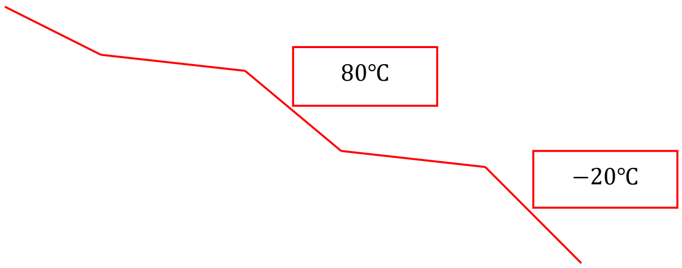
Chemistry 11 - Review - Final Exam

1. A theory is an explanation of why matter behaves the way it does based on circumstantial evidence. A law is a statement about matter based on experimental proof. A theory cannot be directly proven whereas a law can.
2. Qualitative information is obtained by observation and quantitative information is determined by measurement? Qualitative: The colour turned from blue to red and heat was produced.
   Quantitative: The mass of the gas was $22.3\ g$ and the pressure was $98.2\ kPa$.
3. Matter is anything that has mass and occupies volume. Energy can also be measured but is not considered as matter.
4. Properties of matter are characteristics of matter. Intensive properties can be used to identify the substance but extensive properties cannot. Water has a density of $1.0\frac{g}{mL}$; it boils at $100°C$ and freezes at $0°C$.
5. Silver is shiny, ductile and a good conductor of electricity. Sulphur is yellow, not shiny, is brittle, and a poor conductor of electricity.
6. Homogeneous means the same throughout and heterogeneous means not the same throughout. Pure substances are homogeneous because every atom or molecule of the substance is the same as every other atom or molecule. Mixtures can be either homogeneous or heterogeneous depending on the type of mixture. A solution is a homogeneous mixture because it is the same throughout. A mixture of sand and iron filings is heterogeneous because it is not the same throughout.
7. An element cannot be decomposed into simpler substances by simple chemical means whereas a compound can. An atom is the simplest particle of an element that still retains the properties of the element whereas an element is made up of many atoms. A molecule is the smallest particle of a compound that still retains the properties of the compound. It is made up of atoms joined in a specific way. A compound is made up of many molecules of the same kind. An element has no charge whereas ions can have a positive charge if electrons are lost and a negative charge if electrons are gained. It is possible to have a molecule of an element. Many of the non-metals become more stable by forming molecules. Ex: H_2, P_4, & S_8 for example.
8. Separating Mixtures
   a.) Use paper chromatography. Use a solvent that the dyes are soluble in. The more soluble dye dissolves first and moves up the paper faster that the dye of lower solubility.
   b.) Use a distillation apparatus. The alcohol has a lower boiling point than water. It will evaporate first and can be recovered by condensing it.
   c.) Add water and filter the mixture. The salt is soluble in water and sand is not. The sand will remain

in the filter paper and the salt solution will pass through. To recover the salt, evaporate the water.

9. Solid — liquid: melting; liquid — solid: freezing; liquid — gas: evaporating; gas — liquid: condensing; solid — gas: sublimation; gas — solid: deposition.

10. The graph should have flat sections during changes of state at 80℃ and –20℃.



11. Calculate the R_f values for the chromatograph below.

$$\frac{4.50}{6.35} = 0.709; \frac{2.45}{6.35} = 0.386$$

12. The curve is for a mixture. Pure substances have fixed melting and boiling points.

13. a.) a,c, and d are pure substances. b is a mixture.

b.) a and c are elements and d is a compound.

All are in the gaseous state because the particles are far apart.

14. Democritus was the first person to suggest that all matter was made of small particles. Aristotle suggested that all matter was made out of only four elements. The four elements were air, fire, earth and water.

Dalton's Atomic Theory: elements are made up of tiny particles called atoms. All atoms of the same element are identical. Compounds are made up of elements arranged in a specific way. When chemical reactions occur, the atoms arrange themselves to form new compounds. J.J. Thomson discovered the electron. His model was called the Plum Pudding model. The raisins represented the lighter electrons and the heavy batter represented the higher mass positive charge. Rutherford discovered the nucleus in his gold foil experiment. The small dense nucleus scattered the alpha particles. Bohr was the first to suggest that the electrons orbited around the nucleus.

