General Chemistry

NAME ______________________________________ COURSE _____ SECTION _____

ID NO ___________ CLASS SCHEDULE ______________ REVIEWER _______________

MATTER AND ITS PROPERTIES

Matter
- anything that occupies space and has mass.

mass
- a measure of the quantity of matter in that body.

Properties of Matter

Properties are qualities possessed by substances which enable us to distinguish them from one another.

1. Extrinsic properties
    - ♦ properties which depend on how much matter is being considered.

      example: size, length, shape, width, volume, and mass

2. Intrinsic properties
    - ♦ properties which does not depend on how much matter is being considered.

      example: color, odor, taste

    Intrinsic properties can further be classified into:

    (a) physical properties
    - ♦ can be observed and measured without changing the composition or identity of a substance.

      example: color, odor, taste, density, melting point, boiling point.

    (b) chemical properties
    - ♦ properties associated with chemical change or the tendency of the substance to change either alone or by interaction with other substances.

      example: characteristic of iron to rust
      alcohol to burn
      wood to decay
      reaction of H_2 and O_2 to form H_2O

Changes in Matter

All matter are continually undergoing changes either through the efforts of nature or the influence of man.

1. Physical changes
    - ♦ changes that take place without modifying the chemical composition of matter.

      example: melting of ice
      chopping of wood
      tearing of paper

2. Chemical changes

- changes which involve alteration of the chemical composition of matter.
  example: rusting of iron
  digestion of food
  lighting a match

Evidences of a chemical change
(a) evolution of gas
(b) formation of a precipitate
(c) change in the chemical properties
(d) emission of light
(e) production of energy (chemical)

example: (a) When fuel burns, light is emitted and heat is released.
(b) When explosives are ignited, mechanical energy as well as heat and light are produced.

Energy changes in matter
1. endothermic change
   - involves the absorption of heat by the body from the surroundings.
     example: dissolving of sugar (cooling effect)

2. exothermic change
   - involves the release of heat from the body to the surroundings.
     example: burning of wood (heating effect)

Note: Matter and energy are so closely related that it is impossible to deal with one without considering the other. It takes energy to move matter and it takes matter to produce energy.

CLASSIFICATION OF MATTER

A. Pure Substances
- homogeneous materials that have fixed compositions and invariable intrinsic properties.

Types of Pure Substances
1. Elements
   - "simplest" form of matter.
   - a substance that cannot be separated into simpler substances by chemical means.

   Two (2) types of elements:
   (a) metals
   (b) non-metals

2. Compounds
   - a substance composed of atoms of two or more elements chemically united in fixed proportions.
     example: H_2O (composed of 11.19%$_{(w)}$ H_2 and 88.81%$_{(w)}$ O_2
     table salt (NaCl)

B. Impure Substances or Mixtures

- contains two or more distinct substances that can be physically separated from each other.

Types of Impure Substances or Mixtures

1. Homogeneous Mixtures (or SOLUTIONS)
   - mixtures whose composition are uniform throughout the entire region.
     example: sugar and water
     alcohol and water

2. Heterogeneous Mixtures (or SUSPENSIONS)
   - different components occupy distinct regions within the sample.
     example: gravel and sand
     oil and water

ATOMIC NUMBER and MASS NUMBER

Atomic Number (Z) - is the number of protons in the nucleus of an atom.

Mass Number (A) - is the sum of the number of protons and neutrons in the nucleus.

Consider the isotopic symbol:

$$^{A}_{Z}X$$

where: A – mass number
Z – atomic number
X – symbol of any element

For neutral atoms: no. of protons (p) = no. of electrons (e)
*the number of neutrons (n) is calculated using: **n = A – Z**
For ions: no. of protons (p) is **NOT EQUAL** to the no. of electrons (e)

A. For (+) ions: (CATIONS)

$$^{A}_{Z}X^{m+}$$

p = Z
e = Z – m^{+}
n = A – Z

B. For (-) ions: (ANIONS)

$$^{A}_{Z}X^{n-}$$

p = Z
e = Z + n^{-}
n = A – Z

Examples:

1. How many protons, electrons, and neutrons are there in the following species?

(a) $^{63}_{29}Cu$ (b) $^{27}_{13}Al^{3+}$ (c) $^{32}_{16}S^{2-}$

2. What is the isotopic symbol for a specie with 7 protons, 10 electrons and a mass number of 14?

3. Write the nuclide symbol for a specie that has a mass number of 24, atomic number of 12 and the no. of electrons being 10.

