ACCELERATED CHEMISTRY FIRST SEMESTER TEST STUDY GUIDE

This is a list of ideas and concepts you should be able to do on the semester test.

(Formula Sheet) –given to you on the semester test includes solubility rules and activity series.

Other information and tips

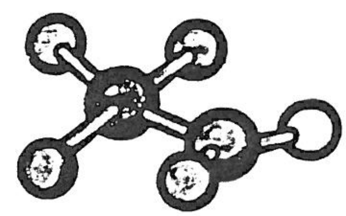
- There are 100 multiple-choice questions on the district semester test.
- **Bring two #2 lead pencils, a good eraser, AND a calculator.**
- Bring something to study or read after you finish the test
- Try to get a good night's sleep
- Eat a good breakfast. This will help put you in a good mood and help your brain function
- Drink water (H_2O). Avoid caffeine ($C_8H_{10}N_4O_2$) and sugar ($C_{12}H_{22}O_{11}$)

Name: ____________

Unit 1 - Chapter 2 – Data Analysis (15 questions)

Be able to:

- determine the correct # of significant digits in a measurement (4)
- interpret and write numbers in scientific notation (2)
- read and use graphs (2)
- use dimensional analysis to solve problems (3)
- determine the appropriate units to use in measurements (4)



1. How many significant figures are in each of these measurements?

a) 143 g ________ b) 0.074 cm ________ c) 57.048 m ________

d) 10 800 cal ______ e) 5.0×10^3 lbs ________

2. Complete the following calculations with the correct number of Sig Figs:

* ***Rule for add and subtract:*** ____________________ * ***Rule for multiply and divide:*** ____________

a. 420.4 + 19.57 = ____________

b. 184.5 / 15.987 = ____________

c. 97.5101 – 14.001 = ____________

d. 9.500 x 15 = ____________

e. $4.34 \times 10^{12} \times 2.8 \times 10^{-7}$ = ____________

f. If 9.2 g are removed from a 12.75 g sample, how many grams remain?____

g. The length of a square is 105.07 cm and the height is 55.64cm, what is the area?____________

3. Round the following numbers to 3 significant figures:

a. 1.2489772×10^{23} = ____________

b. 98451151 = ____________

c. 12.24456 = ____________

d. 0.0021445 = ____________

4. Round the following numbers to 3 sig figs and write in scientific notation:

a. 0.005784 = ____________

b. 6548780000 = ____________

c. 0.004499 = ____________

5. Review the prefixes used to adjust metric base units. Rank these from smallest to largest. (1 = sm, 6=lg)

kilogram   microgram   gram   milligram   centigram   megagram

6. 1 kg = ________ g   1 m = ________ cm

1 L = ________ mL   1 mL = ________ cm^3

7. Give the SI base unit for each quantity:

mass ________   liquid volume ________

length ________   density ________

temperature ________

8. Complete the following conversions.

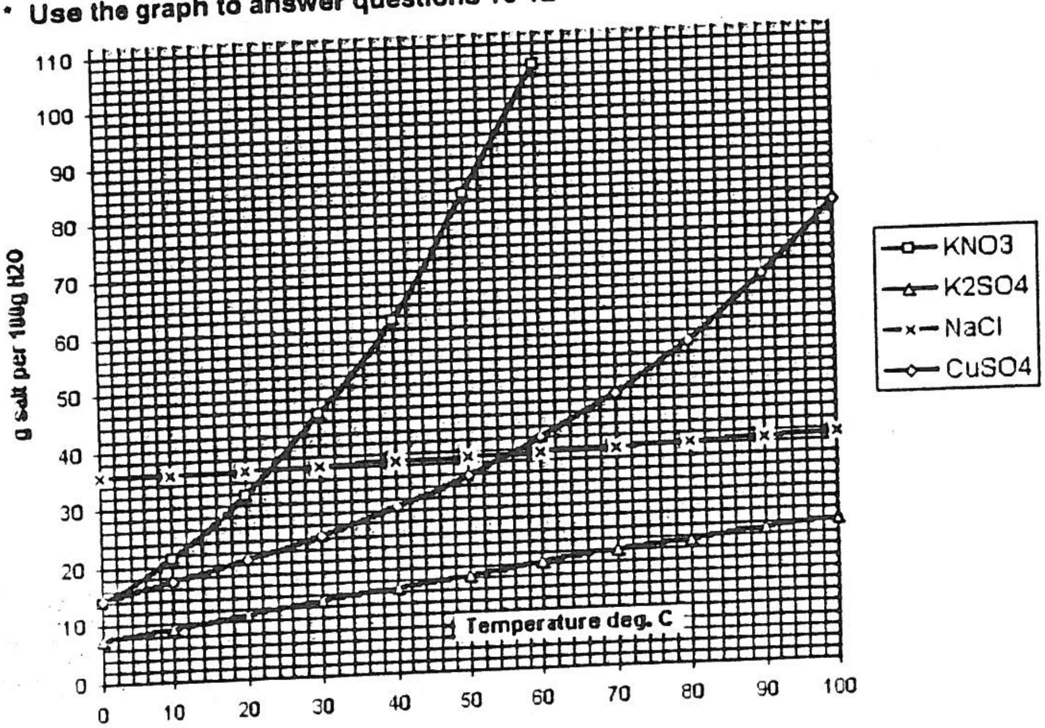
a. How many centimeters are in 2.195 meters? ________

b. How many liters are in 975.0 mL? ________

c. How many grams are in 45.15 kilograms? ________

9. Use dimensional analysis to solve the following problem. If each of your semester tests take 1.5 hrs, how many seconds will you be working on semester tests. (Assume that you have 6 semester tests).

- **Use the graph to answer questions 10-12**



Solubility Graph

10. How many grams of $CuSO_4$ will dissolve at 110.0 deg. C?

____

11. Which chemical has the highest solubility (grams dissolved) at 30 deg C?

____

12. Which chemical will have 10 g dissolved at 20 deg C?

____

Included in Unit 6 - Chapter 3 – Properties and Changes (11 questions)

Be able to:

- distinguish between physical and chemical changes and properties (3)
- determine the indicators of a chemical or physical change (1)
- classify a substance as an element, mixture, or compound (4)

13. Circle the examples of matter in the following list:

concrete, acetone vapor, heat, air, sound, light, steam.

• Remember, matter is anything with mass and volume. It has to be made of atoms!

