**Things to Know For the Chemistry Final**

20. ***Binary compounds*** are substances made up of only *two* kinds of atoms.

“Ternary” compounds contain three (or more) kinds of atoms.

*Which substance is a binary compound?*

*Ammonia magnesium potassium nitrate sucrose*

21. ***Diatomic molecules*** are elements that form two atom molecules in their natural form at STP.

*What are the 7 diatomic elements? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_*

*Which element is a diatomic liquid at STP? chlorine fluorine bromine iodine*

24. ***Solutions*** are the best examples of ***homogeneous mixtures***. They have **two** sets of properties.

25. ***Heterogeneous mixtures*** have discernible components and ***are not*** uniform throughout.

*Air is classified chemically as a(n)*

*Substance compound element solution*

26. A ***solute*** is the substance being dissolved; the ***solvent*** is the substance that dissolves the solute.

*NaCl (s) is added to water.*

*The solute is ….. the solvent is …… the solution is ……..*

30. Calculations bases on quantitative relationships and balanced chemical equations are called ***stoichiometry.***

Stoichiometric calculations are bases on the **mole ratios** between any two substances in the reaction.

*What is the mole ratio between oxygen and water in the following reaction? 2H2 + O2 🡪 2H2O*

31. Use the **mole** ***map*** to help you solve conversions (don’t forget mole ratios!)

between moles, grams, numbers of molecules/atoms, and liters of gases at STP.

*Given the reaction CH4 + 2O2 --> CO2 + 2H2O,  
what amount of carbon dioxide is produced by the reaction of 1 mole of CH4?*

*1 gram 1 liter 1 mole 22 grams*

32. An **empirical formula** is the simplest mole ratio among the elements in a compound.

Use the mole map to convert percent (mass) to moles.

*Find the empirical formula of a compound composed of 75% carbon and 25% hydrogen.*

32.5 A **molecular formula** is the true formula of a molecular compound. Use the molecular (molar) mass of the

compound and its empirical formula mass to find the true molecular formula.

*Find the molecular formula of the compound from problem 32, if its molecular mass is 48.15 g/mol*

33. ***Lewis*** ***Electron dot structures*** represent covalently bonded molecules formed through the sharing

of valence electrons between atoms. It also shows the lone-pair(s) of electrons that may affect the shape and

polarity of the molecules.

*Draw the Lewis structures for these common molecules H2 H2O NH3 CH4*

36. ***Coefficients*** are written in front of the formulas of reactants and products to balance chemical

equations. They give the ratios of reactants and products in a balanced chemical equation.

*……..Na + …….Cl2 🡪 ………NaCl*

37. **Chemical formulas** are written so that the charges of **cations** and **anions** neutralize (cancel) one another.

*calcium phosphate*: Ca2+ PO43- = ………… (criss-cross)

38. When naming **binary ionic compounds**, write the name of the positive ion (cation) first,

followed by the name of the negative ion (anion) with the name ending in “-ide.”

*CaCl2 …………….. MgS ……………..*

39. When naming compounds containing **polyatomic ions**, keep the name of the

polyatomic ion the same as it is written on your reference sheet.

*NH4Cl ………. copper (I) nitrate ……….*

40. **Roman numerals** are used to show the positive oxidation number of the cation if it has more than

one positive oxidation number

*FeO: ……………………. Nickel (III) sulfate: ……………..*

41. When naming a **binary molecular compound,** the first element is named using the name of the element.

The second element always end in “–ide.” Indicate the number of atoms using the prefix…

1 mono 2 di 3 tri 4 tetra 5 penta 6 hexa 7 hepta 8 octa 9 nona 10 deca

If the first element has only one atom, don’t use the mono. *What is the name of the following molecular*

*compounds? CO2 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ N2O­­­­­­­­­­­­­­­\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_*

41. ***Physical changes*** do not form new substances.

They merely change the appearance of the original material. (The melting of ice) H2O (s) 🡪 H2O (l)

42. ***Chemical changes*** result in the formation of new substances or the product(s) of a ***chemical reaction***.

