Molar Mass

- **Molar mass =**
  Mass in grams of one mole of any element, numerically equal to its atomic weight
- Molar mass of molecules can be determined from the chemical formula and molar masses of elements
  - Each H_2O molecule contains 2 H atoms and 1 O atom
  - Each **mole** of H_2O molecules contains 2 moles of H and 1 mole of O
  - One **mole** of O atoms corresponds to 15.9994 g
  - Two **moles** of H atoms corresponds to 2 x 1.0079 g
  - Sum = **molar mass** = 18.0152 g H_2O per mole

Calculation of Molar Masses

- Calculate the molar mass of the following
  - Magnesium nitrate, $Mg(NO_3)_2$
    1 Mg = 24.3050
    2 N = 2x 14.0067 = 28.0134
    6 O = 6 x 15.9994 = 95.9964
    Molar mass of $Mg(NO_3)_2$ = 148.3148 g
  - Calcium carbonate, $CaCO_3$
    1 Ca = 40.078
    1 C = 12.011
    3 O = 3 x 15.9994
    Molar mass of $CaCO_3$ = 100.087 g
  - Iron(II) sulfate, $FeSO_4$
    Molar mass of $FeSO_4$ = 151.909 g

Solutions

- Solution: a homogenous mixture in which the components are evenly distributed in each other
- Solute: the component of a solution that is dissolved in another substance
- Solvent: the medium in which a solute dissolved to form a solution
- Aqueous: any solution in which water is the solvent

Solutions

- The properties and behavior of solutions often depend not only on the type of solute but also on the concentration of the solute.
- Concentration: the amount of solute dissolved in a given quantity of solvent or solution
  - many different concentration units
    - (%, ppm, g/L, etc)
  - often expressed as **Molarity**

Solution Concentrations

- **Molarity** = moles of solute per liter of solution
  - Designated by a capital **M** (mol/L)
- 6.0 M HCl
  - 6.0 moles of HCl per liter of solution.
- 9.0 M HCl
  - 9.0 moles of HCl per liter of solution.

Solution Concentrations

- Molarity can be used as a conversion factor.
- The definition of molarity contains 3 quantities:

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution in liters}}$$

- If you know two of these quantities, you can find the third.

Solution Concentrations

- Determine the molarity of each solution
  - 2.50 L of solution containing 1.25 mol of solute

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution in liters}}$$

  - 225 mL of solution containing 0.486 mole of solute
  - 100. mL of solution containing 2.60 g of NaCl

Strategy:

g → mol → molarity

Solution Preparation

Example: How many moles of HCl are present in 2.5 L of 0.10 M HCl?

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution in liters}}$$

Given: 2.5 L of soln
0.10M HCl
Find: mol HCl

Use molarity as a conversion factor

$$\textbf{Mol HCl} = 2.5\ \cancel{\text{L}}\ \text{soln} \times \frac{0.10\ \text{mol HCl}}{1\ \cancel{\text{L of}}\ \text{soln}}$$

$$= \textbf{0.25 mol HCl}$$

Solution Preparation

Example: What volume of a 0.10 M NaOH solution is needed to provide 0.50 mol of NaOH?

Given: 0.50 mol NaOH
0.10 M NaOH
Find: vol soln

Use M as a conversion factor

$$\text{Vol soln} = 0.50\ \cancel{\text{mol NaOH}} \times \frac{1\ \text{L soln}}{0.10\ \cancel{\text{mol NaOH}}}$$

$$= 5.0\ \text{L solution}$$

Solution Preparation

- Solutions of exact concentrations are prepared by dissolving the proper amount of solute in the correct amount of solvent – to give the desired final volume
  - Determine the proper amount of solute
  - How is the final volume measured accurately?

Solution Preparation

Example: How many grams of $CuSO_4$ are needed to prepare 250.0 mL of 1.00 M $CuSO_4$?