14. Define theses terms:

a) mixture

c) compound

b) pure substance

d) element

e) Which of the above could be easily separated (by physical means)? For example filtration, using a magnet, evaporation…__________________

f) Which could be separated, but only by chemical means (like a chemical reaction)? __________

15. Look at the examples below.

1. Circle the elements
2. Underline the compounds
3. Put boxes around the mixtures
4. Draw arrows to the pure substances
5. Put stars next to the homogeneous mixtures.

water
aluminum
HCl (hydrochloric acid)
solution
sand and water
sugar (sucrose-$C_{12}H_{22}O_{11}$)
CO_2

Ne
Milk
River Water
Glass
Coke with Ice
Nitrogen
Cough syrup

Salt
Salt water
Alloy (steel)
Vinegar and Oil
30% hydrogen peroxide
zinc with water

• ______________________ is a synonym for "homogenous mixture".

16. Define physical property. ____________________

17. Define chemical property. ____________________

18. Define chemical change. ____________________

a) What are the 4 indicators of a Chemical Change?

1. ____________ 2. ____________
3. ____________ 4. ____________

- Why is color a physical property, but color change is an indicator of a chemical change?

19. Define physical change. ____________________

a) __________ __________(like condensation, evaporation, freezing, boiling) are physical changes.

20. In the following examples:

1. label the physical properties with "PP"
2. label the chemical properties with "CP"
3. label the physical changes with "PC"
4. label the chemical changes with "CC"

_____*Apple turns brown (rots)*
_____*The nail is 15 cm long*
_____*rubbing alcohol evaporates*
_____*propane is flammable*
_____*methane is burned*
_____*Sanding wood*
_____*The liquid is yellow*
_____*AlkaSeltzer added to water (lots of bubbling)*
_____*NaCl does not react with water*
_____*Chocolate melts in your pocket*
_____*A white precipitate forms*
_____*Water boils at 100 degrees Celsius*
_____*Copper is ductile*
_____*Statue of Liberty turns green*
_____*metal is pounded flat*
_____*The density of gold is 19.3 g/mL*
_____*hydrochloric acid is corrosive*
_____*water vaporizes*
_____*iron rusts*
_____*clay is rolled into a ball*
_____*iron is hard*

21. List the physical and chemical properties of iron.

Physical -

Chemical -

Unit 2 - Chapter 4 & 5 - Structure of the atom/Electron Configuration (20 questions)

Be able to:

- differentiate between and determine the number of protons, neutrons, and electrons in an atom (5)
- discuss the history of the atomic theory (1)
- define and determine the atomic #, mass #, and the average atomic mass of different isotopes (4)
- write, interpret, and relate electron configuration orbital notation and electron dot structure based on the location on the periodic table (4)
- relate chemical stability and the octet rule (1)
- Describe the wave mechanical view of the hydrogen atom (3)
- Describe the position and velocity of an electron in an atom (2)

22. Describe Rutherford's experiment. What did he discover about the atom?

23. Who discovered the electron? ____________________ What was he using when he discovered it?

24. Draw a picture of what the scientists believe the atom looks like. Try to draw it somewhat in proportion and indicate where the three subatomic particles would be located.

25. The smallest particle of any element is called a(n) ______________________.

26. Select **"P" proton, "N" neutron, and/or "E" electron.** There may be more than one answer.

____ *1+ charge*

____ *1- charge*

____ *no charge*

____ *located in nucleus*

____ *located in "empty space" around nucleus*

____ *relatively "big" particle(s) (~1 amu)*

____ *very small particle(s) (1/1840 amu)*

____ *determines element's identity*

____ *contributes to most of an element's mass*

____ *determines element's reactivity*

____ *changes to gain stability (form an ion)*

____ *varies in different isotopes of an element*

____ *same in different isotopes of the same element*

____ *must be the same in an element and it's ion*

27.

	Protons	Electrons	Neutrons	Mass #	Atomic #
Chromium-53					
Argon-42					
Potassium-42					
Potassium-41					

28. Write both forms (the <u>**symbol and name**</u> forms) of the following isotopes:

a. An atom of oxygen with 8 neutrons.

b. An atom of oxygen with 9 neutrons.

29. Define:

a. atomic number

b. mass number

c. atomic mass

d. isotope

30. An element has 26 p+, 26 e-, and 29 n°. What is the mass #? ____________ What is the element? __

31. Give the **COMPLETE** <u>electron configurations</u> and <u>orbital notations</u> for these elements.

a. Na ____

____

b. C ____

____

c. Cu (exception!!!)____

* configuration only

* remember-electrons fill from lowest to highest ____________(includes energy level and shape!

* Use your periodic table and the s, p, d, and f-blocks to determine the order and labeling.

32. Write the **SHORTHAND** or Noble Gas configuration for the following elements:

a. Si ______________________________

b. Ca ______________________________

33. Draw an electron dot diagram to represent the valance electrons for the following elements:

a. Cl

b. Ne

c. Sr

34. Looking at the periodic table, how would the electron configurations for the following elements end?

a. S ______________

b. Cl ______________

c. Mg ______________

d. Xe ______________

35. Boron has two naturally occurring isotopes, boron-10 and boron-11. The relative abundance of boron-10 is 19.9%; the relative abundance of boron-11 is 80.1%. The atomic mass of boron-10 is 10.01 amu and the atomic mas of boron-11 is 11.01 amu. What is the average atomic mass of boron?

36. What is a line spectra and how is one generated?

37. What is the difference between electrons being in ground state and excited state?

Unit 3 - Chapter 6 & 7 The Periodic Table and the Periodic Law/The Elements (17 questions)

Be able to:

- describe the properties of major groups on the periodic table (4)
- identify properties of metals, non-metals, and metalloids (1)
- identify and use the trends on the periodic table (electronegativity, ionization energy, atomic radius, shielding effect, nuclear charge, oxidation number) (3)
- relate electron configuration and ion formation and oxidation number based on the location on the perioc table (2)

38. Name the period 2 halogen ______________________

39. Name the group 13, period 5 element ______________________

40. Name the 2 period, alkaline-earth metal ______________________

41 Is Cl a representative element, transition metal, or innertransition metal? ______________________

42. Is U a representative element, transition metal, or innertransition metal? ______________________

43. Answer "MT" metal, "MD" metalloid, or "NM" nonmetal.

a. shiny, hard, dense ____________
b. forms negative ions ____________
c. like metals and nonmetals ________
d. border the stairstep ________
e. to the left of the stairstep __________
f. to the right of the stairstep __________
g. form positive ions____________
h. brittle, usually solids and gases ______
i. semiconductors____________
j. poor conductors____________
k. good conductors ________
l. malleable and ductile __________

44.

