Chemistry Review

Part 1: Atomic and Molecular Structure
The Periodic table displays the elements in increasing atomic number and shows how periodicity of the physical and chemical properties of the elements relates to atomic structure.

A. Relate the position of an element in the periodic table to its atomic number and atomic mass.

1) Identify the atomic number, atomic mass (weight), atomic symbol and element name. Draw an arrow pointing to each.

6	12.01
C	
carbon	

2) Use your periodic table to find the atomic number of the following elements:
   - a. Rb
   - b. Nb
   - c. Cr
   - d. Au
   - e. Si
   - f. O
   - g.F

3) Use your periodic table to find the atomic mass of the following elements:
   - a. Sg
   - b. Pa
   - c. Ar
   - d. Kr
   - e. Xe
   - f. Rn
   - g. Pb

B. Use the period if table to identify metals, semi metals, non-metals, and halogens.

4) Draw an outline of the periodic table and use it to color code the metals, semimetals, non-metals, and halogens. Color the metals RED, semimetals BLUE, non-metals YELLOW and halogens PURPLE.

C. Use the periodic table to identify alkali metals, alkaline earth metals, and transition metals, trends in ionization energy, electronegativity, and the relative sizes of ions and atoms.

5) Draw an outline of the periodic table and use it to color code the alkali metals, alkaline metals and transition metals. Color the alkali metals GREEN, the alkaline metals ORANGE, and the transition metals PINK.

6) Define ionization energy

7) Define electronegativity.

8) What is a group?

9) What is a period?

10) Explain what happens, in terms of ionization energy, when you move from left to right across the periodic table. What happens when you move from top to bottom in the groups?

11) Explain what happens, in terms of electronegativity, when you move from the left to the right across the periodic table. What happens when you move from top to bottom in the groups?

12) Explain what happens, in terms of relative sizes of ions and atoms, when you move from left to right across the periodic table. What happens when you move from top to bottom in the groups?

D. Use the periodic table to determine the number of electrons available for bonding.

13) What is a valance electron?

14) How many electrons are available for bonding in the elements in group 1A? 2A? 3A? 4A? 5A? 6A? 7A? 8A?

15) Use your periodic table to determine the number of electrons available for bonding in the following elements

a. Fr
b. Sr
c. Ga
d. Sn
e. Pb
f. Ar
g. Cl
h. P
i. O

E. The nucleus of the atom is much smaller than the atom yet contains most of its mass.

16) Draw a Bohr model of a carbon atom.

17) How many protons, neutrons and electrons are there?

18) What subatomic particles are responsible for most of the mass of the carbon atom?

19) What subatomic particles are in the nucleus of the atom?

20) Describe how the nucleus of the atom relates to the overall atom in terms of size and mass.

Part 2 Chemical Bonds
Biological, chemical, and physical properties of matter result from the ability of atoms to form bonds from electrostatic forces between electrons and protons and atoms and molecules.

A. Atoms combine to form molecules by sharing electrons to form covalent or metallic bonds or by exchanging electrons to form ionic bonds.

21. What is an ionic bond?
22. What is a covalent bond?
23. Describe how a metallic bond works, using a diagram.

B. Chemical bonds between atoms in molecules such as H_2, CH_4, NH_3, H_2CCH_2, N_2, Cl_2, and many large biological molecules are covalent.

24. Explain why many large biological molecules are covalent. What is the chemical formula for glucose? Is the glucose molecule ionic or covalent?

C. Salt crystals, such as NaCl, are repeating patterns of positive and negative ions held together by electrostatic attraction.

25. Describe what a sodium chloride crystal looks like in terms of the positive and negative ions. Which ions are positive and which ions are negative?

26. Describe how a crystal lattice is formed.

D. The atoms and molecules in liquids move in a random pattern relative to one another because the intermolecular forces are too weak to hold the atoms or molecules in a solid form.