*Which process is an example of a chemical change?*

*the melting of ice the electrolysis of water the boiling of water*

43. ***Reactants*** are on the left side of the reaction arrow and ***products*** are on the right.

*What are the names of the reactants in this neutralization reaction?*

*HCl(aq) + NaOH(aq) 🡪 NaCl(aq) + H2O(l)*

44. ***Temperature*** is a measure of average kinetic.

*Which sample has the highest average kinetic energy?*

*H2O (l) at 0oC H2O (s) at 0oC CO2 (g) at STP Mg (s) at 298K*

45. ***Exothermic reactions*** release energy (energy is a product of the reaction) while

***Endothermic reactions*** absorb energy and the *energy is a reactant* in the reaction.

*Given the reaction: CH4 (g) + 2 O2 (g) → 2 H2O (g) + CO2 (g) + heat*

*What is the overall result when CH4 (g) burns according to this reaction?*

*Energy is absorbed Energy is released*

46. *Only* **coefficients** can be changed when balancing chemical equations!

*Given the unbalanced equation: Al + O2 = Al2O3*

*When this equation is balanced using the smallest whole numbers, what is the coefficient of Al?*

*1 2 3 4*

47. ***Synthesis reactions*** occur when two or more reactants combine to form a single product.

*Finish and balance the following synthesis reaction*

*Na (s) + Cl2(g) 🡪 ……*

48. ***Decomposition reactions*** occur when a single reactant forms two or more products

*Finish and balance the following decomposition reaction*

*CaCO3(s) 🡪 CaO(s) + ….…*

49. ***Single replacement reactions*** occur when one element replaces another element in a compound.

*Which equation below represents a reaction classified as a "single replacement" reaction?*

*2 H2 + O2 --> 2 H2O*

*Pb(NO3) 2 + K2CrO4 --> 2 KNO3 + PbCrO4*

*HCl + NaOH --> H2O + NaCl*

*Cu + Zn(NO3) 2 --> Zn + Cu(NO3) 2*

*CaCO3 --> CO2 + CaO*

50. ***Double replacement (displacement) reactions*** occur when two compounds react to form two new compounds.

*Potassium sulfide (K2S) is mixed with lead acetate(Pb(C2H3O2)2). Which of the following products is*

*expected? PbSO4 K2S K3PO4 PbS*

50.5 ***Combustion reactions*** occur when a hydrocarbon burns in the presence of oxygen gas to produce carbon

dioxide gas and water vapor. Write the combustion reaction for the combustion of methane gas (CH4)

*CH4 (g) + \_\_\_\_\_ 🡪 \_\_\_\_\_\_ + \_\_\_\_\_\_\_ \*balance*

51. The masses (and energy) of the reactants in a chemical equation is always equal to

the masses (and energy) of the products. “***Law of Conservation of Mass (and Energy)***.”

*For the reaction: CaCO3 --> CO2 + CaO*

*If 20.0g of CaCO3 decomposes to for 19.5 g of CaO, how many grams of CO­2 gas is formed? \_\_\_\_\_\_*

52. The gram formula mass (**molar mass**) of a substance is the sum of the atomic masses of all the atoms in it. H2SO4 = \_\_\_\_\_\_ g/mole

2 x H = 2 x ………g = ………g

1 x S = 1 x ………g = ………g

4 x O = 4 x………g = ………g

53. Know how to calculate the **percentage composition** of a compound. (Formula is on the last pages.)

*Find the percent by mass of oxygen in CaCO3*.

54. ***6.02 x 1023*** is called ***Avogadro’s number*** and is the number of particles in ***1 mole*** of a substance.

Equal volumes of gases contain an equal number of molecules.