4. Write the isotopic notation for an ion with 26 protons, 24 electrons and 30 neutrons.

5. Write the isotopic notation for an ion with an atomic number of 15, has 16 neutrons in its nucleus, and is a member of group VA of the periodic table.

STOICHIOMETRY: MASS RELATIONSHIPS IN CHEMICAL REACTIONS

Mole (symbol: n)

- the SI quantity that describes an amount of substance by relating it to a number of particles.
- it tells us how many atoms or molecules can an element or a compound have.

Avogadro's Number (N_A)

- is the number of elementary particles per mole.
- is numerically equivalent to 6.022×10^{23} particles → atoms, molecules, formula units

therefore: 1 mol X = 6.022×10^{23} atoms X
1 mol X_2 = 6.022×10^{23} molecules X_2

Examples:

1 mol Si = 6.022×10^{23} **atoms** Si
1 mol N_2O_5 = 6.022×10^{23} **molecules** N_2O_5
1 mol Al = 6.022×10^{23} **atoms** Al
1 mol NH_3 = 6.022×10^{23} **molecules** NH_3
1 mol K_2SO_4 = 6.022×10^{23} **formula units** K_2SO_4
1 mol NaCl = 6.022×10^{23} **formula units** NaCl

Molar Mass (MM)

- the mass of one mole of atoms, molecules, or formula unit.

*for elements, its simply the atomic mass found in the periodic table.
*for compounds, add it up by the element and get the sum.

Sample Problems:

1. How many atoms are in 16.3 g of S? **Answer: 3.06×10^{23} atoms S**
2. How many moles of He are there in 6.46 g He? **Answer: 1.62 moles He**
3. What is the mass of one silver (Ag) atom? **Answer: 1.791×10^{-22} gram Ag**

Molecular Mass

- is the sum of the atomic masses (amu) of all atoms in a molecule.

*the molecular mass in amu is numerically equal to the molar mass in grams.

Examples:

1. Calculate the:
   (a) molecular mass of SO_2. **Answer: 64.07 amu**

   (b) molar mass of $C_6H_{12}O_6$. **Answer: 180.18 g/mole** (grams/mole)

   (c) atomic mass of magnesium. **Answer: 24.31 amu**

   (d) molar mass of Na_3AsO_4. **Answer: 207.89 g/mole**

2. Methane (CH_4) is the principal component of natural gas. How many moles of CH_4 are present in 6.07 g CH_4? **(Answer: 0.378 mole CH_4)**

3. How many hydrogen atoms are present in 25.6 g $(NH_2)_2CO$ which is a compound used as a fertilizer, in animal feed and in the manufacture of plastics. **(Answer: 1.03 x 10^{24} atoms H)**

PERCENT COMPOSITION OF COMPOUNDS

% composition by mass

- is the percent by mass of each element in a compound.

Consider a hypothetical compound $X_aY_bZ_c$ where X, Y, and Z are elements' symbols and a, b, and c are the number of atoms of the respective elements (subscripts). If we assume that the molar mass of $\mathbf{X_aY_bZ_c}$ is **105.0 g/mole**, then:

$$\% X = \frac{a(MM\ of\ X)}{105.0}\ x\ 100$$

$$\% Y = \frac{b(MM\ of\ Y)}{105.0}\ x\ 100$$

$$\% Z = \frac{c(MM\ of\ Z)}{105.0}\ x\ 100$$

Problems:

1. Calculate the percent composition of phosphoric acid, H_3PO_4 which is responsible for the tangy taste in cola drinks. **Answer: 3.09% H, 31.60% P, 65.31% O**
2. Determine the mass percent of water in the hydrate $Cr(NO_3)_3 \bullet 9H_2O$. **Answer: 40.52% H_2O**
3. Adenosine Triphosphate (ATP for short) is the main energy storage molecule in cells. ATP

has the formula $C_{10}H_{11}N_5P_3O_{13}$. What is its percent composition?
Answer: 23.92% C, 2.21% H, 13.95% N, 18.50% P, 41.42% O

EMPIRICAL FORMULA CALCULATIONS

Empirical Formula (EF)
- is the "***simplest***" formula of a compound.
- in this type of formula, the subscripts of the elements in the compound are expressed in ***lowest*** fraction.

Steps in the determination of empirical formula:
1. Assume 100 grams of sample and determine the mass of each element.
2. Convert all masses (in grams) into moles.
3. Write a "tentative" formula based on your answer in step 2.
4. Divide each of the subscripts from the tentative formula by the smallest subscript.
5. Round off any subscripts from step 4 that differ only slightly from whole numbers.
6. If the subscripts will not give a whole number, choose a small whole number which will act as a multiplier to all the subscripts to make them integral.

Sample Problems:
1. Provide the empirical formula for each of the following:
   (a) $C_6H_{12}O_6$
   (b) N_2O
   (c) $C_8H_{10}N_4O_2$
   (d) $K_2Cr_2O_7$
   (e) Si_2Br_6
2. An antiseptic agent contains 24.75% K, 34.77% Mn and 40.51% O. What is the simplest formula (EF) of this antiseptic solution?
3. Ascorbic acid is the generic name of Vitamin C. A molecule of vitamin C is analyzed to contain 40.92% C, 4.58% H and the rest being oxygen. Determine its empirical formula.

Answers:
1. (a) CH_2O
   (b) N_2O
   (c) $C_4H_5N_2O$
   (d) $K_2Cr_2O_7$
   (e) $SiBr_3$
2. $KMnO_4$
3. $C_3H_4O_3$

MOLECULAR FORMULA CALCULATIONS

Molecular Formula (MF)
- is the "***true***" formula of a compound.
- in this type of formula, the exact number of atoms of each element in a molecule are shown, thus, the subscripts are not in the lowest fraction.

Steps in the determination of molecular formula:

1. Follow steps 1 – 6 in the usual empirical formula determination.
2. Calculate the molar mass of the empirical formula.
3. Divide the estimated molar mass (always given in the problem) by the empirical molar mass obtained in step 2 to get the factor, n.

$$n = \frac{\text{estimated molar mass}}{\text{empirical molar mass}}$$

4. Distribute the value of n to the EF to obtain the MF (molecular formula).

$$MF = (EF)_n$$

Sample Problems:

1. A sample of a compound of nitrogen and oxygen contains 1.52 g N and 3.47 g O. The molar mass of this compound is found to be 95 g. Determine the molecular formula of this compound. **Answers: EF = NO_2 , MF = N_2O_4**

2. Nicotine is found in tobacco leaves and is used to manufacture cigarettes. The percentage composition of nicotine was found to be 74.0% C, 8.7% H, and 17.3% N.
   (a) What is the empirical formula? **Answer: C_5H_7N**
   (b) If the molar mass of nicotine was known to be 162.2 g, determine its molecular formula. **Answer: $C_{10}H_{14}N_2$**

LIMITING REACTANT AND REACTION YIELD

Limiting Reagent
- the reactant used up first in a chemical reaction.

Excess Reagent
- the reactant present in quantities greater than necessary to react with the quantity of the limiting reagent.

Theoretical Yield
- the amount of product that would result when all the limiting reagent has already reacted.
- also known as the *"maximum"* obtainable yield.

Actual Yield
- the amount of product actually obtained from a reaction.
- it is always less than the theoretical yield.

Percent Yield
- the ratio of the actual yield to the theoretical yield, expressed as a percentage.

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

where: theoretical yield > actual yield (at ALL times!)

note: % yield will **never** exceed 100.

