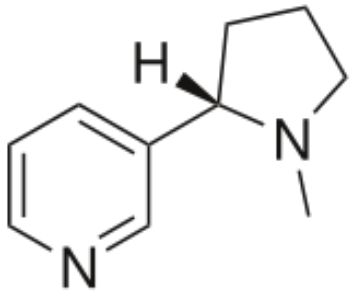
Unit 3 - Presentation A

The Mole, Empirical, and Molecular Formulas



The molecular formula for nicotine is

$C_{10}H_{14}N_2$

The Mole

Recall that 1 mole is defined as 6.022×10^{23} units of a given substance.

1 mol of electrons = 6.022×10^{23} electrons

1 mol of H_2O molecules = 6.022×10^{23} molecules of water

1 mol of NaCl formula units = 6.022×10^{23} formula units NaCl

1 mol of K atoms = 6.022×10^{23} atoms of K

The Mole

Within 1 mole of a compound, there are often differing moles of each element

In 1 mole of $Al(NO_3)_3$

= 1 mol of Al^{3+} ions

= 3 mol of NO_3^- ions

= 3 mol of N atoms

= 9 mol of O atoms

The Mole

Example: How many O atoms are present in 2.0 moles of aluminum nitrate?

move for answer

Molar Mass and Volume

Recall that the mass of 1 mol of a substance is called the molar mass and is measured in g/mol. This can be found on the periodic table.

Molar mass of $CaCl_2$ = 110 g/mol

Molar Mass of Ag = 108 g/mol

Recall also that 1 mol of any gaseous substance will occupy 22.4 L of space at STP.

1 mol of Ar(g) = 22.4 L @STP

1 mol of H_2(g) = 22.4 L @STP

Molar Mass and Volume

Example: What is the volume occupied @STP by 88 grams of carbon dioxide?

move for answer

1

1

Answer

2.4×10^{24} hydroxide ions

2

Natural gas (methane) pipeline.

2

Answer

44,800 mL

Natural gas (methane) pipel

3

- 2 grams of H_2 gas
- 16 grams of methane (CH_4)
- 22.4 L of H_2 gas
- 9 grams of water (H_2O)
- They all contain the same # of H atoms

3

- 2 grams of H_2 g
- 16 grams of m
- 22.4 L of H_2 g
- 9 grams of wat
- They all contain

Answer

B

4

- 0.5 moles
- 78.5 moles
- 1 mole
- 1.5 moles
- 2 moles

4

- 0.5 moles
- 78.5 moles
- 1 mole
- 1.5 moles
- 2 moles

Answer

C

Chemical Formulas

A chemical formula provides the ratio of atoms or moles of each element in a compound.

$H_2O = \frac{2 \text{ atoms H}}{1 \text{ atom O}}$ or $\frac{2 \text{ mol H}}{1 \text{ mol O}}$

$Al(NO_3)_3 = \frac{1\ Al^{3+} \text{ ion}}{3\ NO_3^- \text{ ions}}$ or $\frac{1 \text{ mol } Al^{3+} \text{ ions}}{3 \text{ mol } NO_3^- \text{ ions}}$

Empirical and Molecular Formulas

An empirical formula provides the ***simplest*** whole number ratio of atoms or moles of each element in a compound.

Examples: H_2O, NaCl, C_3H_5O

A molecular formula represents the ***actual*** number of atoms or moles of each element in a compound.

Examples: H_2O, C_3H_5O, $C_6H_{12}O_6$

Empirical and Molecular Formulas

There are two reasons for determining an empirical formula.

Reason 1:

Many compounds are really a gigantic molecule composed of trillions upon trillions of atoms. No one would want to write the actual formula of an NaCl crystal as...

$$Na_{1.8 \times 10^{24}}Cl_{1.8 \times 10^{24}}$$

so we just write the empirical formula instead (NaCl)

Reason 2:

In order to determine the molecular formula, we must first calculate the empirical formula anyway.

Calculating an Empirical Formula

To find an empirical formula:

1. Determine the moles of each element within the compound then.....

Compound "X" consists of 1.2 g C, 0.2 g H, and 1.6 g O

= 0.1 mol C, 0.2 mol H, and 0.1 mol O

2. Find the whole number ratio of these moles by dividing by smallest mole value!

$$\frac{0.1 \text{ mol C}}{0.1 \text{ mol}} = 1 \text{ C} \qquad \frac{0.2 \text{ mol H}}{0.1 \text{ mol}} = 2 \text{ H} \qquad \frac{0.1 \text{ mol O}}{0.1 \text{ mol}} = 1 \text{ O}$$

Empirical formula = CH_2O

Calculating a Molecular Formula

Determining the molecular formula of a compound is easy once the empirical formula and the molecular weight of the compound are known.

1. Determine the ratio of the molecular weight to the empirical formula weight.

MW of Compound "X" = 60 u

Empirical formula weight of CH_2O = 30 u

Ratio = 60/30 = 2/1.

The molecule is twice as heavy as the empirical formula.

