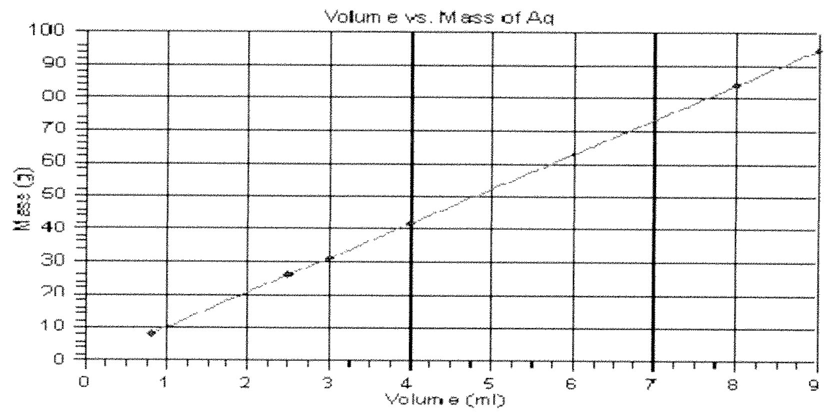
## Accelerated Chemistry Semester 1 Review Sheet    Key

The semester test will be given in two parts. The first part is a performance assessment and will be given the day before the semester test. This will include multiple parts and the questions will focus on interpretation of data, calculations, and short answer explanations for important chemical concepts. The second part given on the semester test day will include approximately 90 multiple choice questions. This part will be given online. No work will need to be shown.

### I. Important Science skills to be incorporated into all standards:

a. Create and interpret graphs




i. What is the expected mass of a silver object if the volume is 3.0 ml? 30.3g

ii. Use the information from the graph above to calculate the density of silver.

D=m/V  D= 30.3g / 3.0ml  10.1g/ml

b. Identify and use SI, base, modified, and derived units

i. What is the base unit of length M

ii. Why does volume not have a base unit. derived  $cm^3 = ml$

iii. Convert from $1.25 \times 10^{11}$ mm to km? Show all work below

$1.25 \times 10^{11} mm \cdot \frac{1 m}{1000 mm} \cdot \frac{1 km}{1000 m} = 125000 Km$ or $1.25 \times 10^5 Km$

iv. Give the value of each of the following prefixies

1. Milli $10^{-3}$  centi $10^{-2}$  kilo $10^{3}$

c. use scientific notation to express large and evaluate large and small measurements

i. Write $1.25 \times 10^6$ in regular notation 1250000

ii. Write $2.68 \times 10^{-5}$ in regular notation .0000268

iii. Write .00000659 in scientific notation $6.59 \times 10^{-6}$

d. Use significant digits to express precision in measurements

i. Identify the number of significant digits in each of the following measurements

1. 2.0040 m 5  .000560 g 3  10300 g 3
   $2.050 \times 10^8$mm 4  20.10 4

2. Perform the following operations:

   6.254 m + 1.3m 7.6m  11.79 g – 6 g 6g

   18.0 cm x 3.658 cm $65.8 cm^2$  21.6g /.00358 $m^3$  6030 $g/m^3$

e. Use dimensional analysis for unit conversions

i. Convert 30 days to seconds. Show all work below in dimensional analysis form

30days $\cdot \frac{24 hrs}{1 days} \cdot \frac{60 min}{1 hrs} \cdot \frac{60 sec}{1 min} = \boxed{2592000 sec}$

II. **UNIT I: Atomic structure and Nuclear**

**For this section review the following article (https://jj104.k12.sd.us/AP%20CHEM%20DOCUMENTS/starformationarticle.pdf ) and the video (https://www.youtube.com/watch?v=uKqvjEE0wFg)**

a. Describe how helium and a small amount of other light nuclei were formed from high energy collisions starting from protons and neutrons in the early universe before stars existed

i. The high pressure inside the star can literally squeeze together two protons, and sometimes, a proton will capture an electron to become a neutron.

