Unit 5: THE MOLE

1. Mole Definition & Background
2. Molar Mass
3. Mole Calculations
4. Percent Composition
5. Empirical Formulas
6. Molecular Formulas

1. Mole Definition & Background

- **The mole was developed from the work of Amadeo Avogadro (1776-1856) who studied the relationship between different gases and their volumes**

He found that a certain number of gas particles, no matter what type, take up the same amount of space. So, by finding the mass of the particles, and comparing it to the mass of hydrogen, one could find out how many particles were in that space.

Wilhelm Ostwald developed & introduced the mole concept in the 1890s, although it didn't catch on until the 1950s.

The Mole

- **The mole is the fundamental unit for measuring amount in chemistry. It's a conversion just like 1 dozen is 12 objects...**
  1 mole = 6.02×10^{23} objects
  602 000 000 000 000 000 000 000
  (Avogadro's Number)
- **The large size of Avogadro's number reflects its use in the counting of very small objects (atoms and molecules)**
- **It is not a perfect conversion, just an approximation, so sig figs count**

A dozen is a convenient way for bakers to talk of buns, whereas a mole is a convenient way for chemists to talk of atoms, molecules, or particles, as one mole of particles would fit in the palm of your hand.

The Mole

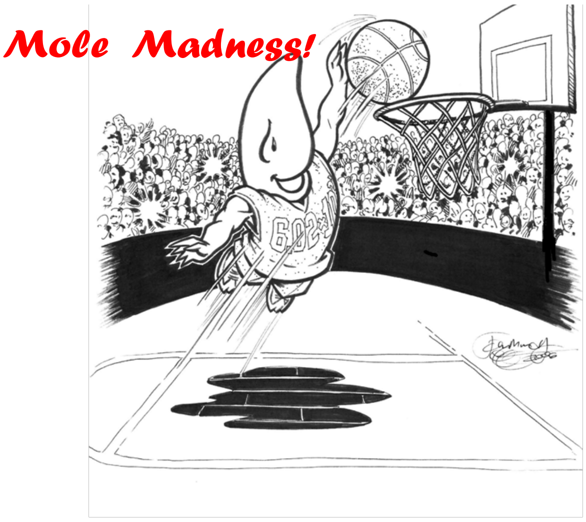
- **By definition, 1 mole is the number of atoms in 12.0 g of Carbon-12**
- **so 12.0 g of carbon contains approx. 6.02214179... x 10^{23} carbon atoms and is defined as 1 mole of carbon**
- **1 mole is further approximated to three sig figs: 6.02 x 10^{23}**

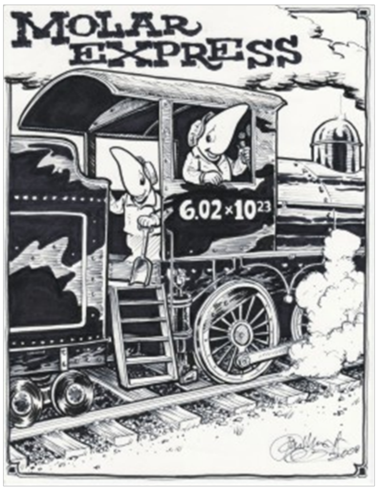
How big is 1 mole?

- **You could stack 1 mole of pennies back and forth from the earth to the sun 500 million times**
- **You could spread 1 mole of unpopped kernels across the U.S.A. and cover it to a depth of 9 miles**
- **If you had 1 mole of dollars and spent it at a rate of one million dollars a second, it would take 19 billion years to spend it all**

- 1 mole of coke cans would cover the surface of the earth to a depth of over 200 miles!

- The Mole Song





2. Molar Mass

Mass Numbers on the Periodic Table take on two meanings

- **Meaning #1:**
- **Atomic Mass: found by adding the number of protons with the number of neutrons – units are atomic mass units (a.m.u.)**
- **A carbon-12 atom has 6 protons and 6 neutrons and a mass of 12 a.m.u.**
- *Recall that masses on the p.t. are not whole numbers due to weighted average of isotopes*

- **Meaning #2**:
- **Molar Mass:** the mass in grams of 1 mole of that element (units are **g/mol**)
- Avogadro figured out how many atoms were in 12.0g of carbon, and called it '1 mole' of carbon
- Then, the mass of one mole of hydrogen atoms is 1.0g because each hydrogen atom has 1/12 the mass of a carbon atom, and so on
- So, the number for **atomic** and **molar** masses is the same, but their definitions and their units are different