27. Draw how close together molecules are in a solid.

28. Draw, in the box, how close together molecules are in a liquid.

29. Draw, in the box, how close together molecules are in a gas.

30. Which state of matter has the most energy, a solid, a liquid or a gas?

31. Which state of matter has the least energy, a solid, a liquid, or a gas?

E. Draw Lewis Dot structures.

32. Use your periodic table to determine the number of valence electrons. Then draw the Lewis Dot Diagrams for the following atoms and molecules.

a. K
b. Ca
c. O
d. Cl
e. Xe
f. CS_2
g. H_2O
h. PH_3

Part 3 Conservation of Matter and Stoichiometry
The conservation of atoms in chemical reactions leads to the principle of conservation of matter and the ability to calculate the mass of products and reactants.

A. Describe chemical reactions by writing balanced equations.

33. Balance the following equations

a. ____$KClO_3$ $\rightarrow$ ____ KCl + ____O_2
b. ____CaC_2 + ____$H_2O$$\rightarrow$____$C_2H_2$ +____ $Ca(OH)_2$
c. ____Cu + ____$AgNO_3$ $\rightarrow$ ____$Cu(NO_3)_2$ + ____Ag
d. ____Zn + ____HCl $\rightarrow$ ____$ZnCl_2$ + ____H_2
e. ____K_2CO_3+ ____HCl $\rightarrow$ KCl + ____H_2O ____ CO_2
f. ____C_4H_{10} + ____O_2 $\rightarrow$ ____CO_2 + ____HO
g. ____Zn + ____H_3PO_4 $\rightarrow$ ____$Zn_3(PO_4)_2$ + ____H_2
h. ____$Hg(NO_3)_2$ + ____NaI $\rightarrow$ ____HgI_2 + ____$NaNO_3$
i. ____Ag + ____HNO_3 $\rightarrow$ ____NO + ____$AgNO_3$ + ____H_2O
j. ____N_2H_4+ ____$H_2O_2$$\rightarrow$ ____N_2 + ____H_2O
k. ____NH_3+ ____$O_2$$\rightarrow$ ____NO_2 + ____H_2O
l. ____Al + ____Cl_2 $\rightarrow$ ____$AlCl_3$

B. The quantity of one mole is set by defining one mole of carbon 12 atoms to have a mass of exactly 12 grams.

34. How many atoms are in one mole of carbon?

C. One mole equals 6.0 x 10^{23} particles (atoms or molecules).

35. What is Avogadro's number?

D. Determine the molar mass of a molecule from its chemical formula. Convert the mass of a molecular substance to moles, number of particles, or volume of gas at standard temperature and pressure.

36. Determine the molar masses of the following substances.
    a. NaCl
    b. $(NH_4)_2SO_4$
    c. $NaHCO_3$

37. Determine the number of moles in each of the following
    a. 25.5 g Ag
    b. 200 g S
    c. 100 g Zn
    d. 1.0 kg F

38. A sample of ethanol (C_2H_5OH) has a mass of 45 grams.
    a. How many carbon atoms does the sample contain?
    b. How many hydrogen atoms does the sample contain
    c. How many oxygen atoms are present

39. What is Avagadro's principle?
40. How many moles of gas are in a sample of nitrogen gas that occupies 4 liters at STP?
41. How much space does 2.06 moles of gas occupy at STP?

E. Calculate the masses of reactants and products in a chemical reaction from the mass of one of the reactants or products and the relevant atomic masses.

42. Given the following equation: $2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$, Show what the following mole rations should be.

    a. $C_4H_{10} : O_2$
    b. $O_2 :\ CO_2$
    c. $O_2 : H_2O$
    d. $C_4H_{10} : CO_2$
    e. $C_4H_{10} : H_2O$

Given the following equation $2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$

43. How many moles of O_2 can be produced by letting 12.00 moles of $KClO_3$ react?

Given the following equation: $2K + Cl_2 \rightarrow 2KCl$

44. a) How many grams of KCl is produced from 2.50 g of K and excess Cl_2? b) From 1.00 g of Cl_2 and excess K?