*Under similar conditions,* *which sample contains the same number of moles of particles as*

*1 liter of O2 (g)? 1 L Ne(g) 0.5 L SO2 (g) 2 L N2 (g) 1 L H2O(l)*

55. Know how to convert an **empirical formula** into a **molecular formula**.

*A compound has the empirical formula NO2. Find its molecular formula if the molar mass = 92g.*

*N2O NO2 N2O4* N*2*O

56. The **kinetic molecular theory** explains the behavior of matter as particles with energy and motion.

57. The particles in a ***solid*** are rigidly held together, closely packed in a **lattice** arrangement.

*Which of the following has a regular geometric arrangement at 298 K and 1.0 atm?*

*Br2 (l) CO2 (g) Mg (s) H2O (l)*

58. ***Solids*** have a definite shape and volume.

#### In what region of the graph below would you only find molecules with definite shape and volume?

59. ***Liquids*** have closely-spaced particles that easily slide past one another; they have no definite shape,

but have a definite volume.

60. ***Gases*** have widely-spaced particles that are in random motion (collide with container to create pressure).

60.5 Fill in the diagram (with high or low) to show how intermolecular forces influence the **volatility**, **vapor**

**pressure**, and **boiling point** of a substance



61. ***Gases*** are easily compressed and have no definite shape or volume.

#### In what region of the graph below would you only find a sample with no definite shape or volume?

|  |  |
| --- | --- |
| 62. Be able to read and interpret heating/cooling curves as pictured below.    *During which* ***interval*** *on the graph are solid and*  *liquid in equilibrium?* |  |

63. Substances that ***sublime*** turn from a solid directly into a gas.

They have very weak attractive or intermolecular forces. (examples include CO2 & I2)

*What type of intermolecular forces exist between molecules of CO2 or I2?*

65. “***STP***” means “***S***tandard ***T***emperature and ***P***ressure.”

#### These conditions define STP P = \_\_\_\_\_atm = \_\_\_\_\_\_kPa T = \_\_\_\_\_\_\_K

66. Degrees Kelvin = C + 273 ***Temperature*** is a measure of the **kinetic energy** of the particles in matter.

*Room temperature = 25oC = …….K Boiling point of helium = 4 K = ……….oC*

71. *Always* *use Kelvins* for temperature when using the ***combined gas law***. P1V1T2 = P2V2T1

*Set up the equation to calculate the volume of 50. mL of methane gas collected at STP*

*when the pressure rises to 2.4 atm and the temperature drops to 240 K.*

72. As the ***pressure*** exerted on a gas increases, the ***volume*** decreases proportionally.

*25 L of a gas is held at 1.2 atm pressure. Find the new volume if pressure drops to 0.80 atm at constant temperature.*

73. As the ***pressure*** on a gas increases, ***temperature*** increases.

*A sample of gas exerts a pressure of 220. kPa at 373 K. Find the pressure at 273 K at constant volume.*

74. As the ***temperature*** of a gas increases, ***volume*** increases.

*15 mL of oxygen gas is collected at 0oC. Find the volume at 50oC at constant pressure.*

75. ***Real gas*** particles have volume and are attracted to one another. They don"t always behave like ***ideal gases***.

Lighter gases (with weaker attractive forces) are often most ideal.

*Which of the following is the most ideal gas?*

*He Ne Ar Kr*

76. Real gases behave more like ideal gases at *low pressures and high temperatures.*

77. According to ***Avogadro’s law*** at constant temperature and pressure, the volume of a gas is directly

proportional to the number of moles. This is true for any gas.

*At STP 22.4 L of any gas = 1 mole, what is this equality called?*

78. The ***Ideal Gas Law*** relates the pressure (atm), volume (L), temperature (K), and amount of gas particles

(moles) of a gas. The formula is PV = nRT where R is 0.0821 L•atm/mol•K

*What is the volume occupied by 36.0g of water vapor at 125˚C and 0.999atm?*

94. The last digit of an element’s group number is equal to its ***number of valence electrons***.

*Which contains the greatest number of valence electrons?*

*Ca Ge Se Kr*

95. Draw one dot for each valence electron when drawing an element’s or ion’s ***Lewis electron dot diagram***.

*Which dot model would contain the fewest dots as valence electrons?*

*Ca Ge Se Kr*

96. **Metallic bonds** can be thought of as a crystalline lattice of kernels surrounded by a “sea” of mobile valence

electrons. *Metallic bonding occurs between atoms of*

*sulfur sodium fluoride sodium carbon*

97. Atoms are most stable when they have 8 valence electrons (an ***octet***) and tend to form ions to obtain such a

configuration of electrons.