	Metal, metalloid, or nonmetal?	Family?
a) Fe	____________	____________
b) Si	____________	____________
c) Na	____________	____________
d) He	____________	____________
e) H	____________	____________
f) W	____________	____________
g) Np	____________	____________
h) Mg	____________	____________

45. Matching

_____ alkali metal
_____ alkaline-earth metal
_____ transition metal
_____ halogen
_____ noble gas

a. $1s^2 2s^2 2p^6$
b. $1s^2 2s^2 2p^6 3s^1$
c. $1s^2 2s^2 2p^5$
d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
e. $1s^2 2s^2 2p^6 3s^2$

46. Elements in the same ____________ have similar properties because ____________________.

47. All ______________ are unreactive because the have a full octet (full s and p orbitals)

48. Define:
a) period

b) group

c) family

d) periodic law

e) atomic radius (what is the trend?)

f) ionization energy (what is the trend?)

g) electronegativity (what is the trend?)

49. Circle the element with the largest ionization energy:

a) P, N   b) Rb, Fr, Li   c) Si, P, Mg

50. Circle the element in each pair that has the highest electronegativity:

a) K or Mg   B) Mg or S   c) F or He

51. Circle the element with the largest atomic radius.

a) Al or B   b) Fe or Cu   c) Br or Cl

Unit 4 - Chapter 8 Ionic Compounds (12 questions)

Be able to:

- determine the properties of ionic and metallic bonds (3)
- determine whether if bond is ionic based on the location on the periodic table (1)
- write formulas and names for ions and ionic compounds (7)

52. Why do ionic compounds form? ______________________

53. State the octet rule.

54. What happens to the valence electrons when an ionic bond is formed? ______________________

55. List 4 properties of compounds that have ionic bonds.

1)

2)

3)

4)

- What causes most of these properties? ______________________(strong network or +/- attractions)

56. An ionic bond is between a ______________________ and a nonmetallic element.

57. What is a cation? ______________________ and an anion? ______________________

58. What is the charge on the following ions?

a) silver ion ______________

b) zinc ion ______________

c) lead (IV) ion ______________

d) chloride ion ______________ (now has same e-config as ______________)

e) potassium ion ______________ (now has same e-config as ____________)

59. Complete the table for these ionic compounds:

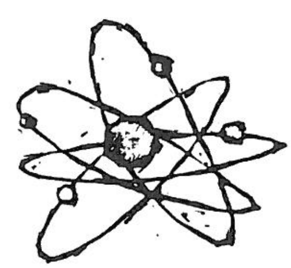
	Name	Formula
a)	ammonium oxide	
b)		Na_2SO_4
c)		K_3PO_4
d)	zinc hydroxide	
e)	iron (II) chloride	

60. Which elements need Roman numerals included in the name? ______________________________
What does the Roman numeral tell us? ____

Unit 5 - Chapter 9 Covalent Bonding (15 questions)

Be able to:

- determine the properties of covalent bonds (2)
- use Lewis Structures to determine shapes of molecules (including expanded octets) (4)
- use Lewis Structures to determine polarity (2)
- use electronegativity to determine the bond type (3)
- Determine the orbital hybridization, sigma and pi bonding in covalent compounds (2)
- determine the type of bond based on electronegativity differences (3)
- write formulas and names for covalent molecules (4)



61. Why do covalent compounds form? ______________________

62. What happens to the valence electrons when a covalent compound forms? ____________________

63. Covalent bonds normally form between 2 ________________________ elements.

64. List 4 properties of covalently bonded compounds:

1)

2)

3)

4)

65. Another name for covelent compounds is ______________________.

Name	Formula	Ionic or Molecular?
66. nitrogen trioxide	______	______
67. calcium nitrate	______	______
68. trisulfur heptaoxide	______	______
69. ______	NH_4Cl	______
70. ______	P_2O_5	______

71. List the prefixes!!!!______

72. What are the 8 shapes of molecules we studied? Complete the table.

	Name of shape	Picture	Example Lewis Structure	Can this shape be polar??
a)				
b)				
c)				
d)				
e)				
f)				
g)				
h)				

	Molecular formula	Lewis Structure	shape	Polar Molecule?
73.	SCl_2			
74.	CO_2			

75. SiF_6

76. N_2

77. What were the Lewis Structure exceptions we learned in Accelerated Chemistry?

78. Describe how electrons are shared/transferred and why in...

nonpolar bonds **polar bonds** **Ionic bonds**

79. Use the electronegativities to determine if the following bonds are nonpolar, polar or ionic.

a) H-H________________ (hydrogen e-neg=2.20)

b) Na-F ________________ (Na e-neg is 0.93 and F e-neg=3.98)

c) H-C ________________ (H e-neg = 2.20 and C e-neg = 2.55)

- More importantly ...ionic is always a ____________ with a ________________.

80. Explain the difference between a sigma and a pi bond. Identify the number of sigma and pi bonds in the CO_2 molecule

Unit 6 - Chapter 10 Chemical Reactions (12 questions)

Be able to:

- distinguish between the five types of chemical reactions (3)
- write and balance chemical reactions (3)
- interpret the law of conservation of mass (3)
- apply the rules of solubility for writing net ionic equations (2)

81. Matching

_____ $A + BY \rightarrow AY + B$ a. combustion

_____ $A + B \rightarrow AB$ b. synthesis

_____ $AY + BX \rightarrow AX + BY$ c. decomposition

_____ C_xH_x(hydrocarbon)$+O_2 \rightarrow CO_2 + H_2O$ d. single displacement

_____ $AB \rightarrow A + B$ e. double displacement

82. *Balance* and *identify* the type of reaction for each of the following:

a. _____C + _____O_2 $\rightarrow$ _____CO_2 ________________

b. ____C_4H_{10} + ____O_2 → ____CO_2 + ____H_2O ____________

c. ____H_2O_2 → ____H_2O + ____O_2 ____________

d. ____Pb + ____Hg_2SO_4 → ____$PbSO_4$ + ____Hg ____________

e. ____NaCl + ____$AgNO_3$ → ____AgCl + ____$NaNO_3$ ____________

h. ____Cr + ____$SnCl_4$ → ____Sn + ____$CrCl_2$ ____________

83. Describe how you might get "No Reaction" for a Single Displacement reaction: ____________
What about Double Displacement: ____________

84. Write a balanced chemical equation for the following word descriptions.

a. Propane gas (C_3H_8) combusts.

b. The synthesis of potassium bromide KBr.

c. Iron reacts with a solution of copper (II) nitrate.

d. NaCl + F_2 →

e.. The reaction of barium chloride solution and sodium carbonate solution.