b. Develop a model to describe how more massive elements up to iron are produced in the cores of stars by a chain of processes of nuclear fusion which also releases energy

i. When two protons and two neutrons band together, they form the nucleus of He, which is the second element in the periodic table. Then, when two nuclei of helium fuse with each other, they form the nucleus of another element, Be. In turn, the fusion of this element with helium produces a C nucleus; the fusion of carbon and helium nuclei leads to an oxygen nucleus, and so on. This way, through successive fusion reactions, the nuclei of most elements lighter than Fe can be formed.

c. Describe how supernovae explosions of massive stars are the mechanism by which elements more massive than iron are produced

i. The iron nucleus is the most stable nucleus in nature, and it resists fusing into any heavier nuclei. When the central core of a very massive star becomes pure iron nuclei, the core can no longer support the crushing force of gravity resulting from all of the matter above the core, and the core collapses under its own weight. It is during the few seconds of the collapse that the very special conditions of pressure and temperature exist in the Star that allow for the formation of elements heavier than Fe.

d. Determine the correlation between a star's mass and stage of development and the type of elements it can create during its lifetime

i. In stars less massive than the sun, the reaction converting hydrogen into helium is the only one that takes place Only in very heavy stars (that are more massive than eight solar masses), the chain reaction continues to produce elements up to iron.

ii. A young star is composed primarily of Hydrogen An example of this would be our sun. It combines Hydrogen to form helium and release a large amount of energy. This is an example of a nuclear Fusion reaction.

e. Analyze electromagnetic emission and absorption spectra to determine a star's composition, motion and distance to Earth

i. To determine which chemical elements are formed inside stars, scientists use a technique known as visible Spectroccopy.

ii. For example, the hydrogen's emission spectrum consists of four lines: purple, blue, green, and red, located at positions that correspond to their wavelengths. The emission spectrum of helium consists of six lines that are purple, cyan, green, yellow, orange, and red. In other words, atoms and molecules produce their own "unique spectrum" when the light they emit is spread in a spectroscope.

f. Describe the basic structure of an atom

i. Complete the following table:

| | Relative mass | Relative charge | location |
|---|---|---|---|
| Protons | 1 amu | + | nucleus |
| Nuetrons | 1 amu | 0 | nucleus |
| Electrons | 0 amu | − | electron cloud |

ii. Describe the nucleus of an atom

very dense tiny positive core

iii. Describe the dual nature of an electron

light can act as both a particle + wave

g. Determine the difference between atomic number and mass number

i. Magnesium with 13 neutrons has a mass number of 25 and an atomic number of 12. Magnesium with 12 neutrons has a mass number of 24 and an atomic number of 12.

h. Describe the development of the current atomic model

i. Dalton was the first with experimental evidence to prove the existence of atom

ii. JJ Thompson discovered the electron using a cathode ray tube

iii. Rutherford discovered the nucleus through the gold foil experiment

i. Identify an element by the number of protons

i. An element with 10 $p^+$, 10 $e^-$, and 11 $n^0$ is Neon

ii. Complete the following table

| | Protons | Neutrons | Electrons | Mass # |
|---|---|---|---|---|
| Aluminum - 27 | 13 | 14 | 13 | 27 |
| Potassium - 40 | 19 | 21 | 19 | 40 |
| Phosphorous - 31 | 15 | 16 | 15 | 31 |

iii. Write the nuclear symbol for Uranium with 144 neutrons $^{236}_{92}U$

j. Determine the number of protons and neutrons in the nucleus before and after decay

i. Identify the following particles

1. Alpha $^{4}_{2}He$ - Relative mass 4 - Charge +
2. Beta $^{0}_{-1}e$ - Relative mass 0 - Charge −
3. Gamma $\gamma$ - Relative mass 0 - Charge 0

k. Identify the emitted particles from alpha, beta, and gamma decay

i. Write the equation for the alpha decay of Uranium-242

$^{242}_{92}U \rightarrow \, ^{238}_{90}Th + \, ^{4}_{2}He$

1. Uranium begins with 92 protons and the resulting particle after the alpha decay has 90 protons

ii. Write the equation for the beta decay of Nickel – 65

$^{65}_{28}Ni \rightarrow \, ^{65}_{29}Cu \quad ^{0}_{-1}e$

1. Nickel begins with 37 neutrons and the resulting particle has 36 neutrons

iii. The two equations above are examples of nuclear Fission. Nuclear Fusion is the combination of two or more particles. Both processes can release a tremendous amount of energy

iv. Just recently the 7th period of the periodic table was completed. Scientists used a process of nuclear Fusion to create these elements.

