Topic IV-2 Study Guide

<u>II. The Mole Concept:</u>

A. Mole Amount Calculations

1. Definitions
2. Molar amounts of elements & compounds

<u>mole:</u> SI unit for quantity

$\frac{6.02 \times 10^{23} \text{ particles}}{1 \text{ mole}}$ <-----molecules (H_2O), atoms (H_2), ions (Ca^{2+})

$\boxed{\text{PARTICLES}}$ <---------------> $\boxed{\text{MOLE}}$<-------------> $\boxed{\text{MASS (g)}}$

I — **Avo. #** $\frac{6.02 \times 10^{23} \text{ particles}}{1 \text{ mole}}$

I — **Molar Mass**

<u>Examples:</u>

10.0 g Al
1) ? mole
2) ? particles

1) $10.0\text{g Al} \times \frac{1 \text{ mol Al}}{27.0 \text{ g Al}} =$ **0.371 mol Al**

2) $0.371 \text{ mol Al} \times \frac{6.02 \times 10^{23} \text{ particles}}{1 \text{ mole}} =$ **2.23×10^{23} atoms Al**

A silicon chip has a mass of 5.68 mg. How many silicon atoms are present in this chip?

$5.68 \text{ mg Si} \times \frac{1\text{g}}{1 \times 10^3 \text{ mg}} \times \frac{1 \text{ mol Si}}{28.1 \text{ g Si}} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol}} =$ **1.22×10^{20} Si atoms**

Chromium is added to steal to improve resistance to erosion. Calculate the <u># of moles</u> and <u>mass</u> of Cr contained in 5.00×10^{20} atoms.

1) 5.00 x 1020 atoms Cr x <u>1 mol Cr</u>_________ = **8.30×10^{-4}_mole Cr**

6.02 x 1023 atoms

2) 8.30×10^{-4} mol Cr x $\frac{52.0 \text{ g Cr}}{1 \text{ mol}}$ = **4.32×10^{-2} g Cr**

3. Volume of gases: STP & Avogadro's Hypotheses

Avogadro's Hypothesis: Equal volumes of gases at the same temperature & pressure contain equal numbers of particles.

1 mole of any gas at STP occupies a volume of 22.4 L (molar volume)

STP conditions: Standard Temperature & Pressure

T: 0^{o} C = 273 K
P: 1 atm= 760 mm

molar volume: (STP)

$\frac{22.4 \text{ L}}{1 \text{ mol}}$ **or** $\frac{1 \text{ mol}}{22.4 \text{ L}}$

Example:

Nitrogen gas
N_2
1.75 L at STP
? grams

1.75 L N_2 X x $\frac{1 \text{ mol } N_2}{22.4 \text{ L } N_2}$ x $\frac{28.0 \ N_2}{1 \text{ mol } N_2}$ = **2.19 g N_2**

152 g CO_2 @ STP
? L CO_2

152 g CO_2 x $\frac{1 \text{ mol } CO_2}{44.0 \text{ g } CO_2}$ =3.45 mol CO_2

3.45 mol CO_2 x $\frac{22.4 \text{ L}}{1 \text{ mol}}$ = **77.3 CO_2**

B. Empirical & Molecular Formulas

1. From laboratory data
2. From percent composition

"Percent to mass, mass to mole, divide by small, multiply till whole."

Empirical: simplest whole # ratio of atoms present in a molecule "simplest"

Molecular: Actual composition of molecules present. "Actual"

Molecular	**Empirical**
C_6H_6	CH
$C_{12}H_4Cl_4O_2$	$C_6H_2Cl_2O$
$C_6H_{16}N_2$	C_3H_8N

Calculations of Empirical Formula:

1. Obtain the mass of each element in grams.
2. Determine # moles of each type of atom in the compound.
3. Divide $\frac{\text{\# moles of each element}}{\text{smallest \# of moles}}$
4. If all numbers are not whole numbers, multiply by smallest integer that will convert to whole #'s.