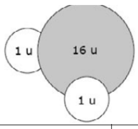
Molar Mass of Elements

- 1 mol of C is 12.0g. 1 mol of Fe is 55.8g
- this is because a C atom has less mass than an Fe atom
- a C atom has 12p and 12-14n, whereas an Fe atom has 26p, and 28-34n, so an Fe atom is much heavier
- thus a mole of Fe atoms would be much heavier than a mole of C atoms
-1mol of O_2 is 32.0g (2 x 16), because it's diatomic, so there are actually 2 mol of singular O atoms. This is the case for the 7 diatomics

http://www.wiredchemist.com/anim-mole

Molar Mass of a Compound

- The total mass of all atoms in the molecule
- Multiply the mass of each element by the subscript on that element in the formula
- E.g. Molar mass for H_2O
- = 2 (molar mass of H) + 1 (molar mass of O)
- = 2(1.0) + 16.0 = 18.0 g/mol
- so, 1 mole of water (6.02 x 10^{23} molecules) has a mass of 18.0g



Example: H_2O	1 mol = 18.0 g	$\frac{18.0\ g}{1\ mol}$	$\frac{1\ mol}{18.0\ g}$

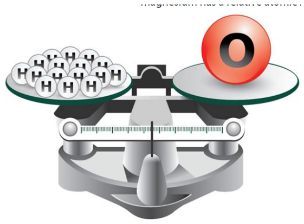


Figure 3.1c *The mass of an oxygen atom is equal to the mass of 16 hydrogen atoms.*

Calculate the Molar Mass

- Ne
- just look at periodic table   20.2 g/mol
- **Round all molar masses to the nearest TENTH (one decimal place)**
- SF_6
-   32.1 + 6(19.0) = 146.1 g/mol
- C_2H_5OH
-   2(12.0) + 6(1.0) + 16.0 = 46.0 g/mol

- H_2SO_4
-   2(1.0) + 32.1 + 4(16.0) = 98.1 g/mol
- $Cu(NO_3)_2$
-   63.5 + 2(14.0) + 6(16.0) = 187.5 g/mol
- Iron II sulfate trihydrate
- $FeSO_4$ $3H_2O$
- 55.8 + 32.1 +7(16.0) +6(1.0)= 205.9 g/mol

<u>HOMEWORK:</u>
Mole Problems #0
Worksheet - Part 1

3. MOLE CALCULATIONS

- Version 2 - Mole Song

Moles to Mass

- Find the mass of one mole of NaCl
- 58.5g
- Find the mass of two moles of NaCl
- 117.0g
- How did you get the answer?
- 2 moles x 58.5g/mol = 117.0g

- What is the mass of 0.500 moles of NaCl?
- 0.500mol x 58.5g/mol = 29.3g
- Find the mass of 0.35 moles of NaCl.
-

0.35mol	58.5g
	1 mol

= 20.475g = 2.0×10^1g

- Notice the mole unit cancels (one on top, one on the bottom), leaving only the gram unit!

- Using a table:

starting value	unit we want for answer
	unit we have for starting value

- Top left and bottom right units cancel, so we are left with unit we want in answer (the top right unit)

- Do all calculations for moles using a table as it will be very helpful for more complicated problems.
- Calculate the mass of 0.25mol of $MgBr_2$.
- Molar mass $MgBr_2$ = 24.3 + 2(79.9) = 184.1g/mol

0.25mol	184.1g
	1 mol

= 46g $MgBr_2$

- Calculate the mass of 2.40 mol of ammonium carbonate.
- NH_4^+ CO_3^{2-}
- NH_4^+ (to balance charges)
- $(NH_4)_2CO_3$
- MM = 2(14.0) + 8(1.0) + 12.0 + 3(16.0) =
- 96.0g/mol
-

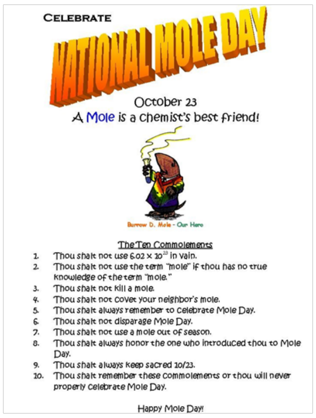
2.40mol	96.0g
	1 mol

= 230.4g = 2.30×10^2g

Calculate the mass of each

- 1.20 mol of HNO_3
- 75.6g HNO_3

- 0.60 mol of acetic acid
- 36g of acetic acid



CELEBRATE

NATIONAL MOLE DAY

October 23

A Mole is a chemist's best friend!