l. Distinguish the difference between alpha beta and gamma decay

i. When Copper – 65 undergoes Alpha decay it produces Cobalt – 61

ii. When Copper – 65 undergoes Beta decay it produces Zinc - 65

m. Relate radioactive decay to element formation and transformation

i. When an unstable nuclei forms elements will Break apart to form more stable nuclei. This stable nuclei may be a new element or a different Isotope of the same element. This particle has the same number of protons and a different number of neutrons

**n.** Compare average atomic mass to mass number

i. Argon with 21 neutrons has a mass number of 39 and an atomic mass of 39.9 amu. Argon with 22 neutrons has a mass number of 40 an atomic mass of 39.9 amu. These two particles are Isotopes of each other.

ii. Calculate the average atomic mass for Nitrogen if it has two naturally occurring isotopes. Nitrogen – 14 which has a mass of 13.998 amu and a percent of 98.71 percent and Nitrogen 15 which has a mass of 14.998 amu and a percent of 1.29 percent.

(13.998 amu × .9871)
+ (14.998 amu × .0129)

$\boxed{14.01\ \text{amu}}$

iii. Boron has two naturally occurring isotopes B-10 and B-11 which is more abundant and why? Boron - 11 the avg mass of 10.81 amu is closer to 11

## III. Unit II: The Periodic Table

a. Describe the development of the periodic table and how it is arranged according to atomic structure

i. The modern periodic table was arranged by Mendeleev. The current version which closely resembled his version is arranged by increasing atomic #. Often times elements properties repeat every 8 elements.

ii. Elements in the same groups have Same # of valence e⁻ or similar properties

iii. Elements in the same period have Same # of energy levels

iv. Sodium has properties that are more similar to Potassium or Magnesium (Circle one)

v. Elements in the "D" block of the periodic table contain some of the first discovered elements. These metals are generally unreactive

vi. Elements In group Noble gases were among the last discovered because they are colorless and are rarely found in compounds

vii. Rare earth metals located in the f block contain both naturally occurring elements and radioactive/Man made elements. Many of these elements are Synthetic

viii. The S block contains only metals. The P block contains metals nonmetals and metalloids

b. Predict properties of elements based on their location on the periodic table. (reactivity, electronegativity, atomic size, ionic radius, ionization energy)

i. Metals become more reactive down a group and Nonmetals become less reactive down a group. Reactivity of both decrease as you get closer to the center of the periodic table

1. Pick the more reactive particle: 1. K or Ca 2. K or Rb 3. O or S 4. O or F 5. Na or Ne 6. Ca or Fe
2. Based on the electron configuration pick the more reactive particle
   a. [Ne] $3s^2$ or [Ne] $3s^2 3p^1$
   b. [Ne] $3s^2$ $3p^4$ or [Ne] $3s^2 3p^5$

ii. Electronegativity is the want for electrons in a chemical bond. Electronegativity increases across a period and decreases down a group.

1. Pick the element with a higher electronegativity: 1. F or Cl 2. Na or K 3. Cl or Ar 4. S or Cl 5. Na or Mg

iii. Atomic size Increase down a group due to shielding + more energy levels and decreases across a period due to increased nuclear pull

1. Pick the element with a larger radius: 1. Li or Na 2: Be or C 3. Br or I 4. Se or Br

iv. Ionic radius for metals Increases down a group and decreases across a period. Ionic radius for nonmetals Increases down a group and decreases across a period.

