Chapter 7 – Chemical Quantities
Part 1 – Notes: Mole Calculations

Objectives: Identify, define, and explain: mole, Avogadro's number, representative particle, gram atomic mass, gram molecular mass, gram formula mass, molar mass, standard temperature and pressure, and molar volume.
Describe how Avogadro's number is related to a mole of any substance.
Calculate the mass of a mole of any substance.
Convert between various units: mass, volume at STP, molecules, atoms of a specific element in a compound, and mole, using unit analysis.

Text Reference: Section 7.1 (pages 171-181) and Section 7.2 (pages 182-186)

Start-up Problem: We go to the store to buy apples. We know that the medium-sized apples we wish to buy have a mass of 2.0-kg per dozen apples. What would be the mass of the 90 apples we wish to purchase? We need to use unit analysis!!!

Previously you learned that matter is composed of different component particles. One method of measuring a substance is to count the number of component particles in the substance but since atoms, ion, and molecules are very small, counting them is very impractical. So we introduce a unit that allows us to count large quantities of component particles. Due to the size of atoms, ions, and molecules, this unit will need to account for a LARGE number of component particles. Let's begin exploring the ***MOLE***.

The ***MOLE*** is unit used to count large numbers of representative particles in compounds. The number associated with the mole was experimentally determined.

The experimentally determined number is 6.02×10^{23}. This number is known as ***Avogadro's number.***

Officially, the number of particle present in **1 mole** of a substance is equal to the number of carbon-12 atoms in exactly 12 grams of a sample of carbon-12. Since it is a number, it can also be applied to other substances with various representative particles.

Representative particle: the smallest unit into which a substance may be broken down without a change in composition; it refers to the component particles of various types of substance: atoms, ions, or molecules.

The representative particle of an element is an atom. Remember that there are **seven** elements which exist as **diatomic elements** in their free (uncombined) state. These elements are:

Let's solve some problems using our new-found **mole knowledge**:

TYPE I: You know that the mole relationship can be used to determine the number of atoms or molecules in a given number of moles and vice versa:

MOLES ↔ ATOMS of an element

MOLES ↔ MOLECULES or FORMULA UNITS of a compound

Example 1: How many atoms of calcium are present in 5.376 moles of calcium?

TYPE II: You know that since a mole is a specified number of objects and you can take the mass of a specified number of things, then a mole of things has a certain mass. You use the PERIODIC TABLE to determine the mass of a given element. You can use the relationship between mole and mass from the periodic table to perform the following calculations:

MOLE ↔ MASS (grams) OF AN ELEMENT

Example 2: How many moles of copper are present in a sample of copper that has a mass of 213.025 g?

Note every substance you encounter will be an element. You will have to be able to determine the mass of a compound. How do you accomplish this? And what do we call this?

Gram Atomic Mass:

Gram Molecular Mass:

Gram Formula Mass:

TYPE III: Since you can determine the mass of a single element and you can correctly write formulas of compounds, then you can determine the molar mass of various compounds and use the determined molar mass in the following calculations:

MOLE ↔ MASS (grams) OF A COMPOUND

Example 3: What is the mass of a 2.7654 mol sample of copper (I) oxalate?

Example 4: How many moles are present in a sample of copper (II) acetate that has a mass of 213.564?

TYPE IV: You know you can relate various quantities to the mole. You may relate mass to moles and we may relate number of particles to the mole. But can you relate mass directly to the number of particles? NO!!! WE MUST GO THROUGH THE MOLE. It is a two-step process. MOLECULES ↔ MOLES ↔ MASS

Example 5: How many molecules of CO_2 are present in a sample with a mass of 187.564 g?

TYPE V: You must also be able to use the formula of a compound to determine information about the elements that make up the element. For example, water is made up of hydrogen and oxygen. If you have 1 mole of water molecules, you have two moles of hydrogen atoms and 1 mole of oxygen atoms. Be sure to write the correct formula of the compound to receive full credit.

Example 6: How many atoms of hydrogen are present in a 4.25 mol sample of lithium acetate?

Example 7: How many moles of carbon dioxide could be made by using 5.46×10^{24} atoms of oxygen?

You may also have mass in these problems. Remember – to get out of mass – you need to go through moles. Show the formula and molar mass to receive full credit.

Example 8: How many atoms of hydrogen are in 164.987 g aluminum acetate?

