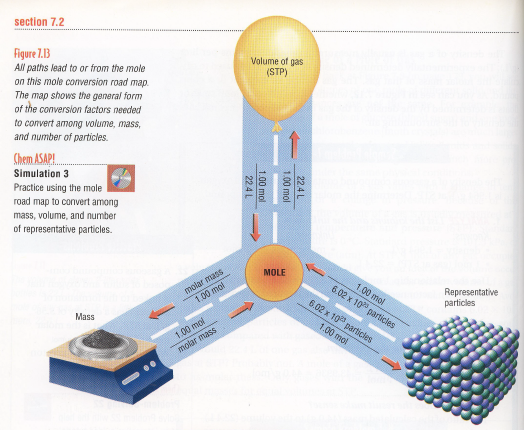
**UNIT 7 – Math of Chemistry - Test March 18, 2016**   
Kavanah pp. 45 - 56 Name \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_



# Formula Mass

## Atomic mass

### Weighted average of the masses of the existing isotopes of an element

### On periodic table, OK to round to make the math easier

### Measured in a.m.u., atomic mass units

## Gram atomic mass HW worksheet #1

### The mass of one mole of an element

### Measured in grams

### Same number as atomic mass

### Ex. What is mass of one atom of iron? What is the mass of one mole of iron atoms?

## Formula Mass

### Sum of the atomic masses of all the atoms in a compound

### Mulitply each atomic mass by the subscript

### Measured in a.m.u.

### Ex. H2O H – 2 x 1 = 2 O – 1 x 16= 16 18

### Ex. CaCO4 formula mass is 100 amu

### Ex. H2SO4 formula mass is 98 amu

## Gram formula mass (gfm) HW Kavanah p.47 # 11-13

### The formula mass of a substance expressed in grams instead of amu

### The mass of one mole of a substance

## Gram molecular mass and molar mass (MM)

### Molecule – a covalently bonded compound

### Some books (and tests) say “gram molecular mass” instead of “gram formula mass” when talking about molecules.

### Molar mass (MM) is a general term for gram atomic mass, gram formula mass, or gram molecular mass

# Percent Composition

## Part/whole HW Kavanah p. 48

### What is the percent mass of your head?

### Look for formula in Table T

### % composition by mass = mass of part/mass of whole x 100

### Ex. What is the percent by mass of carbon in CO2?

#### Step one – determine the formula mass of CO2 C – 1 x 12 = 12 amu O - 2 x 16 = 32 amu 44 amu

#### Step two - What is the mass of the “part” (in this case carbon)? C – 12 amu

#### Step three – Divide part over whole to find percent composition

#### Answer – 27%

### Ex. What is the percent by mass of nitrogen in NH4NO­3?

#### formula mass is 80 g

#### mass of N is 2x14=28 g

#### 28/80 x 100 = 35%

## Percent water in a hydrate HW Kavanah p 49 #24-26

### Review: Ionic substances often include definite amounts of water as part of the crystal structure

### Water molecules are shown as part of the formula, separated by a dot

### ex. CuSO4 · 5H2O means there are 5 water molecules per unit of copper II sulfate

### Anhydrous means that no water is incorporated into the crystal

### Treat the water molecule as a single unit with a formula mass of 18 amu when calculating percentage of water in a hydrate

### Ex. What is the percentage by mass of water in CuSO4 · 5H2O?

#### formula mass is 159.5 + 90 = 249.5 amu

#### mass of water is 5 x 18 = 90 amu

#### percent water is 90/249.5 x 100 = 36%

### Ex. What is the percent of water in sodium carbonate with ten waters

#### formula mass is 286 g

#### Mass of water is 180 g

#### percent water is 180/286 x 100 = 62.9% water

## Hydrate Lab HW Lab Report

# Mole / Mole Wheel HW Kavanah p. 50 show your work

## Mole

### Analogy: how many pencils in a dozen pencils? 12 How many eggs in a dozen eggs? 12 A dozen always means 12 of whatever,

### A mole is always 6.02 x 1023 of whatever

### Ex. How many Neon atoms in a mole of neon? How many electrons in a mole of electrons?

### Mole definition: the number of atoms of carbon-12 in a 12.000 gram sample

### Abbreviation: mol

## Particles to moles HW Mole Conversion Mini Workout

### Draw the mole wheel on the board

### Use Avogadro’s number 6.02 x 1023

### DIMO – divide in, multiply out

### Find the EE button on your calculator. Practice entering Avogadro’s number

### Ex. How many atoms in 2.6 mol Na? 1.6 E24

### Ex. How many molecules in 3.8 mol CO2? 2.3 E24

## Grams to moles

### Avogadro’s number is so awkward. Why do we use it?

### Avo is the conversion from a.m.u. to grams.

### The gfm of any substance is the mass of one mole of that substance.

### DIMO using the gfm

### Ex. How many grams are in 40.5 mol of sulfuric acid? 3970 g

### Look at example of NO on homework worksheet. How many moles are equivalent to 45 grams of NO?

## Liters to moles

### Review: gases are rapidly moving molecules.

### Volume of actual molecule is very small compared to size of the container.

### Therefore, all gases take up the same amount of volume, regardless of what gas it is.

### Molar volume is 22.4 liters of gas at STP

### Volume depends on temperature and pressure, so STP is 0 celcius and 1 atmosphere.

### Use the mole wheel.

### Ex. from worksheet: CO2 – What is volume of 2 moles of CO2 at STP?

### Ex. from worksheet H2O2 – How many moles of hydrogen peroxide is 11.2 liters at STP?

# Finding the Molecular Formula HW Kavanah p. 51 Show your work

## Find the molecular formula if the empirical formula and molecular mass are both known

### A molecule as the empirical formula CH2 and molecular mass of 42 amu. What is the molecular formula? Guess: it could be CH2, C2H4, etc.