15. e: 1-, ≈ 0 u; p: 1+, 1 u; and n: 0. 1 u

16. Complete the following table for the first four subshells.

Subshell	Starts at level #	Number of orbitals	Number of electrons
s	1	1	2
p	2	3	6
d	3	5	10
f	4	7	14

17. Complete the following table for the first five energy levels.

Energy level	Number of subshells	Number of electrons
1	1	2
2	2	8
3	3	18
4	4	32
5	4	32

18. $2n^2$, where n = the number of subshells. 128

19. OMIT

20. Electron configurations for the following:

a.) Mg: $1s^2 2s^2 2p^6 3s^2$ b.) Mn: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$ c.) Se: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$

d.) As: $ls^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

21. Core electron configurations for:

a.) $_{42}$Mo: $[Kr]5s^1 4d^5$ b.) $_{51}$Sb: $[Kr]5s^2 4d^{10} 5p^3$ c.) $_{28}$Ni: $[Ar]4s^2 3d^8$ d.) $_{55}$Cs $[Xe]6s^1$

22. The valence electrons are those found in the **outermost s & p** orbitals. As you go across a row of representative elements the number of valence electrons increases. As you go down a group the number of valence electrons remains the same. The valence electrons are the same in each group. Because chemical properties are determined by the valence electrons, elements in the same group have similar properties.

23. Draw Lewis (Electron Dot) symbols for the **elements** across the second period.

Li	Be	B	C	N	O	F	Ne
1+	2+	3+	4±	3-	2-	1-	

24. As you go across a row of RE the size of the atom gets smaller because the ENC increases drawing the electrons in closer to the nucleus. Down a group the size increases because electrons are placed in new energy levels farther from the nucleus.

25. IE is the amount of energy needed to remove the most loosely held electron from an atom in the gaseous state. Across a row on the periodic table IE increases because it gets harder to remove an electron as the atom gets smaller. Down a group IE decreases because it gets easier to remove an electron as the size gets larger. The exceptions in period 3 are Al and S. A filled s-block and a half-filled p-block have special stability. As a result, the electrons that follow are easier to remove.

26. Na. Its second electron is a core electron, which is harder to remove than Mg's second electron, which is a valence electron.

27. When electrons are lost the size gets smaller because each electron experiences a stronger pull. When electrons are gained the size gets larger because each electron experiences a weaker pull.

28. Mg^{2+}, Na^{1+}, F^{1-}, O^{2-} ; They all have the same electron configuration as Ne: $ls^2 2s^2 2p^6$

29. Arrange the following groups in order from smallest size to largest:

a.) F, N, C b.) Na, K, Rb c.) Cl, S, Se d.) K^{1+}, Ar, Cl^{1-} e.) F^{1+}, F, F^{1-} f.) Co^{3+}, Co^{2+}, Co

30. Arrange the following groups from smallest ionization energy to largest.

a.) Sr, Ca, Mg b.) B, C, N

31. The d-block metals use the outermost and unpaired d electrons as valence electrons. Short form electron configurations for: a.) Cr^{3+}: [Ar]$3d^3$ b.) Co^{2+}: [Ar]$3d^7$ c.) Cu^{1+}: [Ar]$3d^{10}$

32. Metallic properties decrease across a row and increase down a group. Fr is the largest and has the most valence electrons. Metals are ductile, shiny, and good conductors of electricity.

33. The atomic number identifies the number of protons and electrons. To find the number of neutrons subtract the atomic number from the atomic mass. Isotopes of an element have different mass because the number of neutrons can vary slightly.

34. Mendeleev arranged the elements in groups according to similar properties and in rows according to increasing atomic mass. The modern periodic table is arranged in rows according to increasing atomic number and in columns according to similar electron configuration.

35. The rows are called periods and the columns are called groups or families. Group: 1 alkali metals, 2 alkaline earth metals, 17 halogens, 18 noble gases; d-block metals: transition metals; s & p block elements: representative elements.

36. Complete the following table:

Particle	Atomic Number	Atomic Mass	Number of Protons	Number of Electrons	Number of Neutrons
$^{37}_{17}Cl$	17	37	17	17	20
$^{40}_{19}K^{1+}$	19	40	19	18	21
$^{33}_{16}S^{2-}$	16	33	16	18	17
$^{79}_{14}Si$	14	79	14	14	15
$^{79}_{35}Br^{1-}$	35	79	35	36	44
$^{31}_{15}P^{3-}$	15	31	15	18	16

37. Ionic compounds formed when there is a transfer of valence electrons from a metal to a nonmetal in an attempt to make all atoms obey the octet rule.