Illustrative Example:

A recipe calls for 3.5 cups of flour and 2 well-beaten eggs to make 60 cookies. The chef has 8.75 cups of flour and 9 well-beaten eggs.

(a) Write the representation of this cookie-making process in chemical equation form.
(b) What is the maximum number of cookies that can be prepared? **Answer: 150**
(c) Identify the limiting ingredient. **Answer: Flour**
(d) Which of the two ingredients is in excess and by how much? **Answer: eggs, 4 eggs leftover**

Problems:

1. A commercial method of manufacturing hydrogen gas involves the reaction of iron with steam (or water gas). The balanced equation is:

   $$\mathbf{3}Fe(s) + \mathbf{4}H_2O(g) \rightarrow Fe_3O_4(s) + \mathbf{4}H_2(g)$$

   In one process, 16.8 g Fe are allowed to react with 10.0 g H_2O.
   (a) Identify the limiting reagent. **Answer: Fe**
   (b) What maximum amount of Fe_3O_4 (in grams) can be obtained? **Answer: 23.2 g**
   (c) If 21.8 g Fe_3O_4 are actually obtained, calculate the percent yield. **Answer: 94.0%**

2. Urea, $(NH_2)_2CO$, is used as fertilizer and as animal feed and is also an important source in the manufacture of plastics. It is prepared between ammonia (NH_3) and carbon dioxide (CO_2) according to the balanced equation:

   $$\mathbf{2}NH_3(g) + CO_2(g) \rightarrow (NH_2)_2CO(aq) + H_2O(l)$$

   In one process, 637.2 g NH_3 are made to react with 1142 g CO_2.
   (a) Which of the two reactants is the limiting reagent? **Answer: NH_3**
   (b) Calculate the mass of urea formed. **Answer: 1123 g $(NH_2)_2CO$**
   (c) How much of the excess reagent (in grams), is left at the end of the reaction? **Answer: 319 g CO_2 left unreacted**

QUANTUM THEORY AND ELECTRON CONFIGURATION

<u>Quantum Numbers</u>

- are used to describe the distribution of electrons in atoms.
- there are four (4) quantum numbers, the first three (3) are used to describe atomic orbitals and to label electrons that reside in them. The fourth quantum number describes the

behavior of a specific electron and completes the description of electrons in atoms.

<u>The Four Quantum Numbers</u>

I. The Principal Quantum Number (n)
- determines the energy level of an orbital.
- relates to the average distance of the electron from the nucleus in a particular orbital.
- can have integral values of 1, 2, 3, ….., n.

Note: The **<u>larger</u>** the value of n, the **<u>larger</u>** the orbital.

II. The Angular Momentum Quantum Number (l)
- tells us the "shape" of the orbitals.
- the values of **l** depend on the value of the principal quantum number, n.
- can have integral values of **0** to **(n – 1)**.

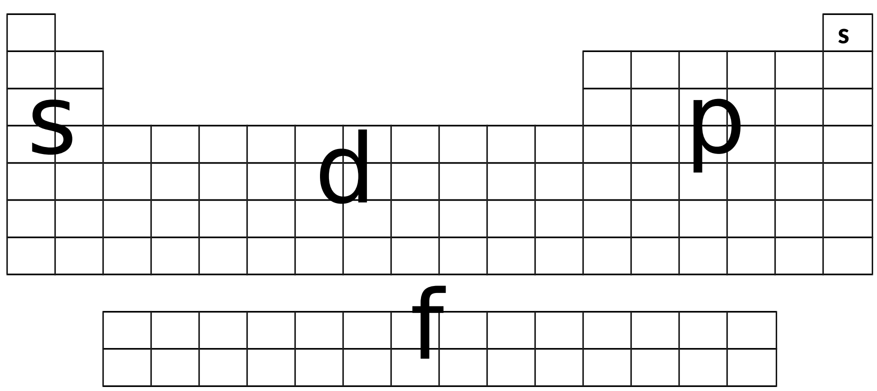
Table relating the value of "l" to the orbital's symbol (must be memorized!)

l	0	1	2	3	4	5
symbol of orbital	s	p	d	f	g	h

Shell
- a collection of orbitals with the same value of "n".