Given: 250.0 mL solution
1.00 M $CuSO_4$
Find: g $CuSO_4$

Conversion factors: Molarity, molar mass

Strategy:
mL → L → mol → grams

Solution Preparation

Given: 250.0 mL solution, 1.00 M $CuSO_4$
Find: g $CuSO_4$

Strategy:
mL → L → mol → grams

$$\text{g CuSO4} = 250.0\ \cancel{\text{mL}}\ \text{soln} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{1.00\ \cancel{\text{mol}}}{1\ \cancel{\text{L}}\ \text{soln}}$$

$$\times \frac{159.6\ \text{g CuSO4}}{1\ \cancel{\text{mol}}}$$

$$= 39.9\ \text{g CuSO4}$$

Solution Preparation

- Describe how to prepare the following:
  - 500. mL of 1.00 M $FeSO_4$

Strategy:

mL → L → mol → grams

  - 100. mL of 3.00 M glucose
  - 250. mL of 0.100 M NaCl

Solution Preparation

- Steps involved in preparing solutions from pure solids

(a) (b) (c) (d)

Solution Preparation

- Steps involved in preparing solutions from pure solids
  - Calculate the amount of solid required
  - Weigh out the solid
  - Place in an appropriate volumetric flask
  - Fill flask about half full with water and mix.
  - Fill to the mark with water and invert to mix.

Solution Preparation

- Solutions of exact concentrations can also be prepared by diluting a more concentrated solution of the solute to the desired ***concentration***

Solution Preparation

- Many laboratory chemicals such as acids are purchased as concentrated solutions (stock solutions).
  - 12 M HCl
  - 12 M H_2SO_4
- More dilute solutions are prepared by taking a certain quantity of the stock solution and diluting it with water.

Solution Preparation

- A given volume of a stock solution contains a specific number of moles of solute.
  - 25 mL of 6.0 M HCl contains 0.15 mol HCl

    (How do you know this???)
- If 25 mL of 6.0 M HCl is diluted with 25 mL of water, the number of moles of HCl present does not change.
  - Still contains 0.15 mol HCl

Solution Preparation

Moles solute before dilution	=	moles solute after dilution

- Although the number of moles of solute does not change, the volume of solution does change.
- The concentration of the solution will change since Molarity = $\dfrac{\text{mol solute}}{\text{volume solution}}$

Solution Preparation

- When a solution is diluted, the concentration of the new solution can be found using:

$$M_c \bullet V_c = \textbf{moles} = M_d \bullet V_d$$

Where, M_c *= concentration of concentrated solution* (mol/L)

Vc = volume of concentrated solution

M_d *= concentration of diluted solution* (mol/L)

V_d *= volume of diluted solution*

Solution Preparation

Example: What is the concentration of a solution prepared by diluting 25.0 mL of 6.00 M HCl to a total volume of 50.0 mL?

Given: $V_c = 25.0$ mL
$M_c = 6.00$ M
$V_d = 50.0$ mL

$M_c \times V_c = M_d \times V_d$

Find: M_d

Solution Preparation

$$M_c \bullet V_c = M_d \bullet V_d$$

$$6.00 \text{ M} \times 25.0 \text{ mL} = M_d \times 50.0 \text{ mL}$$

$$M_d = \frac{6.00 \text{ M} \times 25.0 \cancel{\text{mL}}}{50.0 \cancel{\text{mL}}} = 3.00 \text{ M}$$

Note: V_c and V_d do not have to be in liters, but they must be in the same units.

Solution Preparation

- Describe how to prepare 500. mL of 0.250 M NaOH solution using a 6.00 M NaOH solution.
  Given: $M_c = 6.00$ M
  $M_d = 0.250$ M
  $V_d = 500.0$ mL

  Find: V_c

  $$M_c \bullet V_c = M_d \bullet V_d$$

- What volume of 2.30 M NaCl should be diluted to give 250. mL of a 0.90 M solution?

