Notes - Chemical Quantities

All Roads Lead to the Mole

Count Amedeo Avogadro

- 1776-1856
- Lawyer who became interested in math and physics
- Discovered that equal volumes of different gases contained an equal
- number of particles
- 9 years after his death, Joseph Loschmidt determined a constant and named it after Avogadro.

AVOGADRO'S CONSTANT = 6.02×10^{23}

1 mole = 6.02×10^{23} particles

1 mole = mole mass (grams)

1 mole = 1 mole (of a gas @ STP) = 22.4 L

MOLE : CHEMISTS as DOZEN : BAKERS

Carry your units and your units will carry you!

VOLUME ⟷ MOLE ⟷ MASS

↕

R.P.

Formula units Ions

Molecules Atoms

R.P.: Representative Particles

Types of representative particles:

- Molecules (breaks down into atoms)
- atoms
- formula units
- ions

Naming Representative Particles:

- Pure Substance
  - Element
    - Monotomic
      - Charged ion
      - Neutral Atom
    - Polyatomic
      - Molecule
  - Compound
    - Molecular
      - Molecule
    - Ionic
      - Formula Unit

Examples: Name the representative particle for each substance given. For each molecule, state how many atoms are present. For each formula unit, state how many ions make up the ionic compound.

H^+: ion

Cl_2: molecule (2 atoms)

C_2H_6: molecule (8 atoms)

$Cu(NO_3)_2$: formula unit (3 ions)

Al: atom

NaCl: formula unit (2 ions)

R.P. Example 1:
How many moles are in 1.4×10^{22} molecules of H_2O?

$$1.4 \times 10^{22} \text{ molecules } H_2O \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ molecules}} = 0.023 \text{ mol } H_2O$$

R.P. Example 2:
How many representative particles are in 2.6 mol CO_2?

$$2.6 \text{ mol } CO_2 \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 1.6 \times 10^{24} \text{ molecules } CO_2$$

R.P. Example 3:
How many atoms are in 5.2 mol CO_2?

$$5.2 \text{ mol } CO_2 \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \times \frac{3 \text{ atoms}}{1 \text{ molecule}} = 9.4 \times 10^{24} \text{ atoms}$$

MOLAR MASS ... a.k.a. Molecular Weight (MW)

- Molar mass = mass of 1 mole of a substance
- Molar mass can be determined by adding up the atomic masses from the periodic table.

MW Example 1: Find the MW of CH_4.

1 C + 4 H
12.0 + 4(1.0)
$\boxed{16.0 \text{ g/mol}}$

MW Example 2: Find the MW of $Mg(OH)_2$

1 Mg + 2 O + 2 H
24.3 + 2(16.0) + 2(1.0)
58.3 g/mol

MW Example 3: Find the MW of $MgSO_4 \cdot 7H_2O$.

1 Mg + 1 S + 4 O + 7(H_2O)
24.3 + 32.1 + 4(16.0) + 7(18.0)
246.4 g/mol

Mass Example 1: How many grams are in 7.20 moles of dinitrogen trioxide?

$$7.20 \text{ mol } N_2O_3 \times \frac{76.0 \text{ g}}{1.00 \text{ mol}} = \boxed{547 \text{ g } N_2O_3}$$

Mass Example 2: Find the number of moles in 92.2 g of iron(III) oxide, Fe_2O_3.

$$92.2 \text{ g } Fe_2O_3 \times \frac{1.00 \text{ mol}}{159.6 \text{ g}} = \boxed{0.578 \text{ mol } Fe_2O_3}$$

<u>Volume Example 1:</u> Determine the volume, in liters, of 0.600 mol of SO_2 gas at STP.

$$0.600 \text{ mol } SO_2(g) \times \frac{22.4 \text{ L}}{1 \text{ mol}} = \boxed{13.4 \text{ L } SO_2}$$

<u>Volume Example 2:</u> Determine the number of moles in 33.6 L of He gas at STP.

$$33.6 \text{ L He}(g) \times \frac{1 \text{ mol}}{22.4 \text{ L}} = \boxed{1.50 \text{ mol He}}$$

<u>Density:</u>

Density = Mass/Volume

When given the density of an unknown gas, one can multiply by the molar volume to find the MW. The MW can allow for identification of the gas from a list of possibilities.