Example:

65.02 % Pt= 65.02 g Pt
9.34 % N= 9.34 g N
2.02 % H= 2.02 g H
23.63 % Cl= 23.63 g Cl

Find Emp. Formula:

$$65.02 \text{ g Pt} \times \frac{1 \text{ mol Pt}}{1951.1 \text{ g Pt}} = \frac{0.3333 \text{ mol Pt}}{0.3333} = 1$$

$$9.34 \text{ g N} \times \frac{1 \text{ mol N}}{14.0 \text{ g N}} = \frac{0.6670 \text{ mol N}}{0.3333} = 2$$

$$2.02 \text{ g H} \times \frac{1 \text{ mol H}}{1.0 \text{ g H}} = \frac{2.02 \text{ mol H}}{0.3333} = 6$$

$$23.63 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.5 \text{ g Cl}} = \frac{0.6666 \text{ mol Cl}}{0.3333} = 2$$

$PtN_2H_6Cl_2$

Molecular Formula
*Given molar mass of molecular formula

$$n = \frac{\text{molecular mass}}{\text{empirical mass}}$$

Example:

71.65 % Cl = 71.65 g Cl
24.27 % C= 24.27 g C
4.07 % H= 4.07 g H

Molar mass (molecular)=98.96 g
Find: 1) Emp. Formula
2) Molec. Formula

$$71.56 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.5 \text{ g Cl}} = \frac{2.00 \text{ mol Cl}}{2.00} = 1$$

$$24{,}27 \text{ g C} \times \frac{1 \text{ mol C}}{12.0 \text{ g C}} = \frac{2.00 \text{ mol C}}{2.00} = 1$$

$$4.07 \text{ g H} \times \frac{1 \text{ mol H}}{1.0 \text{ g H}} = \frac{4.07 \text{ mol H}}{2.00} = 2$$

$ClCH_2$(empirical formula)

$$n = \frac{\text{molecular mass}}{\text{empirical mass}} = \frac{98.96 \text{ g}}{49.5 \text{ g}} = 2$$

n=2 (multiply formula by 2)

$Cl_2C_2H_4$(molecular formula)

C. The Mole in solutions

1. Solution concentrations-mass percent
2. Molarity
3. Dilutions

Solutions:

Concentrated: large amount of solute is dissolved

dilute: relatively small amount of solute is dissolved mass percent: $\frac{\textbf{mass solute}}{\textbf{mass solution}} \times \mathbf{100}$

(mass solute + mass solvent)

Example:

1.00 g CH_3CH_2OH (ethanol) with 100g H_2O ?
% mass CH_3CH_2OH ?

100 g CH_3CH_2OH

$\frac{1.00 \text{ g } CH_3CH_2OH}{101.0 \text{ g Solution}} = \mathbf{1.0\ \%}$

175 g milk
4.4 % mass Lactose
? mass solute

mass solute= (mass solm.) (% mass)
mass solute= (175 g) (0.044)
=**7.7 g Lactose**

Solution Molarity (M)

molarity= $\frac{\text{moles of solute}}{\text{L Solution}}$ $\left(M = \frac{\text{mol}}{\text{L}} \right)$

Example:

Calculate the molarity of a solution by dissolving 115 g sodium hydroxide in 1500 mL solution.

11.5 g NaOH x $\frac{1 \text{ mol NaOH}}{40.0\text{g NaOH}}$= 0.228 mol

1500 mL= 1.5 L

M= $\frac{2.88 \text{ mol}}{1.5 \text{ L}}$ = **0.192 M**

Molarity by Diluting:

$M_1V_1=M_2V_2$

M_1= Initial Molarity
V_1 = Initial volume

M_2 = Final Molarity
V_2= Final Volume

M_1= 0.3 M M_2 = 0.1 M
V_1= ? V_2= 100 mL

$$V_1 = \frac{M_2 V_2}{M_1} = \frac{(0.1\ M)(100\ mL)}{0.3\ M} = \mathbf{33.3\ mL}$$