85. Label the coefficient and the subscript: $3\ H_2O$

- Know what each stands for!!!

86. Which side of the reaction arrow are the reactants on? ____________

87. List the 7 diatomic elements. ____________

88. We balance chemical equations because ____________ cannot be created or destroyed according to the law of conservation of ____________.

89. Write a) the balanced reaction and b) the net ionic equation for the reaction that occurs when a solution of sodium hydroxide reacts with a solution of copper (II) sulfate.

a)

b)

Unit 7 - CH 11 – The Mole (10 questions)

90. 1 mole Fe = ______________ atoms = ______________ grams

91. Calculate the number of moles of a gold sample containing 3.33×10^{24} atoms.

92. Calculate the grams of 1.25 moles of magnesium.

93. Calculate the grams of 1.34×10^{25} atoms of Lead.

94. Determine the molar mass of ammonium dichromate $(NH_4)_2Cr_2O_7$. ______________

95. Circle the empirical formulas, square the molecular formulas.

$C_3H_6O_3$, CH_2O, $C_6H_{12}O_6$

96. Determine the empirical formula for a compound having 80.68% mercury, 12.87% oxygen, and 6.45% sulfur.

97. Caffeine is a compound that was found to have the empirical formula – $C_4H_5N_2O$. If its molar mass is 194.19 g/mol calculate its molecular formula.

98. If the molecular compound of glucose is $C_6H_{14}O_6$, what is the empirical formula?__________

99. Describe how you could get the molecular formula from the empirical formula.

100. Calculate the percent composition of Lead (II) chloride $PbCl_2$.

ACCELERATED CHEMISTRY
1st SEMESTER TEST

FORMULA SHEET

Element	Reaction
Li, Rb, K, Ba, Ca, Na	React with cold H_2O and acids, replacing hydrogen
Mg, Al, Mn, Zn, Fe	React with acids or steam, but usually not liquid water to replace hydrogen
Ni, Sn, Pb	React with acids but not water, to replace hydrogen
H_2, Cu, Hg	React with oxygen to form oxides
Ag, Pt, Au	Mostly unreactive

Halogen	Reaction
F_2, Cl_2, Br_2	Listed from most reactive to least

Solubility Rules

NO_3^{1-}	All nitrates are **soluble**.
Cl^{1-}	All chlorides are **soluble** except AgCl, Hg_2Cl_2, $PbCl_2$
SO_4^{2-}	Most sulfates are **soluble**; exceptions include: $SrSO_4$, $BaSO_4$, and $PbSO_4$. $CaSO_4$ is slightly soluble.
CO_3^{2-}	All carbonates are **insoluble** except those in Group 1 elements and NH_4^+
OH^{1-}	All hydroxides are **insoluble** except those of Group 1 elements, $Sr(OH)_2$ and $Ba(OH)_2$; $Ca(OH)_2$ is slightly soluble.
S^{2-}	All sulfides except those of Group 1 and 2 elements and NH_4^+ are **insoluble.**

JAWS

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- Eat a good breakfast. This will help put you in a good mood and help your brain function
- Drink water (H_2O). Avoid caffeine ($C_8H_{10}N_4O_2$) and sugar ($C_{12}H_{22}O_{11}$)

Name: KEY

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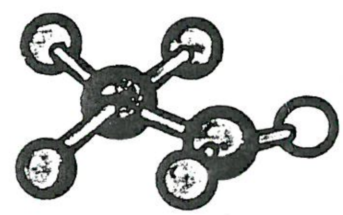


1. How many significant figures are in each of these measurements?

a) 143 g 3 b) 0.074 cm 2 c) 57.048 m 5

d) 10 800 cal 3 e) 5.0×10^3 lbs 2

2. Complete the following calculations with the correct number of Sig Figs:

*** Rule for add and subtract:*** least Decimals *** Rule for multiply and divide:*** least # Sig Figs

a. 420.4 + 19.57 = 440.0

b. 184.5 / 15.987 = 11.54

c. 97.5101 – 14.001 = 83.509

d. 9.500 x 15 = 140

e. $4.34 \times 10^{12} \times 2.8 \times 10^{-7}$ = 1.2×10^6

f. If 9.2 g are removed from a 12.75 g sample, how many grams remain? 3.55 3.6g

g. The length of a square is 105.07 cm and the height is 55.64cm, what is the area? 420.28 cm^2?

3. Round the following numbers to 3 significant figures:

a. 1.2489772×10^{23} = 1.25×10^{23}

b. 98451151 = 98500000

c. 12.24456 = 12.2

d. 0.0021445 = .00214

4. Round the following numbers to 3 sig figs and write in scientific notation:

a. 0.005784 = 5.78×10^{-3}

b. 6548780000 = 6.55×10^9

c. 0.004499 = 4.50×10^{-3}

5. Review the prefixes used to adjust metric base units. Rank these from smallest to largest. (1 = sm, 6=lg)

kilogram	microgram	gram	milligram	centigram	megagram
5	1	4	2	3	6
10^3	10^{-6}	—	10^{-3}	10^{-2}	10^6

6. 1 kg = 1000 g　　1 m = 100 cm

1 L = 1000 mL　　1 mL = 1 cm^3

7. Give the SI base unit for each quantity:

mass Kg　　liquid volume NA cm^3?

length m　　density NA g/cm^3?

temperature K

8. Complete the following conversions.

a. How many centimeters are in 2.195 meters? $2.195\,m \cdot \frac{100\ cm}{1\ m} = \boxed{219.5\,cm}$