Given the following equation $Na_2O + H_2O \rightarrow 2\ NaOH$

45. How many grams of NaOH is produced from 1.20 X 10^2 grams of NaO? How many grams of Na_2O are required to produce 1.60 X 10^2 grams of NaOH

Given the following equation $8\ Fe + S_8 \rightarrow 8\ FeS$

46. What mass of iron is needed to react with 16.0 grams of sulfur? How many grams of FeS are produced?

Given the following equation: $2NaClO_3 \rightarrow 2NaCl + 3\ O_2$

47. 12.00 moles of $NaClO_3$ will produce how many grams of O_2? How many grams of NaCl are produced when 80.0 grams of O_2 are produced?

Given the following equation: $Cu + 2\ AgNO_3 \rightarrow C\ (NO_3)_2 + 2\ Ag$

48. How many moles of Cu are needed to react with 3.50 moles of $AgNO_3$? If 89.5 grams of Ag were produced, how many grams of Cu reacted?

49. Molten iron and carbon monoxide are produced in a blast furnace by the reaction of iron (III) oxide and coke (pure carbon). If 25.0 kilograms of pure Fe_2O_3 is used, how many kilograms of iron can be produced? This reaction is: $Fe_2O_3 + 3\ C \rightarrow 2\ Fe + 3\ CO$.

50. The average human requires 120.0 grams of glucose ($C_6H_{12}O_2$) per day. How many grams of CO_2 (in the photosynthesis reaction) are required for this amount of glucose?

    The photosynthetic reaction is $6\ CO_2 + 6\ H_2O \rightarrow C_6H_{12}O_2 + 6\ O_2$

Part 4 Gasses and their Properties
The Kinetic molecular theory describes the motion of atoms and molecules and explains the properties of gases.

A. The random motion of molecules and their collisions with a surface create the observable pressure on that surface.

51. Draw a diagram that helps to explain how the random motion of molecules creates pressure inside a balloon.
52. Explain why molecules move randomly and which kind of molecule moves the most, those in a solid, liquid, or gas?

B. The random motion of molecules explains the diffusion of gases.

53. Your favorite chemistry teacher got a new fragrance for their birthday. When this chem teacher put that fragrance on in the morning, they got a bit overzealous with the quantity applied. You sit in the back of the room in the back and notice that all the kids in front are commenting on how strong the odor is. You can't smell the fragrance until about a minute later than they do. What is happening? Describe how molecules move throughout the room.

C. Apply the gas laws to the relations between the pressure, temperature, and volume of any amount of an ideal gas or any mixture of ideal gases.

54. What is Boyle's law? Describe in words and list the equation.
55. What is Charles' law? Describe it in words, and list the equation.
56. What is Gay-Lussac's law? Describe it in words and list the equation.
57. What is the ideal gas law? Describe it in words and list the equation.
58. What is the combined gas law? Describe it in words and list the equation.
59. Solve the following Boyle's law problems.
    a. A gas occupies 12.3 liters at a pressure of 40.0 mm Hg. What is the volume when the pressure is increased to 60.0 mm Hg?
    b. If a gas at 25.0 degrees C occupies 3.60 liters at a pressure of 1.00 atm, what will be its volume at a pressure of 2.50 atm?
    c. To what pressure must a gas be compressed in order to get into a 3.00 cubic foot tank the entire weight of a gas that occupies 400.0 cu. Ft. at standard pressure?
    d. A gas occupies 1.56 L at 1.00 atm. What will be the volume of this gas if the pressure becomes 3.00 atm?

60. Solve the following Charles' law problems.
    a. Calculate the decrease in temp when 2.00 L at 20.0 degrees C is compressed to 1.00 L
    b. 600.0 mL of air is at 20.0 degrees C. What is the volume at 60 degrees C?
    c. A gas occupies 900.0 mL at a temperature of 27.0 degrees C. What is the volume at 132.0 degrees C?
    d. Solve the following Gay-Lussac's law problems.