Concentrations of Acids

- The pH scale

$pH = -\log [H^+]$

pH of vinegar = -log (1.6 x 10^{-3} M) = - (-2.80) = 2.80

pH of pure water = -log (1.0 x 10^{-7} M) = - (-7.00) = 7.00

pH of blood = -log (4.0 x 10^{-8} M) = - (-7.40) = 7.40

pH of ammonia = -log (1.0 x 10^{-11}M) = - (-11.00) = 11.00

- **Lower the concentration of H^+, higher the pH**

acidic substance, pH< 7
Basic substance, pH >7
neutral, pH = 7

Concentrations of Acids

- On serial dilution of acid solution, the pH increases because the concentration of H^+ ions decreases with dilution

Concentration	pH
0.1M HCl	1
0.01 M HCl	2
0.001 M HCl	3

- The H^+ concentration of a solution of known pH can be calculated using the following equation:

$$[H^+] = 10^{-pH}$$

Concentrations of Acids

- Calculate pH of 0.0065 M HCl solution.
- Calculate the concentration of H^+ ion in a solution of pH 7.5.

Solution Preparation Review

- If you dissolve 9.68 g of potassium chloride in 1.50 L, what is the final molar concentration?
- How many grams of sodium sulfate are contained in (dissolved in) 45.0 mL of 3.00 M solution?

Solution Preparation Review

- What volume of 8.00 M sulfuric acid should be diluted to produce 0.500 L of 0.250 M solution?
- What's the pH of the final solution?

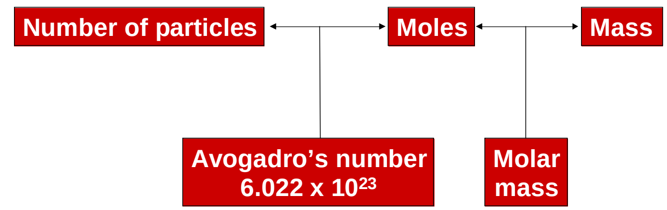
Chapter 5

Solution Preparation Review

- What volume of 2.06 M potassium permanganate contains 322 g of the solute?

Chapter 3

Conversion Factors



Chapter 3

Molar Conversions

- Determine the following:
  - The moles of potassium atoms in a 50.0 g sample

$$\text{Grams} \times \frac{1\text{ mol}}{\text{grams}} = \text{moles}$$

  - The mass of Mg in a 1.82 mole sample

$$\text{Moles} \times \frac{\text{grams}}{1\text{ mol}} = \text{grams}$$

Chapter 3

Molar Conversions

- Determine the following:
  - The moles of $FeCl_3$ in a 50.0 g sample
  - The mass of $MgCl_2$ in a 2.75 mole sample

Molar Conversions

- glucose = $C_6H_{12}O_6$ (Molar mass = 180.1 g)
  - How many moles of glucose are contained in 70.0 g of $C_6H_{12}O_6$?

$$\text{Grams} \times \frac{1\ \text{mol}}{\text{grams}} = \text{moles}$$

  - How many moles of oxygen are contained in 70.0 g of $C_6H_{12}O_6$?
  - How many grams of H are contained in 70.0 g of $C_6H_{12}O_6$?

Molar Ratios

- The relative number of moles of each element in a substance can be used as a conversion factor called the molar ratio.
- Molar ratio = $\frac{\text{moles element A}}{\text{mole of substance}}$

or

- Molar ratio = $\frac{\text{moles element A}}{\text{moles element B}}$

Molar Ratios

- H_2O:
  - Molar Ratio = $\frac{\text{2 moles of H}}{\text{mole } H_2O}$
  - Molar Ratio = $\frac{\text{1 mole O}}{\text{mole } H_2O}$
  - Molar Ratio = $\frac{\text{2 moles H}}{\text{mole O}}$

Molar Ratios

- **$C_6H_{12}O_6$:**
  - Molar Ratio = $\frac{\text{6 moles of C}}{\text{mole } C_6H_{12}O_6}$
  - Molar Ratio = $\frac{\text{12 moles H}}{\text{mole } C_6H_{12}O_6}$
  - Molar Ratio = $\frac{\text{6 moles O}}{\text{mole } C_6H_{12}O_6}$
  - Molar Ratio = $\frac{\text{6 moles C}}{\text{12 moles H}} = \frac{\text{1 mol C}}{\text{2 moles H}}$

Molar Conversions

- Molar ratios can be used to determine the number of moles of a particular element in a given substance.
- How many moles of oxygen are contained in 70.0 g of $C_6H_{12}O_6$?