*Which of the following atoms forms a stable ion that does* ***not*** *have an octet structure?*

*Li F Na Cl*

98. ***Covalent bonds*** form when two atoms ***share*** a pair of electrons.

*How many covalent bonds are found in a nitrogen (N2) molecule?*

99. ***Ionic bonds*** form when one atom ***transfers*** an electron to another atom when

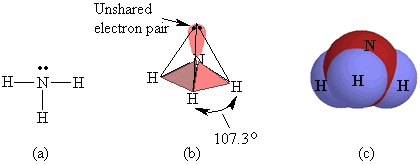
forming a bond with it.

*Which substance exhibits ionic bonding rather than covalent bonding?*

###### CO2 N2O4 SiO2 CaBr2 C6H12O6

100. **Lewis** **Dot models** may be used to represent the formation of ions or covalent molecules and help

determine their shape using VSEPR theory and their polarity.



*Given the Lewis structure in figure (a) and the diagram figure (b), what is the shape of the ammonia*

*molecule? Is this molecule polar or nonpolar?*

101. ***Nonpolar covalent bonds*** form when two atoms of the *same element* bond together.

102. ***Polar covalent bonds*** form when the electronegativity difference between two bonding atoms is between

0.4 and 1.7.

*Which of the following combinations would form a polar covalent bond?*

*H and H Na and N H and N Na and Br*

103. ***Ionic bonds*** form when the electronegativity difference between two bonding atoms is *greater than* 1.7.

*Which of the following combinations would form an ionic bond?*

*H and H Na and N H and N Na and Br*

104. Substances containing mostly covalent bonds are called ***molecular substances***.

They are attracted to each other by weak van der Waals or stronger hydrogen attractions

*Which of the following is a molecular substance?*

*Lithium chloride carbon monoxide sodium nitrate aluminum oxide*

105. **London Dispersion**  attractive forces are the attractive force between nonpolar molecules.

Nonpolar molecules are molecules that have structural symmetry.

106. **Dipole-dipole** attractions occur between all polar moleucles.

*Which of the following samples has the dipole-dipole forces of attraction?*

*F2 CH4 NH3 I2*

107. **Polar molecules** have stronger forces of attraction. The lack structural symmetry.

*Which of the following is a polar molecule?*

*CO2 H2O C4H10 N2*

108. ***Hydrogen bonds*** are attractive forces that form when hydrogen bonds to the elements N, O, or F and

gives the compound unexpectedly high melting and boiling points.

*The strongest forces of attraction occur between molecules of*

*HCl HBr HF HI*

109. Substances containing mostly ionic bonds are called ***ionic compounds***.

They are made of metal and nonmetallic ions. They are held together by electrostatic (ionic) forces.

110. Complete and memorize this table.

|  |  |
| --- | --- |
| Substance Type | Properties |
| **Ionic** | Hard  *(Low/high) melting and boiling points*  Conduct electricity when molten or aqueous |
| **Covalent (Molecular)** | Soft  *(Low/high) melting and boiling points*  Do not conduct electricity (insulators) |

111. Remember: substances tend to be soluble in solvents with similar molecular properties.

**“Like dissolves like”**

*Pentane does not dissolve in water because pentane is ………. and water is ………..*

112. As temperature increases, solubility increases for most solids.

*For which substance does increasing temperature have the least effect on solubility?*

*Sodium chloride calcium carbonate oxygen sodium bicarbonate*

113. At low temperatures and high pressures solubility *increases* for most gases.

*Carbon dioxide gas is* ***least*** *soluble in water at conditions of …. temperature and .… pressure.*

116. ***Molarity*** is a way to measure the *concentration* of a solution.

Molarity is equal to the number of moles of solute divided by the number of liters of solution.