Burrow D. Mole - Our Hero

The Ten Commolements

1. Thou shalt not use 6.02×10^{23} in vain.
2. Thou shalt not use the term "mole" if thou has no true knowledge of the term "mole."
3. Thou shalt not kill a mole.
4. Thou shalt not covet your neighbor's mole.
5. Thou shalt always remember to celebrate Mole Day.
6. Thou shalt not disparage Mole Day.
7. Thou shalt not use a mole out of season.
8. Thou shalt always honor the one who introduced thou to Mole Day.
9. Thou shalt always keep sacred 10/23.
10. Thou shalt remember these commolements or thou will never properly celebrate Mole Day.

Happy Mole Day!

Mass to Moles

- How many moles are in 100.0g of CaS?
- MM of CaS = 40.1 + 32.1 = 72.2g/mol

- $\dfrac{100.0\text{g}}{} \,\Big|\, \dfrac{1 \text{ mol}}{72.2\text{g}}$

  = 1.39 mol CaS

- Note: Molar mass conversion can be written as 72.2g in one mole (72.2g/mol) or that one mole has a mass of 72.2g (1 mol/72.2g)

- How many moles are in 22.5g of aluminum oxide?
- Al_2O_3

- $\dfrac{22.5\text{g}}{} \,\Big|\, \dfrac{1 \text{ mol}}{102.0\text{g}}$

  = 0.221 mol Al_2O_3

- Calculate the number of moles in each of the following:
- 50.0g of copper (II) oxide

- 0.629 mol CuO
- 24mg of NaCl (Hint: change to g first)

- 4.1×10^{-4} mol NaCl

HOMEWORK: Mole Problems #0 Worksheet

Finding Molar Mass Using Experimental Data

- If you do not know the chemical formula of a substance, its molar mass can be found if you know the mass of a sample and the number of moles it contains.

Example

- If a 0.35 mole sample of a compound has a mass of 44.4 g, what is its molar mass?
- Molar mass = $\frac{\text{grams}}{\text{moles}} = \frac{44.4\text{ g}}{0.35\text{ mol}} = 130\text{ g/mol}$
- If you know this compound is either $FeCl_3$ or $FeCl_2$ you can compare molar masses with these compounds and find a match.
- $FeCl_3$ = 55.9 + 3(35.5) = 162.4 g/mol
- $FeCl_2$ = 55.9 + 2(35.5) = 126.9 g/mol
- The unknown compound is $FeCl_2$

Moles to Molecules

- We must use Avogadro's number
- There are 6.02×10^{23} particles in 1 mole
- How many molecules are in 1.00 mole of NaCl?
- 6.02×10^{23} molecules
- How many molecules are in 2.00 mole of NaCl?
- 1.20×10^{24} molecules

- How many molecules are in 0.500 mole of NaCl?
- 3.01×10^{23} molecules
- How many molecules are in 0.500 mole of CuS?
- 3.01×10^{23} molecules
- http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/stoichiometry/solid_atoms.html
- Keep in mind that Avogadro's number is ***independent*** of the type of molecule or object we are talking about.
- Mass is ***dependent*** because every type of atom has a different mass!

Name	Equivalence Statement	Conversion Factors	
Avogadro's number	1 mol = 6.02×10^{23} items	$\frac{6.02 \times 10^{23}\text{ items}}{1\text{ mol}}$	$\frac{1\text{ mol}}{6.02 \times 10^{23}\text{ items}}$

- Use the TABLE for all questions!
- How many molecules are in 1.20 mol of $PbSO_4$?
-
- $\frac{1.20\cancel{\text{mol}}}{} \Big| \frac{6.02 \times 10^{23}\text{ molecs}}{1\cancel{\text{mol}}}$

$= 7.22 \times 10^{23}$ molecs

- How many molecules in 1.30 mol of potassium chloride?
-
- $\frac{1.30\text{ mol}}{} \Big| \frac{6.02 \times 10^{23}\text{ molecs}}{1\text{ mol}}$

