Name: ___________________________________ Period ____

KEY: Pre-AP Chemistry Final Exam Review: Spring

Given Information: Periodic Table, Pressure Conversions, Solubility Rules, Molecular Geometry Chart, Electronegativities, Heats of Formation, Specific Heats

Chemical Reactions/Redox

1. What symbols are used in chemical equations to indicate heat added? A catalyst?
   Triangle over the arrow; chemical formula for catalyst over the arrow
2. Balance the equation: _2_ C_3H_6 + _9_ O_2 → _6_ CO_2 + _6_ H_2O What type of reaction is this?
   Combustion
3. Balance the equation: ___ P_4O_{10} + _12_ KOH → _4_ K_3PO_4 + _6_ H_2O
4. What is the general form for a synthesis reaction? Give an example of a synthesis reaction.
   A + X → AX; 4 Na + O_2 → 2 Na_2O
5. What is the general form for a decomposition reaction? Give an example of a decomposition reaction.
   AX → A + X; 2 H_2O → 2 H_2 + O_2
6. What is the general form for a single replacement reaction? Give an example of a single replacement reaction.
   Metal replacement: A + BC → AC + B; Mg + $CuCl_2$ → $MgCl_2$ + Cu
   Halogen replacement: D + BC → BD + C; F_2 + 2 KCl → 2 KF + Cl_2
7. What is the general form for a double replacement reaction? Give an example of a double replacement reaction.
   AB + CD → AD + CB; Na_2SO_4 + $Pb(NO_3)_2$ → 2 $NaNO_3$ + $PbSO_4$
8. What is the general form for a combustion reaction? Give an example of a combustion reaction.
   C_xH_y + O_2 → H_2O + CO_2; CH_4 + O_2 → CO_2 + H_2O
9. Metal oxides + water produce? Give an example of this reaction type.
   Metal hydroxide (base); CaO + H_2O → $Ca(OH)_2$
10. Nonmetal oxides + water produce? Give an example of this reaction type.
    Acid; SO_3 + H_2O → H_2SO_4
11. Nitrogen gas reacts with hydrogen gas to form? Write and balance this reaction.
    Ammonia; N_2 + 3 H_2 → 2 NH_3
12. Metallic carbonates decompose to form ? Give an example of this reaction type.
    Metallic oxide and carbon dioxide; K_2CO_3 → K_2O + CO_2
13. Metallic chlorates decompose to form ? Give an example of this reaction type.
    Metallic chloride and oxygen gas; 2 $NaClO_3$ → 2 NaCl + 3 O_2
14. Write a balanced equation for the reaction between phosphoric acid and calcium hydroxide.
    2 H_3PO_4 + 3 $Ca(OH)_2$ → $Ca_3(PO_4)_2$ + 6 H_2O
15. Predict whether a reaction between potassium and copper (II) nitrate will occur. If it does, write a balanced equation.
    YES; 2 K + $Cu(NO_3)_2$ → 2 KNO_3 + Cu
16. Predict whether a reaction between iodine and sodium bromide will occur. If it does, write a balanced equation.
    NO (iodine is less reactive than bromine)
17. Predict whether a reaction between chlorine and sodium bromide will occur. If it does, write a balanced equation.
    YES; Cl_2 + 2 NaBr → 2 NaCl + Br_2
18. Fluorine in a compound always takes an oxidation number of _- 1_.
19. The oxidation number for oxygen is typically _- 2_, unless it is part of a peroxide, in which case it has an oxidation number of _- 1_. Of course, when it is uncombined, its oxidation number is _0_.
20. Hydrogen has an oxidation number of _+1_, unless it is a hydride (combined with a metal), in which case it has an oxidation number of _-1_.
21. What are the oxidation numbers of each element in H_2SO_4? +1, +6, -2

22. What are the oxidation numbers in the ion NO_2^-? +3, -2
23. In oxidation, an atom's oxidation number ___increases___ and it ____loses_____ electrons.

