

## Chapter 6 – Chemical Composition and Quantities

### 1) Sec 6.1 – Counting by Weighing (Measuring Matter)

Chemistry is a quantitative science. We will have to analyze the composition of samples of matter and to learn how to be able to measure the amount of matter we have.

Think about my friend and the driving range from the Chapter 5 test. How much easier is it to weigh out a quantity of golf balls than to try to count out, say, 8000 balls?

We are returning to (or, better, continuing with) conversion factors, and will be dealing with counts, mass and volume relationships. In chemistry, we call the unifying factor we will use a “MOLE.”

### 2) Sec 6.2 – Atomic Masses

Because atoms are so small, the usual units of mass – the gram and kilogram – are entirely too large to use in normal circumstances. For example, the mass of a single carbon atom is just  $1.99 \times 10^{-23}$  g, but do we want to worry about all the  $10^{-23}$ 's, etc.? To make things easier, chemists have defined a smaller unit for mass, the **atomic mass unit** or **amu**, to use.

Think back to Chapter 3 when we talked about protons, neutrons, electrons, and atoms. At that time we mentioned a “relative” mass for the proton and set that value as equal to 1. The **amu** is related to that concept, so that the “average atomic mass” listed in the Periodic Table for each element is like adding up contributions from all the protons, neutrons and electrons in an atom. In actuality, however, it is based on the carbon-12 isotope ( $^{12}\text{C}$ ) which most chemists define as having a mass of exactly 12.000 amu and, thus, 1 amu is 1/12 of the mass of the carbon-12 isotope and is the mass of a proton.

In actuality, we can convert by using the relationship:  $1 \text{ amu} = 1.66 \times 10^{-24} \text{ g}$  but most chemists do not do that and we will learn why as this chapter develops.

All the other atomic masses in the Periodic Table are treated the same way and so, in relation to each other, are consistent. This is similar to the idea of multiplying each side of an equation by the same number; this does not change the real value of the equation. The mass of an atom, in amu, is the atomic mass as listed on the Periodic Table. Therefore we can use this number to determine the amu mass of a certain number of atoms or how many atoms are present if we know the amu's.

Examples:

- Calculate the mass, in amu, of a sample of sodium that contains 75 atoms. The mass of a sodium atom = 22.99 amu.
  
- Calculate the number of aluminum atoms present in a sample that has a mass of 1375.98 amu, if 1 Al atom = 26.98 amu

- 3) Sec 6.3 – The Mole A mole refers to the amount of substance but it is not something that has a meaning all by itself. (Read: One mole is a lot of things but it is also the ONE way we can talk about all quantities in chemistry!)
- a. Representative Particle or unit refers to whether a substance exists as atoms, ions, or molecules.
- 1) The representative particle of most elements is the atom.
  - 2) There are seven elements whose representative particle is a diatomic molecule. ( $H_2$ ,  $N_2$ ,  $O_2$ ,  $F_2$ ,  $Cl_2$ ,  $Br_2$ ,  $I_2$ )
  - 3) The representative particle of molecular compounds is the molecule.
  - 4) The representative particle of ionic compounds is the formula unit.
- b. Avogadro's Number: One mole of a substance represents  $6.02 \times 10^{23}$  representative particles of that substance.
- ==> If we have some number of representative particles, we **divide** by  $6.02 \times 10^{23}$  in order to determine the number of moles.
- ==> If we have some number of moles, we **multiply** by  $6.02 \times 10^{23}$  in order to determine the number of representative particles.

#### Examples

- 1) How many moles are there in  $1.20 \times 10^{25}$  atoms of phosphorus?
- 2) How many molecules are in 0.40 moles of  $CO_2$ ?
- 3) How many ammonium ions are in 0.036 moles of ammonium phosphate,  $(NH_4)_3PO_4$ ?

#### 4) Sec 6.4 – Molar Mass or The Mass of One Mole

- a) Molecular Mass or Formula Mass (traditionally called "molecular weight" or "formula weight") is calculated by adding the atomic masses (weights) of the atoms in the molecule or formula unit. Always retain at least three significant digits in the molecular or formula mass. More commonly we now call this the "**molar mass**" and do not worry about the specifics of the representative particle. (Remember: we are adding to get these values so that decimal places RULE!)

Examples – Calculate the molar mass of the following:

- 1) oxygen,  $O_2$
- 3) glucose,  $C_6H_{12}O_6$

2) aluminum sulfate,  $\text{Al}_2(\text{SO}_4)_3$

4) hydrated cupric sulfate,  $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$

b) The mass of one mole of an element that exists whose representative particle is the atom, is equal to the atomic mass (weight) of the element in grams. (This is called a "gram atomic mass" in your book.)

(1 mole of Mg = 24.3 g)

c) The mass of one mole of an element or compound whose representative particle is the molecule, is equal to the molecular mass (weight) of the element or compound in grams. (This is called a "gram molecular mass" in your book.)

(1 mole of  $\text{N}_2$  = 28.0 g and 1 mole of CO = 44.0 g)

d) The mass of one mole of an ionic compound whose representative particle is the formula unit, is equal to the formula mass (weight) of the compound in grams. (This is called a "gram formula mass" in your book.)

(1 mole of NaCl = 58.5 g)

==> If we have some number of grams of a substance, we **divide** by the molar mass of that substance in order to determine the number of moles.

==> If we have some number of moles of a substance, we **multiply** by the molar mass of that substance in order to determine the mass of that substance particles.

### Examples

1) How many grams are there in 5.8 moles of calcium carbonate,  $\text{CaCO}_3$ ?

2) How many moles are there in 8.8 g of carbon dioxide,  $\text{CO}_2$ ?

3) How many molecules are there in 500 g of sulfur dioxide,  $\text{SO}_2$ ?

4) What is the mass in grams of 100 molecules of chlorine,  $\text{Cl}_2$ ?

5) The Volume of 1 Mole of a Gas

The volume of 1 mole of any gas measured at  $0^\circ\text{C}$  (**standard temperature**) and 1 atmosphere pressure (**standard pressure**) is 22.4 liters. (The term Molar Volume is often used for the volume of 1 mole of a gas at STP.)

[1 mole of a gas = 22.4 L at STP]

==> If we have some volume of a gas at STP, we **divide** by 22.4 liters in order to determine the number of moles.

==> If we have some number of moles of a gas at STP, we **multiply** by 22.4 liters in order to determine the volume of the gas at STP.

Examples

a) What is the volume at STP of 0.48 moles of methane,  $\text{CH}_4$ ?

b) How many moles are in 89.6 L of Neon, Ne?

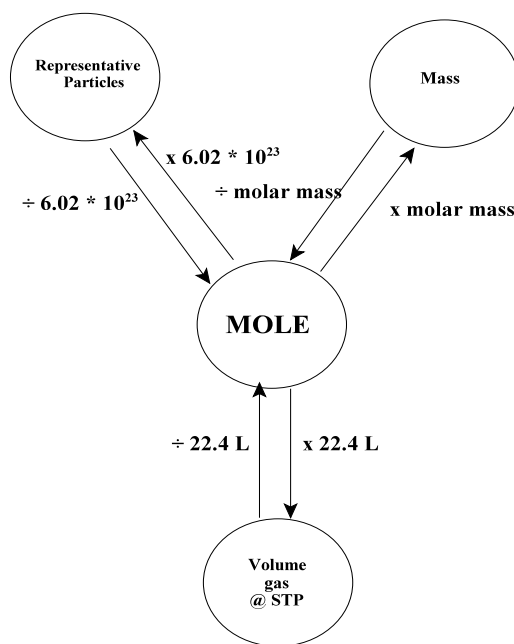
c) What is the mass of 0.448 L of oxygen gas,  $\text{O}_2$  (volume measured at STP)?

d) What is the density of  $\text{CO}_2$  gas at STP in g/L?

e) If the density of a gas is 0.714 g/L at STP, what is its molecular mass?

6) Mr Weinkauff's Summary of Mole Calculations

There are all kinds of ways to simplify thinking about calculating to and from moles. This is the simplest I have seen. At times it has helped me remember some basic concepts.



One of the beauties of this particular visual model is the idea that whenever you are at MOLES you multiply to get to the other value.

Likewise, if you have grams, particles, or volume at STP you always divide to get to MOLES.

7) Sec 6.5 – Calculating Percent Composition

a) Percent Composition is the percent by mass of each element in a compound.

b) % mass of element E =  $\frac{\text{grams of element E}}{\text{grams of compound}} \times 100$

Example - Compounds containing chlorates are known to decompose by giving off oxygen when heated. A student heated 0.800 g of a chlorate and obtained a residue that had a mass of 0.600 g. What is the percent of oxygen in the chlorate?

- c) Occasionally you may want to calculate the percentage composition of a known compound. The percentage of an element in a compound can be calculated from atomic masses if the formula of the compound is given, using the following relationship.

$$\% \text{ mass of element A} = \frac{\# \text{ of atoms of element A} \times \text{molar mass of A}}{\text{molar mass of compound}} \times 100$$

#### Examples

- 1) What is the percentage of phosphorus in calcium phosphate,  $\text{Ca}_3(\text{PO}_4)_2$ ?

- 2) How many grams of iron are there in 250 g of  $\text{Fe}_2\text{O}_3$ ?

#### 8) Sec 6.7 – Calculating Empirical Formulas

- a) The empirical formula gives the lowest whole number ratio of the elements in a compound (lowest whole number ratio of moles of atoms in a compound).

#### b) Steps to Calculating an Empirical Formula

- 1) If the percentage composition of the compound is given, use the percentages to find the mass of each element that would be present in a 100 g sample of the compound. (In some cases you may be given the mass of each element.)
- 2) Determine the number of moles of each element present in the compound by dividing the mass of each element by its atomic mass.
- 3) To determine the simplest whole number ratio, divide the number of moles of each element by the smallest number of moles that is present. If this does not give a whole-number ratio (within 0.1 of a whole number), multiply by a number like 2 or 3 to make the numbers whole numbers.

#### Examples

- 1) Calculate the empirical formula of a compound that contains 2.0 g of calcium and 8.0 g of bromine.

- 2) Calculate the empirical formula of a compound given the following percentage composition: 15.8% Al, 28.1% S, and 56.1% O.

9) Sec 6.8 – Calculating Molecular Formulas

- a) Molecular Formula of a compound that exists as a molecule (molecular compound) is the actual formula of a molecule of the compound. It is either the same or a whole-number multiple of the empirical formula.
- b) To determine the molecular formula, you must be given (or be able to calculate) the empirical formula and the molecular mass of the compound.

$$\text{M.F.} = a * (\text{E.F.}) \qquad a = \frac{\text{Molecular Mass}}{\text{Formula Mass of E.F.}}$$

Example – What is the molecular formula of a compound that is 92.3% carbon and 7.7% hydrogen. The compound's molecular mass is 78.

10) Evaluating Measurements

- a) Error = |Accepted Value - Experimental Value| (always positive)
- b) Percentage Error =  $\frac{|\text{Error}|}{\text{Accepted Value}} \times 100$

Examples: What is the error and percentage error in each of the following?

- a) A student sold \$10 worth of candy, but only collected \$6.
- b) A clerk sold \$10,000 worth of goods, but collected only \$9996.