*What is the molarity of an NaCl solution if 2.0 mol NaCl is present in 0.50 L solution?*

117. ***Molality*** = (moles of solute/ kg of sovent)

*A solution of glucose is prepared by added 10. g glucose (C6H12O6) to 40. g water.*

*What is its molality?*

118. ***Dilutions*** 🡪 *M1V­1 = M2V2*  is the way to determine how to make a less concentrated solution.

*A student needs 250 mL of a 1.0 M HCl solution, how many mL of 6.0M HCl do they need to make this*

*solution?*

119. Solutes **raise** the boiling points and **lower** the melting points of solvents. ***Colligative Properties***

*Which of the following will have the highest boiling point?*

*1 mol NaCl in 100 g water 1 mole CH3OH in 100 g water 1mol CaCl2 in 100 g water*

120. The ***freezing-point depression*** is the change in the lowering of the freezing point of a substance when a

solute is added to it. The ***boiling-point elevation*** is the change is the substance’s boiling point under the

same conditions. *What is the freezing-point depression of a 0.020 m aqueous CaCl2 solution?*

123. Covalently bonded substances form **molecules**. Molecular compounds are made up of nonmetals only.

*Which of the following are molecular compounds? H­2O CH4 NaCl CuSO4  CO2*

123.5 The formula of a **molecular compound** represents one molecule of the substance, also one mole of

molecules and the number of atoms for each nonmetal element making up the molecule.

*The formula for methane gas, CH4, tells us the make-up of one molecule of methane is 1 carbon atom and 4 hydrogen atoms covalently bonded together. We can also figure out the molar mass or the mass in one mole of methane*. *What is the molar mass of CH4? 47 g/mol 24 g/mol 16 g/mol 10 g/mol*

124. Increasing the **concentration** of reactants will increase ***reaction rate***.

*Which sample of HCl (aq) will react most rapidly with magnesium metal?*

*0.50 M HCl 1.0 M HCl 3.0 M HCl 6.0 M HCl*

125. ***Reaction rate*** increases with an increase in temperature (and pressure for gases).

126. ***Catalysts*** speed up reactions by lowering their ***activation energies***.

They are not changed themselves and can be reused many times over.

153. ***Acids*** and ***bases*** are both ***good electrolytes***. Their solutions conduct electricity well.

*Which of the following is a nonelectrolyte?*

*LiOH HBr HC2H3O2 C2H5OH*

154. **Weak acids** taste *sour* and *react with metals*.

155. **Weak bases** taste *bitter* and *feel slippery*.

156. Acids and bases turn ***indicators*** different colors.

*Which solution will change red litmus to blue?*

*HCl(aq) NaCl(aq) CH3OH(aq) NaOH(aq)*

157. **pH** is the negative log (exponent) of the hydronium [H+] ion concentration.

*What is the pH of a 1.0 x 10-5 molar HCl solution?*

*1 9 5 4*

158. **Acids** have a pH < 7. **Bases** have a pH > 7.

159. Every 1 pH number **decrease** represents a ten-fold [H+] **increase**.

162. ***Arrhenius*** model of acids and bases states:

“Acids give off H+ to form H3O+ ions in aqueous solution as their only (+) ion.”

“Bases give off OH- ions in aqueous solution as their only (-).”

*Which of the following is neither an Arrhenius acid nor an Arrhenius base?*

*KOH CH3COOH CH3OH HNO3*

163. ***Brønsted*** model of acids and bases states:

“Acids *donate* protons.” “Bases *accept* protons.”

*Identify one Bronsted acid and one Bronsted base in the reaction below:*

*H2O + NH3 ⬄ NH4 + + OH*-

164. Bronsted acids become Bronsted bases; Bronsted bases become Bronsted acids; forming conjugate pairs.

#### Identify one conjugate acid-base pair from question #163

165. Acids and bases react in ***neutralization*** reactions to make ***water*** and a ***salt***.

*Name the salt produced when sulfuric acid is neutralized by potassium hydroxide.*

166. Acids are ionic formulas in which the positive ion is H+. Use as many H+ ions as the charge on the

negative anion.