24. In reduction, an atom's oxidation number ___decreases___ and it ____gains_____ electrons.

$$\overset{+1\ -1}{2NaCl} + \overset{0}{F_2} \rightarrow \overset{+1\ -1}{2NaF} + \overset{0}{Cl_2}$$

25. For the reaction $2NaCl + F_2 \rightarrow 2NaF + Cl_2$, label all oxidation numbers. What element is reduced? What element is oxidized? F is reduced; Cl is oxidized

$$\overset{+6\ -2}{Cr_2O_7^{2-}} + \overset{+4\ -2}{SO_2} + \overset{+1}{H^+} \rightarrow \overset{+3}{Cr^{3+}} + \overset{+1+6\ -2}{HSO_4^-} + \overset{+1\ -2}{H_2O}$$

26. For the reaction $Cr_2O_7^{2-} + SO_2 + H^+ \rightarrow Cr^{3+} + HSO_4^- + H_2O$, label all oxidation numbers. What element is reduced? What element is oxidized? Cr is reduced; S is oxidized

Stoichiometry

27. Which of the following is conserved in a chemical reaction: atoms, moles, mass atoms and mass
28. Interpret the following equation in 2 different ways, and write all mole relationships:
$N_2 (g) + 3 H_2 (g) \rightarrow 2 NH_3 (g)$
1 molecule of nitrogen reacts with 3 molecules of hydrogen to produce 2 molecules of ammonia.
1 mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia.
1 mol N_2 : 3 mol H_2 1 mol N_2 : 2 mol NH_3 3 mol H_2 : 2 mol NH_3
29. What is the mole ratio of oxygen to water in the reaction that represents the decomposition of water?
1:2

30. When solving a stoichiometry problem, if the given quantity is in (1) grams, (2) molecules, or (3) liters, what conversion factor would you first apply <u>in each case</u>?
1 mol X / molar mass in grams ; 1 mol X / 6.02×10^{23} particles ; 1 mol X / 22.4 L (of gas at STP)

31. For the reaction of zinc with hydrochloric acid, how many moles of zinc are needed to produce 10.0 mol of hydrogen gas?
10.0 mol Zn

32. For the reaction of copper with silver nitrate (use Cu^{2+}), how many grams of silver can be produced from 1.40 g silver nitrate and excess copper?
0.889 g Ag

33. How many moles of hydrochloric acid react with 0.350 mol potassium hydroxide to form water and potassium chloride?
0.350 mol

34. How many moles of magnesium must react with 8.75 moles of oxygen gas to form magnesium oxide?
17.5 mol Mg

35. In the synthesis reaction between zinc and iodine, how many molecules of iodine react with 0.850 mol zinc?
5.12×10^{23} molecules I_2

36. For the combustion of methane, how many liters of carbon dioxide (at STP) are formed from 4.00 mol methane and excess oxygen?
89.6 L CO_2

37. What is the definition of the limiting reactant in a chemical reaction?
The reactant that is completely used up.

38. What is the definition of the excess reactant in a chemical reaction?
The reactant that has some left over.

39. Consider the reaction $2 Mg + O_2 \rightarrow 2 MgO$. If magnesium is the limiting reactant, what substance(s) will be

present when the reaction stops? O_2 and MgO

40. In the synthesis reaction between potassium and fluorine, how many grams of potassium fluoride can be formed if you have 10.0 g of each reactant?
    14.9 g KF

41. Consider the double replacement reaction between barium nitrate and sodium phosphate for the next 3 questions. Assume 0.500 mol of each reactant.
42. What is the limiting reactant?
    $Ba(NO_3)_2$

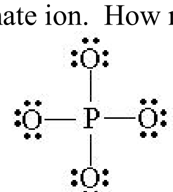
43. How many moles of barium phosphate precipitate are formed?
    0.167 mol

44. How many moles of the excess reactant (Na_3PO_4) are used up? Left over?
    0.333 mol ; 0.167 mol

45. In a reaction that produces silver, if only 2.25 g of silver are produced for a reaction with a theoretical yield of 3.00 g, what is the percent yield?
    75.0%