Subshell
- one or more orbitals with the same value for "n" but different values of " l ".

<u>Relative positions of the s, p, d, f orbitals in the periodic table</u>



III. The Magnetic Quantum Number (m)
- describes the orientation of the orbital in space.
- the values of m depends on the value of **l**.
- for a given value of **l**, m would have integral values of –**l** to +**l**.

–**l**, (–**l** + 1), …..0…..(+**l** – 1), +**l**

IV. The Electron Spin Quantum Number (s)

- describes the possible spinning motion of an electron.
- there are two (2) possible spinning motions of an electron, one **clockwise** and the other **counterclockwise**.
- to account for the difference in spin, we adopt s = +½ and s = –½.

Note: Every orbital can contain **two (2) different electrons**.

These four (4) quantum numbers enable us to completely label an electron in any orbital in an atom. These four quantum numbers can be regarded as the "**address**" of an electron in an atom.

Atomic Orbital

- a region in space close to an atom's nucleus in which one or two electrons can reside.

Since every orbital can contain 2 different electrons, then:

1 orbital = 2 electrons

it follows that:

Orbital Symbol	Maximum electrons
s	2
p	6
d	10
f	14

Sample Problems:

1. List different ways on how to write the four quantum numbers that designate an electron in a **3p** orbital.
2. Write the four quantum numbers for an electron in a **5s** orbital.
3. Indicate which of the following sets of quantum numbers in an atom are **unacceptable**. ***Encircle*** the specific quantum number which makes it incorrect and explain briefly why.
   (a) (1, 0, ½, ½)
   (b) (3, 0, 0, ½)
   (c) (2, 2, 1, ½)
   (d) (4, 3, –2, –½)
   (e) (3, 2, 1, 1)

Answers:

1. (3, 1, -1, ½) ; (3, 1, -1 -½) ; (3, 1, 0, ½) ; (3, 1 0, -½) ; (3, 1, 1, ½) ; (3, 1, 1, -½) **6 possible sets.**

2. (5, 0, 0, ½) and (5, 0, 0, -½) only… therefore **2 possible sets**.

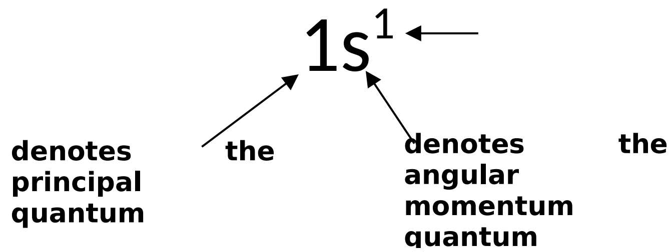
Electron Configuration

- tells us how the electrons are distributed among the various atomic orbitals.

Convention:

Using the hydrogen-atom ground state as example:

denotes the number of electrons in the orbital or



Orbital Diagram (OD)

- shows us the distribution of electrons in an orbital by the rectangular box-and-arrow convention.

For hydrogen,

H [↑] $1s^1$

Rules in filling atomic orbitals

I. The Aufbau Principle
   - derived from the German word "***aufbau***" which means **building up**.
   - the principle which follows the filling of energy subshells/sublevels in the order of **increasing** energy.

Examples:

Write the electron configuration for each of the following:

(a) Be **Answer: $1s^2 2s^2$**
(b) F **Answer: $1s^2 2s^2 2p^5$**
(c) Fe **Answer: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$**
(d) V **Answer: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$**
(e) Ge **Answer: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$**

II. The Pauli's Exclusion Principle
   - "No two electrons can occupy the same orbital/subshell."
   - "No two electrons can have the same set of four quantum numbers."