38. An ionic bond is the force of attraction between ions of opposite charge.

39. Ionic bonds are strong because the force of attraction between ions of opposite charge is strong. As a result, all ionic compounds are solids at room temperature and have high melting points.

40. The higher the charge on the ion and the smaller the ion the stronger the bond.

41. BeF_2. Be^{2+} is smaller and has a higher charge than Li^+.

42. The ions are too close together and can't move around.

43. When dissolved in water the ions are free to move around.

44. Use a conductivity apparatus. If the bulb lights up the compound is ionic.

45. $K_2Cr_2O_{7\,(s)} \rightarrow 2\ K^+_{(aq)} + Cr_2O_7{}^{2-}{}_{(aq)}$; $NH_4Br_{(s)} \rightarrow NH_4{}^+{}_{(aq)} + Br^-_{(aq)}$; $Mg(NO_2)_{2\,(s)} \rightarrow Mg^{2+}_{(aq)} + 2\ NO_2{}^-{}_{(aq)}$ $Fe_2(CO_3)_{3\,(s)} \rightarrow Fe_2(CO_3)_{3\,(s)}$

46. Covalent compounds are formed when non-metals share valence electrons in an attempt to make all atoms obey the octet rule.

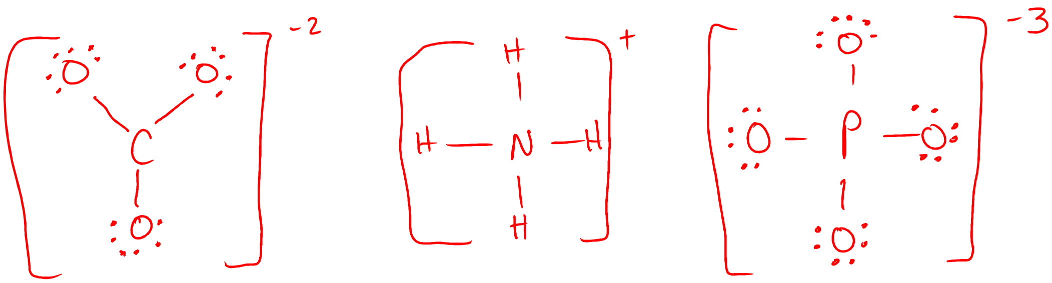
47. A covalent bond is the force of attraction between the shared valence electrons and the nuclei of the two atoms.

48. Draw Lewis structures for the following compounds, name their shapes, & predict which are polar or non-polar. Which of these compounds will be expected to dissolve in water? Explain.

    CH_4, NH_3, SO_2, Cl_2, BH_3, H_2S, CO_2

    NH_3, SO_2, & H_2S will dissolve water. They are polar like water. Likes dissolves likes.

49. Draw Lewis diagrams for the following: $CO_3{}^{2-}$, $NH_4{}^{1+}$, $PO_4{}^{3-}$



50. Names: lithium sulphite, Cobalt (II) sulphate, Carbon Tetrafluoride, dinitrogen pentoxide, hydrochloric acid, Aluminium Sulphate, Iron (III) phosphate pentahydrate, mercury (I) bicarbonate, carbon monoxide, methane.

51. Formulae: NH_3, MnO_2, $CoCl_2{\cdot}6H_2O$, HCl, HNO_3, SO_3, $Hg_3(PO_4)_2$, $Zn(HPO_4)_2$, H_2O_2, Al_4C_3

52. Equations

    a. Balance the following equations

        i. $CaC_2 + 2\ O_2 \rightarrow Ca + 2\ CO_2$

        ii. $5\ C + 2\ SO_2 \rightarrow CS_2 + 4\ CO$

        iii. $2\ BN + 2\ F_2 \rightarrow 2\ BF_2 + N_2$

        iv. $Al_2C_6 + 6\ H_2O \rightarrow 2\ Al(OH)_3 + 3\ C_2H_2$

v. $3\ NO_2 + H_2O \rightarrow 2\ HNO_3 + NO$

b. Write the products for the following reactions and balance them.

vi. $Cu + FeSO_4 \rightarrow$ no reaction

vii. $2\ C_5H_{10} + 15\ O_2 \rightarrow 10\ CO_2 + 10\ H_2O$

viii. $2\ Al + 3\ I_2 \rightarrow 2\ AlI_3$

ix. $Mg(OH)_2 + 2\ HBr \rightarrow MgBr_2 + 2\ H_2O$

x. $Al_2(SO_4)_{3\ (aq)} + 3\ Na_2CO_{3\ (aq)} \rightarrow Al_2(CO_3)_{3\ (s)} + 3\ Na_2SO_{4\ (aq)}$

c. Write the ionic and net ionic equation for vi & x in 52b.) above.