### Step 1 - Calculate gfm of the empirical formula. What is gfm of CH2? 14

### Step 2 - Divide molecular mass of the compound by the formula mass of the empirical formula. Divide 42 amu by 14. 42/14=3

### Step 3 - Multiply each subscript by that number Molecular formula is C3H6

### Ex. A compound has a molecular mass of 180 amu and an empirical formula of CH2O. What is its molecular formula? Step 1 – Find gfm of empirical formula: C 1x12 = 12 H 2x1=2 O 1x16=16 30 amu Step 2 – Divide molecular mass by empirical mass: 180/30=6 Step 3 – Multiply all subscripts of empirical formula by 6: C6H12O6

# Stoichiometry

## Balancing review HW Kavanah p. 36 SHOW YOUR WORK

### Change coefficients in order to conserve mass in the reaction

### Atoms of reactants are same as atoms of products

### Some equations can be balanced just by looking at them.

### If harder, list all elements underneath and the number of atoms of each element.

### Hint: Keep polyatomic ions together if they are unchanged

### Use examples from worksheet

### Ex. Ca(OH)2 + (NH4)2SO4 🡪 CaSO4 + NH3 + H2O

### Ex. C3H8 + O2 🡪 H2O + CO2

## Mole ratios in balanced equations

### Balance this equation: 2 C2H6 + 7 O2 🡪 4 CO2 + 6 H2O

### Qualitative information: formulas of compounds. From the formula you can know the behavior, color, smell, etc.

### Quantitative information: coefficients. How much of each compound.

### Coefficients can represent molecules, moles, or ratios of moles.

### Ex. When 2 molecules of ethane react, how many molecules of water are formed?

### Ex. When 2 moles of ethane react, how many moles of water are produced?

### Ex. When 4 moles of ethane react, how many moles of water are formed?

### The numbers can change, but the ratio has to stay the same.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| 2 C2H6 + | 7 O2 | 🡪 | 4 CO2 + | 6 H2O |
| Moles C2H6 | Moles O2 |  | Moles CO2 | Moles H2O |
| 2 | 7 |  | 4 | 6 |
| 4 | 14 |  | 8 | 12 |
| 1 | 3.5 |  | 2 | 3 |

### Ratio of moles is always the same!

## Solving Mole-Mole Stoichiometry problems

### Solve use ratios! Easy!

### Ex. What amount of oxygen is needed to completely react with 3 moles of methane in a combustion reaction?

#### Step 1 – Always start with a balanced equation: CH4 + 2O2 🡪 CO2+ 2H2O

#### Step 2 – Write the given number of moles aver that coefficient.

#### Write x over the coefficient of the unknown quantity.

#### Solve the ratio for x. That is the number of moles you are looking for.

### Ex. Kavanah p. 54 #54

### Ex. Kavanah p. 54 #55 HW Kavanah p. 54 #54-61 SHOW YOUR WORK

## Solving volume-volume stoichiometry problems

### Volume is proportional to number of moles for any gas at STP

### So you can still use ratios!

### Ex. Haber process Balance this equation N2 + 3H2 🡪 2NH3 What volume of hydrogen is necessary to react with 5 liters of nitrogen to produce ammonia?

## Solving Harder Stoichiometry problems

### Solve using ratios. Still easy!

### Include the factor to convert to moles when needed.

#### Grams 🡪 moles. Use gfm

#### Liters of gas 🡪 moles. Use 22.4

#### molecules of compound to moles 🡪 use 6.02 x 1023

### Over / Under / Solve HW Worksheet #\_\_

### Examples from worksheet

## Extra Credit “Extension” Topics HW MgO Lab

# Find Empirical Formula from data about mass

## HW Lab Report

## Finding empirical formula

### If you analyze a compound in the lab and find out the mass of each component, you can determine the empirical formula

### Divide the mass of each element by the gfm of that element to find moles

### Find the smallest whole-number ratio by dividing each mole value by the smallest number of moles

### Ex. What is empirical formula of a compound that is 75.0% carbon by mass and 25.0% hydrogen by mass? Step 1 – Assume you have 100 grams of the compound. Step 2 – Divide each mass by the gfm. Use 3 sig figs: C – 75 g/12.0 g = 6.25 mol H – 25 g/1.01 g = 24.75 mol Step 3 – Divide each value by the smaller value: C – 6.25/6.25=1.00 H – 24.75/6.25=3.96 Step 4 – Round to whole numbers, and write the empirical formula! CH4

# Limiting Reagent HW Kavanah p. 54 #63-64

## Cheese sandwich analogy

### Recipe is 2 slices of bread, 1 slice of cheese

### If I have 4 slices of bread and 4 cheese, how many sandwiches can I make? Answer: 2 sandwiches Which ingredient was limiting? bread

### If I have 10 slices of bread and 4 slices of cheese, how many sandwiches? 4 sandwiches Which ingredient was limiting? cheese

## Limiting reagent stoichiometry

### Reagent and reactant are synonyms

### Ex. Hydorgen gas and chlorine gas react to form hydrogen chloride. If 2 mol of hydrogen gas are mixed with 4 mol of chlorine gas, how many mols of hydrogen chloride will be produced?

#### Step 1 – Start with a balanced equation.

#### Step 2– Solve the stoichiometry problem twice, once for each reactant..

#### Step 3 – Which reactant gives you less product? That is the limiting reagent.

#### Step 4 – The other reactant is in excess.

# Percent Yield