1. Pick the larger particle: 1. Na or $Na^+$ 2. O or $O^{2-}$ 3. $Cl^-$ or $S^{2-}$ 4. K or $K^{+}$ 5. $Ca^{2+}$ or $S^{2-}$ 6. $Al^{3+}$ or $P^{3-}$

v. Ionization energy is the ~~want for~~ energy to remove an electron it increases across a period and decreases down a group.

1. Pick the particle with the higher ionization energy: 1. Na or K 2. O or F 3. Cl or Ar 4. C or Si
2. Remember that there is an exception to ionization energies as you go across the periodic table. This happens after groups 2 and 5
   a. Knowing the exception which has a higher ionization energy: 1. Mg or Al 2. P or S

c. Using valence electrons students will be able to describe the behavior of atoms

i. Magnesium will lose 2 electrons to become a +2 ion
ii. Aluminum will lose 3 electrons to become a +3 ion
iii. Chlorine will gain 1 electron to become a −1 ion
iv. Nitrogen will gain 3 electrons to become a −3 ion

d. Determine the number and charges of stable ions within groups on the periodic table

i. Group 2 forms ions with a +2 charge
ii. Group 17 forms ions with a −1 charge
iii. Noble gasses do not form ions because stable octet / oribits are full
iv. Transition metals have multiple oxidation states. So to determine the charge a roman Numeral must be given. This tells us the number of electrons lost. For example Copper (I) has lost one $e^-$ and Copper (II) has lost 2 $e^-$

e. Correlate the electron configuration with the number of valence electrons within groups on the periodic table.

i. Write the electron configuration for Ca $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
ii. Write the orbital notation for Phosphorus

1s 2s 2p 3s 3p

iii. Write the noble gas configuration for Cu [Ar]- $4s^1 3d^{10}$ (exception)

iv. Write the noble gas configuration for Sn [Kr]- $5s^2 4d^{10} 5p^2$

v. Write the electron configuration of $Ca^{2+}$ $1s^2 2s^2 2p^6 3s^2 3p^6$

vi. Write the electron configuration of $S^{2-}$ $1s^2 2s^2 2p^6 3s^2 3p^6$

vii. Writhe the electron configuration of $Fe^{3+}$ $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$

viii. [Ne} $3s^2$ $3p^3$ is in the 5 or 15 group and the 3rd period

ix. [Kr] $5s^2$ $4d^{10}$ $5p^1$ is in the 3 or 13 group and the 5th period

x. Noble gases have full s and p sublevels giving them 8 electrons in their outer shells

xi. Beryllium is in the metal region, s block, 2nd group and has a noble gas configuration of [He]-$2s^2$

## IV. Unit III: Bonding

a. Describe the chemical and physical properties of covalent and ionic compounds. (volatility, melting point, boiling point, surface tension) due to interactions at the molecular and macroscopic level (intermolecular forces, shapes, expanded octets, polarity)

i. Ionic compounds consist of 2 ions. To identify an ionic compound they consist of a metal and a Nonmetal

ii. Crystal Lattice is the model of an ionic compound. As a result they have very high melting points, are hard and brittle, form electrolytes in water. They conduct electricity in solution but not as a solid

iii. Lewis Structur is the model of a covalent molecule. As a result they have low melting points because Intermolecular Forces have to be broken to melt these molecules. They Never conduct electricity, the stronger the intermolecular force the less volatile the compound, and the higher the surface tension.