61. Solve the following Gay-Lussac's law problems.
    a. Determine the pressure change when a constant volume of gas at 1.00 atm is heated from 20.0 degrees C to 30.0 degrees C.
    b. A gas has a pressure of 0.370 atm at 50.0 degrees C. What is the pressure at standard temperature?
    c. A gas has a pressure of 699.0 mm Hg at 40.0 degrees C. What is the temperature at standard pressure?
    d. If a gas is cooled from 323.0 K to 273.15 K and the volume is kept constant what final pressure would result if the original pressure was 750.0 mm Hg?

62. Solve the following Ideal Gas Law Problems
    a. How many moles of gas are contained in 890.0 mL at 21.0 C and 750.0 mm Hg pressure?
    b. 1.09 g of H_2 is contained in a 2.00 L container at 20.0 C what is the pressure in this container in mm Hg?
    c. Calculate the volume of 3 moles of a gas wil occupy at 24.O C and 762.4 mm Hg.
    d. What volume will 20.0 of Argon occupy at STP?

63. Solve the following combined Gas Law problems
    a. A gas has a volume of 800 mL at minus 23.00 C and 300 torr. What would the volume of the gas be at 227.0 C and 600 torr of pressure?
    b. 500 liters of a gas are prepared at 700.0 mm Hg and 200.0 degrees C. The gas is placed into a tank under pressure. When the tank cools to 20.0 C, the pressure of the gas is 30 atm. What is the volume of the gas?
    c. What is the final volume of a 400.0 mL gas sample that is subjected to a temperature change from 22.0 C to 30 C and a pressure change from 760 mm Hg to 360 mm Hg?
    d. What is the volume of gas at 2.00 atm and 200 K if its original volume was 300 L at .250 atm and 400 K.

D. The value of standard temperature in both temp and pressure (STP).

64. What is the value of standard temperature in both Celsius and Kelvin?
65. What is the value of standard pressure in mmHg? In Torr? In atm? In kPa? In Hg?

E. Convert between Celsius and Kelvin temperature scales.

66. Convert the following Celsius temperatures to Kelvin.
    a. 67 °C
    b. 102 °C
    c. 2 °C
    d. 0 °C
    e. 175 °C

67. convert the following Kelvin temperatures to Celsius
    a. 375 K
    b. 456 K
    c. 0 K
    d. 216 K
    e. 13 K

F. There is no temperature lower than 0 Kelvin.

68. Explain what is happening at the molecular level when something is at zero Kelvin.
69. How does this help explain why there is no temperature lower than zero Kelvin?
70. What kind of energy is being measured when we measure in Kelvin?

Part 5 Acids and Bases

Acids, bases, and salts are three classes of compounds that form ions in water solutions.

A. The observable properties of acids, bases, and salt solutions:

71. List the physical and chemical properties of acids.
72. List the physical and chemical properties of bases
73. List the physical and chemical properties of salts.
74. When an acid and a base combine, what does it form?

B. Strong acids and bases fully dissociate and weak acids and bases partially dissociate.

75. Describe what happens at the molecular level when a weak acid is dissolved in water.
76. Describe what happens at the molecular level when a strong acid is dissolved in water.

C. Use the pH scale to characterize acid and base solutions.

77. Draw a pH scale from pH 0 to pH 14. Label the acidic values and the basic solutions.
78. Use the pH scale to classify the following solutions as either acidic or basic.
    a. Lemonade pH 5.2
    b. Pepsi pH 3.4
    c. Acetic Acid pH 3.0
    d. Hydrofluoric Acid pH 1.0
    e. Draino pH 13.4
    f. Water pH O

Part 6 Solutions

Solutions are homogenous mixtures of two or more substances.

A. The definitions of solute and solvent.

79. Define solute
80. Define solvent

B. Describe the dissolving processes at the molecular level by using the concept of random molecular motion.

81. Define random molecular motion
82. Describe in detail what is happening to the molecules when sugar is dissolved in water. Draw a diagram that shows the sugar dissolving in water. Use different colors to represent the molecules, to help you explain more clearly.