*Three rules for naming:* if the **anion** ends with: the acid is named:

–ite \*\*\*\*\*\*\*\*ous acid –ate \*\*\*\*\*\*\*\*ic acid –ide hydro\*\*\*\*\*\*\*\*ic acid

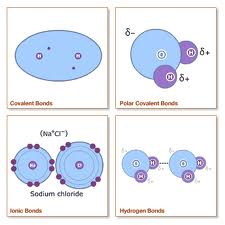
*Which of the following acids would be called sulfuric acid?*

*H2S H2SO3 H2SO4  H2CO3*

167. A ***buffer*** is a solution that stabilizes the pH of a solution when small amounts of acids or bases

are added. A buffer is a solution of a weak acid and one of its salts, or a solution of a weak base

and one of its salts.



**Some Helpful Stuff to Look Over**

**Rate of Solution**

|  |  |  |
| --- | --- | --- |
| **Factor** | **Effect on Solid Solute** | **Effect on Gaseous Solute** |
| Particle Size | Reducing particles size by crushing increases the rate by increasing **surface area** | Not applicable |
| Stirring | Increases the rate by exposing fresh solvent to solute and increasing **kinetic energy** | Decreases the rate by increasing **kinetic energy**, thereby reducing solubility |
| Amount of dissolved solute | As the amount of dissolved solute increases, the rate decreases | As the amount of dissolved solute increases, the rate decreases |
| Temperature | As the temperature increases, the kinetic energy increases, and the rate increases | As the temperature increases, the rate decreases |











**Ideal Gas Problems**

Gases at low pressures obey the ideal gas law,

*p V* = *n R T* (1)

where *R* is a constant (known as the *gas constant*) that has the value

*R* = 0.08206 L atm K-1 mol-1 (2)

Appropriate units to use for *p*, *V*, *n*, and *T* in the ideal gas equation are those used for *R* above. Thus the pressure (*p*) should be in atm, the volume (*V*) in L, the temperature (*T*) in degrees K, and the amount of gas (*n*) should be in moles. Useful conversion factors are

Pressure: 1 atm = 760 Torr = 760 mmHg = 101.3 kPa = 1.013 bar

Temperature: K = 273 + oC

Volume: 1 L = 1000 mL = 1000 cm3

Since , and *R* is a constant, it follows that

 (3)

where the subscript “1” represents one set of conditions, and the subscript “2” represents another set of conditions. More specialized equations may be derived from Eq(3) when one or more of the variables is held constant. For example, you can easily derive the familiar equations given below in this manner (convince yourself that this works!):

Boyle’s law:  (obtained when *n1* = *n2* and *T1* = *T2*)

Charles’s law:  (obtained when *n1* = *n2* and *p1* = *p2*)

Avogadro’s Principle:  (obtained when *T1* = *T2* and *p1* = *p2*)

***STP***

Often you will see gas volumes reported at STP (*standard temperature and pressure*). STP is defined as *T* = 273 K (0oC) and *p* = 1 atm. Substitution of these values into Eq(1) shows that the *volume of 1 mol of any gas is approximately 22.4 L at STP*. (You should verify this for yourself using Eq(1)!).

|  |  |
| --- | --- |
| **Molecular Crystals**  Examples: | One or more of the following: |
| (a) Need H bonded to O, N or F: H2O, HF, NH3. | (a) *Hydrogen bonding*: Hydrogen bonds are weaker than covalent bonds, but stronger than (b) or (c) below. |
| (b) C6H6 (benzene), polyethylene, I2, F2, and all the compounds from (a) above. | (b) *Dispersion forces* (induced dipole – induced dipole or London dispersion forces): universal force of attraction between instantaneous dipoles. These forces are weak for small, low-molecular weight molecules, but large for heavy, long, and/or highly *polarizable* molecules. They usually dominate over (c) below. |
| (c) CHF3, CH3COCH3 (acetone) and H2O, HF, NH3. | (c) *Dipole-dipole forces*: these forces act between *polar* molecules. They are much weaker than hydrogen bonding. |
| Note: ***Van der Waals Forces*** is a category which includes *both* categories (b) and (c) above. | |