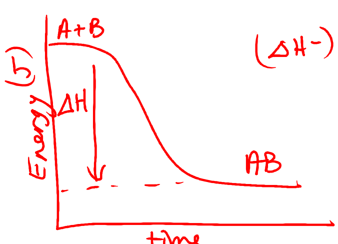
x. $2\ Al^{3+} + 3\ SO_4^{2-} + 6\ Na^{1+} + 3\ CO_3^{2-} \rightarrow Al_2CO_{3\ (s)} + 6\ Na^{+} + 3\ SO_4^{2-}$

$2Al^{3+} + 3\ CO_3^{2-} \rightarrow Al_2CO_{3\ (s)}$

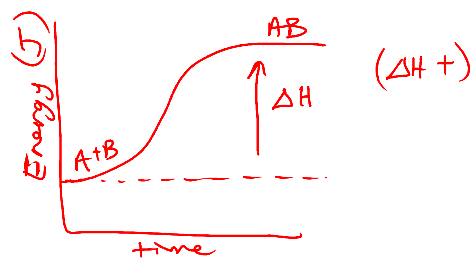
53. Exothermic reactions produce energy and endothermic reactions use energy.

Exothermic produces (releases) more energy than the reaction absorbs, and endothermic absorbs more energy than the reaction produces.

exothermic



endothermic



54. What is the molar mass of $Ni_2(SO_3)_3$? $357.7\frac{g}{mol}$

55. What is the mass of $3.65\ mol$ of CO_2? $161\ g$

56. How many moles of SO_2 are there in $12.6\ g$ SO_2? $0.197\ mol$

57. How many moles of O are there in 2.45×10^{24} molecules of H_2O? $4.07\ mol$

58. How many atoms of Ca are there in $1.34\ mol$ of Ca? $8.07 \times 10^{23}\ atoms$

59. How many moles of $N_{2\ (g)}$ are there in $46.1\ L$ of N_2 at STP? $2.06\ mol$ N_2

60. What is the mass of $16.9\ L$ $CH_{4\ (g)}$ at STP? $12.1\ g$

61. What volume of $F_{2\ (g)}$ at STP would 6.19×10^{22} molecules of F_2 have? $2.30\ L$

62. How many molecules of NO_2 are there in $4.87\ g$ of NO_2? $6.37 \times 10^{22}\ molecules$

63. What is the mass in grams of 1 atom of K? $6.50 \times 10^{-23}\ g$

64. A certain amount of P_2O_3 has $3.98\ g$ of P. How many grams of oxygen are there? $3.08\ g$ O

65. How many molecules of CO_2 are there in $68.2\ g$ CO_2? $9.33 \times 10^{23}\ molecules$

66. What is the density of $NH_{3\ (g)}$ at STP? $0.759\frac{g}{L}$

67. The density of $CCl_{4\ (l)} = 1.59\frac{g}{mL}$. How many moles of CCl_4 are there in $87.1\ mL$ of CCl_4? $0.899\ mol$

68. What is the % composition of $Ca(NO_3)_2$? $24.44\%\ Ca, 17.1\%\ N, 58.5\%\ O$

69. What is the % of H_2O in $CoSO_4 \cdot 6H_2O$? 41.1%

70. Determine the empirical formula for a compound consisting of 50.0% C, 10.0% H, and 40% O. $C_5H_{12}O_3$

71. $1.99\ g$ of Al combined with oxygen to form an oxide of aluminium with a mass of $3.75\ g$. What is the empirical formula for this compound? OMIT

72. A compound has an empirical formula of C_2H_4O and a molecular mass of $220.0\frac{g}{mol}$. What is its molecular formula? $C_{10}H_{20}O_5$

73. A compound containing only C and H was burned in a combustion reaction to produce $11.0\ g$ CO_2 and $5.40\ g$ H_2O. What is the empirical formula of the compound? If the molar mass of the compound is $144.0\frac{g}{mol}$ what is the molecular formula? C_5H_{12} , $C_{10}H_{24}$