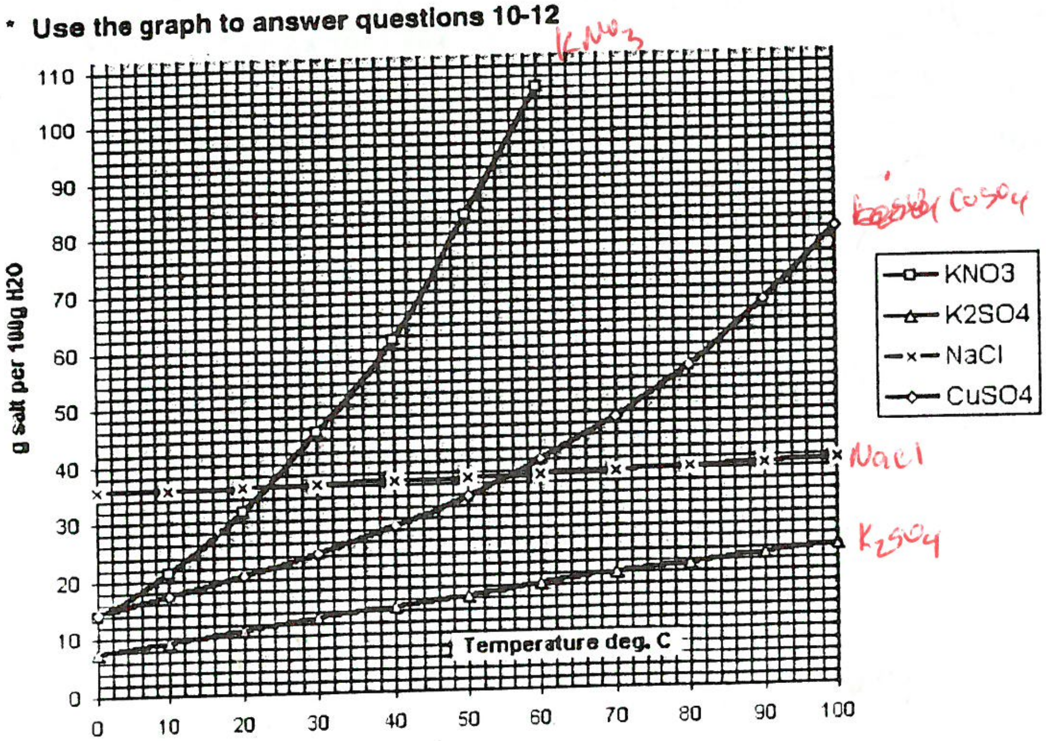
b. How many liters are in 975.0 mL? $975.0\,ml \cdot \frac{1\ L}{1000\ ml}$ $\boxed{.9750\,L}$

c. How many grams are in 45.15 kilograms? $45.15\,Kg \cdot \frac{1000\ g}{1\ Kg} = 45150\,g$

9. Use dimensional analysis to solve the following problem. If each of your semester tests take 1.5 hrs, how many seconds will you be working on semester tests. (Assume that you have 6 semester tests).

$6\ \text{Semester Tests} \cdot \frac{1.5\ hr}{1\ ST} \cdot \frac{60\ min}{1\ hr} \cdot \frac{60\ sec}{1\ min} = \boxed{32400\,sec}$

* **Use the graph to answer questions 10-12**



Solubility Graph

10. How many grams of $CuSO_4$ will dissolve at 110.0 deg. C? ~90g

11. Which chemical has the highest solubility (grams dissolved) at 30 deg C? KNO_3

12. Which chemical will have 10 g dissolved at 20 deg C? K_2SO_4

Included in Unit 6 - Chapter 3 – Properties and Changes (11 questions)

Be able to:

- distinguish between physical and chemical changes and properties (3)
- determine the indicators of a chemical or physical change (1)
- classify a substance as an element, mixture, or compound (4)

13. Circle the examples of matter in the following list:

concrete, acetone vapor, heat, air, sound, light, steam.

* Remember, matter is anything with mass and volume. It has to be made of atoms!

14. Define theses terms:

a) mixture Two or more pure substances physically combined

b) pure substance one element or one compound

c) compound Two or more pure substances chemically combined

d) element simplist form of pure substance - 1 type of atom

e) Which of the above could be easily separated (by physical means)? For example filtration, using a magnet, evaporation... Mixture

f) Which could be separated, but only by chemical means (like a chemical reaction)? Compound

15. Look at the examples below.

1. Circle the elements
2. Underline the compounds
3. Put boxes around the mixtures
4. Draw arrows to the pure substances
5. Put stars next to the homogeneous mixtures.

water

aluminum

HCl (hydrochloric acid)

solution *

sand and water

sugar (sucrose-$C_{12}H_{22}O_{11}$)

CO_2

Ne

Milk * ?

River Water

Glass *

Coke with Ice

Nitrogen

Cough syrup

Salt

Salt water *

Alloy (steel)

Vinegar and Oil

30% hydrogen peroxide *

zinc with water

- Solution is a synonym for "homogenous mixture".

16. Define physical property. Describes physical characteristic of substance

17. Define chemical property. Describes how substance interacts w/ another

18. Define chemical change. Substance physical & chemical properties change due to rearrangement of bonds

a) What are the 4 indicators of a Chemical Change?

1. Δ color
2. Δ Temp
3. Δ smell
4. produce precipitate or gas

- Why is color a physical property, but color change is an indicator of a chemical change?

19. Define physical change. Change in appearance, no change in bonds

a) State Changes (like condensation, evaporation, freezing, boiling) are physical changes.

20. In the following examples:

1. label the physical properties with "PP"
2. label the chemical properties with "CP"
3. label the physical changes with "PC"
4. label the chemical changes with "CC"

CC ***Apple turns brown (rots)***
PP ***The nail is 15 cm long***
PC ***rubbing alcohol evaporates***
CP ***propane is flammable***
CC ***methane is burned***
PC ***Sanding wood***
PP ***The liquid is yellow***
CC ***AlkaSeltzer added to water (lots of bubbling)***
CP ***NaCl does not react with water***
PP ***Chocolate melts in your pocket***
CC ***A white precipitate forms***
PP ***Water boils at 100 degrees Celsius***
PP ***Copper is ductile***
CC ***Statue of Liberty turns green***
PC ***metal is pounded flat***
PP ***The density of gold is 19.3 g/mL***
CP ***hydrochloric acid is corrosive***
PC ***water vaporizes***
CC ***iron rusts***
PC ***clay is rolled into a ball***
PP ***iron is hard***

21. List the physical and chemical properties of iron.

Physical - All metallic properties

Chemical - Reacts with acid to form H_2 gas

Unit 2 - Chapter 4 & 5 - Structure of the atom/Electron Configuration (20 questions)

Be able to:

- differentiate between and determine the number of protons, neutrons, and electrons in an atom (5)
- discuss the history of the atomic theory (1)
- define and determine the atomic #, mass #, and the average atomic mass of different isotopes (4)
- write, interpret, and relate electron configuration orbital notation and electron dot structure based on the location on the periodic table (4)
- relate chemical stability and the octet rule (1)
- Describe the wave mechanical view of the hydrogen atom (3)
- Describe the position and velocity of an electron in an atom (2)

22. Describe Rutherford's experiment. What did he discover about the atom?

Nuclear Atomic model (nucleus) Gold Foil Experiment

23. Who discovered the electron? JJ Thompson What was he using when he discovered it?
Cathode Ray Tuble

24. Draw a picture of what the scientists believe the atom looks like. Try to draw it somewhat in proportion and indicate where the three subatomic particles would be located.