C. Temp, pressure, and surface area affect the dissolving process.

83. You need to add 150 grams of sugar cubes to 500 ml of water and you need the sugar to dissolve quickly. Describe what you can do, in terms of temperature, pressure, and surface area to affect the dissolving rate of the sugar.

D. Calculate the concentration of a solute in terms of grams per liter, molarity, parts per million, and percent composition.

84. Determine the **grams per liter** of sugar in a 250 ml of soda that contains 47 grams of sugar.
85. Calculate the **percent composition by mass** for a solution that contains 24 grams of NaCl dissolved in 100g of water?
86. What is the **molarity** of a solution that contains 2.7 grams of NaCl dissolved in 500 ml of water?
87. What kind of things are measured in **parts per million**? Describe at least two.

Parts 7 Chemical Thermodynamics
Energy is exchanged or transformed in all chemical reactions and physical changes of mater.

A. Describe temperature and heat flow in terms of the motion of molecules (or atoms).

88. When you heat a metal cute and put it into water, the water will warm up. Describe what is happening in terms of the heat flow. Is heat flowing from the water to the metal? Is heat flowing from the metal to the water?
89. You are 98.6 degrees Fahrenheit. When you sit on a metal bench in the winter, your bum feels cold. Explain what is happening in terms of heat flow? Where is the heat moving? Why does your bum feel chilly?

B. Chemical processes can either release (exothermic) or absorb (endothermic) thermal energy.

90. When NaOH is combined with water in a beaker, the solution gets hot. Is this an exothermic or endothermic reaction?
91. When energy is released from a reaction is the value for enthalpy (change H) positive or negative?
92. When energy is absorbed during a reaction is the value for enthalpy (change h) positive or negative?

C. Energy is released when a material condenses or freezes and is absorbed when a material evaporates or melts.

93. You are making ice cream in a cup that is sitting in an ice bath. The liquid in the cup freezes over time. Is the ice cream absorbing or releasing energy?
94. You are melting crayons in a pot. Is the new "crayon soup" releasing or absorbing heat?

D. Solve problems involving heat flow and temperature changes using known values of specific heat and latent heat of phase change. Use your book to look up the values for ΔH (fusion and vaporization) and the values for Cp (specific heat).

95. Calculate the heat required to melt 25.7 grams of solid methane at its melting point.
96. How much heat is evolved when 175 grams of ammonia gas condenses to a liquid at its boiling point.
97. If the temp of a sample of water increases from 20 C to 46.6 C as it absorbs 5650 J of heat. What is the mass of the sample?
98. If the temp of 34.4 grams of ethanol increases from 25 C to 78.8 C, how much heat has been absorbed by the ethanol?

98-2. How much heat is required to completely melt a 7.8 g piece of copper metal from a 25.0 C solid to a liquid with a temp of 1083 C?

99. From your notes, draw the heating curve of water here and explain what is happening at each point in the curve, (phases/phases changes) also, explain what equation you would use at each part of the curve to calculate heat then, explain why we need to use 2 different equations to calc heat.

Part 8 Reaction rates
Chemical rates depend on factors that influence the frequency of collision of reactant molecules.

A. The rate of reaction is the decrease in concentration of reactants or the increase in concentration of products with time.

100. Define reaction rate
101. Define collision theory
102. Define activation energy
103. What is the general equation for calculating reaction rates?

B. Reaction rates depend on factors such as concentration, temperature, and pressure.

104. Describe how concentration affects reaction rates.
105. Describe how temperature affects reaction rates.
106. Describe how pressure affects reaction rates.

C. A catalyst plays a role in increasing reaction rates.

107. Define Catalyst. What does a catalyst do to the reaction rate?

Part 9 Chemical Equilibrium
Chemical equilibrium is a dynamic process at the molecular level.