iv. Arrange the following in order of highest to lowest melting point – $MgCl_2$, $C_{diamond}$, $CH_4$, $H_2O$, Fe Cdiamond, $MgCl_2$, Fe, $H_2O$, $CH_4$

v. Why does water have a higher melting point then $H_2S$ and $CO_2$
has Hydrogen bonds (H is bonded to N-O-or F), $H_2S$ is dipole/dipole $CO_2$ has dispersion

b. Experimentally determine what type of bond is present in a compound

i. A compound is hard, brittle and conducts electricity when dissolved in water. This most likely contains Ionic bonds

ii. After doing an experiment a student concluded a compound contained covalent bonds because when their group dissolved the compound in water it conducted an electrical current. Why is this student incorrect covalent does not conduct in solution

iii. A compound is shiny, is malleable and conducts electricity as a solid. It must contain metallic bonds

iv. A compound has extremely high melting points, it does not dissolve in water, it is very hard. It must be a covalent network solid

c. Determine the changes compounds undergo during a chemical change and a physical change

i. The five indicators of a chemical change are
Δ in color, Δ in temp Δ in smell

ii. Identify the following as a chemical or physical change:

1. Melting wax physical
2. Burning wood chemical

3. Leaf changing colors in the fall Chemical
4. An antacid tablet giving off heat and gas when placed in water chemical
5. Smashing a rock into smaller pieces physical

d. Determine the type of bond formed due to valance electrons and electronegativity (ionic vs. covalent)

i. Element X has an electronegativity of 1.0 and element Y has a value of 2.2. If they are in the same group which is more likely to be further left on the periodic table? X. What type of bond would form based on their electronegativity difference. Polar covalent difference of 1.2

ii. Two elements have an electronegativity difference of 0.0 it is a Non polar covalent bond. The electrons are shared equally Two elements have an electronegativity difference of 1.3. It is a polar bond. The electrons are shared unequally

iii. Fluorine has an electronegativity of 4.0 and magnesium has an electronegativity of 1.3. Give two reasons why this is an ionic bond
Metal + Non metal combined Difference of 2.7 is greater than 2.0

iv. In a covalent bond electrons are Shared in a ionic bond electrons are transfered - metals lose electrons, nonmetals want to gain electrons and in a metallic bond the behavior of an electron is described as sea of electrons

e. Describe the differences in structure in different types of compounds. (Lewis structure vs. crystal lattice)

i. Draw the Lewis structure of $H_2O$, $BH_3$, $CO_2$, $NH_3$, $SO_3$, and $CCl_4$ – Describe its shape, draw a 3-D image, identify its angle, its hybridization, and its bond and molecular polarity

| | | | | | | | |
|---|---|---|---|---|---|---|---|
| $H_2O$ | H–Ö–H | O, H, H | Bent | 105° | $sp^3$ | δ- δ+ O–H polar Bond | polar Molecule |
| $BH_3$ | H–B–H, H | H, B, H, H | trigonal planer | 120° | $sp^2$ | B–H polar δ- δ+ | Non polar Molecule |
| $CO_2$ | O=C=O | O–C–O | linear | 180 | sp | δ+ δ- C–O polar | Non polar Molecule |
| $NH_3$ | H–N–H, H | N, H, H, H | trigonal pyramidal | 107 | $sp^3$ | δ- δ+ N–H polar | polar Molecule |
| $SO_3$ | [O–S=O, O] ↔ [O=S–O, O] ↔ [O–S–O, O] | O, S, O, O | trigonal Planar | 120 | $sp^2$ | δ+ δ- S–O Polar Bond | Non polar Molecule |
| $CCl_4$ | Cl–C–Cl, Cl, Cl | Cl, C, Cl, Cl, Cl | tetrahedral | 109.5 | $sp^3$ | δ+ δ- C–Cl polar Bond | Non polar Molecule |

ii. Draw the lewis structure of $MgCl_2$. :Cl–Mg–Cl: . Why would a crystal lattice be a better model for this compound?
Ionic compounds form network of bonds

f. Name and write formulas for chemical compounds, including those with polyatomic ions and multivalent metals

i. Write the names of the following compounds

1. CaS Calcium Sulfide  $Rb_2O$ Rubidium Oxide
2. $CoCl_2$ Cobalt(II) Chloride  $Fe_2S_3$ Iron (III) Sulfide
3. $CCl_4$ Carbon Tetrachloride  $S_2Cl_5$ disulfur pentachloride
4. $Zn_3N_2$ Zinc Nitride  $Ni(NO_3)_2$ Nickel(II) Nitrate
5. $NH_4Cl$ Ammonium Chloride  $NH_4ClO_3$ Ammonium Chlorate
6. $K_2CO_3$ Potassium Carbonate  $NaHCO_3$ Sodium Bicarbonate
7. NaOH Sodium Hydroxide  $Al_2(SO_4)_3$ Aluminum Sulfate

ii. Write the formulas of the following compounds.

