Honors Chemistry Chapter 3 Section 3 and Chapter 7 Sections 3 and 4

Chapter 3 Section 3: Counting Atoms pgs. 82-87. These notes also cover pgs. 237-242 of Chapter 7 Section3

Objectives:

1. **Define mole, Avogadro's number, and molar mass, and state how all three are related.**
2. **Solve problems involving mass in grams, amount in moles, and number of atoms of an element.**
3. **Calculate the formula mass or molar mass of any given compound.**
4. **Use molar mass to convert between mass in grams and amount in moles of a chemical compound.**

Vocabulary: Define the following.

1. **mole--**
2. **Avogadro's number--**
3. **formula mass--**
4. **molar mass--**

Relating Mass to Numbers of Atoms

Atomic Mass and Formula Mass

The mass of an atom is so small that a unit of measure was created. This unit is called the atomic mass unit, (amu).

Relative Scale -- is a scaled based on a comparison with a standard. The relative standard in chemistry is the carbon-12 atom. This means that every element in the periodic table was compared to the mass of the carbon-12 atom, and called the atomic mass of that particular element. For example, the carbon atom is about 12 times more massive than the hydrogen atom. Therefore, since the atomic mass of carbon is 12 amu, then the mass of hydrogen is 1 amu.

Atomic mass -- is the weighted average of the isotopes of that element as listed in the periodic table. In our calculations we will round the atomic mass to the nearest hundreth, (0.01).

Since the amu is so small, it is impractical to use in laboratory experiments. It was agreed by chemists over the last 50 years to enlarge the amu to grams when needed. This was accomplished by establishing another unit called the mole, which relates the amu to the gram through the mole (to be discussed shortly).

To avoid confusion at this point, all that is needed to know is that the mass of any element in the periodic table can be discussed as either amu or grams. For example the atomic mass of oxygen can be discussed as either 16 amu or 16 grams.

Formula Mass

The formula mass is the term used when dealing with mass of compounds. It is usually expressed in amu.

This is obtained by multiplying the mass of each element in the compound by its subscript and then adding them together.
For instance the formula mass of CO_2 is obtained as follows:
C = 12.01amu x 1 = <u>12.01amu</u>   O = 16.00amu x 2 = <u>32.00amu</u>   and   12.01amu + 32.00amu = 44.01amu
The formula mass of H_2SO_4 is:
H = 1.01amu x 2 = <u>2.02amu</u>   S = 32.01amu x 1 = <u>32.01amu</u>   O = 16.00amu x 4 = <u>64.00amu</u>.
When added together the formula mass equals 2.02 + 32.01 + 64.00 = 98.03amu
Remember: The units can be either amu or grams depending on what is asked for.
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Formula mass may be called gram formula mass for ionic compounds and gram molecular mass for molecular compounds.
<u>Molar mass</u> as used in this text may mean the mass of a mole of atoms, ions, molecules, or formula units expressed as grams/mole of atoms, grams/mole of ions, grams/mole of molecules, or grams/mole of formula units.
Confused? Simply write the mass of the element from the periodic table over 1 mole of that element. Example: The molar mass of carbon atoms is 12.0g/mole of carbon atoms. The molar mass of CO_2 is 44.0g/mole of CO_2 molecules.
The only difference between formula mass and molar mass is that formula mass is usually expressed as amu and molar mass is expressed as grams per mole.
Answer the following.

1. What is the formula mass of $(NH_4)_2CO_3$?

2. What is the molar mass of $(NH_4)_2CO_3$?

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<u>The Mole</u>

The textbook definition of a mole of any element is the <u>number of atoms</u> of that element equal to the number of atoms in exactly 12.0 grams of carbon-12.
This number of atoms is the same for all elements and is called Avogadro's number. <u>Avogadro's number means that there are 6.022 x 10^{23} atoms</u> of that atom in one mole of that element.
*******Memorize this number!!!!**
The mole (mol) was the chemists' way of increasing atomic mass units in a relative scale to grams. This was accomplished by using Avogadro's number. For instance if 6.022 x 10^{23} atoms of carbon, each with a mass of 12.01 amu, were placed on a balance they would have a mass of 12.01 grams, and if 6.022 x 10^{23} atoms of oxygen, each with a mass of 16.00 amu, were placed on a balance they would have a mass of 16.00 grams.
So what is the mass of a mole of carbon atoms? From the periodic table--Answer: 12.01g/mole. What is the mass of oxygen? From the periodic table--Answer: 16.00g/mol. These values are the <u>molar masses</u> of carbon and oxygen.
Examples:
How many atoms are in 1.00 mole of carbon? Answer: 6.022 x 10^{23} atoms of carbon in 1.00 mole of carbon. What is the mass of 6.022 x 10^{23} atoms of carbon (1.00 mole)? Answer: 12.01 grams.
or
1 mole C = 6.022x 10^{23} atoms C = 12.01grams C

How many atoms in 1 mole of oxygen atoms? Answer: 6.022×10^{23} atoms of oxygen in 1 mole of oxygen atoms. What is the mass of 6.022×10^{23} atoms of oxygen? Answer: 16.00 grams.

or

1.00 mole O = 6.022×10^{23} atoms O = 16.00 grams O

*********Summary: This is how these quantities will appear when using them as conversion factors in dimensional analysis. Memorize the following 2 conversion factors. This will summarize most of the prior discussion and make it fairly easy to perform mole conversions. This is the main theme of these sections in your text.**

Memorize:

$$\frac{6.022 \times 10^{23} \text{ atoms of any element}}{1 \text{ mole of that element}} \quad \text{or} \quad \frac{\text{Mass of the element(from the PT)}}{1 \text{ mole of the element}}$$

Problems:

1. How many atoms are in 2.50 mols Fe?

Answer: $\frac{2.50 \text{ mol Fe}}{1} \times \frac{6.022 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}} = 15.06 \times 10^{23} \text{ atoms Fe}$

$= 1.51 \times 10^{24}$ atoms Fe

2. What is the mass of 1.60 mols of Fe?

Answer: $\frac{1.60 \text{ mol Fe}}{1} \times \frac{55.85 \text{ grams Fe}}{1 \text{ mol Fe}} = 89.4$ grams F

3. How many mols of calcium (Ca) are there in 50.80 grams of Ca?

Answer: $\frac{50.80 \text{ g Ca}}{1} \times \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} = 1.270$ mols Ca

4. How many grams are in 1.51×10^{24} atoms of sulfur (S)?

Answer:

$\frac{1.51 \times 10^{24} \text{ atoms S}}{1} \times \frac{1 \text{ mol S}}{6.022 \times 10^{23} \text{ atoms S}} \times \frac{32.01 \text{g S}}{1 \text{ mol S}} = 8.01 \times 10^{1}$ g S

or

$= 80.1$ g S

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USE ONLY THE 2 CONVERSION FACTORS GIVEN ABOVE WHICH CONTAIN THE MOLE IN EACH. OTHER CONVERSION FACTORS WILL CAUSE CONFUSION.

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<u>Tricks of the Trade</u>

Beware of the 7 diatomic molecules (the honorable 7): H_2, O_2, N_2, F_2, Cl_2, Br_2, and I_2.

Example:

What is the mass of a mole of oxygen atoms (O)? Answer: 16.00 grams.

What is the mass of a mole of oxygen molecules (O_2)? Answer: 32.00 grams.

Examples:

1. How many moles are in 42.60 grams of oxygen?

$$\frac{42.60\text{g O}}{1} \times \frac{1.00 \text{ mol O}}{16.00 \text{ g}} = 2.66 \text{ mol O}$$

2. How many mols of oxygen gas are there in 42.60 grams of oxygen?

Since oxygen can only exist as a gas as a diatomic molecule, the problem uses O_2.

$$\frac{42.60\text{g } O_2}{1} \times \frac{1.00 \text{ mol } O_2}{32.00\text{g}} = 1.33 \text{ mol } O_2$$

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The mole is a type of comparison is called a <u>size independent relationship</u>, which means that the same physical laws will apply to both measures regardless of the size of the sample. Meaning an amu will become a gram when using Avogadro's number.

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The meaning of the mole extends to include molecules and formula units.

For instance:

1. How many formula units are there in 2.00 moles of NaCl?

Answer:

$$\frac{2.00 \text{ mol NaCl}}{1} \times \frac{6.022 \times 10^{23} \text{ formula units}}{1.00 \text{ mol NaCl}} = 12.04 \times 10^{23} \text{ formula units NaCl}$$

$$= 1.20 \times 10^{24} \text{ formula units NaCl}$$

2. What is the mass of 3.000 moles of NaCl formula units?

Answer: First obtain the formula mass.

Na = 22.99 g/mol, Cl = 35.45 g/mol = Mass of NaCl = 53.44 g/mol

Then solve for the number of grams.

$$\frac{3.000 \text{ mol NaCl}}{1} \times \frac{53.44\text{g NaCl}}{1 \text{ mol NaCl}} = 160.3\text{g NaCl}$$

3. What is the mass of 1.500 moles of CO_2 molecules.

Answer: First obtain the formula mass of CO_2

C = 12.01 g/mol x 1 = <u>12.01</u> g/mol O = 16.00 g/mol x 2 = <u>32.00</u> g/mol

CO_2 = 12.01 + 32.00 = 44.01 g/mol

Then solve for the number of grams.

$$\frac{1.50 \text{ mol } CO_2}{1} \times \frac{44.01 \text{ g}}{\text{mol } CO_2} = 66.02 \text{ grams}$$

Hint: Though this may seem confusing, you are only using 2 conversion factors. Remember that dimensional analysis uses prompts. These prompts and using the appropriate conversion factor makes these problems far easier than you may think.

Prompts:

1. When these words appear go to the <u>periodic table</u> and get the mass.

Words: <u>mass</u>, or <u>gram</u>, or <u>kilogram</u>, or <u>molar mass</u>

2. When these words appear use <u>6.022×10^{23}</u>.

Words: <u>atoms</u>, or <u>molecules</u>, or <u>formula units</u>.

Examples.

A. How many formula units are in 68.30 grams of NaCl?

There are 2 prompts in the problem, formula units and grams.

Remember, the label with the number is the beginning and the label without the number is the end. The beginning of the solution is the most important part and should look like this.

$$\frac{68.30\text{ g}}{1} = \frac{\text{formula units}}{1}$$

Step 2

The work gram is in the beginning→ Gram is a prompt to go to the Periodic table.

Na = 22.99 g/mol Cl = 35.45 g/mol. The formula mass of NaCl therefore equals

$$\frac{58.44\text{ g NaCl.}}{1\text{ mol NaCl}}$$

Place this value in the first space remembering to cancel from top to bottom

$$\frac{68.30\text{gNaCl}}{1} \times \frac{1\text{ mol NaCl}}{58.44\text{ g NaCl}} = \frac{\text{formula units}}{1}$$

Step 3

Cancel gNaCl with g NaCl and you are correct so far.

Now you are left as a prompt with 1 mol NaCl on the left. The only other prompt remaining is formula units on the right.

Formula unit means use $\frac{6.022 \times 10^{23}}{1\text{ mol}}$**. So place it in the next space so that it cancels mol.**

Step 4

$$\frac{68.30\text{g NaCl}}{1} \times \frac{1\text{ mol NaCl}}{58.44\text{ g}} \times \frac{6.022 \times 10^{23}\text{ formula units}}{1\text{ mol NaCl}} = \text{formula units}$$

Since the only remaining label on the left and right are the same and both are on top, you must be correct.

Step 5

Do the math.

$$\frac{68.30\text{g NaCl}}{1} \times \frac{1.00\text{ mol NaCl}}{58.44\text{ g}} \times \frac{6.022 \times 10^{23}\text{ formula units}}{1\text{ mol NaCl}} = 7.038 \times 10^{23}\text{ formula units}$$

Try the following:

1. What is the molar mass of $C_6H_{12}O_6$?

2. What if the mass of 2.50 mols of $Ca(NO_3)_2$?

3. How many molecules are there in 0.340 mols if CO_2?

4. How many atoms are in 0.32 mols of iron?

5. How many mols are in 5.38 x 10^{24} formula units of magnesium bromide?

6. What is the mass of 3.62 x 10^{24} molecules of CH_3OH?

7. A chemical reaction produces 3.20 x 10^{24} molecules of CO_2. What is the mass of the CO_2 produced?

8. How many moles are in 6.60 g of $(NH_4)SO_4$?

9. How many moles are in 4.5 kg of $Ca(OH)_2$?

Chapter 7 Formulas and Chemical Compounds
Section 3: Percentage Composition pgs. 242-244.
Objectives:
1. Calculate the percentage composition of a given chemical compound.

Vocabulary: Define the following.
1. percentage composition---

Percentage composition is useful to find the percent of a particular element by mass in a chemical compound.
The percentage composition of a compound is obtained by taking the mass of each element in a compound and dividing each mass by the entire mass of the compound. The decimal is then converted into a percent.
Examples:
1. Determine the percentage composition of CO_2.
If C = 12.01 g/mol O = 16.00 g/mol x 2 = 32.00 g/mol, then CO_2 = 44.01 g/mol

% composition C = $\frac{12.01}{44.01}$ = 0.2729 = 27.29%

% composition O = $\frac{32.00}{44.01}$ = 0.7271 = 72.71%

2. If 8.200g of magnesium combines with 5.400g of oxygen to form a compound, what is the percent composition of this compound?
8.200g Mg + 5.400g O_2 → 13.60g MgO

% composition Mg = $\frac{8.20g}{13.60g}$ = .6029 = 60.29%

% composition O = $\frac{5.40g}{13.60g}$ = .3970 = 39.70%

3. Calculate the mass of carbon in 82.00g of C_3H_8 if the percentage composition of carbon is 81.80 %.

82.00g C x 0.8180 = 67.08g C

Answer the following.
1. Calculate the percentage composition of $(NH_4)_2CO_3$.

2. Calculate the percentage composition of $AgNO_3$.

3. Determine the percentage composition of $Mg(OH)_2$.

Chapter 7 Section 4 Determining Chemical Formulas pgs. 245-249
Objectives:
1. **Define empirical formula and explain how the term applies to ionic and molecular compounds.**
2. **Determine an empirical formula from either a percentage or a mass composition.**
3. **Explain the relationship between the empirical formula and the molecular formula of a given compound.**
4. **Determine a molecular formula from an empirical formula.**

Vocabulary: Define the following.
1. **empirical formula--**

2. **molecular formula--**

Newly discovered compounds are first analyzed by means of percentage composition from which an empirical formula can be determined. If the substance is covalent in nature the molecular formula can then be determined.

Calculation of Empirical Formula
A formula that gives the simplest whole number ratio of the atoms of the element is the empirical formula. There are different ways in which the empirical formula may be obtained.

Example 1: Empirical Formula from Percentage Composition
Step 1: If the problem contains a percentage composition, consider the percent of each element to be the same as grams.
Example: What is the empirical formula of a compound that is 25.9% nitrogen and 74.1 % oxygen?
The percents could be written as 25.9g and 74.1g respectively.
Step 2: Since the ratio of atoms in a chemical formula is the ratio of moles of the atoms, we can next change the grams of the nitrogen and oxygen to moles.

$$\frac{25.9\text{g N}}{1} \times \frac{1 \text{ mol N}}{14.01\text{g}} = 1.85 \text{ mol N}$$

$$\frac{74.1\text{g O}}{1} \times \frac{1 \text{ mol O}}{16.00\text{g}} = 4.63 \text{ mol O}$$

Step 3: The lowest whole number ratio is obtained by dividing each mole value by the smallest mole value.

$$N = \frac{1.85}{1.85} = \underline{1} \qquad O = \frac{4.63}{1.85} = \underline{2.5}$$

Step 4: Since this has to be a whole number ratio both numbers are multiplied by 2.

$$N = \underline{1} \times 2 = 2 \qquad O = \underline{2.5} \times 2 = 5$$

The empirical formula is therefore N_2O_5.

Example 2: Empirical Formula from Grams of a Substance

Determine the empirical formula for a compound containing 2.128g Cl and 1.203g Ca.

Solution:

Step 1: Change each gram measurement to the mole.

$$\frac{2.128\text{g Cl}}{1} \times \frac{1 \text{ mol Cl}}{35.45\text{g Cl}} = 0.0600$$

$$\frac{1.203\text{g Ca}}{1} \times \frac{1 \text{ mol Ca}}{40.08\text{g Ca}} = 0.0300$$

Step 2: Divide each mole by the smallest mole value.

$\frac{0.0600 \text{ Cl}}{0.0300} = 2 \qquad \frac{0.0300 \text{ Ca}}{0.0300} = 1$ **The empirical formula is $CaCl_2$**

Example 3: Empirical Formula from Given Masses.

Analysis of a 10.150 g sample of a compound known to contain only phosphorus and oxygen shows the phosphorus content to be 4.433 g. What is the empirical formula of the compound?

Solution:

Step 1: Find the mass of the missing element oxygen.

10.150 g total sample minus 4.433 g, the mass of phosphorus equals 5.717 g of oxygen.

Step 2: Change the grams of phosphorus and oxygen to moles.

$$4.433 \text{ g P} \times \frac{1 \text{ mol P}}{30.97 \text{ g P}} = 0.1431 \text{ mol P}$$

$$5.717 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.3573 \text{ mol O}$$

Step 3: Divide each mole by the smallest mole.

$$\frac{0.1431 \text{ mol P}}{0.1431} = 1 \text{ mol P} \qquad \text{and} \qquad \frac{0.3573 \text{ mol O}}{0.1431} = 2.497 \text{ mol O}$$

Round the 2.497 mol of O to 2.5

Step 4: Multiply both numbers to obtain a 2:5 ratio or an empirical formula of P_2O_5.

Answer the following:

1. A chemical analysis of a liquid shows that it is 60.0 % carbon, 13.4 % hydrogen, and 26.6% oxygen by mass. Calculate the empirical formula.

2. Analysis of a compound indicates that it contains 1.04 g K, 0.70 g Cr, and 0.86 g O. Find its empirical formula.

3. Determine the empirical formula of a compound containing 63.50% silver, 8.25% nitrogen, and 28.25% oxygen.

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Determining Molecular Formula

The molecular formula of a compound may be different than its empirical formula by some whole number ratio.

Example: Calculate the molecular formula of the compound whose molecular mass is 60.0g and its empirical formula is CH_4N.

Solution: Find the mass of the empirical formula.

C = 12, H_4 = 4, N = 14 = 30.0 g/mol

$$\frac{\text{molar mass } 60.0\text{g}}{\text{empirical mass } 30.0\text{g}} = 2$$

This means the molecular formula is twice that of the empirical formula.

CH_4N becomes $C_2H_8N_2$.

Answer the following.

1. The formula of a drug is $C_3H_6N_2$. The experimental molar mass is 210 g/mol. What is its molecular formula?

2. A compound has the empirical formula CH_2O. Its experimental molar mass is 90.0 g/mol. What is its molecular formula?

If a given compound can be evenly divided by some whole number than it is a molecular not an empirical formula.

Example: Which of the following is an empirical formula and which is a molecular formula?

1. **HgCl, Hg_2CL and Hg_2Cl_2.** **Answer:** ___________________

2. **C_4H_6O, $C_6H_{12}O_6$, and $C_9H_2O_4$.** **Answer:** ___________________