Example:

Consider a Helium (He) atom, the possible ways of drawing the orbital diagram are as follows:

OD no. 1 **OD no. 2**

[↑↑] $1s^2$ [↓↓] $1s^2$ [↑↓] $1s^2$

Notice that the first two orbital diagrams (OD no. 1 and OD no. 2) have **EXACTLY** the same four quantum numbers, as such, these are out-ruled and hence, violated Pauli's Exclusion Principle. Therefore, the only correct orbital diagram is OD no. 3 since the two sets of quantum numbers written are distinct.

Sample Problems:

1. Draw the orbital diagram for each of the following:
   (a) nitrogen (Z = 7)
   (b) sodium (Z = 11)

2. Draw the orbital diagram for the <u>valence electrons</u> of each of the following:
   (a) arsenic (As)
   (b) tellurium (Te)
   (c) antimony (Sb)

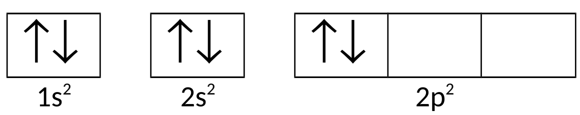
III. Hund's Rule
- states that: "The most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins."

Example:

Carbon, C, has an atomic number of 6. The electron configuration for carbon in its ground state is:

$$EC = 1s^2 2s^2 2p^2$$

The possible orbital diagrams are:



Remarks: ______________________
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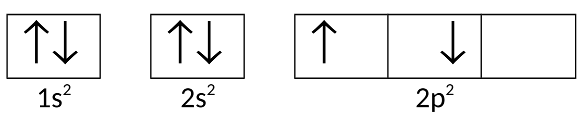

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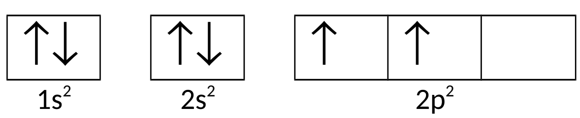

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<u>Diamagnetism and Paramagnetism</u>

Diamagnetic
- electron spins are completely paired or are anti-parallel to each other.

Paramagnetic
- unpaired electrons are still present and are considered to be parallel in nature.

Examples:

For each of the following:
(a) write the electron configuration.
(b) draw the orbital diagram and classify whether the electrons are ***paramagnetic*** or ***diamagnetic***.

1. Li (Z = 3)
2. Ne (Z = 10)
3. K (Z = 19)

Partial Answers:
1. paramagnetic (odd atomic number, 1 unpaired electron)
2. diamagnetic (fully filled highest shell, $1s^2 2s^2$ **$2p^6$**, NO unpaired electrons)
3. paramagnetic (odd atomic number, 1 unpaired electron)

<u>**Noble Gas Core (NGC)**</u>
- shows in square brackets, the noble gas element that most nearly precedes the element being considered, followed by the symbol of the highest filled subshells in the outermost shells.

Examples:

Write the:
(a) complete electron configuration
(b) noble gas core configuration (NGC)

1. sulfur (Z = 16)
2. calcium (Z = 20)
3. krypton (Z = 36)
4. rubidium (Z = 37)

Answers:
1. NGC = $[Ne]3s^2 3p^4$
2. NGC = $[Ar]4s^2$
3. NGC = $[Ar]4s^2 3d^{10} 4p^6$
4. NGC = $[Kr]5s^1$

SOLUTIONS AND CONCENTRATION UNITS

Concentration is a general term expressing the amount of solute obtained in a given amount of material.

1.) **Molarity**

It is the number of moles of solute divided by the number of liters of solution. It can also be expressed as millimoles (mmoles) over milliliters (mL).

$$\text{Molarity} = \frac{\text{moles of the solute}}{\text{liter of solution}} = \frac{\text{mmoles of the solute}}{\text{milliliter of solution}}$$

Note: It is important to first convert the given mass into number of moles by dividing the mass of the substance by its molar mass.

$$\text{no. of moles} = \text{g of solute} \times \frac{\text{1 mole of solute}}{\text{molar mass of solute (g)}} = \text{mg of solute} \times \frac{\text{1 mmole of solute}}{\text{molar mass of solute (mg)}}$$

Sample Problem:

Calculate the molarity of a 0.0105 L aqueous solution containing 0.82 g ethanol (C_2H_5OH, MM = 46.08 g/mole).