1. Lithium Bromide $LiBr$  Calcium Phosphide $Ca_3P_2$
2. Aluminum Nitrate $Al(NO_3)_3$  Beryllium Iodide $BeI_2$
3. Dinitrogen pentoxide $N_2O_5$  Trichlorine heptasulfide $Cl_3S_7$
4. Ammonium Hydroxide $NH_4OH$  Silver Phosphate $Ag_3PO_4$
5. Iron (III) Nitrite $Fe(NO_2)_3$  Chromium (II) Acetate $Cr(C_2H_3O_2)_2$
6. Hydrochloric acid $HCl$  Copper (II) Fluoride $CuF_2$

iii. Identify the cation and anion is the following compounds

1. $CaI_2$  $Ca^{2+}$  $I^-$
2. $Fe_3(PO_4)_2$  $Fe^{3+}$  $PO_4^{3-}$

## V. UNIT IV: Reactions

a. Differentiate between the different types of chemical reactions

i. A + B → AB is a Synthesis reaction

ii. XY → X + Y is a Decomposition reaction

iii. X + ZY → XY + Z is a Single replacement reaction. X is a metal
This will only occur If X is higher than Z on the activity series

iv. XY + WZ → XZ + WY is a double replacement reaction. If two aqueous solutions are formed then. No reaction occurs, If XZ is a solid write the net ionic equation. (All charges are +1 or -1`) ~~X + W~~ $X^+ + Z^- \rightarrow XZ_{(s)}$

v. The burning of Methane is classified as a Combustion reaction

b. Use solubility rules to determine ionization of salts in solution to write net ionic equations

i. Write the balanced chemical equation for the following. Be sure to include states of matter.

1. Solutions of Sodium Sulfide and Silver Nitrate are mixed.
$Na_2S_{(aq)} + 2AgNO_{3(aq)} \rightarrow Ag_2S_{(s)} + 2NaNO_{3(aq)}$
Write the net ionic equation
$2Ag^+_{(aq)} + S^{2-}_{(aq)} \rightarrow Ag_2S_{(s)}$

2. Solutions of Ammonium Chloride and Lead (II) nitrate are mixed
$2NH_4Cl + Pb(NO_3)_2 \rightarrow 2NH_4NO_{3(aq)} + PbCl_{2(s)}$
Write the net ionic equation
$Pb^{2+}_{(aq)} + 2Cl^-_{(aq)} \rightarrow PbCl_{2(s)}$

3. Copper is placed in a solution of Zinc Nitrate
$Cu + Zn(NO_3)_2 \rightarrow$ ~~$Cu(NO_3)_2 + Zn$~~ (NR)

4. Zinc is placed in a solution of Copper (II) Acetate
$Zn_{(s)} + Cu(C_2H_3O_2)_{2(aq)} \rightarrow Zn(C_2H_3O_2)_{2(aq)} + Cu_{(s)}$

5. Chlorine is bubbled through a solution of Sodium Iodide
$Cl_{2(g)} + 2NaI_{(aq)} \rightarrow 2NaCl_{(aq)} + I_{2(s)}$

c. Use the placement of elements on the periodic table to predict the type of reaction that occurs

i. Why would Nickel not replace Magnesium in a solution of Magnesium Chloride?
Nickel is below Mg on the activity series

ii. Write the balanced chemical equation for when Calcium is burned in the presence of Oxygen $2Ca_{(s)} + O_{2(g)} \rightarrow 2CaO_{(s)}$

iii. Write the balanced chemical equation for the synthesis of Aluminum Chloride from its elements. $2Al_{(s)} + 3Cl_{2(g)} \rightarrow 2AlCl_{3(s)}$

iv. Write the net ionic equation for the formation of water
$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

v. Write the reaction of water from its elements
$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

d. Describe the law of conservation of mass and how it relates to chemical equations

i. In a chemical reaction the total mass of the reactants must equal the total mass of the products.