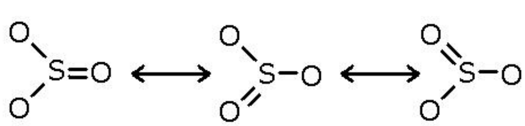
Chemical Bonding

46. Which of the following contain(s) covalent bonds? How can you tell? CO_2, K_2O, NaCl, $SO_4{}^{2-}$
    CO_2 and $SO_4{}^{2-}$, because they involve bonding of nonmetals
47. Correctly represent a formula unit of: sodium fluoride; potassium phosphate; magnesium sulfate
    NaF, K_3PO_4, $MgSO_4$
48. Which of the following do <u>not</u> have the same electron configuration: Mg^{2+} and Mg, Mg^{2+} and Ne, Cl^- and Ar, Br^- and Xe.
    Mg^{2+} and Mg, Br- and Xe
49. How many protons and electrons are in the following ions: F^-, Li^+, S^{2-}, Al^{3+}
    9 and 10, 3 and 2, 16 and 18, 13 and 10
50. In what state(s) can an ionic compound conduct an electric current?
    When dissolved in water (aqueous) and when melted (molten/liquid)
51. Describe the electron sea model of metals. What properties of metals can be accounted for by this model?
    Metals have vacant d and/or f orbitals that allow electrons to "roam" throughout the electron cloud (they are "delocalized"). Because of this, metals exhibit high electrical and thermal conductivity, and they are shiny, malleable, and ductile.
52. Draw a Lewis structure for the phosphate ion. How many <u>unshared pairs</u> of electrons are in the structure?



12 unshared e- pairs

53. How many valence electrons (A, available) are in a molecule of ammonia?
    8
54. Draw all possible resonance forms for the sulfur trioxide molecule.



55. What is the molecular geometry (shape) of carbon tetrachloride? tetrahedral

56. What is the molecular geometry (shape) of carbon dioxide? linear

57. What is the molecular geometry (shape) of phosphorus trichloride? trigonal pyramidal
58. What is the molecular geometry (shape) of the nitrate ion? trigonal planar

59. Which of the following molecules is/are polar: CO_2, H_2O, NH_3
    Remember to consider both bond polarity and molecular geometry; H_2O and NH_3 are polar

60. Which of the following molecules is/are nonpolar: CCl_4, C_3H_6, SCl_2
    Remember to consider both bond polarity and molecular geometry; NP, NP, P

Gas Laws

61. Convert 1.50 atm to: mm Hg ; kPa (1 atm = 760 mm Hg = 101.3 kPa)
    1140 mm Hg, 152 kPa
62. What is the relationship between pressure and temperature if volume is held constant?
    P and T are directly proportional at constant V
63. How high (in mm) is a column of mercury in a barometer if the barometer is at sea level?
    760 mm
64. How are moles (molecules) and temperature related to pressure?
    Both are directly proportional to pressure
65. Convert 100ºC to K. Convert 298 K to ºC.
    373 K, 25ºC
66. What is the relationship between pressure and volume if temperature is held constant?
    P and V are inversely proportional at constant T.
67. What is the name of the gas law that relates P and V? P and T? T and V?
    Boyle's, Gay-Lussac's, Charles'
68. Under what conditions of temperature and pressure do real gases exhibit ideal behavior?
    High temperature, low pressure (fast-moving and not crowded)
69. What variable is held constant in Charles's Law? Boyle's? Gay-Lussac's?
    Pressure; Temperature; Volume (Big Tony Can Play Great Violin)
70. Define absolute zero. What is this temperature in K? In ºC?
    The temperature at which all molecular motion ceases; 0 K or – 273ºC
71. If the temperature of a gas goes down, what happens to its density?
    Density increases; the volume of the gas decreases as it cools. (or think about the formula $D = \mathcal{M}P/RT$)
72. 4 gases are mixed in a container. The total pressure of the mixture is 760 mm Hg. The partial pressures of 3 of the gases are: Gas A = 150 mm Hg, Gas B = 25 mm Hg, Gas C = 300 mm Hg. What is the partial pressure of Gas D?
    285 mm Hg
73. When a gas volume is specified, why must temperature and pressure also be stated?
    Volume of a gas is dependent upon temperature and pressure
74. A sample of nitrogen gas at 22ºC has a volume of 15.0 L and a pressure of 720. mm Hg. What is the mass of the gas?
    16.4 g

