The Mole

Objectives

- State the value of Avogadro’s number (6.02×10^{23}).
- Use Avogadro’s number as a unit factor to determine the number of particles, given the number of moles.
- Use Avogadro’s number as a unit factor to determine the number of moles, given the number of particles.
- Calculate the molar mass of a substance given its chemical formula.
- Use the molar mass as a unit factor to calculate the number of moles of a substance, given the mass of the substance.
- Use the molar mass as a unit factor to calculate the mass of a sample, given the number of moles of the substance.

Objectives

- State the value for the molar volume of a gas at STP (22.4 L/mol).
- Use the molar volume as a unit factor to calculate the number of moles of a gas, given the volume of the gas.
- Use the molar mass as a unit factor to calculate the volume of a gas, given the number of moles of the gas.
- Calculate the percent composition of a compound given its chemical formula.
- Determine the empirical formula of a compound, given its composition by mass.
- Determine the empirical formula of a compound, given its percent composition.

Objectives

- Determine the molecular formula for a compound, given its empirical formula and molar mass.

Measuring Things in Chunks

- Paper: 1 ream = 500 sheets
- Eggs: 1dozen = 12 eggs
- Atoms: 1 mole = 6.02×10^{23} atoms
- The quantity 6.02×10^{23} is known as Avogadro's number.

Mole – Particle Calculations

- Avogadro's number can be use as a unit factor to convert from the number of particles to the corresponding number of moles.
- Example:

.25 mol O_2 X 6.02 X 10^{23} molecules of O_2/1 mol O_2
= 1.51 X 10^{23} molecules O_2

Mole – Mass Calculations

- **Molar Mass** – The mass in grams of an element that corresponds to 1 mole of the substance.
- The molar mass of an element is equal to its atomic mass in grams.
- Molar mass of a compound is calculated by adding together the molar masses of each of the elements that form it.
- Example:
  - Molar mass of Fe_2O_3:
    $2(55.85\ g) + 3\ (16.00\ g) = 159.70\ g$

Mole – Mass Calculations

- Examples:

Molar mass of NH_3

$14.01\ g + 3(1.01\ g) = 17.04\ g$

Molar mass of $Mg(NO_3)_2$

$24.31\ g + 2(14.01\ g) + 6(16.00\ g) = 148.33\ g$

Mass of 2.01×10^{22} atoms of Hg

2.01×10^{22} atoms Hg X 1 mol Hg/6.02×10^{23} atoms Hg X 200.59 g Hg/ 1 mol Hg = 6.70g Hg

Number of atoms in .470 g O_2

.470 g O_2 X 1 mol O_2/32.00 g O_2 X

6.02×10^{23} molecules O_2 /1 mol O_2 = 8.84×10^{21} molecules O_2

Mole – Volume Calculations

- **Molar Volume** – The volume occupied by 1 mole of a gas at **STP** (Standard Temperature and Pressure), 273 K, 1 atmosphere pressure.
- The molar volume of a gas at STP is 22.4 L.
- Density:

  Density = Molar Mass (g)/Molar Volume (L)

Mole – Volume Calculations

- Examples:
  - Density of ammonia:
    - D = M/V = 17.04 g/22.4 L = .761 g/L
  - Mass of 3.36 L of O_3 at STP:
    - 3.36 L O_3 X 1 mol O_3/22.4 L O_3 X 48 g O_3/1 mol O_2 = 7.2 g O_3
  - Number of molecules in 50.0 mL H_2:
    - 50.0 mL X 1 L/1000 mL X 1 mol H_2/22.4 L H_2 X 6.02 X 10^{23} molecules H_2/1 mol H_2 = 1.34 X 10^{21} molecules H_2
  - Molar mass of a gas with a density of 1.96 g/L:
    - D = M/V
    - M = DV
    - M = (1.96 g/L)(22.4(L/mol) = 43.9 g/mol

Percent Composition

- The **percent composition** is a list of the mass percent of each element in a compound.
- To find the percent composition, assume that the sample contains one mole of the compound.
- The mass percent of each element is found by dividing the molar mass of the element present in the compound by the molar mass of the entire compound and multiplying by 100%.

Percent Composition

- Examples:
  - Water H_2O
    - 1 mole of water contains 2 moles of hydrogen and 1 mole of oxygen.
    - Molar mass of water: 2(1.01 g) + 1(16.00 g) = 18.02 g
    - Hydrogen: [2(1.01 g)/18.02 g] X 100% = 11.2%
    - Oxygen: (16.00 g/18.02 g) X 100% = 88.8%
  - TNT $C_7H_5(NO_2)_3$
    - C: 37.01%
    - H: 2.22%
    - N: 18.50%
    - O: 42.26%

Determining Empirical Formulas

- From masses of reactants:
  1. Find the number of moles of each element.
  2. Find the smallest whole number ratio of the number of moles.

  Example: 0.500 g Sc, 0.767 g Sc_xO_y
  - 0.500 g Sc X 1 mol Sc/44.96 g Sc = 0.0111mol Sc
  - 0.267 g O X 1 mol O/16 g O = 0.0167 mol O
  - Smallest whole number ratio: 2:3
  - Empirical formula: Sc_2O_3

Determining Empirical Formulas

- From percent composition:
  1. Assume the sample has a mass of 100 g.
  2. Find the number of moles of each element.
  3. Find the smallest whole number ratio of the number of moles.

Example: 92.2% carbon, 7.83% hydrogen

92.2 g C X 1 mol C/12.01 g C = 7.68 mol C

7.83 g H X 1 mol H/1.01 g H = 7.75 mol H

Smallest whole number ratio: 1:1

Empirical formula: CH

Determining Molecular Formulas

- From the molar mass and empirical formula.
  1. Find the molar mass experimentally.
  2. Divide the molar mass of the empirical formula into the experimentally determined molar mass.
  3. Multiply the number of atoms of each element in the empirical formula by the ratio found in step 2.

Example: Fructose Empirical formula CH_2O, Molar mass 180 g/mol

Empirical molar mass: 12.00 g + 2(1.01 g) + 16.00 g = 30 g/mol

Ratio: 180 g/mol/30 g/mol = 6

Molecular formula: $C_6H_{12}O_6$