Moles $C_6H_{12}O_6$ $\xleftrightarrow{\text{Molar Ratio}}$ Moles O

Molar Conversions

Given: 70.0 g of $C_6H_{12}O_6$
Find: Moles of O

Strategy:

grams of glucose $\xrightarrow{\text{Molar mass}}$ moles of glucose $\xrightarrow{\text{Molar ratio}}$ moles of oxygen

$$\text{Moles of O} = 70.0\text{ g } \cancel{\text{of}}\ C_6H_{12}O_6 \times \frac{1\,\cancel{\text{mol}}\ C_6H_{12}O_6}{180.1\ \cancel{\text{g}}} \times \frac{6\text{ mol O}}{\cancel{\text{mol}}\ C_6H_{12}O_6}$$

$$= 2.33\text{ mol O}$$

Molar Conversions

- Once the number of moles of a substance present is known, we can use:
  - Molar mass to find the number of grams
  - Avogadro's number to find the number of atoms, ions, or molecules

Moles A $\xleftrightarrow{\text{Molar Ratio}}$ Moles B

Grams B $\xleftrightarrow{\text{Molar mass}}$ Moles B $\xleftrightarrow{N_A}$ Atoms B

Molar Conversions

- How many grams of H are contained in 70.0 g of $C_6H_{12}O_6$?

Moles $C_6H_{12}O_6$ $\xleftrightarrow{\text{Molar Ratio}}$ Moles H

Grams H $\xleftrightarrow{\text{Molar mass}}$ Moles H $\xleftrightarrow{N_A}$ Atoms H

Molar Conversions

Given: 70.0 g of $C_6H_{12}O_6$
Find: gram H

$$\text{g H} = 70.0\ \text{g}\ \cancel{C_6H_{12}O_6} \times \frac{1\text{mol}\ \cancel{C_6H_{12}O_6}}{180.1\ \cancel{\text{g}}} \times \frac{12\ \cancel{\text{mol}}\ \text{H}}{\cancel{\text{mol}\ C_6H_{12}O_6}} \times \frac{1.0079\text{g}}{\cancel{\text{mol}}\ \text{H}}$$

=4.70g H

Molar Conversions

- How many moles of nitrogen are contained in 70.0 g of $C_6H_5NO_3$?
- How many grams of oxygen are contained in 1.5 moles of $C_6H_5NO_3$?
- How many atoms of C are contained in 70.0 g of $C_6H_5NO_3$?

Molar Conversions

- How many moles of H^+ ions are present in 2.5 moles of H_2SO_4 ?

Given: 2.5 moles H_2SO_4
Find: moles H^+

Conversion factor: molar ratio

$$\text{Mol } H^+ = 2.5\ \cancel{\text{mol } H_2SO_4} \times \frac{2\ \text{mol } H^+}{1\ \cancel{\text{mol } H_2SO_4}}$$

$$\text{Mol } H^+ = 5.0\ \text{mol } H^+$$

Empirical and Molecular Formulas

- Empirical formula: smallest whole-number ratio of atoms present in a compound
- Molecular formula: actual number of each type of atom present in a given compound
- Many molecular compounds have different empirical and molecular formulas

Molecular Formula	Empirical Formula
C_6H_6	CH
$C_6H_{12}O_6$	CH_2O
N_2H_4	NH_2
N_2O_4	NO_2

Chapter 3

Percent Composition

- Empirical formulas are generally obtained by determining the percent composition of a compound:
- Percent composition:
  - the percentage of the mass contributed by each element in a substance
- % Element X = (# atoms of X)(AW) x 100% / FW of compound

$$\%\text{ Element X} = \frac{(\#\text{ atoms of X})(\text{AW})}{\text{FW of compound}} \times 100\%$$

Formula Weight (FW): the sum of the atomic weights of all of the atoms in a chemical formula

Note: Molar mass is a **mass** in grams that is numerically the same as the **formula weight**

Chapter 3

Percent Composition

Calculate the % composition of H_2O (i.e find %H and %O).