75. A 25.0 mL sample of a gas is at a pressure of 3.0 atm and a temperature of 29ºC. What is the new temperature <u>in ºC</u> if the volume is decreased to 15.0 mL and the pressure is changed to standard pressure?
    – 210 ºC

76. Write the equation for the synthesis of water vapor from hydrogen and oxygen gas. Assuming constant conditions of temperature and pressure, how many liters of oxygen gas are needed to produce 10.0 L of water vapor?
    $2\ H_{2\,(g)} + O_{2\,(g)} \rightarrow 2\ H_2O_{(g)}$; $5.00\ L\ O_2$

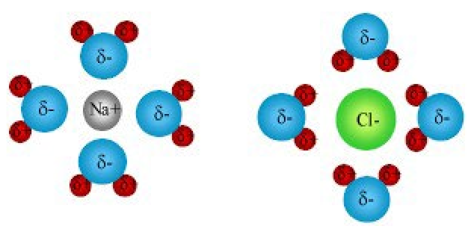
77. Write the equation for the single replacement reaction between magnesium and hydrochloric acid. How many grams of Mg must be used (with excess HCl) to produce 2.0 L of hydrogen gas at STP?
    $Mg + 2\ HCl \rightarrow MgCl_2 + H_2$; 2.17 g

78. Find the molar mass of a gas with a mass of 1.60 g and a volume of 500. mL. The gas is at – 18ºC and has a pressure of 2.17 atm.
    30.9 g/mol

IMFs/Solutions

79. List three properties of water that can be explained by hydrogen bonding.
    High melting/boiling points, high surface tension, high specific heat and heats of fusion/vaporization, liquid more dense than solid
80. Compare the density of water in its liquid vs. solid form. What explains this difference?
    The liquid phase is more dense (ice floats on water). Water molecules are more closely packed in the liquid phase, due to maximum hydrogen bonding.
81. List the type(s) of intermolecular forces present in: C_2H_6, HCl, NH_3
    C_2H_6: London dispersion forces only; HCl: London dispersion and dipole-dipole forces; NH_3: London dispersion forces, dipole-dipole and hydrogen bonding
82. How is boiling point related to strength of intermolecular forces?
    The stronger the IMFs, the higher the BP

83. Define solubility.
    The maximum amount of solute that will dissolve in a given amount of solvent at a specified temperature.
84. How does temperature affect the solubility of most solids? Gases?
    Increase in temperature increases solubility of most solids; increase in temperature decreases the solubility of gases.
85. When a solution is diluted, which of the following change(s): moles of solute, moles of solvent, volume of solution, concentration (molarity) of solution.
    Moles of solvent, volume of solution, and concentration of solution all change. Moles of solute does not change.
86. Explain the rule "like dissolves like" in terms of polarity and solubility.
    Substances with like polarities (ex. both polar/ionic or both nonpolar) will dissolve in each other, but substances with opposite polarities will not dissolve in each other.
87. Give an example of a substance that will dissolve in water, and one that will not, based on "like dissolves like."
    HCl (polar covalent) dissolves in water (polar covalent); CCl_4 (nonpolar covalent) does not dissolve in water.
88. Draw a picture of the solvation of sodium chloride by water. Be sure to orient the water molecules in the correct direction.



89. How many moles of solute and how many mL of total solution are in a 0.25 M solution of lead (II) nitrate?
    *This question is missing information about the mass of the solute. Assume a **100. g** mass of lead (II) nitrate.*
    0.302 mol $Pb(NO_3)_2$; 1.2 L or 1200 mL total solution volume
90. Calculate the molarity of a solution made by diluting 50.0 mL of a 6.0 M solution to a new volume of 1000. mL.
    0.30 M
91. How many mL of a 12.0 M stock solution of HCl are needed to make 2.0 L of a 1.0 M HCl solution?
    170 mL
92. Find the molarity of a 300. mL solution containing 16.0 g KCl.
    0.715 M
93. What is the volume of a 0.60 M solution of NaOH if it contains 80.0 g NaOH?
    3.3 L
94. Define an unsaturated solution. How could you test a solution to determine whether it is unsaturated?
    An unsaturated solution holds less than the maximum amount of dissolved solute for that temperature. To test, add a crystal of salt. If it dissolves, the solution was unsaturated.