ii. In an experiment a student heated 5.00 grams of a compound and recorded the mass of the product to be 3.88 grams. Describe what happened in this reaction and how it appears that mass was lost? Decomposition reaction in which 1.12g was released to the atmosphere.

iii. In a different experiment a student burned .750 grams of Magnesium and recorded the mass of the product to be .883 grams. Describe what happened in this reaction and how it appears that mass was gained. Synthesis reaction that combined with .133g of gas in the air to form products

## VI. Chemical Quantities

a. Calculate the molar mass of elements and compounds

i. What is the molar mass of Sulfur 32.1 g/mol what is the atomic mass of Sulfur 32.1 amu

1. Differentiate between the two
molar mass is mass of 1 mole of S - atomic mass is mass of 1 atom.

ii. What is the molar mass of Silver Sulfate 311.9 g/mol. What is the molar mass of Nitrogen tri iodide 394.7 g/mol $Ag_2SO_4$

b. Identify and describe the quantities of chemical reactants and products of a chemical reaction in terms of atoms, moles, mass and liters.

i. Calculate the mass in grams produced when 2.20 grams of Aluminum is combined with excess fluorine to produce Aluminum fluoride. $\boxed{\ \ g}$

$$\overset{2.20g}{2Al} + 3F_2 \rightarrow 2AlF_3$$

$$2.20g \cdot \frac{1\ mol}{27.0g} \cdot \frac{2\ mol}{2\ mol} \cdot \frac{84\ g}{1\ mol} \quad \boxed{6.84g}$$

ii. What volume of $CO_2$ would be produced when 1.50 moles of sodium carbonate reacts with excess HCl according to the following reaction. The density of $CO_2$ is 1.55 g/L.

$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$

1.50 mol

$1.50 mol \cdot \frac{1 mol}{1 mol} \cdot \frac{44 g}{1 mol} \cdot \frac{1 L}{1.55 g} = \boxed{42.6 L}$

iii. What is the number of molecules of lead (II) chloride that precipitates from 2.75 grams of lead (II) Nitrate solution and Sodium Chloride solution.

2.75g

$Pb(NO_3)_2 + 2NaCl_{(aq)} \rightarrow PbCl_{2(s)} + 2NaNO_{3(aq)}$

$2.75 g \cdot \frac{1 mol}{331.2 g} \cdot \frac{1 mol}{1 mol} \cdot \frac{6.02 \times 10^{23} cules}{1 mol} = \boxed{5.00 \times 10^{21} cules}$

iv. What is the mole relationship between $C_2H_6$ and $H_2O$ in a combustion equation

$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$ $\boxed{2:6}$

c. Describe how the mass of a substance can be used to determine the number of atoms, molecules or ions using mole relationships and avogadro's number.

i. Perform the following calculations

1. What is the number of atoms in 2.67 grams of Copper

$2.67 g \cdot \frac{1 mol}{63.5 g} \cdot \frac{6.02 \times 10^{23} atoms}{1 mol} = \boxed{2.53 \times 10^{22} atoms}$

2. What is the number of molecules in 2.22 moles of Sodium Oxalate

$2.22 mol \cdot \frac{6.02 \times 10^{23} cules}{1 mol} = \boxed{1.34 \times 10^{24} cules}$

3. Calculate the number of grams of 2.65 x $10^{25}$ molecules of water.

$2.65 \times 10^{25} cules \cdot \frac{1 mol}{6.02 \times 10^{23} cules} \cdot \frac{18.0 g}{1 mol} = \boxed{792 g}$

4. Calculate the number of atoms of Oxygen in 1.90 gram's of Aluminum Oxide

$Al_2O_3$ $1.90 g \cdot \frac{1 mol}{102 g} \cdot \frac{6.02 \times 10^{23} cules}{1 mol} \cdot \frac{3 atoms}{1 cules} = \boxed{3.36 \times 10^{22} atoms}$

5. For the reaction: $2KClO_3 \rightarrow 2KCl + 3O_2$ How many grams of potassium chlorate must be decomposed to yield 30.0 grams of oxygen.