<u>Density Example:</u> The density of an unknown gas is 2.054 g/L.
(a) What is the molar mass?

$$\frac{2.054 \text{ g}}{\text{L}} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = \boxed{46.01 \text{ g/mol}}$$

(b) Identify the gas as either nitrogen, fluorine, nitrogen dioxide, carbon dioxide, or ammonia.

Nitrogen Dioxide

Always convert to units of moles first when converting between grams, liters, and representative particles.

1 mole = 6.02 x 1023 RP's = MW = 22.4 L of gas @STP

<u>Mixed Mole Example 1:</u> How many carbon atoms are in a 50.0-carat diamond that is pure carbon? Fifty carats is the same as 10.0 g.

$$\frac{10.0 \cancel{\text{g}} \text{ C}}{} \times \frac{1 \text{ mol}}{12.0 \cancel{\text{g}}} \times \frac{6.02 \times 10^{23} \text{ atom}}{1 \text{ mol}} = \boxed{5.02 \times 10^{23} \text{ atoms C}}$$

<u>Mixed Mole Example 2:</u> How many atoms are in 22.0 g of water?

$$22.0 \cancel{\text{g}} \text{ } H_2O \times \frac{1 \text{ mol}}{18.0 \cancel{\text{g}}} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \times \frac{3 \text{ atoms}}{1 \text{ molecule}} = \boxed{2.2 \times 10^{24} \text{ atoms } H_2O}$$

Notes: Percent Composition

The chemical composition of a compound can be expressed as the mass percent of each element in the compound.

Example: Determine the percent composition of C_3H_8.

3 C = 3(12.0) = 36.0 g % C = $\frac{36.0\,g}{44.0\,g}$ x 100 = 81.8%

8 H = 8(1.0) = 8.0 g % H = $\frac{8.0\,g}{44.0\,g}$ = 0.18 = 18%

Example: Determine the percent composition of iron (III) sulfate.

$Fe_2(SO_4)_3$

2 Fe = 2(55.85) = 111.7 g % Fe = $\frac{111.7\,g}{399.9\,g}$ x100 = 27.9%

3 S = 3(32.07) = 96.21 g % S = $\frac{96.21\,g}{399.9\,g}$ = 24.1%

12 O = 12(16.00) = 192.0 g % O = $\frac{192.0\,g}{399.9\,g}$ x 100 = 48.0%

Notes: Hydrated Compounds

Some compounds exist in a "hydrated" state; some specific # of water molecules are present for each molecule of the compound.

Example: oxalic acid $(COOH)_2$ can be obtained in the laboratory as $(COOH)_2{\cdot}2H_2O$.
(The dot shows that the crystals of oxalic acid contain 2 water molecules per $(COOH)_2$ molecule.)

The molar mass of $(COOH)_2$ = 90.0 g/mol

% anhydrous molecule = $\frac{90.0\,g}{126.8}$ x 100 = 71.4%

The molar mass of $(COOH)_2{\cdot}2H_2O$ = 126.0 g/mol

%

% mass of anhydrous salt = $\frac{90.0\,g}{126.0\,g}$ x 100 = 71.4%

% mass of water = $\frac{36.0}{126.0}$ x 100 = 28.6%

Water can be driven out of a hydrated compound by heating it to leave an "anhydrous" (without water) compound.

Example: A 7.0 g sample of calcium nitrate, $Ca(NO_3)_2{\cdot}4H_2O$, is heated to constant mass. How much anhydrous salt remains?

% mass $Ca(NO_3)_2$ = 40.08 + 2(14.01) + 6(16.00) = 164.1 g/mol $\frac{164.1\,g}{236.1\,g}$ x 100 = 69.5%

$4H_2O$ = 8(1.01) + 4(16.00) = 72.0

Hydration Number:

Some molecules attach themselves to water molecules. This is done in set numbers, depending on the molecule. For example, Magnesium sulfate attaches to 7 water molecules. We say it's hydration number is 7.

$MgSO_4{\cdot}7H_2O$ name: magnesium sulfate heptahydrate

Anhydrides:

A compound that is normally a hydrate and has lost its hydration water is said to be anhydrous and is called an anhydride.

$BaCl_2{\cdot}2H_2O$ name: Barium Chloride Dihydrate

$BaCl_2$ name: Barium Chloride Anhydrate

or Anhydrous Barium Chloride

Finding the Hydration Number:

The hydration number can be conveniently found by heating the compound and measuring its mass loss. This mass loss is usually due to the hydration water molecules being driven off.