95. Define a supersaturated solution. What would you expect to happen if you cooled a supersaturated solution and then added a seed crystal?
A supersaturated solution is an unstable solution in which more than the maximum amount of solute has been dissolved for that temperature. If a supersaturated solution is cooled and then stirred or a seed crystal is added, the excess solute will crystallize out of solution.

96. Use the solubility curve in your solutions notes. If a saturated solution of KNO_3 made with 100 g water is cooled from 60ºC to 50ºC, how much solute will crystallize out of solution? 25 g

97. Using the same solubility curve, how many grams of $KClO_3$ can be dissolved in 250 g of water at 20ºC? 25 g

98. Using the same solubility curve, would a solution of 100 g NH_4Cl in 200 g water at 50ºC be unsaturated, saturated, or supersaturated? saturated

99. Why are some solutions electrolytes while others are non-electrolytes? What type(s) of substances would not produce an electrolyte when dissolved in water and what type(s) would?
Not all substances produce ions in solution. Ions in solution are required in order to conduct an electric current. Substances that are electrolytes are: soluble ionic compounds (ex. NaCl), acids (ex. HCl), and bases (ex. NaOH). Strong acids and bases are strong electrolytes because they ionize/dissociate 100%. Weak acids and bases are weak electrolytes because the ionization is only partial. Molecular compounds contain covalent bonds (shared electrons), and do not produce ions when dissolved in water. They are nonelectrolytes (example: sugar).

100. Which of the following act(s) as a <u>strong</u> electrolyte when dissolved in water: HCl, $HC_2H_3O_2$, Na_2SO_4, Ag_2S, CO_2. HCl (strong acid); Na_2SO_4 (soluble ionic compound)

101. How does the conductivity of tap water compare to the conductivity of pure water? Explain.
Tap water conducts electricity because it contains dissolved minerals/salts (ionic compounds). Pure water contains only water molecules, so there are no dissolved ions and it does not conduct electricity.

102. For the following combinations of reactants: write the double replacement reaction if it occurs, identify the precipitate, write the overall ionic equation, identify the spectator ions, and write the net ionic equation.

a. $Cu(NO_3)_2 + NaOH$ b. $AgNO_3 + K_2S$ c. $Na_2CO_3 + Li_3PO_4$

a. $Cu(NO_3)_{2(aq)} + 2\ NaOH_{(aq)} \rightarrow 2\ NaNO_{3(aq)} + \mathbf{Cu(OH)_{2(s)}}$; spectators $Na^+_{(aq)}$ and $NO_3^-{}_{(aq)}$;
$Cu^{2+}_{(aq)} + 2\ OH^-_{(aq)} \rightarrow Cu(OH)_{2(s)}$

b. $2\ AgNO_{3(aq)} + K_2S_{(aq)} \rightarrow 2\ KNO_{3(aq)} + \mathbf{Ag_2S_{(s)}}$; spectators $K^+_{(aq)}$ and $NO_3^-{}_{(aq)}$;
$2\ Ag^+_{(aq)} + S^{2-}_{(aq)} \rightarrow Ag_2S_{(s)}$

c. No precipitate/no reaction

Acids and Bases

103. Describe the properties of an acid and of a base (i.e. taste, litmus color, ions present, pH, etc.)
Acids: taste sour, turn litmus red, produce H+ in solution, pH<7, react with bases and metals
Bases: taste bitter, feel slippery, turn litmus blue, produce OH- in solution, pH>7, react with acids

104. Name the following acids:

HCl hydrochloric acid	$HClO_3$ chloric acid	HNO_3 nitric acid
HClO hypochlorous acid	$HClO_4$ perchloric acid	
$HClO_2$ chlorous acid	HNO_2 nitrous acid	

105. If a solution has a $[OH^-]$ greater than 1×10^{-7}, then the solution is __basic__, less than is ___acidic___, and equal to is __neutral__.