$30.0 g \cdot \frac{1 mol}{32 g} \cdot \frac{2 mol}{3 mol} \cdot \frac{122.6 g}{1 mol} = \boxed{76.6 g}$

d. Determine the empirical and molecular formulas

i. Determine the percent composition of $Mg(NO_3)_2$ = 148.3g

24.3/148.3g x100 = 16.4% Mg

28.0g/148.3g x100 = 18.9% N

96.0g/148.3g x100 = 64.7% O

ii. Find the empirical formula for a compound that contains 36.84 % Nitrogen and 63.16% Oxygen.

$36.84 g N \cdot \frac{1 mol}{14.0 g} = 2.63 mol / 2.63 mol = 1$

$63.16 g O \cdot \frac{1 mol}{16.0 g} = 3.94 mol / 2.63 mol = 1.5$

$N_1O_{1.5} = \boxed{N_2O_3}$

iii. Find the molecular formula for a compound that contains 40.68 % C 5.08% H and 54.24% Oxygen. Its molar mass is 118.1 g.

$40.68 g C \cdot \frac{1 mol}{12 g} = 3.39 mol$ | $\frac{}{3.39 mol}$ $=1$

$5.08 g H \cdot \frac{1 mol}{1.0 g} = 5.08 mol$ $=1.5$

$54.24 g O \cdot \frac{1 mol}{16.0 g}$ $3.39 mol$ $=1$

$C_1O_1H_{1.5}$

| | |
|---|---|
| $C_2H_3O_2$ | 59 g |
| $C_4H_6O_4$ | 118 g |

e. Determine the limiting reactant and percent yield for a reaction

i. In the production of Lead (II) chloride 24.6 grams were produced. It was calculated that the theoretical yield of the reaction was 28.9 grams. Calculate the percent yield.

$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100$ $\frac{24.6 g}{28.9 g} \times 100 = \boxed{85.1\%}$

ii. For the reaction $\underset{10.0 g}{NaCl_{aq}} + \underset{10.0 g}{AgNO_{3aq}} \rightarrow NaNO_{3aq} + AgCl_s$. If 10.0 grams of both Sodium Chloride and Silver Nitrate react, identify the limiting and excess reactants. How many grams of Silver Chloride are theoretically produced? If 8.06 grams of precipitate were collected in the lab after filtration and proper dying time what is the percent yield and percent error.

$10.0 g\ NaCl \cdot \frac{1 mol}{58.55} \cdot \frac{1 mol}{1 mol} \cdot \frac{143.4 g}{1 mol} = 24.5 g$

$AgNO_3\ 10.0 g \cdot \frac{1 mol}{169.9 g} \cdot \frac{1 mol}{1 mol} \cdot \frac{143.4 g}{1 mol} = \boxed{8.44 g}$ theoretical

$AgNO_3$ = Limiting Reactant

NaCl = excess Reactant

$\% \text{ yield} = \frac{act}{theor} \times 100$ $\frac{8.06 g}{8.44 g} \times 100 = \boxed{95.5\%}$ $4.5\%$ error

f. Student will be able to connect real world reactions and actual output with experimental calculations.

i. 120 g **acetic acid** $HC_2H_3O_2$ (60 g/mol,) was reacted with 230 g **ethanol** $C_2H_5OH$(46 g/mol), yielding 132 g **ethyl acetate** (88 g/mol,). The percent yield was 75%. Without doing stoichiometry calculate the theoretical yield.

$\% \text{ yield} = \frac{act}{theor} \times 100$

$75\% = \frac{132 g}{x}$

$= \frac{132 g}{.75}$

$\boxed{= 176 g}$