74. How many grams of K_2SO_4 are needed to prepare $250.0\ mL$ of $0.350\ M$ K_2SO_4? $15.2\ g$

75. What is the molarity of a solution in which $45.6\ mL$ of solution contained $8.34\ g$ $MgCl_2$? $1.92\ M$

76. A chemist diluted $27.9\ mL$ of $12.00\ M$ HCl to a volume of $500.0\ mL$. What is the new concentration of HCl? $0.670\ M$

77. How many mL of $5.34\ M$ NaCl are needed to produce $234\ mL$ of $1.74\ M$ NaCl? $76.2\ mL$

78. If $34.8\ ml$ of $0.74\ M$ KBr are mixed with $51.2\ mL$ of $0.63\ M$ KBr, what is the resulting concentration of KBr? $0.68\ M$

79. What are the concentrations of each of the ions in $0.35\ M$ $Al_2(SO_4)_3$? Include the dissociation equation.
$Al_2(SO_4)_3 \rightarrow 2\ Al^{3+} + 3\ SO_4^{2-}$; $[Al^{3+}] = 0.70\ M$, $[SO_4^{2-}] = 1.05\ M$

80. Given: $4\ Al_{(s)} + 3\ O_{2\,(g)} \rightarrow 2\ Al_2O_{3\,(s)}$
    a. How many grams of Al_2O_3 can be produced from $34.2\ g$ Al? $64.6\ g$
    b. How many grams of Al are needed to completely react with $13.8\ L$ of O_2 at STP? $22.2\ g$

81. Given - $Sr(OH)_{2\,(aq)} + 2\ HCl_{(aq)} \rightarrow SrCl_{2\,(aq)} + 2\ H_2$
    a. If it takes $28.39\ mL$ of $0.760\ M$ HCl to neutralize $25.00\ mL$ of $Sr(OH)_2$ what is the concentration of $Sr(OH)_2$? $0.431\ M$
    b. What volume of $0.450\ M$ $Sr(OH)_2$ is needed to neutralize $245.0\ ml$. of $0.800\ M$ HCl?

82. Given – $4\ NiS_{(s)} + 7\ O_{2\,(g)} \rightarrow 2Ni_2O_{3\,(s)} + 4\ SO_{2\,(g)}$ $0.218\ L$
    a. If $45.2\ g$ NiS is reacted with $37.6\ g$ O_2 how many g of Ni_2O_3 will be produced? $40.9\ g$
    b. What is the limiting reactant? NiS
    c. How many grams of the excess reactant will remain unconsumed? $9.72\ g$

83. Given - $16\ Fe + 3\ S_8 \rightarrow 8\ Fe_2S_3$
    a. If $35.0\ g$ of Fe were reacted with excess S_8 to produce $49.4\ g$ Fe_2S_3 what is the % yield for the reaction? 75.8%

84. What is the standard temperature in $°C$ and K? $0°C$ $273\ K$

85. What is the standard pressure in kPa? $101.325\ kPa$

86. What is the standard molar volume? $\frac{22.41\ L}{1\ mol}$

87. How many litres of $O_{2\ (g)}$, at STP, can be produced if $25.0\ g$ of $KClO_3$ decomposed according to the following equation: $2\ KClO_{3\ (s)} \rightarrow 2\ KCl_{\ (s)} + 3\ O_{2\ (g)}$ $6.86\ L\ O_2$

88. If $43.6\ L$ of O_2 at STP reacted completely according to the following equation, how many grams of Al_2O_3 would be produced? $4\ Al_{\ (s)} + 3\ O_{2\ (g)} \rightarrow 2\ Al_2O_{3\ (s)}$ $132\ g\ Al_2O_3$

89. Nitrous oxide (N_2O) can be produced by the thermal decomposition of ammonium nitrate:

$$NH_4NO_{3\ (s)} \rightarrow N_2O_{\ (g)} + 2\ H_2O_{\ (l)}$$

At STP what volume of N_2O can be produced by the reaction of $3.6\ g$ NH_4NO_3? $1.0\ L\ N_2O$

90. How many litres of H_2 and O_2 gas can be collected by the electrolysis of $76.2\ g$ of H_2O at STP? The gases are collected over water. $2\ H_2O_{\ (l)} \rightarrow 2\ H_{2\ (g)} + O_{2\ (g)}$ $94.7\ L\ H_2$ and $47.4\ L\ O_2$