25. The smallest particle of any element is called a(n) electron ? atom ?.

26. Select **"P" proton, "N" neutron, and/or "E" electron.** There may be more than one answer.

- P ***1+ charge***
- e⁻ ***1- charge***
- n ***no charge***
- $p^+ + n$ ***located in nucleus***
- e^- ***located in "empty space" around nucleus***
- $p^+ + n^o$ ***relatively "big" particle(s) (~1 amu)***
- e^- ***very small particle(s) (1/1840 amu)***
- P ***determines element's identity***
- $p^+ n$ ***contributes to most of an element's mass***
- e ***determines element's reactivity***
- e ***changes to gain stability (form an ion)***
- n ***varies in different isotopes of an element***
- p^{+}s ***same in different isotopes of the same element***
- P ***must be the same in an element and it's ion***

27.

	Protons	Electrons	Neutrons	Mass #	Atomic #
Chromium-53	24	24	29	53	24
Argon-42	18	18	24	42	18
Potassium-42	19	19	23	42	19
Potassium-41	19	19	22	41	19

28. Write both forms (the **symbol and name** forms) of the following isotopes:

a. An atom of oxygen with 8 neutrons. $^{16}_{8}O$ Oxygen-16

b. An atom of oxygen with 9 neutrons. $^{17}_{8}O$ Oxygen-17

29. Define:

a. atomic number # of p^+

b. mass number $p^+ + n^o$

c. atomic mass avg mass of all naturally occuring isotopes of that element

d. isotope 2 ~~different~~ elements – same # of p^+ different # of n^o

30. An element has 26 p+, 26 e-, and 29 n°. What is the mass #? 55 What is the element? Fe

31. Give the **COMPLETE** electron configurations and orbital notations for these elements.

a. Na $1s^2 2s^2 2p^6 3s^1$
[↑↓] [↑↓] [↑↓ ↑↓ ↑↓] [↑]
1s 2s 2p 3s

b. C $1s^2 2s^2 2p^2$
[↑↓] [↑↓] [↑ | ↑ |]
1s 2s 2p

c. Cu (exception!!!) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
* configuration only

- remember-electrons fill from lowest to highest Energy (includes energy level and shape!
- Use your periodic table and the s, p, d, and f-blocks to determine the order and labeling.

32. Write the **SHORTHAND** or Noble Gas configuration for the following elements:

a. Si [Ne] - $3s^2 3p^2$

b. Ca [Ar] - $4s^2$

33. Draw an electron dot diagram to represent the valance electrons for the following elements:

a. Cl ·C: b. Ne :Ne: c. Sr Sr·

34. Looking at the periodic table, how would the electron configurations for the following elements end?

a. S ~~4~~ · $3p^4$

b. Cl $3p^5$

c. Mg $3s^2$

d. Xe 5

35. Boron has two naturally occurring isotopes, boron-10 and boron-11. The relative abundance of boron-10 is 19.9%; the relative abundance of boron-11 is 80.1%. The atomic mass of boron-10 is 10.01 amu and the atomic mass of boron-11 is 11.01 amu. What is the average atomic mass of boron?

(.199 x 10.01amu)
+
(.801 x 11.01amu)

(.801 x 11.01)
+
(.199 x 10.01)

36. What is a line spectra and how is one generated?

Energy transitions for a specific element

37. What is the difference between electrons being in ground state and excited state?

ground state lowest energy possible
excited state when excited by energy

Unit 3 - Chapter 6 & 7 The Periodic Table and the Periodic Law/The Elements (17 questions)

Be able to:

- describe the properties of major groups on the periodic table (4)
- identify properties of metals, non-metals, and metalloids (1)
- identify and use the trends on the periodic table (electronegativity, ionization energy, atomic radius, shielding effect, nuclear charge, oxidation number) (3)
- relate electron configuration and ion formation and oxidation number based on the location on the periodic table (2)

38. Name the period 2 halogen F

39. Name the group 13, period 5 element IN

40. Name the 2 period, alkaline-earth metal Be

41 Is Cl a representative element, transition metal, or innertransition metal? ____________

42. Is U a representative element, transition metal, or innertransition metal? ____________

43. Answer "MT" metal, "MD" metalloid, or "NM" nonmetal.

a. shiny, hard, dense MT
b. forms negative ions NM
c. like metals and nonmetals MD
d. border the stairstep MD
e. to the left of the stairstep MT
f. to the right of the stairstep NM
g. form positive ions MT
h. brittle, usually solids and gases NM
i. semiconductors MD
j. poor conductors NM
k. good conductors MT
l. malleable and ductile MT

44.

	Metal, metalloid, or nonmetal?	Family?
a) Fe	MT	Transition
b) Si	MD	Group 14
c) Na	MT	Alkali
d) He	NM	Noble Gas
e) H	NM	Hydrogen/Alkali
f) W	MT	Transition
g) Np	MT	Inner Transition/Actinides
h) Mg	Metal	Alkali Earth

45. Matching

B ~~E~~ alkali metal
e alkaline-earth metal
d transition metal
C halogen
A noble gas

a. $1s^2 2s^2 2p^6$
b. $1s^2 2s^2 2p^6 3s^1$
c. $1s^2 2s^2 2p^5$
d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
e. $1s^2 2s^2 2p^6 3s^2$

46. Elements in the same group have similar properties because same # of valence e.

47. All Noble Gases are unreactive because the have a full octet (full s and p orbitals)

48. Define:

a) period
row

b) group
column

c) family
group

d) periodic law
Elements have ~~same~~ periodic repititions in their chemical + physical properties based on Atomic #

e) atomic radius (what is the trend?)

Decreases →
Increases ↓

f) ionization energy (what is the trend?)

Increases →
Decreases ↓

g) electronegativity (what is the trend?)

