}{\text{\# of charges n the ion.}}$

pH = - log $[H_3O^+]$ $[H_3O^+] = 10^{-pH}$

Kw = $[H_3O^+][OH^-] = 1.00 \times 10^{-14}$

Ka = $\dfrac{[H_3O^+][A^-]}{[HA]}$

v/v% = $\dfrac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}}$ x 100%

w/w% = $\dfrac{\text{weight of solute (g)}}{\text{weight of solution (g)}}$ x 100%

w/v% = $\dfrac{\text{weight of solute (g)}}{\text{volume of solution (mL)}}$ x 100%

After decay amt = original amt $(½)^n$

Before decay amt = end amt $(2)^n$

Boiling Point Elevation
$\Delta T_b = k_b \cdot m$

Freezing Point Depression
$\Delta T_f = k_f \cdot m$

Radiation Intensity
$\dfrac{I_a}{I_b} = \dfrac{d_b^2}{d_a^2}$

Solubility Rules for Some Ionic Compounds in Water

Soluble Ionic Compounds use state symbol (aq)	Except Those Containing: use state symbol (s)
1. All sodium (Na^+), potassium (K^+) and ammonium (NH_4^+) salts are SOLUBLE.	None
2. All nitrate (NO_3^-), acetate ($C_2H_3O_2^-$), chlorate (ClO_3^-), and perchlorate (ClO_4^-) salts are SOLUBLE	None
3. All chloride (Cl^-), bromide (Br^-), and iodide (I^-) salts are SOLUBLE.	Pb^{+2}, Ag^+, & Hg_2^{+2} are NOT soluble. Mercury (II) iodide (HgI_2) is also NOT soluble.
4. All fluoride (F^-) salts are SOLUBLE.	Mg^{+2}, Ca^{+2}, Sr^{+2}, Ba^{+2}, & Pb^{+2} are NOT soluble.
5. All sulfate (SO_4^{-2}) salts are SOLUBLE.	Ca^{+2}, Ag^+, Hg_2^{+2}, Sr^{+2}, Ba^{+2}, Pb^{+2} are NOT soluble
Not Soluble Ionic Compounds use state symbol(s)	Except Those Containing: use state symbol (aq)
6. Hydroxide (OH^-) and oxide (O^{-2}) compounds are NOT SOLUBLE	Li^+, Na^+, K^+, Rb^+, Cs^+, NH_4^+, & Ba^{+2}, are soluble. (Ca^{2+} and Sr^{2+} are moderately soluble)
7. Sulfide (S^{-2}) salts are NOT SOLUBLE	Li^+, Na^+, K^+, Rb^+, Cs^+, NH_4^+, & Ba^{+2} are soluble.
8. Carbonate (CO_3^{-2}), phosphate (PO_4^{-3}) and chromate (CrO_4^{-2}), & oxalate ($C_2O_4^{-2}$) salts are NOT SOLUBLE	Li^+, Na^+, K^+, Rb^+, Cs^+, & NH_4^+ are soluble.

Soluble compounds are defined as those that dissolve to the extent of 1 g or more per 100 g water.
NOT Soluble compounds are further classified as:
- Slightly soluble, which dissolve to the extent of 0.01 g to 1 g per 100 g water.
- Difficultly soluble or Insoluble, for which less than 0.01 g per 100 g water will dissolve.

SOLUTIONS MADE FROM THE ABOVE SPECIES, WHEN SOLUBLE, ARE FOUND TO EXIST AS CHARGED PARTICLES AND THUS CONDUCT ELECTRIC CURRENT. THEY ARE CONSIDERED ELECTROLYTES. WRITE THEM IN IONIZED FORM IN AQUEOUS SOLUTIONS.

SUMMARY OF STRONG AND WEAK ELECTROLYTES	
RULE	**EXCEPTIONS**
1. Most acids are weak electrolytes	Common strong acids (strong electrolytes) are HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_3$, and $HClO_4$
2. Most bases are weak electrolytes	Strong base hydroxides (strong electrolytes) are those of Li, Na, K, Rb, Ca, Sr, and Ba.
3. Most soluble salts are strong electrolytes.	Important weakly ionized salts are $HgCl_2$, $Hg(CN)_2$, $CdCl_2$, $CdBr_2$, CdI_2, and $Pb(C_2H_3O_2)_2$.