For example...

A 15.35 g sample of Strontium nitrate $Sr(NO_3)_2{\bullet}nH_2O$ is heated to a constant mass of 11.45 g. Calculate the hydration number.

Sample Data:

Mass Hydrate	15.35g
Mass Anhydride	11.45g
Mass of Water (mass loss)	3.90g

Calculations:

Moles Anhydride

$$11.45\ g\ Sr(NO_3)_2 \times \frac{1\ mol\ Sr(NO_3)_2}{211.6\ g\ Sr(NO_3)_2} = 0.0541\ mol\ Sr(NO_3)_2$$

Moles Water

$$3.90\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.0\ g\ H_2O} = 0.217\ mol\ H_2O$$

Divide both by smallest...

$$\frac{mol\ H_2O}{mol\ Sr(NO_3)_2} = \frac{0.217\ H_2O}{0.0541\ mol\ Sr(NO_3)_2} = 4.01 \approx 4$$

Hydration Number is 4

$Sr(NO_3)_2{\cdot}4H_2O$ name: Strontium Nitrate Tetrahydrate

Notes - Empirical and Molecular Formulas

Empirical Formulas: The empirical formula is the simplest whole number ratio of the atoms of each element in a compound. *Note: it is not necessarily the true formula of the compound.* For example, the molecular formula for glucose is $C_6H_{12}O_6$, but its empirical formula is CH_2O. Empirical formulas are like fractions reduced to lowest terms.

Molecular Formulas: The *molecular formula* gives the actual numbers of each element, and thereby represents the true formula of the compound.

Calculating the Empirical Formula:

Example 1: A compound is found to contain 2.199 g of copper and 0.277 g of oxygen. Calculate its empirical formula.

Step 1: convert masses to moles

Step 2: Divide all the moles by the smallest #. This gives the "mole ratio"

Step 3: Round off these numbers, they become the subscripts for the elements.

Copper:
$$2.199\,\cancel{g}\,Cu \times \frac{1\,mol\,Cu}{63.5\,\cancel{g}\,Cu} = 0.03460\,mol\,Cu \qquad \frac{0.03460\,mol\,Cu}{0.01731\,mol} = 1.999\,Cu \approx 2$$

Oxygen:
$$0.277\,\cancel{g}\,O \times \frac{1\,mol\,O}{16.00\,\cancel{g}\,O} = 0.01731\,mol\,O \qquad \frac{0.0173\,mol\,O}{0.0173} = 1\,O$$

Empirical formula: Cu_2O

Example 2: A material is found to be composed of 38.7% Carbon, 51.6% Oxygen, and 9.7% Hydrogen. By other means, it is known that the molecular weight is 62.0 g/mol. Calculate the empirical and molecular formula for the compound. **Note: If you assume a sample weight of 100grams, then the percents are really grams.**

Carbon: $38.7\,\cancel{g} \times \frac{1\,mol}{12.0\,\cancel{g}} = \frac{3.23\,mol}{3.23} = 1.00$

Oxygen: $51.6\,\cancel{g} \times \frac{1\,mol}{16.0\,\cancel{g}} = \frac{3.23\,mol}{3.23} = 1.00$

Hydrogen: $9.7\,\cancel{g} \times \frac{1\,mol}{1.0\,\cancel{g}} = \frac{9.7\,mol}{3.23} = 3.00$

Empirical formula: CH_3O

The molecular weight of the empirical formula is....

C 12x1 = 12
H 1x3 = 3
O 16x1 = 16
31g/mol

Remember, the empirical formula is not necessarily the molecular formula!
MW of the empirical formula = 31

MW of the molecular formula = 62

$$\frac{\text{Molecular}}{\text{Empirical}} = \frac{62}{31} = 2$$

Remember, the molecular formula represents the actual formula.

$$2 \times (CH_3O) = C_2H_6O_2$$

What if the mole ratios don't come out even?

Example 3: A compound is analyzed and found to contain 2.42g aluminum and 2.15g oxygen. Calculate its empirical formula.

Aluminum $2.42\cancel{g} \times \frac{1\ mol}{27\cancel{g}} = \frac{0.0896\ mol}{0.0896} = 1.0 \times 2 = 2$

Oxygen $2.15\ g \times \frac{1\ mol}{16\ g} = \frac{0.134\ mol}{0.0896} = 1.495 \times 2 = 3$

Empirical formula: Al_2O_3