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 = \( 1.66 \times 10^{-24} \) 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

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
Chemical Packages - Moles

- We use a package for atoms and molecules called a mole.
- A mole is the number of particles equal to the number of Carbon atoms in 12 g of C-12.
- One mole = $6.022 \times 10^{23}$ units.
- The number of particles in 1 mole is called Avogadro’s Number.
- 1 mole of C atoms weighs 12.01 g and has $6.02 \times 10^{23}$ atoms.

The Mole

- When two chemicals react with each other, we want to know how many atoms or molecules of each are used to form a formula of the product can be established.
- This means we need a way to count atoms, but HOW?
- The solution to this problem is to define a convenient unit of matter that contains a known number of particles.
- 1 mole = $6.022136736 \times 10^{23}$ particles
  - This value is commonly known as Avogadro’s number.
- Remember that one mole always contains the same number of particles, no matter what the substance.

Molar Mass

- The mass of one mole of atoms of any element is the molar mass, which is numerically equal to the atomic mass, but in grams.
- Example:
  - 1 mole of sodium weighs 22.99 g.
- This not only true for atoms, but also for molecules, but we use the molecular mass instead of the atomic mass.
- To determine the molecular mass of a molecule, we add up the atomic masses of the elements that make up the molecule.
  - Example:
    - $\text{CCl}_4$
      - We have 1 carbon and 4 chlorines
      - $12.001 \text{ g/mol} + 4(35.4527 \text{ g/mol}) = 153.81 \text{ g/mol}$
      - Therefore, if we have 1 mole of $\text{CCl}_4$ it would weigh 153.81 g.

Molecules, Compounds, and The Mole

- Thinking about moles with respect to molecules and compounds is not all that different than with atoms.
- If I have 1 mole of $\text{CH}_4$:
  1. How many $\text{CH}_4$ molecules do I have?
     - I have $6.022 \times 10^{23}$ molecules of $\text{CH}_4$
  2. How much does it weigh?
     - 16.032 g
  3. How many carbon atoms do I have?
     - $6.022 \times 10^{23}$
  4. How many hydrogen atoms do I have?
     - $4 \times 6.022 \times 10^{23} = 2.409 \times 10^{24}$
EXAMPLES

At first these concepts are difficult to understand, but working problems is the best way to become comfortable with these concepts.

- What is the mass, in grams, of 1.5 mols of silicon? 42.1283 g
- How many moles are represented by 454 g of sulfur? How many atoms?
  14.2 mols
  8.53 x 10^{24} atoms
- Calculate the molar mass of CaCO₃ (limestone).
  100.08 g/mol
- If you have 454 g of CaCO₃, how many moles do you have? How many molecules? How many atoms of oxygen?
  4.54 mols
  2.73 x 10^{23} molecules

At first these concepts are difficult to understand, but working problems is the best way to become comfortable with these concepts.

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

Example #3

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

- Use Avogadro’s Number to determine the number of atoms in 1 mole
  1 mole Al = 6.02 x 10^{23} atoms
- Use this as a conversion factor for moles-to-atoms

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
  water = H₂O = 2(1.008 amu) + 16.00 amu = 18.02 amu
  1 mole of H₂O will weigh 18.02 g, therefore the molar mass of H₂O is 18.02 g
  1 mole of H₂O 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

\[ \text{Percentage} \times 100\% = \text{whole part} \]

**Example**
- Eugenol is the active component of oil of cloves. It has a molar mass of 164.2 g/mol and is 73.14% C and 7.37% H; the remainder is oxygen. What are empirical and molecular formulas for eugenol?

1. What is the % of oxygen?
   - 100% - (73.14% + 7.37%) = 19.49%
2. How many grams of C, H, and O?
   - 73.14 g of C
   - 7.37 g of H
   - 19.49 g of O
3. Change the grams into moles.
   - 73.14 g of C / 12.011 g of C = 6.09 mols of C
   - 7.37 g of H / 1.0079 g of H = 7.31 mols of H
   - 19.49 g of O / 15.9994 g of O = 1.22 mols of O
4. Find the mole ratios (divide the large amount(s) of moles by the smallest amount of moles).
   - 6.09 mols of C / 1.22 mols of O = 5 mols of C to 1 mol of O
   - 7.31 mols of H / 1.22 mols of O = 6 mols of H to 1 mol of O
5. So what does this mean?
   - C₅H₆O is the empirical formula
6. How to get the molecular formula?
   - Molecular weight of the empirical formula.
   - 82 g/mol
   - Divide the molecular weight of eugenol by the molecular weight empirical formula.
   - 164.2 g/mol / 82 g/mol = 2
   - So there are 2 units of the empirical formula.
   - (C₅H₆O)₂ = C₁₀H₁₂O₂

**Empirical and Molecular Formulas**

- We have discussed molecular formulas
- Empirical formulas
  - A molecular formula showing the simplest possible ratio of atoms in a molecule.
- From percent compositions we can determine empirical and molecular formulas by the following procedure.

\[ \frac{\% A}{x \text{ mol A}} \times \frac{x \text{ mol A}}{\% B} \times \frac{y \text{ mol B}}{\% \text{ ratio}} = A_B \]

**Determining the Units of Hydration**

- Example: The following reaction takes place.
  - 1.023 g of CuSO₄ • x H₂O + heat → 0.654 g CuSO₄ + ? g H₂O

1. Find the units of hydration.
2. Step 1: determine the mass of water
   - 1.023 g of CuSO₄ • x H₂O – 0.654 g of CuSO₄ = 0.369 g of water
3. Step 2: determine the number of moles
   - 0.369 g of water / 1 mol of water = 0.0205 mol of water
   - 0.654 g of CuSO₄ / 1 mol CuSO₄ = 1.00 mol CuSO₄
4. Step 3: determine the molar ratio
   - So, we know that we started with 1.023 g of CuSO₄ • 5 H₂O
Example #4
Determine the Percent Composition from the Formula C₂H₅OH

- 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 C₂H₅OH = 46.07 g

Example #4
Determine the Percent Composition from the Formula C₂H₅OH

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

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

\[
\begin{align*}
\text{mass A (g)} & = \text{MM}_A \times \% \text{A} \\
\text{mass B (g)} & = \text{MM}_B \times \% \text{B}
\end{align*}
\]

Example #5
Determine the Empirical Formula of Benzopyrene, C₂₀H₁₂

- 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
  \(C_{5}\text{H}_3\)

Empirical Formula = \(C_{5}\text{H}_3\)

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

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
**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

<table>
<thead>
<tr>
<th>C</th>
<th>H</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>3.9 mol</td>
<td>6.0 mol</td>
<td>2.9 mol</td>
</tr>
</tbody>
</table>

= 1.3 2.9 3.9

= 1 2.9 2.9

= 2 2.9 6.0


- 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

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

C₄H₆O₃

**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 C₅H₃

- Determine the empirical formula
  - May need to calculate it as previous
  - C₅H₃
- Determine the molar mass of the empirical formula

\[ 5 \text{ C} = 60.05 \text{ g}, \ 3 \text{ H} = 3.024 \text{ g} \]

\[ \text{C}_5\text{H}_3 = 63.07 \text{ g} \]
Example #7
Determine the Molecular Formula of Benzopyrene if it has a molar mass of 252 g and an empirical formula of C₅H₃

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

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

Homework

• Questions and Problems:
  • 6, 12, 14, 16, 18, 20, 22, 24, 26, 28, 30, 32, 34, 36, 38, 40, 42, 44, 46, 48, 50, 54, 56, 58, 60, 62, 64, 66, 68, 70, 72, 74, 76, 78, 80, 82, 84