$= 7.83 \times 10^{23}$ molecs

Exercises

- Calculate the number of molecules in each of the following:
- 0.75 mol H_2SO_4
- Ans: 4.5 x 10^{23} molecs of H_2SO_4
- 2.35 x 10^{-1} mol $Ca(OH)_2$
- Ans: 1.41 x 10^{23} molecs of $Ca(OH)_2$

- Happy Mole Day

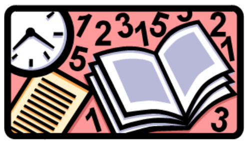
Molecules to Moles

- How many moles in 5.42 x 10^{22} molecules of Na_2S ?
-
- $\frac{5.42 \times 10^{22} \text{ molecs}}{} \Big| \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{molecs}}$

$= 9.00 \times 10^{-2}$ mol Na_2S

Exercise

- How many moles in 3.21 x 10^{23} molecules of KBr ?
- 0.533 mol KBr



Moles to Atoms

- If we are dealing with an element, do moles to atoms just as you did moles to molecules, because an element is made up of singular atoms (except the 7 diatomics)
- If we are dealing with compounds, we must do a 'moles to molecules' step and then a 'molecules to atoms' step

- How many atoms are in 0.10mol of H_2SO_4
- 1st -How many molecules are in 0.10 mol of H_2SO_4?
- 2nd -How many atoms in one molecule of H_2SO_4?
- $\frac{0.10 \text{ mol}}{} \Big| \frac{6.02 \times 10^{23} \text{ molec}}{1 \text{ mol}} \Big| \frac{7 \text{ atoms}}{1 \text{ molec}}$

$= 4.2 \times 10^{23}$ atoms

- Note: the number of atoms in a molecule is a counting number, not a measurement, so sig figs don't apply!

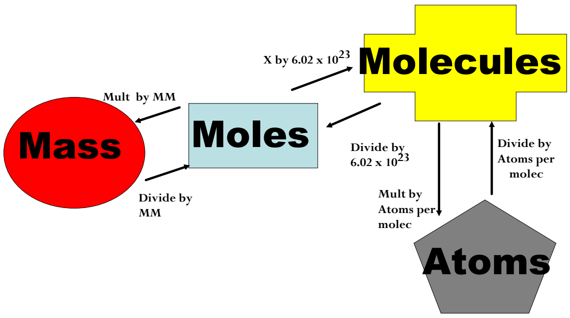
- How many atoms are in 1.50 mol of calcium carbonate?
- $CaCO_3$
- $\frac{1.50\,\text{mol}}{} \Big| \frac{6.02 \times 10^{23}\,\text{molecs}}{1\,\text{mol}} \Big| \frac{5\,\text{atoms}}{1\,\text{molec}}$

$= 4.52 \times 10^{24}$ atoms

Exercise

- Determine the number of oxygen atoms in 0.65 mol $Mg(OH)_2$.
- 7.8×10^{23} atoms

MOLE MAP



<u>HOMEWORK:</u>
Mole Problems #1 - Parts 1 & 2

Multi Step Mole Problems

- Many times problems will involve more than one conversion in the same question.
- These problems are solved by using multiple conversion factors.
- Pay attention to cancelling units to aid in putting the numbers in the correct positions

Mass to Molecules

- How many molecules are present in 60.0g of NaOH?
- Find the molar mass of NaOH and set up your table

60.0 g	1mol	6.02×10^{23} molec
	40.0g	1 mol

= **9.03×10^{23} molec**

- How many molecules are present in 5.00g of magnesium sulphate?
- $MgSO_4$ MM = 24.3 + 32.1 + 4(16.0) = 120.4 g/mol
-
-

5.00g	1mole	6.02×10^{23} molecs
	120.4 g	1mole

$= 2.50 \times 10^{22}$ molecs

Mass to Atoms

- How many atoms would be present in 4.50g of copper (II) chloride
- Find the formula and MM for copper (II) chloride
-