First, find the FW of H_2O:

FW = 2(1.0079 amu) + 1(15.9994 amu)
= 18.0152 amu

$$\%\text{ H} = \frac{2(1.0079\text{ amu})}{18.0152\text{ amu}} \times 100\% = 11.2\%\text{ H}$$

$$\%\text{ O} = \frac{1(15.9994\text{ amu})}{18.0152\text{ amu}} \times 100\% = 88.8\ \%\text{ O}$$

Chapter 3

Mass Percentage

- **What is the mass percentage of each element in urea (CH_4N_2O)?**

FW = 12.011 amu + 4(1.0079 amu) + 2 (14.0067 amu) + 1(15.9994 amu)
= 60.0554 amu

$$\%\text{ C} = \frac{1(12.011\text{ amu})}{60.0554\text{ amu}} \times 100\% = 20.00\%\text{ C}$$

$$\%\text{ H} = \frac{4(1.0079\text{ amu})}{60.0554\text{ amu}} \times 100\% = 6.71\ \%\text{ H}$$

$$\%\text{ N} = \frac{2(14.0067\text{amu})}{60.0554\text{ amu}} \times 100\% = 46.65\%\text{ C}$$

$$\%\text{ O} = \frac{1(15.9994\text{ amu})}{60.0554\text{ amu}} \times 100\% = 26.64\ \%\text{ H}$$

Chapter 3

Calculating Empirical Formulas

- Percent composition data is commonly used to determine the empirical formula of a compound.

Empirical formula: smallest whole-number ratio of atoms present in a compound

Percent composition: the percentage of the mass contributed by each element in a substance

- Four steps:
  - percent to mass
  - mass to mole
  - divide by smallest
  - multiple 'til whole

Calculating Empirical Formulas

- A sample of nicotine has the following mass composition: 74.03% C; 8.70% H; 17.27% N. What is the empirical formula of nicotine?

Step 1: % to mass

If you assume 100.00 g of substance, then

74.03 % C → 74.03 g C

8.70 % H → 8.70 g H

17.27 % N → 17.27 g N

Calculating Empirical Formulas

Step 2: Mass to moles

$$\text{\# moles C} = 74.03\ \cancel{\text{g C}} \times \frac{1\ \text{mole C}}{12.01\ \cancel{\text{g C}}} = 6.164\ \text{mol C}$$

$$\text{\# moles H} = 8.70\ \cancel{\text{g H}} \times \frac{1\ \text{mole H}}{1.008\ \cancel{\text{g H}}} = 8.631\ \text{mol H}$$

$$\text{\# moles N} = 17.27\ \cancel{\text{g N}} \times \frac{1\ \text{mole N}}{14.00\ \cancel{\text{g N}}} = 1.234\ \text{mol N}$$

Calculating Empirical Formulas

Step 3: Divide by smallest

(this gives the molar ratio of the elements)

$$\text{C:}\ \frac{6.164\ \text{moles C}}{1.234\ \text{moles N}} = 4.995$$

$$\text{H:}\ \frac{8.631\ \text{mol H}}{1.234\ \text{mol N}} = 6.994$$

$$\text{N:}\ \frac{1.234\ \text{mol N}}{1.234\ \text{mol N}} = 1.000$$

Calculating Empirical Formulas

Step 4: Multiply 'til Whole

(this step is necessary only when step 3 does not give whole number molar ratios)

$$\text{C:}\ \frac{6.164\ \text{moles C}}{1.234\ \text{moles C}} = 4.995 \times 1 = 5.000\ \text{mol C}$$

$$\text{H:}\ \frac{8.631\ \text{mol H}}{1.234\ \text{mol C}} = 6.994 \times 1 = 7.000\ \text{mol H}$$

$$\text{N:}\ \frac{1.234\ \text{mol N}}{1.234\ \text{mol C}} = 1.000 \times 1 = 1.000\ \text{mol N}$$