Increases →
Decreases ↓

49. Circle the element with the largest ionization energy:

a) ~~P, N~~ N b) Rb, Fr, Li c) Si, P, Mg

50. Circle the element in each pair that has the highest electronegativity:

a) K or Mg B) Mg or S c) F or He

51. Circle the element with the largest atomic radius.

a) Al or B b) Fe or Cu c) Br or Cl

Unit 4 - Chapter 8 Ionic Compounds (12 questions)

Be able to:

- determine the properties of ionic and metallic bonds (3)
- determine whether if bond is ionic based on the location on the periodic table (1)
- write formulas and names for ions and ionic compounds (7)

52. Why do ionic compounds form? ↑ stability achieve a stable octet

53. State the octet rule. atoms gain lose or share e⁻ to get 8 in their outer shell.

54. What happens to the valence electrons when an ionic bond is formed? Transfer

55. List 4 properties of compounds that have ionic bonds.

1) ↑ M.P.
2) Brittle
3) Good conductors as a solution/melted
4) Solid

- What causes most of these properties? Strong network (strong network or +/- attractions)

56. An ionic bond is between a Metal and a nonmetallic element.

57. What is a cation? positive ion and an anion? negative ion

58. What is the charge on the following ions?

a) silver ion Ag^{+}

b) zinc ion Zn^{2+}

c) lead (IV) ion Pb^{4+}

d) chloride ion Cl^{-} (now has same e-config as Ar)

e) potassium ion K^{+} (now has same e-config as Ar)

59. Complete the table for these ionic compounds:

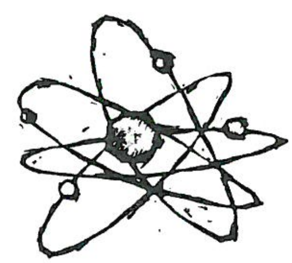
	Name	Formula
a)	ammonium oxide	$(NH_4)_2O$
b)	Sodium Sulfate	Na_2SO_4
c)	Potassium Phosphate	K_3PO_4
d)	zinc hydroxide	$Zn(OH)_2$
e)	iron (II) chloride	$FeCl_2$

60. Which elements need Roman numerals included in the name? Transition
What does the Roman numeral tell us? Oxidation State, Charge, # of e- lost.

Unit 5 - Chapter 9 Covalent Bonding (15 questions)

Be able to:

- determine the properties of covalent bonds (2)
- use Lewis Structures to determine shapes of molecules (including expanded octets) (4)
- use Lewis Structures to determine polarity (2)
- use electronegativity to determine the bond type (3)
- Determine the orbital hybridization, sigma and pi bonding in covalent compounds (2)
- determine the type of bond based on electronegativity differences (3)
- write formulas and names for covalent molecules (4)



61. Why do covalent compounds form? ↓ Energy be Stable, Share to get 8 in outer shell

62. What happens to the valence electrons when a covalent compound forms? Share

63. Covalent bonds normally form between 2 nonmetallic elements.

64. List 4 properties of covalently bonded compounds:

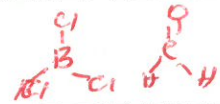
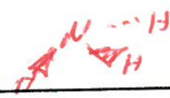
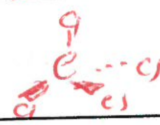
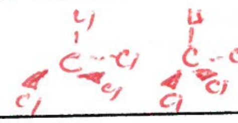
1) Soft
2) Poor conductors
3) ↓ B.P.
4) All 3 states

65. Another name for covelent compounds is Molecular Compound.

	Name	Formula	Ionic or Molecular?
66.	nitrogen trioxide	NO_3	Molecular
67.	calcium nitrate	$Ca(NO_3)_2$	Ionic
68.	trisulfur heptaoxide	S_3O_7	Trisulfur heptaoxide
69.	Ammonium Chloride	NH_4Cl	Ionic
70.	diphosphorous pentoxide	P_2O_5	Molecular

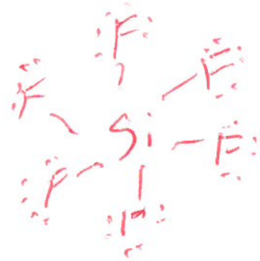
71. List the prefixes!!!! mono, di, tri, tetra, penta, hexa, hepta, octa, non, dec

72. What are the 8 shapes of molecules we studied? Complete the table.

	Name of shape	Picture	Example Lewis Structure	Can this shape be polar??
a)	Linear	O—O—O	Ö=C=Ö H−C≡N:	Yes
b)	Bent			Yes-always
c)	Trigonal planer			Yes
d)	Trigonal Pyramidal			Yes-always
e)	Tetrahedral			~~Polar~~ Yes
f) optional	Bent			Yes
g)				
h)				

	Molecular formula	Lewis Structure	shape	Polar Molecule?
73.	SCl_2	:C̈l−S̈−C̈l:	 Bent	Yes
74.	CO_2	O=C=O	Linear	NO

75. SiF_6



76. N_2 :N≡N:

77. What were the Lewis Structure exceptions we learned in Accelerated Chemistry?
Less then 8, odd # More then 8

78. Describe how electrons are shared/transferred and why in...

nonpolar bonds	**polar bonds**	**ionic bonds**
equal	unequal	Transfered

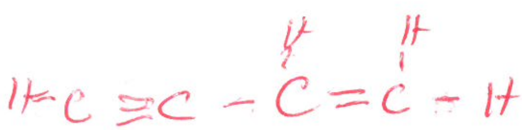
79. Use the electronegativities to determine if the following bonds are nonpolar, polar or ionic.

a) H-H ___0 N.P.___ (hydrogen e-neg=2.20)
b) Na-F ___3.0 Ionic___ (Na e-neg is 0.93 and F e-neg=3.98)
c) H-C ___.35 polarish___ (H e-neg = 2.20 and C e-neg = 2.55)

* More importantly ...ionic is always a ___Metal___ with a ___Nonmetal___.

80. Explain the difference between a sigma and a pi bond. Identify the number of sigma and pi bonds in the CO_2 molecule.

direct overlap (1st overlap) → parallel overlap (2nd & 3rd)



Unit 6 - Chapter 10 Chemical Reactions (12 questions)

Be able to:

- distinguish between the five types of chemical reactions (3)
- write and balance chemical reactions (3)
- interpret the law of conservation of mass (3)
- apply the rules of solubility for writing net ionic equations (2)

81. Matching

___d___ $A + BY \rightarrow AY + B$
___b___ $A + B \rightarrow AB$
___e___ $AY + BX \rightarrow AX + BY$
___a___ C_xH_x(hydrocarbon)$+O_2 \rightarrow CO_2 + H_2O$
___c___ $AB \rightarrow A + B$

a. combustion
b. synthesis
c. decomposition
d. single displacement
e. double displacement

82. *Balance* and *identify* the type of reaction for each of the following:

a. _____C + _____O_2 → _____CO_2 ___Synthesis___

b. $\underline{2}\,C_4H_{10}$ + $\underline{13}\,O_2$ → $\underline{8}\,CO_2$ + $\underline{10}\,H_2O$ combustion

c. $\underline{2}\,H_2O_2$ → $\underline{2}\,H_2O$ + ____ O_2 decomp

d. ____ Pb + ____ Hg_2SO_4 → ____ $PbSO_4$ + $\underline{2}$ Hg single replacement

e. ____ NaCl + ____ $AgNO_3$ → ____ AgCl + ____ $NaNO_3$ double replacement

h. $\underline{2}$ Cr + ____ $SnCl_4$ → ____ Sn + $\underline{2}\,CrCl_2$ single

83. Describe how you might get "No Reaction" for a Single Displacement reaction: single element lower
What about Double Displacement: two aq products

84. Write a balanced chemical equation for the following word descriptions.

a. Propane gas (C_3H_8) combusts.

$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$

b. The synthesis of potassium bromide KBr.

$2K_{(s)} + Br_{2(l)} \rightarrow 2KBr_{(s)}$

c. Iron reacts with a solution of copper (II) nitrate.

$Fe + Cu(NO_3)_2 \rightarrow Fe(NO_3)_2 + Cu$

d. NaCl + F_2 →

$2NaCl_{(aq)} + F_{2(g)} \rightarrow 2NaF_{(aq)} + Cl_{2(g)}$

e.. The reaction of barium chloride solution and sodium carbonate solution.

$BaCl_2 + Na_2CO_3 \rightarrow 2NaCl + BaCO_{3(s)}$

$Ba^{2+} + CO_3^{2-} \rightarrow BaCO_{3(s)}$

85. Label the coefficient and the subscript:

3 H_2O

coefficient — # of molecules

subscript — # of atoms in a molecule

- Know what each stands for!!!

86. Which side of the reaction arrow are the reactants on? left

87. List the 7 diatomic elements. H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2

88. We balance chemical equations because Mass (atoms) cannot be created or destroyed according to the law of conservation of Mass.

89. Write a) the balanced reaction and b) the net ionic equation for the reaction that occurs when a solution of sodium hydroxide reacts with a solution of copper (II) sulfate.

a) $2NaOH + CuSO_4 \rightarrow Na_2SO_{4(aq)} + Cu(OH)_{2(s)}$

b) $Cu^{2+} + 2OH^- \rightarrow Cu(OH)_2$

Unit 7 - CA 11 – The Mole (10 questions)

90. 1 mole Fe = 6.02x10^{23} atoms = 55.8 grams

91. Calculate the number of moles of a gold sample containing 3.33x10^{24} atoms.

$3.33 \times 10^{24} \text{ atoms} \cdot \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ atoms}} = \boxed{5.53 \text{ mol}}$

92. Calculate the grams of 1.25 moles of magnesium.

$1.25 \text{ mol} \cdot \frac{24.3 \text{ g}}{1 \text{ mol}} = \boxed{30.4 \text{ g}}$

93. Calculate the grams of 1.34x 10^{25} atoms of Lead.

$1.34 \times 10^{25} \text{ atoms} \cdot \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ atoms}} \cdot \frac{207.2 \text{ g}}{1 \text{ mol}} = \boxed{4610 \text{ g}}$

94. Determine the molar mass of ammonium dichromate $(NH_4)_2Cr_2O_7$. 252g/mol

95. Circle the empirical formulas, square the molecular formulas.

$\boxed{C_3H_6O_3}$, CH_2O, $\boxed{C_6H_{12}O_6}$

96. Determine the empirical formula for a compound having 80.68% mercury, 12.87% oxygen, and 6.45% sulfur.

$80.68 \text{ g Hg} \cdot \frac{1 \text{ mol}}{200.59 \text{ g}} = .402 \text{ mol}$ ÷ .201 mol = 2

$12.87 \text{ g O} \cdot \frac{1 \text{ mol}}{16.0 \text{ g}} = .804 \text{ mol}$ ÷ .201 mol = 4

$6.45 \text{ g S} \cdot \frac{1 \text{ mol}}{32.1 \text{ g}} = .201 \text{ mol}$ ÷ .201 mol = 1

$\boxed{Hg_2SO_4}$

97. Caffeine is a compound that was found to have the empirical formula – $C_4H_5N_2O$. If its molar mass is 194.19 g/mol calculate its molecular formula.

$C_4H_5N_2O$	97g
$C_8H_{10}N_4O_2$	194g

* –1

98. If the molecular compound of glucose is $C_6H_{14}O_6$, what is the empirical formula? $C_3H_7O_3$

99. Describe how you could get the molecular formula from the empirical formula.

Find molar Mass of Empirical & compare it to Molar Mass of Molecular

100. Calculate the percent composition of Lead (II) chloride $PbCl_2$.

$M = PbCl_2 = 278.2$

$Pb = \frac{207.2\,g}{\cancel{278.2}\,g} \times 100 = \boxed{74.5\%}$

$Cl = \frac{71\,g}{278.2\,g} \times 100 = \boxed{25.5\%}$

ACCELERATED CHEMISTRY
1st SEMESTER TEST

FORMULA SHEET

Element	Reaction
Li, Rb, K, Ba, Ca, Na	React with cold H_2O and acids, replacing hydrogen
Mg, Al, Mn, Zn, Fe	React with acids or steam, but usually not liquid water to replace hydrogen
Ni, Sn, Pb	React with acids but not water, to replace hydrogen
H_2, Cu, Hg	React with oxygen to form oxides
Ag, Pt, Au	Mostly unreactive

Halogen	Reaction
F_2, Cl_2, Br_2	Listed from most reactive to least

Solubility Rules

NO_3^{1-}	All nitrates are **soluble**.
Cl^{1-}	All chlorides are **soluble** except AgCl, Hg_2Cl_2, $PbCl_2$
SO_4^{2-}	Most sulfates are **soluble**; exceptions include: $SrSO_4$, $BaSO_4$, and $PbSO_4$. $CaSO_4$ is slightly soluble.
CO_3^{2-}	All carbonates are **insoluble** except those in Group 1 elements and NH_4^+
OH^{1-}	All hydroxides are **insoluble** except those of Group 1 elements, $Sr(OH)_2$ and $Ba(OH)_2$; $Ca(OH)_2$ is slightly soluble.
S^{2-}	All sulfides except those of Group 1 and 2 elements and NH_4^+ are **insoluble.**