Example 9: Calculate the mass of carbon in 82.0 g of propane (C_3H_8)?

TYPE VI: Frequently you will encounter a problem with a gas. It is not very helpful to refer to the mass of a gas since a sample of a gas can expand to fill its container. It is more helpful to work with the ***VOLUME*** of a gas. Since the volume of a gas varies with temperature and pressure, it is useful to have a set of standard conditions with which to work; we will work with ***STP***.

STP – Standard temperature and Pressure – the condition where the *pressure = 1 atmosphere* and the *temperature = 273 K*

Standard Molar Volume – the volume of one mole of a gas at STP

One mole of any gas at STP conditions has a volume of 22.4 L. This information may be used in mole problems.

Example 10: Assuming conditions of STP, what is the volume (in Liters) of 145.0 g of chlorine gas?

Chapter 7 – Chemical Quantities
Part 1 – Assignment: Mole Calculations

Solve the following problems. Use unit analysis and show all work and set-ups.

1. You have one mole of each of the following substances:
   cobalt (II) chloride   gold (III) chloride   potassium dichromate   ammonium acetate

   (a) Which sample contains the greatest number of formula units?

   (b) Which sample has the greatest mass?

   (c) Which sample contains the greatest number of atoms?

2. You have 11.78 g of lithium. How many moles of lithium do you have?

3. You have 2.75 moles of calcium acetate. What is the mass of your sample?

4. How many moles are present in a sample that contains $8.25x10^{24}$ molecules of sulfur trioxide?

5. You have a 147.35-g sample of potassium sulfate. How many moles if this sample?

6. What is the mass of 0.001254 moles of tin (IV) carbonate?

7. How many moles are present in a sample of 5.743×10^{24} molecules of water?

8. How many formula units are in 0.0236 moles of barium phosphate?

9. How many formula units are in a 13.098 g sample of tungsten (V) oxide?

10. What is the mass of 5.468×10^{23} "molecules" of iron (III) oxide?

11. How many formula units are in 225.0 g of calcium permanganate?

12. How many atoms of calcium are present in a sample that has a mass of 64.25 g calcium?

13. What is the mass of a sample containing 8.756×10^{22} formula units of ammonium oxide?

14. How many atoms of hydrogen are present in 1.9083 moles of aluminum bicarbonate?

15. What mass of tin (II) hydroxide may be formed using 4.738×10^{23} atoms of oxygen?

16. How many atoms of hydrogen are found in a 14.576 g sample of ammonium acetate?

17. How many atoms of nitrogen are in a 2.39 mole sample of aluminum nitrate?

18. Calculate the mass of nitrogen in 125 g of $CO(NH_2)_2$.

19. Which contains more molecules: 1.00 mol H_2O_2, 1.00 mol C_2H_6, or 1.00 mol CO?

20. Which contains more atoms: 1.00 mol H_2O_2, 1.00 mol C_2H_6, or 1.00 mol CO?

21. Calculate the volume of 7.6 mol of argon gas at STP.

22. Calculate the volume of 135.25 g carbon dioxide at STP.

23. What is the density of helium gas at STP?

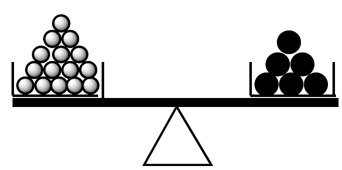
24. What is the density of fluorine gas at STP?

25. The densities of gases A, B, and C are 1.25 g/L, 2.86 g/L, and 0.714 g/L, respectively. Calculate the molar mass of each substance. Then identify each substance as ammonia (NH_3), sulfur dioxide, chlorine, nitrogen, or methane (CH_4).

26. What is the total mass of a mixture of 3.50×10^{22} formula units of sodium sulfate, 0.500 mole of water, and 7.23 g of AgCl?

27. A typical virus is 5.00×10^{-6} cm in diameter. If Avogadro's number of these virus particles were laid in a row, how many kilometers long would the line be?

28. An imaginary "atomic balance" is shown to the right. Fifteen atoms of boron on the left side are balanced by six atoms of an unknown element E on the right side.



a. What is the atomic mass of element E?

b. What is the identity of element E?

Chapter 7 – Chemical Quantities
Part 2 – Notes: Percent Composition and Molarity

Objectives: Identify, define, and explain: percent composition, molarity, and concentration.
Calculate the percentage by mass of a specific component in a compound.
Calculate the molar concentration of a solution. Calculate other components of a solution using molarity.
Text Reference: Section 7.2 (Pages 188-191)

Percent Composition by Mass

A whole compound is made of two or more elements. Each element is a part of the whole compound. You can determine the percentage of a whole compound that is a specific element part. This is the percent composition by mass of a compound.

Example 1: What is the percent composition by mass of each element in aluminum oxalate?

Molarity

Concentration refers to. . .

Molarity is a unit of concentration. The ***MOLARITY*** of a solution is the number of moles of solute in every 1 liter of solution. A 1 molar (1 *M*) solution of a pure substance may be prepared by adding enough water to one mole of substance to make a total of one liter of solution. Note: the volume is the volume of the total solution, not the volume of the solvent (water added).

So the Molarity is expressed in the formula: **Molarity = *M* = moles of solute / liters of solution**

Example 1: What is the concentration in molarity of a solution prepared by dissolving 35.45 g calcium chloride in enough water to make 1.75 liters of solution?

Example 2: What is the mass of sodium hydroxide needed to make 3.25 liters of a 1.75*M* solution?

Example 3: What is the volume of a 0.825*M* solution that may be made with 123.45 g calcium chloride?

Example 4: How many moles of tin (II) hydroxide are required to make 0.875 L of a 2.50*M* solution?

Chapter 7 – Chemical Quantities
Part 2 – Assignment: Percent Composition and Molarity

Solve the following. You must show all work, set-ups, formulas, units, etc.

1. What is the percent composition by mass of each element in ammonium carbonate?

2. An organic compound is decomposed into its elements and it produces 8.20 g C and 1.44 g H from the original sample whose mass was 32.80 g. The remainder of the sample escapes as oxygen gas during the decomposition. What is the mass of the oxygen gas? What is the percent composition, by mass, of each element in this compound?

3. Consider separate 100.0-g samples of each of the following:

   H_2O N_2O $C_3H_6O_2$ CO_2

   Rank them from highest to lowest percent oxygen by mass. Be sure to show your work.

4. What ***mass*** of NaCl is contained in 806 mL of a 2.48*M* solution?

5. Calculate the mass of sulfuric acid in 2.00 L of a 0.100*M* solution?

6. The density of nickel is 8.91 g/cm^3. How large a cube, in cm^3, would contain $2.00x10^{24}$ atoms of nickel?

7. In your own words, describe how to make a 1.00*M* solution of NaCl.

8. Calculate the percent composition by mass of hydrogen in (a) calcium acetate and (b) hydrocyanic acid.

9. Use the results from question 8, calculate the amount of hydrogen in 124 g calcium acetate.

10. Dry air is about 20.95% oxygen by volume. Assuming STP, how many oxygen molecules are in a 75.0-g sample of air? The density of air is 1.19 g/L?

Chapter 7 – Chemical Quantities
Part 3 – Notes: Empirical and Molecular Formulas

Objectives: Identify, define, explain, and differentiate between: empirical formula, molecular formula.
Determine empirical formulas from lab data or percentage by mass data.
Determine molecular formulas from an empirical formula and the molar mass.
List the common fractions that will be used in empirical formulas to determine the smallest whole number ratio.
Text Reference: Section 7.3 (Pages 192-195)

Molecular formulas indicate the actual number of atoms present in a molecule or ionic unit of a substance.

A molecule of glucose has _______ atoms of C, _______ atoms of H, and _______ atoms of O.

Empirical formulas represent the elements present in a substance in the simplest *whole number ratio* of the atoms of these elements.

The formula of glucose shows the actual number of atoms in the molecule. But the empirical formula of glucose shows the atoms in a ratio of _______ atoms of C to _______ atoms of H to _______ atoms of O.

The empirical formula and the molecular formula are related in the following way:

$(\text{empirical formula})_n$ = molecular formula

n = (molar mass of molecular formula) / (molar mass of empirical formula0

In other words, an empirical formula times some whole number is equal to the molecular formula.

DETERMINING EMPIRICAL FORMULAS

When scientists analyze an unknown substance to determine of what it is composed, they determine the empirical formula of the substance. They use the "**SMMRF**" method.

Substances: determine the substances present
Mass: determine the mass of each substance
Moles: determine the moles of each substance
Ratio: determine the ration of moles of each substance with the smallest whole numbers
Formula: use the numbers from the ratio as the subscripts in the empirical formula

Example 1: A sample of a compound with a mass of 8.66 g is decomposed and found to be 0.17 g hydrogen, 2.82 g sulfur, and 5.67 g oxygen. Determine the empirical formula.

Example 2: Analysis of 100.0 g of a compound shows that it is composed of carbon, hydrogen, and oxygen. The sample is 40.7 g carbon and 5.0 g hydrogen. Determine the empirical formula.

NOTE: In this example, the ratio is not close to a whole number. This will occur from time to time and you must round the decimal to one of the following places: 0.25, 0.33, 0.50, 0.66, or 0.75. Always round to the closest decimal. Once you have rounded, convert the decimal to a fraction.

0.25 = 1/4 0.33 = 1/3 0.50 = 1/2 0.66 = 2/3 0.75 = 3/4

It is not possible to have a fraction of an atom in a chemical formula. You must multiply all of the atoms by the denominator of the fraction to cancel the denominator and give the smallest whole number ratio.

Example 3: 100.0 g of an unknown compound is analyzed and found to be composed of carbon, hydrogen, and oxygen. 51.5 g are carbon and 8.7 g are hydrogen. Determine the empirical formula.

Some examples will give you percentages instead of masses of a compound. When this occurs, assume you have 100.0 g of the sample and use the percentages as masses in grams.

Example 4: A compound is found to consist of 46.0% carbon and 53.9% nitrogen. What is the empirical formula?

DETERMINING MOLECULAR FORMULAS

Calculating the molecular formula is a rather easy one once you have determined the empirical formula. You will be given the same type of information as in an empirical formula question but you will also be given the molar mass of the molecular formula. This allows you to determine "n".

Remember: **(empirical formula mass)$_n$ = molecular formula mass** then **(empirical formula)$_n$ = molecular formula**

When the problem states that the "molar mass is," it is referring to the molar mass of the molecular formula.

Example 5: Determine the molecular formula of the compound is *Example 5* given the fact that the molar mass is 52.04 g/mol.

Example 6: A compound has an empirical formula of CH_2O and a molar mass of 240.24 g/mol. Determine the molecular formula of the compound.

Chapter 7 – Chemical Quantities

Part 3A – Assignment: Empirical and Molecular Formulas – 1

Determine the empirical and/or molecular formula for each substance represented below. Show all steps, set-ups, and work, etc.

1. Which of the following molecular formulas are also empirical formulas?

   (a) ribose ($C_5H_{10}O_5$)

   (b) ethyl butyrate ($C_6H_{12}O_2$)

   (c) chorophyll ($C_{55}H_{72}MgN_4O_5$)

   (d) DEET ($C_{12}H_{17}ON$)

2. A compound is 9.93% carbon, 58.64% chlorine, and 31.34% fluorine. What is its empirical formula?

3. A compound is 35.00% nitrogen, 5.05% hydrogen, and 59.95% oxygen. What is its empirical formula?

4. Analysis of an organic compound shows that it is 40.7% C, 5.00% H, and 54.3% O. The molar mass of the compound is 118.10 g. What is its molecular formula?

5. An unknown compound has a composition of 51.8% C, 8.72% H, and 39.5% O. The molar mass of the compound is 324.42 g. What is its molecular formula?

6. You find that 7.36 g of a compound has decomposed to give 6.93 g of oxygen. The only other element in this compound is hydrogen. If the molar mass of the compound is 34.0 g/mol, what is the molecular formula?

Chapter 7 – Chemical Quantities
Part 3B – Assignment: Empirical and Molecular Formulas – 2

Determine the empirical and/or molecular formula for each substance represented below. Show all steps, set-ups, and work, etc.

1. A sample of an organic compound has a mass of 5.678 g. It is analyzed and found to contain 3.780 g C, 0.318 g H, 0.840 g O, and 0.736 g N. The molar mass of the compound is 108.13 g. What is the molecular formula of the compound?

2. A sample of an unknown compound has a mass of 3.58 g; it is analyzed and found to contain 3.18 g carbon and 0.40 g hydrogen. The molar mass of the compound is 135.25 g. What is the molecular formula of the compound?

3. Analysis of a gaseous compound shows that it consists of 89.92% carbon and 10.08% hydrogen. The molar mass of the compound is 120.21 g. What is the molecular formula of this substance?

4. A sample of unidentified compound has a mass of 8.366 g; it is analyzed and found to contain 2.180 g carbon, 0.366 g hydrogen, and 5.820 g sulphur. The molar mass of the substance is 230.45 g? What is the molecular formula?

5. Calculate the empirical formula for a compound consisting of 0.40 mol Cu per 0.80 mol Br.

6. Textbook – page 199 – Question 68. Complete the question and show your work here.
   a.

   b.

   c.

Chapter 7 – Chemical Quantities
Part 4 – Notes: Hydrates

Objectives: Identify, define, and explain: water of hydration, hydrate, anhydrous salt, hydration, and dehydration.
Determine the empirical formula of a hydrated salt using lab data.
Name and write formula for hydrated salts.
Calculate the molar mass of a hydrate and the percent of water in a hydrate.
Text Reference: Section 17.3 (Pages 485-488)

A **HYDRATE** is . . .

The loosely attached water molecules are called ____.

A hydrate is formed when . . .

When writing formulas for hydrates, a dot is used to separate the ionic salt from the water of hydration. A prefix is used to indicate how many water molecules are loosely attached. This number of water molecules comes after the regular name of the ionic salt.

$Na_2B_4O_7 \bullet 10\ H_2O$	Sodium tetraborate decahydrate	Borax
$CaCl_2 \bullet 2\ H_2O$	???	---
$MgSO_4 \bullet 7\ H_2O$	???	Epsom salts
$Na_2SO_4 \bullet 10\ H_2O$	???	---
???	Iron (II) sulfate heptahydrate	Green vitriol
???	Barium hydroxide octahydrate	---

Waters of hydration may be removed by ______________________________. This process is called dehydration and is due to a ____________________________ in energy, making it an ____________thermic process.

The compound without the attached water molecules is called a(n) ____.

The water molecules may be re-introduced into the anhydrous compound by adding water. This re-hydration process would cause a ___________________________ in energy, making it an ___________thermic process.

To calculate the molar mass of a hydrated salt, you must also take into account the amount of water attached to the compound.

For example:

Calcium chloride dihydrate = $CaCl_2 \bullet 2H_2O$ = 40.08 g + 2(35.45 g) + 2{ 2(1.01 g) + 16.00 g}
= 40.08 g + 70.90 g + 2(18.02 g) = 147.02 g/mol

What is the percent by mass of water in calcium chloride dihydrate?

The molar mass of iron (II) sulfate heptahydrate is ____.

The percent of water in barium hydroxide octahydrate is ____.

We will be performing a lab where the empirical formula of a hydrate will be determined. We will be heating the hydrate to drive off the water. The ***substances are ________________________________ & _________________________________. Once the water is driven off, we can determine the ***mass*** of the water and the moles of the water and salt. Then we can determine the ratio of ***moles***, from the ***ratio of moles*** we can determine the ***empirical formula***.

Let's look at some lab data:

Mass of the beaker	52.25 g
Mass of beaker + hydrate	65.50 g
Mass of hydrate	???
Mass of beaker + anhydrous salt	58.72 g
Mass of anhydrous salt	???
Mass of water (driven off)	???
Moles of water (driven off)	???
Molar mass of anhydrous salt	120.27
Moles of anhydrous salt	???

If I wanted to find the formula of the hydrate, I need to know how many moles of water are attached to 1 mole of the anhydrous salt. I need to set up the ratio - with the smallest whole numbers.

The ratio of anhydrous salt to water

The formula of the hydrate

Chapter 7 – Chemical Quantities
Part 4 – Assignment: Hydrates

Solve the following problems. Show all your work, set-ups, units, etc. Be complete.

1. One of the new superconducting materials discovered in 1987 has the following composition by mass: 41.23% barium, 28.62% copper, 13.35% yttrium, and 16.81% oxygen. What is its empirical formula?

2. Find the percent composition by mass of each element in cobalt (II) arsenate.

3. A student in Ms. Anderson's chemistry class performed an experiment and obtained the following results:

mass of hydrate + evaporating dish	37.69 g	(before heating)
mass of anhydrous compound + evaporating dish	37.21 g	(after heating)
mass of evaporating dish	34.11 g	
Identity of the anhydrous salt	barium sulfate	

(a) Calculate the empirical formula of the hydrate.

(b) What is the percentage of water in the hydrated salt?