Calculating Empirical Formulas

- Use these numbers to write the empirical formula:

5.000 mol C
7.000 mol H
1.000 mol N

$\left.\begin{array}{l}\textbf{5.000 mol C}\\ \textbf{7.000 mol H}\\ \textbf{1.000 mol N}\end{array}\right\}\ \mathbf{C_5H_7N}$

Calculating Empirical Formulas

- An iron compound contains 69.943 % Fe and 30.057 % O. Calculate its empirical formula.

- Step 1: % to mass

  – 69.943 % Fe → 69.943 g Fe
  – 30.057 % O → 30.057 g O

Calculating Empirical Formulas

- Step 2: Mass to Moles

$$\text{mol Fe} = 69.943\ \cancel{\text{g Fe}} \times \frac{\text{mol Fe}}{55.845\ \cancel{\text{g Fe}}} = 1.2524\ \text{mol Fe}$$

$$\text{mol O} = 30.057\ \cancel{\text{g O}} \times \frac{\text{mol O}}{15.9994\ \cancel{\text{g O}}} = 1.8786\ \text{mol O}$$

Calculating Empirical Formulas

- Step 3: Divide by smallest

$$\text{O:}\quad \frac{1.8786\ \text{mol O}}{1.2524\ \text{mol Fe}} = 1.5000$$

$$\text{Fe:}\quad \frac{1.2524\ \text{mol Fe}}{1.2524\ \text{mol Fe}} = 1.0000$$

Calculating Empirical Formulas

- Step 4: Multiply 'til whole

$FeO_{1.5}$ doesn't make sense.

1.5000 mol O x 2 = 3.000 mol O

1.0000 mol Fe x 2 = 2.000 mol Fe

Fe_2O_3

Using Empirical Formulas

- A sample of nicotine has the following mass composition: 74.03% C; 8.70% H; 17.27% N
- The molar mass of nicotine is 162.23 g
- What is the molecular formula of nicotine?

Using Empirical Formulas to Find Molecular Formulas

- Steps:
  - Find the empirical formula
  - Calculate the formula weight for the empirical formula.
  - Get the whole number ratio between MW and FW
  - Multiply the subscripts in the empirical formula by the whole number ratio.

Using Empirical Formulas to Find Molecular Formulas

Step 1:

Empirical formula of nicotine: C_5H_7N

Step 2:

Formula Weight = 81.06

Molar mass = Molecular weight = 162.23

Step 3:

Ratio= $\frac{MW}{FW} = \frac{162.23}{81.06} = 2.00$

Step 4:

Molecular formula: $C_{(5x2)}H_{(7x2)}N_{(1x2)} = C_{10}H_{14}N_2$

Chapter 3

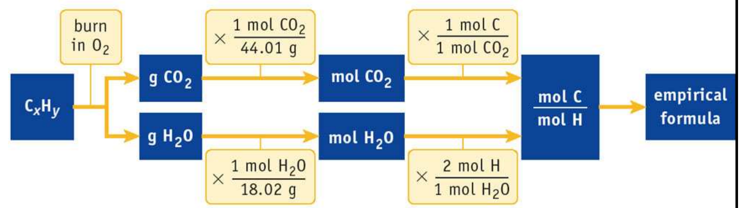
Using Mass Percent Data

- Ethylene glycol analyzes as 38.70 %C, and 9.74 %H with the remainder being oxygen.
- Determine the empirical formula of ethylene glycol

Chapter 4

Combustion Analysis

- A 1.125 g sample of a hydrocarbon was burned to produce 3.447 g of CO_2 and 1.647 g of H_2O. Determine the empirical formula.



Chapter 3

Empirical and Molecular Formulas

Molecular weight : Mass spectrometer

% C, H, O, N : Combustion analysis

Empirical and Molecular Formulas of an unknown compound