Solution:

First: Calculate for the number of moles of ethanol

$$\text{moles of } C_2H_5OH = 0.82 \text{ g } C_2H_5OH \times \frac{1 \text{ mole } C_2H_5OH}{46.08 \text{ g } C_2H_5OH} = 0.018 \text{ mole } C_2H_5OH$$

Second: Use the formula for molarity and substitute the number of moles of ethanol into the equation.

$$\text{Molarity} = \frac{0.018 \text{ moles } CaCl_2}{0.0105 \text{ L solution}} = 1.7 \text{ M } C_2H_5OH$$

2.) **Molality**

It is the number of moles of solute divided by the number of kilograms of solvent. It can also be expressed as millimoles (mmoles) over grams (g).

$$\text{Molality} = \frac{\text{moles of the solute}}{\text{kilograms of the solvent}} = \frac{\text{mmoles of the solute}}{\text{kilograms of the solvent}}$$

Sample Problem:

What is the molality of a solution prepared by dissolving 32.0 grams of calcium chloride, $CaCl_2$ (MM = 110.98 g/mole) in 271 g of water?

Solution:

First: Calculate for the number of moles of $CaCl_2$

$$\text{moles of } CaCl_2 = 32.0 \text{ g } CaCl_2 \times \frac{1 \text{ mole } CaCl_2}{110.98 \text{ g } CaCl_2} = 0.288 \text{ mole } CaCl_2$$

Second: Use the formula for molality and substitute the number of moles of $CaCl_2$ into the equation. Use the 271 g of water but convert it first into kilograms (kg)

$$\text{Molality} = \frac{0.288 \text{ moles } CaCl_2}{271 \text{ g} \times \frac{1 \text{ kg}}{1000 \text{ g}}} = 1.06 \text{ m } CaCl_2$$

3.) **Percent concentrations**

It is the most common manner of reporting the final result of a quantitative determination is by percent (parts per hundred) of the analyte. The percentage can be in terms of weight or volume.

$$\%\text{ weight} = \frac{\text{mass of solute}}{\text{mass of solution}} \text{ x } 100\%$$

Note: The above equation do not carry any unit except %.

Practice Problems:

MOLARITY

1.) How many moles of Na_2CO_3 are present in 50.00 mL of 0.1204 M Na_2CO_3.
2.) How many mL of HNO_3 must be measured in order to obtain 4.36 millimoles of HNO_3 from a solution of 0.1815 M HNO_3.
3.) Calculate the concentration of a 250.0 mL solution that contains 1.348 g of $Al_2(SO_4)_3$. The molar mass (MM) of $Al_2(SO_4)_3$ is 342.17 g/mole.

MOLALITY

1.) Calculate the molality of a sulfuric acid solution containing 24.4 g of sulfuric acid, H_2SO_4 (MM = 98.09 g/mole) in 198 g of water.
2.) What is the molality of a solution containing 7.78 g of urea [$(NH_2)_2CO$, molar mass, MM = 60.07 g/mole] in 203 g of water.

PERCENTAGE CONCENTRATIONS

1.) Calculate the concentration of a 250.0 grams solution that contains 1.348 grams of $HClO_3$.
2.) How many grams of the solution must be measured in order to obtain 4.36 grams of HNO_3 from a solution of 18.15% HNO_3?
3.) How many grams of Na_2CO_3 are present in 50.00 mL of a 12.04 % (w/v) Na_2CO_3?

Answers:

MOLARITY

1.) 6.02×10^{-3} mole

2.) 24.0 mL HNO_3
3.) 0.01576 M

MOLALITY

1.) 1.26 m
2.) 0.638 m

PERCENTAGE CONCENTRATIONS

1.) 0.5392%(w/w) $HClO_3$
2.) 24.0 g solution
3.) 6.02 g Na_2CO_3

ADDITIONAL LECTURE NOTES: