Chemical Composition

Chapter 8

Atomic Masses

- Balanced equation tells us the relative numbers of molecules of reactants and products

$$C + O_2 \rightarrow CO_2$$

1 atom of C reacts with 1 molecule of O_2

to make 1 molecule of CO_2

- If I want to know how many O_2 molecules I will need or how many CO_2 molecules I can make, I will need to know how many C atoms are in the sample of carbon I am starting with

Atomic Masses

- Dalton used the percentages of elements in compounds and the chemical formulas to deduce the **relative** masses of atoms
- Unit is the **amu**.
  - **atomic mass unit**
  - **1 amu = 1.66×10^{-24}g**
- We define the masses of atoms in terms of atomic mass units
  - **1 Carbon atom = 12.01 amu,**
  - **1 Oxygen atom = 16.00 amu**
  - **1 O_2 molecule = 2(16.00 amu) = 32.00 amu**

Atomic Masses

- Atomic masses allow us to convert weights into numbers of atoms

If our sample of carbon weighs 3.00×10^{20} amu we will have 2.50×10^{19} atoms of carbon

C+O2----------CO2 , C +1/2O2-------CO

$$3.00 \times 10^{20} \text{ amu} \times \frac{1 \text{ C atom}}{12.01 \text{ amu}} = 2.50 \times 10^{19} \text{ C atoms}$$

Since our equation tells us that 1 C atom reacts with 1 O_2 molecule, if I have 2.50×10^{19} C atoms, I will need 2.50×10^{19} molecules of O_2

Example #1

Calculate the Mass (in amu) of 75 atoms of Al

- Determine the mass of 1 Al atom

  1 atom of Al = 26.98 amu

- Use the relationship as a conversion factor

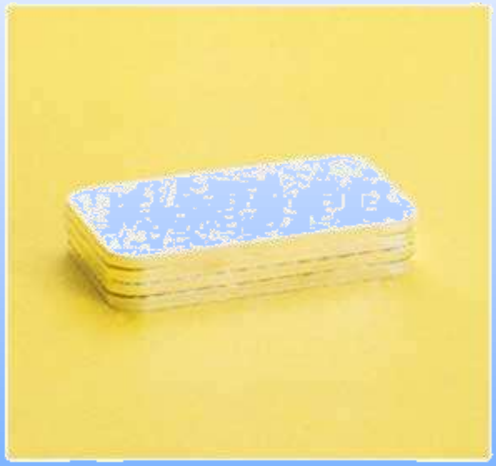
$$75\,\text{Al atoms x}\,\frac{26.98\,\text{amu}}{1\,\text{Al atom}} = 2024\,\text{amu}$$

Chemical Packages - Moles

- We use a package for atoms and molecules called a **mole= wt/ M.W**
- A mole is the number of particles equal to the number of Carbon atoms in 12 g of C-12
- One mole = 6.022×10^{23} units
- The number of particles in 1 mole is called **Avogadro's Number**
- 1 mole of C atoms weighs 12.01 g and has 6.02×10^{23} atoms
- One carbon atoms wieght= 12/ 6.02 10^{23}

Figure 8.1: All these samples of pure elements contain the same number (a mole) of atoms: 6.022×10^{23} atoms.

Lead bar
207.2 g



Silver bars
107.9 g

Pile of copper
63.55 g

Figure 8.2: One-mole samples of iron (nails), iodine crystals, liquid mercury, and powdered sulfur.

Example #2

Compute the number of moles
and number of atoms in 10.0 g of Al

- Use the Periodic Table to determine the mass of 1 mole of Al

1 mole Al = 26.98 g

- Use this as a conversion factor for grams-to-moles

$$10.0\ \mathrm{g\ Al} \times \frac{1\ \mathrm{mol\ Al}}{26.98\ \mathrm{g}} = 0.371\ \mathrm{mol\ Al}$$

Example #2

Compute the number of moles
and number of atoms in 10.0 g of Al

- Use Avogadro's Number to determine the number of atoms in 1 mole

  1 mole Al = 6.02×10^{23} atoms

- Use this as a conversion factor for moles-to-atoms

$$0.371 \text{ mol Al x } \frac{6.02 \text{ x } 10^{23} \text{ atoms}}{1 \text{ mol Al}} = 2.23 \text{ x } 10^{23} \text{ Al atoms}$$

Example #3

Compute the number of moles
and mass of 2.23×10^{23} atoms of Al

- ✶ Use Avogadro's Number to determine the number of atoms in 1 mole

1 mole Al = 6.02×10^{23} atoms

- ✷ Use this as a conversion factor for atoms-to-moles

$$2.23 \mathrm{x} 10^{23} \text{ Al atoms x } \frac{1 \text{ mol Al}}{6.02 \mathrm{x} 10^{23} \text{ atoms}} = 0.370 \text{ mol Al}$$

Example #3

Compute the number of moles
and mass of 2.23×10^{23} atoms of Al

- Use the Periodic Table to determine the mass of 1 mole of Al

1 mole Al = 26.98 g

- Use this as a conversion factor for moles-to-grams

$$0.370 \text{ mol Al x } \frac{26.98 \text{ g}}{1 \text{ mol Al}} = 9.99 \text{ g Al}$$

Molar Mass

- The **molar mass** is the mass in grams of one mole of a compound
- The relative weights of molecules can be calculated from atomic masses

$$\text{water} = H_2O = 2(1.008\text{ amu}) + 16.00\text{ amu}$$
$$= 18.02\text{ amu}$$

- 1 mole of H_2O will weigh 18.02 g, therefore the molar mass of H_2O is 18.02 g
- 1 mole of H_2O will contain 16.00 g of oxygen and 2.02 g of hydrogen

Percent Composition

- Percentage of each element in a compound
  - By mass
- Can be determined from
- the formula of the compound or
- the experimental mass analysis of the compound
- The percentages may not always total to 100% due to rounding

$$\text{Percentage} = \frac{\text{part}}{\text{whole}} \times 100\%$$

Example #4

Determine the Percent Composition from the Formula C_2H_5OH

- Determine the mass of each element in 1 mole of the compound

  2 moles C = 2(12.01 g) = 24.02 g

  6 moles H = 6(1.008 g) = 6.048 g

  1 mol O = 1(16.00 g) = 16.00 g

- Determine the molar mass of the compound by adding the masses of the elements

  1 mole C_2H_5OH = 46.07 g

Example #4

Determine the Percent Composition from the Formula C_2H_5OH

- Divide the mass of each element by the molar mass of the compound and multiply by 100%

$$\frac{24.02\text{g}}{46.07\text{g}} \times 100\% = 52.14\%\text{C}$$

$$\frac{6.048\text{g}}{46.07\text{g}} \times 100\% = 13.13\%\text{H}$$

$$\frac{16.00\text{g}}{46.07\text{g}} \times 100\% = 34.73\%\text{O}$$

Empirical Formulas

- The simplest, whole-number ratio of atoms in a molecule is called the **Empirical Formula**
  - can be determined from percent composition or combining masses
- The Molecular Formula is a multiple of the Empirical Formula

$$\%\,A \xrightarrow{100g} \text{mass A (g)} \xrightarrow{MM_A} \text{moles A}$$

$$\%\,B \xrightarrow{100g} \text{mass B (g)} \xrightarrow{MM_B} \text{moles B}$$

$$\longrightarrow \frac{\text{moles A}}{\text{moles B}}$$

Example #5

Determine the Empirical Formula of Benzopyrene, $C_{20}H_{12}$

- Find the greatest common factor (GCF) of the subscripts

factors of 20 = (10 x 2), (5 x 4)

factors of 12 = (6 x 2), (4 x 3)

GCF = 4

- Divide each subscript by the GCF to get the empirical formula

$C_{20}H_{12} = (C_5H_3)_4$

Empirical Formula = C_5H_3

Example #6

Determine the Empirical Formula of Acetic Anhydride if its Percent Composition is 47% Carbon, 47% Oxygen and 6.0% Hydrogen

- ✶ Convert the percentages to grams by assuming you have 100 g of the compound
  - – Step can be skipped if given masses

$$100\text{g} \times \frac{47\text{gC}}{100\text{g}} = 47\text{gC}$$

$$100\text{g} \times \frac{47\text{gO}}{100\text{g}} = 47\text{gO}$$

$$100\text{g} \times \frac{6.0\text{gH}}{100\text{g}} = 6.0\text{gH}$$

Example #6

Determine the Empirical Formula of Acetic Anhydride if its Percent Composition is 47% Carbon, 47% Oxygen and 6.0% Hydrogen

- Convert the grams to moles

$$47\mathrm{g\,C} \times \frac{1\,\mathrm{mol\,C}}{12.01\mathrm{g}} = 3.9\,\mathrm{mol\,C}$$

$$6.0\,\mathrm{g\,H} \times \frac{1\,\mathrm{mol\,H}}{1.008\mathrm{g}} = 6.0\,\mathrm{mol\,H}$$

$$47\,\mathrm{g\,O} \times \frac{1\,\mathrm{mol\,O}}{16.00\mathrm{g}} = 2.9\,\mathrm{mol\,O}$$

Example #6

Determine the Empirical Formula of Acetic Anhydride if its Percent Composition is 47% Carbon, 47% Oxygen and 6.0% Hydrogen

- Divide each by the smallest number of moles

$$3.9 \text{ mol C} \div 2.9 = 1.3$$

$$6.0 \text{ mol H} \div 2.9 = 2$$

$$2.9 \text{ mol O} \div 2.9 = 1$$

Example #6

Determine the Empirical Formula of
Acetic Anhydride if its Percent Composition is
47% Carbon, 47% Oxygen and 6.0% Hydrogen

- If any of the ratios is not a whole number, multiply all the ratios by a factor to make it a whole number
  - If ratio is ?.5 then multiply by 2; if ?.33 or ?.67 then multiply by 3; if ?.25 or ?.75 then multiply by 4

Multiply all the Ratios by 3 Because C is 1.3

$$3.9 \text{ mol C} \div 2.9 = 1.3 \text{ x } 3 = 4$$

$$6.0 \text{ mol H} \div 2.9 = 2 \text{ x } 3 = 6$$

$$2.9 \text{ mol O} \div 2.9 = 1 \text{ x } 3 = 3$$

Example #6

Determine the Empirical Formula of Acetic Anhydride if its Percent Composition is 47% Carbon, 47% Oxygen and 6.0% Hydrogen

- Use the ratios as the subscripts in the empirical formula

$$3.9 \text{ mol C} \div 2.9 = 1.3 \text{ x } 3 = 4$$

$$6.0 \text{ mol H} \div 2.9 = 2 \text{ x } 3 = 6$$

$$2.9 \text{ mol O} \div 2.9 = 1 \text{ x } 3 = 3$$

$$C_4H_6O_3$$

Molecular Formulas

- The molecular formula is a multiple of the empirical formula
- To determine the molecular formula you need to know the empirical formula and the molar mass of the compound

Example #7

Determine the Molecular Formula of Benzopyrene if it has a molar mass of 252 g and an empirical formula of C_5H_3

- ✶ Determine the empirical formula
  - May need to calculate it as previous

$$C_5H_3$$

- ✸ Determine the molar mass of the empirical formula

$$5\ C\text{x}12 = 60.05\ g,\ 3\ H\text{x}1 = 3.024\ g$$

$$C_5H_3 = 63.07\ g$$

Example #7

Determine the Molecular Formula of Benzopyrene if it has a molar mass of 252 g and an empirical formula of C_5H_3

- Divide the given molar mass of the compound by the molar mass of the empirical formula
  - Round to the nearest whole number

$$\frac{252g}{63.07g} = 4$$

Example #7

Determine the Molecular Formula of Benzopyrene if it has a molar mass of 252 g and an empirical formula of C_5H_3

- Multiply the empirical formula by the calculated factor to give the molecular formula

$$(C_5H_3)_4 = C_{20}H_{12}$$