4.50g	1 mole	$6.02. \times 10^{23}$ molecs	3 atoms
	134.5 g	1 mole	1 molec

$= 6.04 \times 10^{22}$ atoms

Molecules to Mass

- What is the mass of 1.2×10^4 molecules of copper (II) chloride?
-

1.2×10^4 molecs.	1 mol	134.5 g
	6.02×10^{23} molecs	1 mol

$= 2.7 \times 10^{-18}$ g

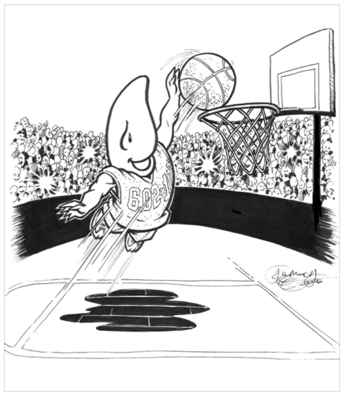
Exercise

- How many carbon atoms would be present in 795.0mg of acetic acid?
-

795mg	1 g	1 mole	6.02×10^{23} molecs	2 atoms
	1000mg	60.0 g	1 mol	1 molec

$= 1.60 \times 10^{22}$ carbon atoms

HOMEWORK:
Mole Problems #1 - Parts 3 & 4



Moles and Volumes of Gases

- 1.00 mole of <u>any gas</u> will occupy a volume of 22.4 L at STP
-
- STP = <u>S</u>tandard <u>T</u>emperature and <u>P</u>ressure
- Standard Temperature = 0 degrees C
- Standard Pressure = 1 atmosphere

Moles to Volume

- How much space does 0.563 moles of H_2 gas take up at STP?
- $\frac{0.563 \text{ mol}}{} \Big| \frac{22.4 \text{ L}}{1 \text{ mol}} = 12.6 \text{L}$

Volume to Moles

- How many moles in 18.7L of CH_4 gas (methane) at STP?
- $\frac{18.7\text{L}}{} \Big| \frac{1 \text{ mole}}{22.4 \text{ L}} = 0.835 \text{ moles}$

Mass to Volume

- What would be the volume of 90.0g of H_2 gas at STP?
- $\frac{90.0 \text{ g}}{} \Big| \frac{1 \text{ mol}}{2.0\text{g}} \Big| \frac{22.4 \text{ L}}{1 \text{ mol}} = 1008 \text{ L}$

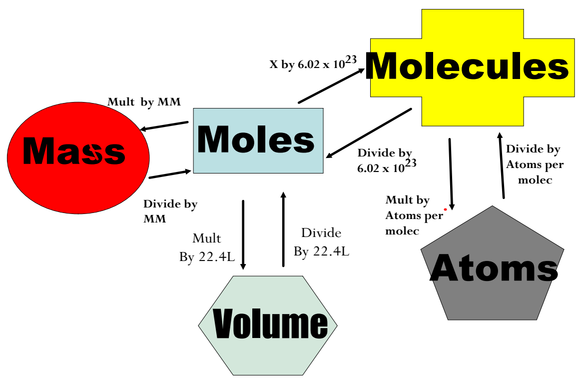
$= 1.0 \times 10^3 \text{ L}$

Molecules to Volume

- What volume (in litres) is occupied by 8.14 x 10^{22} molecules of CO_2 gas (at STP)?
-
- $\frac{8.14 \times 10^{22} \text{ molecs}}{} \Big| \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ molecs}} \Big| \frac{22.4 \text{ L}}{1 \text{ mol}}$

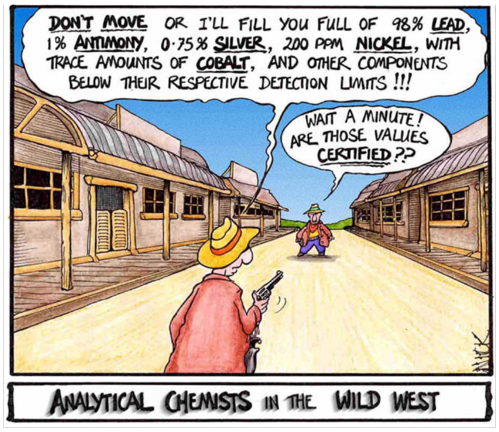
$= 3.03 \text{ L}$

MOLE MAP



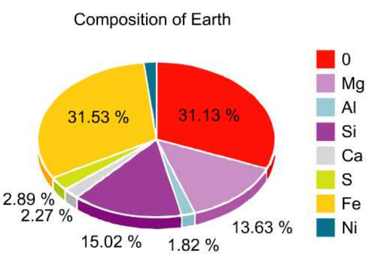
HOMEWORK:
Mole Problems #1 - Part 5
Mole Problems #2 Worksheet

- Official Mole Day Website



59

4. Percent Composition



Percent composition tells us the amount, by mass, of a certain component within an item.

Real-World Connections

What is it?