106. Define a Bronsted-Lowry acid and a Bronsted-Lowry base.
A Bronsted-Lowry acid is a H+ (proton) donor, a Bronsted-Lowry base is a H+ (proton) acceptor.

107. Label the Bronsted-Lowry acid, the Bronsted-Lowry base, the conjugate acid, and the conjugate base. Identify each conjugate acid/base pair by connecting them with brackets.

$$\underset{A}{HCl} + \underset{B}{H_2O} \rightarrow \underset{CA}{H_3O^+} + \underset{CB}{Cl^-}$$

HCl and Cl- make a pair ; H_2O and H_3O^+ make a pair

$$\underset{B}{NH_3} + \underset{A}{H_2O} \leftrightarrow \underset{CA}{NH_4^+} + \underset{CB}{OH^-}$$

NH_3 and NH_4^+ make a pair ; H_2O and OH^- make a pair

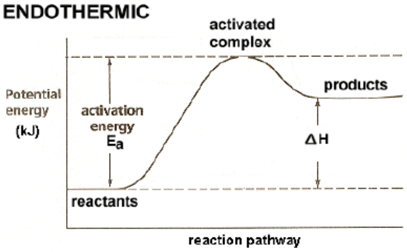
108. Complete the following neutralization reaction: $Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + 2\ H_2O$

109. If an aqueous solution has a $[OH^-]$ concentration of 3.5×10^{-8} M, what is the $[H_3O^+]$ concentration?
2.9×10^{-7} M
110. What is the pH of a 10^{-4} M HBr solution?
4
111. What is the pH of a 10^{-3} M KOH solution?
11
112. If $[H_3O^+]$ is 5.77×10^{-6} M, what is the pH of the solution?
5.239
113. Define titration.
Titration: a laboratory technique in which the concentration of an unknown acid (or base) solution is determined by adding a measured volume of a base (or acid) of known concentration to a measured volume of the unknown.
114. 39.4 mL of $Ba(OH)_2$ requires 42.2 mL of 0.550 M HCl for complete neutralization. What is the molarity of the $Ba(OH)_2$? 0.295 M

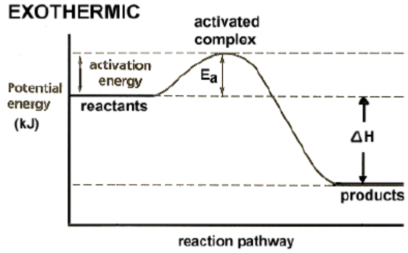
Thermochemistry

115. Draw energy diagrams for both an endothermic and an exothermic reaction. Include reactants, products, activation energy (E_a) and heat of reaction (ΔH).

ENDOTHERMIC



EXOTHERMIC



116. Using heats of formation in your thermochemistry notes, find ΔH (change in enthalpy or heat of reaction) for the following chemical reactions:

$4\,NO_{(g)} + 6\,H_2O_{(g)} \rightarrow 4\,NH_{3\,(g)} + 5\,O_{2\,(g)}$
+904.4 kJ, endothermic

$HCl_{(g)} + NaOH_{(s)} \rightarrow NaCl_{(s)} + H_2O_{(l)}$
-177.8 kJ, exothermic

$2\,H_2S_{(g)} + 3\,O_{2\,(g)} \rightarrow 2\,H_2O_{(g)} + 2\,SO_{2\,(g)}$
-1035.6 kJ, exothermic

117. What quantity of heat is required to raise the temperature of 10.0 g of lead from 25ºC to 100ºC?
98 J

118. A 50.0 g sample of aluminum at 90.0ºC is added to a 250.0 g sample of water at 15.0ºC. What is the final temperature of the mixture?
18.1ºC

119. List 3 phase changes that are endothermic.
Melting, boiling, sublimation
120. List 3 phase changes that are exothermic.
Freezing, condensation, deposition
121. Is energy absorbed or released during boiling? During condensation?
Absorbed (endo) ; released (exo)
122. What physical process does **vapor → liquid + energy** represent?
condensation