A. Use LeChatelier's principle to predict the effect of changes in concentration, temperature, and pressure.

108. Explain what LeChatelier's principle's is.
109. Solve the following problems by LeChatelier's principle.

$N_2 + 3\ H_2 \leftrightarrow 2\ NH_3$ + heat
a. remove NH_3 gas
b. decrease pressure (volume increases)
c. add N_2 gas
d. Increase temperature

$2\ SO_2 + O_2 \leftrightarrow 2\ SO_3$ + heat
e. increase SO_2 concentration
f. increase temperature
g. remove O_2

CO_2 + C + Heat $\leftrightarrow$ 2 CO
h. increase temperature
i. increase CO concentration
j. decrease pressure

$H_2 + Cl_2 \leftrightarrow 2\ HCl$ + heat
k. increase H_2 Concentration
l. increase pressure

$N_2 + O_2$ + heat $\leftrightarrow 2\ NO$
m. decrease O_2 concentration
n. add catalyst

$PCl_3 + Cl_2 \leftrightarrow PCl_5$ + Heat
o. increase Cl_2 concentration
p. decrease pressure

$CO_2 + H_2$ + heat $\leftrightarrow CO + H_2O$
q. decrease pressure
r. add a catalyst

N_2O_4 + Heat $\leftrightarrow 2\ NO_2$
s. decrease pressure
t. remove N_2O_4

B. Equilibrium is established when forward and reverse reaction rates are equal.

110. Define Equilibrium
111. Explain what is happening in terms of reaction rates when a reaction is at equilibrium

Part 10 organic chemistry and biochemistry
The bonding characteristics of carbon allow the formation f many different organic molecules of varied sizes, shapes, and chemical properties and provide the biochemical basis of life.

A. Large molecules (polymers) such as proteins, nucleic acids, and starch, are formed by repetitive combinations of simple subunits.

112. Define polymer.
113. Define monomer.
114. What is the monomer of protein?
115. What is the monomer of nucleic acids (DNA and RNA)?
116. What is the monomer of starch?

B. The bonding characteristics of carbon result in the formation of a large variety of structures ranging from simple hydrocarbons to complex polymers and biological molecules.

117. Why is carbon able to form so many different compounds?
118. How many valence electrons are in carbon?
119. What is a hydrocarbon
120. What is an organic compound and how is it different from an inorganic compound?

C. Amino acids are the building blocks of proteins.

Part 11 Nuclear processes

Nuclear processes are those in which an atomic nucleus changes, including radioactivity decay of naturally occurring and human made isotopes, nuclear fission, and nuclear fusion.

A. Protons and neutrons in the nucleus are held together by nuclear forces that overcome the electromagnetic repulsion between the protons.

121. What kind of force holds protons and neutrons together in the nucleus?
122. What is the charge on a proton? What is the mass?
123. What is the charge on a neutron? What is the mass?
124. What is the charge on an electron? What is the mass?

B. The energy release per gram of material is much larger in nuclear fusion or fission reactions than in chemical reactions. The change in mass (calculated by $E = mc^{2)}$ is small but significant in nuclear reactions.

125. How are nuclear fusion and fission reactions different from chemical reactions? What subatomic particles are involved in these kinds of reactions?
126. Describe what happens in a fission reaction.
127. Describe what happens in a fusion reaction.
128. Define the variables in the equations $E=mc^2$.

C. Some naturally occurring isotopes of elements are radioactive, as are isotopes formed in nuclear reactions.

129. Define isotope.
130. Define radioactivity.
131. What are daughter products?
132. What is an isotope of carbon? How many neutrons does it have?

D. The three most common forms of radioactivity decay (alpha, beta, and gamma) and know how the nucleus changes in each type of decay.

133. What is alpha decay? How does the nucleus change during alpha decay?
134. What is beta decay? How does the nucleus change during beta decay?
135. What is gamma decay? How does the nucleus change during gamma decay?

E. Alpha, beta, and gamma radiation produce different amounts and kinds of damage in matter and have different penetrations.

136. What kind of material can alpha radiation penetrate? What kind of damage can it do to matter and living tissue?
137. What kind of material can beta radiation penetrate? What kind of damage can it do to matter and living tissue?
138. What kind of material does gamma and alpha radiation penetrate? What kind of damage can it do to matter and living tissue?