2. Multiply each subscript of empirical formula by the ratio determined in step 1

CH_2O x 2 = **$C_2H_4O_2$** = Molecular Formula

Calculating Empirical and Molecular Formulas

Class Example: Given the following data, calculate the empirical formula of phosphine gas. Phosphine gas is created by reacting solid phosphorus with $H_2(g)$.

Mass of P(s) initial	Mass of P(s) unreacted
1.45 g	1.03 g

Mass of $H_2(g)$ initial	Mass of $H_2(g)$ unreacted
0.041 g	0.000 g

move for answer

Calculating Empirical and Molecular Formulas

Class example 2: Black iron oxide (aka magnetite) is used as a contrast agent in MRI scans of human soft tissue. To determine the empirical formula, a student reacted solid iron with $O_2(g)$.

Fe(s) reacted	Mass of iron oxide obtained.
3.05 g	4.22 g

What is the empirical formula?

move for answer

Calculating Empirical and Molecular Formulas

Class example 3: Lactic acid is produced when our muscle cells run out of oxygen.

Fe(s) reacted	Mass of iron oxide obtained.
3.05 g	4.22 g

What is the empirical formula?

move for answer

Calculating Empirical and Molecular Formulas

Class Example 4: Butane gas can be produced when solid carbon is reacted with hydrogen gas. If 0.45 grams of carbon were found to react with 1.05 L of H_2 gas @STP, what is the molecular formula of butane given it has a molar mass of 58 g/mol.

move for answer

5

- Fe_2O_3
- H_2NNH_2
- CH_3OH
- CH_3CH_2Cl
- All are empirical formulas

5

- Fe_2O_3
- H_2NNH_2
- CH_3OH
- CH_3CH_2Cl
- All are empirical

Answer

B

6

- Na_2N_3
- Na_3N
- NaN_3
- NaN
- None of these

6

- Na_2N_3
- Na_3N
- NaN_3
- NaN
- None of these

Answer

C

7

- C_2H_5Cl
- CH_2Cl_2
- C_2H_6Cl
- CH_3Cl
- None of these

7

- C_2H_5Cl
- CH_2Cl_2
- C_2H_6Cl
- CH_3Cl
- None of these

Answer

B

8

- NH_2
- NH_3
- N_2H_6
- N_2H_4
- None of these

8

- NH_2
- NH_3
- N_2H_6
- N_2H_4
- None of these

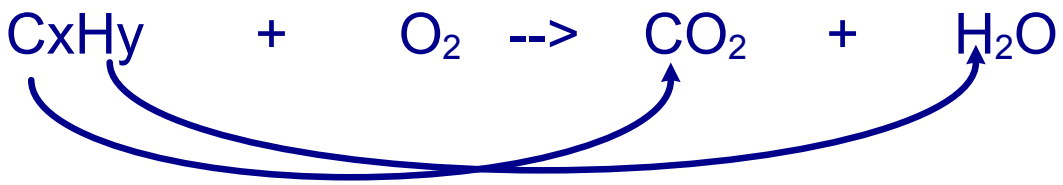
Answer

D

Combustion and Elemental Analysis

The mass amount of each element in a compound can often be found by decomposing the compound into its elements or by reacting it with another substance.

Organic compounds (containing C) can be combusted with oxygen to determine the mass amounts of carbon and hydrogen.



All C will be converted to CO_2 and all H will be converted to H_2O.

Combustion and Elemental Analysis

C_xH_y	+	O_2	-->	CO_2	+	H_2O
0.97 g		excess		2.67 g		2.18 g

The empirical formula can be determined by determining the mass of C in CO_2 and H in H_2O

$$2.67\ g\ CO_2 \times \frac{12\ g\ C}{44\ g\ CO_2} = 0.73\ g\ C = 0.061\ mol\ C$$

$$2.18\ g\ H_2O \times \frac{2\ g\ H}{18\ g\ H_2O} = 0.24\ g\ H = 0.24\ mol\ H$$

$0.061/0.061 = C_1$ $\quad$ $0.24/0.061 = H_4$

Empirical Formula = CH_4

Combustion and Elemental Analysis

Elements other than hydrogen or carbon must be determined either by direct analysis or by subtraction from the total mass of the organic compound.

$C_xH_yO_z$	+	O_2	-->	CO_2	+	H_2O
1.52 g		excess		2.91 g		1.78 g

2.91 g CO_2 --> 0.79 g C

1.78 g H_2O --> 0.20 g H

1.52 g $C_xH_yO_z$ - (0.79 g C + 0.20 g H) = 0.53 g O

0.79 g C --> 0.066 mol C	0.20 g H --> 0.20 mol H	0.53 g O --> 0.033 mol O
0.033	0.033	0.033
C_2	H_6	O_1

Empirical Formula = C_2H_6O

9

- CH_4
- C_2H_6
- C_7H_{14}
- C_3H_4
- None of these

9

- CH_4
- C_2H_6
- C_7H_{14}
- C_3H_4
- None of these

Answer

D

10

- CH_3F
- C_2H_5F
- CH_2F_2
- $C_4H_{14}F_2$
- None of these

10

- CH_3F
- C_2H_5F
- CH_2F_2
- $C_4H_{14}F_2$
- None of these

Answer

B