- The percentage (by mass) of each element in a chemical formula
- AND
- The percentage (by mass) of each component in a mixture

From a Chemical Formula:

- Find atomic mass of each element and molar mass of entire compound
- Assume 1 mole of substance
- Calculate mass of each element by multiplying atomic mass by subscript
- Divide mass of each element by mass of compound (molar mass) then express as a %

Find the % composition of Fe_2O_3

- Atomic mass of Fe = 55.8 g/mol
- Atomic mass of O = 16.0 g/mol
- Molar mass of Fe_2O_3
- = 2 x 55.8 g/mol + 3 x 16.0 g/mol
- = 159.6 g/mol

- Mass Fe = 55.8 g/mol x 2 mol Fe
- = 111.6 g
- % Fe = $\frac{111.6\ g}{159.6\ g}$ x 100 = 69.92 %

- Mass of O = 16.0 g/mol x 3 m
- = 48.0 g
- % O = $\frac{48.0\ g}{159.6\ g}$ x 100 = 30.1 %

Fe_2O_3 is 69.92 % iron and 30.1 % oxygen by mass

What is the % composition of $C_6H_{12}O_6$

- Mass of 1 mole $C_6H_{12}O_6$ =
- = 6(12.0 g/mol) + 12(1.0 g/mol) + 6(16.0 g/mol)
- = 180.0 g
- % C = $\frac{(6 \times 12.0\ g)}{180.0\ g}$ x 100 = 40.0 %
- % H = $\frac{(12 \times 1.0\ g)}{180.0\ g}$ x 100 = 6.7 %
- % 0 = $\frac{(6 \times 16.0\ g)}{180.0\ g}$ x 100 = 53.3 %

From experimental data:

- A sample of a gas compound is analyzed and found to contain 23.5 g nitrogen and 53.9 g oxygen. What is the % composition?
- Total mass 23.5 + 53.9 = 77.4 g
- % N = $\frac{23.5\ g}{77.4\ g}$ x 100 = 30.4 %
- % O = $\frac{53.9\ g}{77.4\ g}$ x 100 = 69.6 %

What is the percent water in calcium chloride hexahydrate ?

$CaCl_2 \cdot 6H_2O$

Mass of H_2O= 6(1.0 +1.0+16.0) = 108.0 g

- Formula mass $CaCl_2 \cdot 6H_2O$
- = 40.1 + (2 x 35.5) + (108.0)
- = 219.1 g
- % H_2O = $\frac{108.0\ g}{219.1\ g}$ x 100 = 49.29 %

% Composition of a Mixture

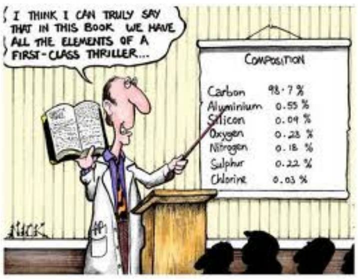
- Brass is made by melting together copper, zinc and other metals
- What is the percent composition of the alloy if the starting materials are 350 kg copper, 180 kg zinc and 25 kg tin?

- Total mass of mixture
- = 350 kg + 180 kg + 25 kg
- = 555 kg

- % copper = $\frac{350 \text{ kg}}{555 \text{ kg}}$ x 100 = 63 %
- % zinc = $\frac{180 \text{ kg}}{555 \text{ kg}}$ x 100 = 32 %
- % tin = $\frac{25 \text{ kg}}{555 \text{ kg}}$ x 100 = 4.5 %

71

HOMEWORK:
Percent Composition
Worksheet



5. Empirical Formulas

What is Empirical Formula?

- The formula of a compound determined from experimental data (mass or percent)
- Gives whole number ratio of elements
- Simplest or "reduced" form of the molecular formula

Examples

- Molecular Formula: C_2H_6
- Empirical Formula: CH_3
-
- Molecular Formula: $C_6H_{12}O_6$
- Empirical Formula: CH_2O
-
- Molecular Formula: $C_4H_9O_3$
- Empirical Formula: $C_4H_9O_3$

1. Mass of component elements

- Data may be in mass or as a percent
- If data is in percent, assume 100 g total and use percent as mass

- 25.3 % element A and 74.7 % element B becomes
- 25.3 g element A and 74.7 g element B if we work with 100.0 g total

2. Convert Mass to Moles

- Convert each element to moles by dividing by molar mass of the element

3. Find Mole Ratio

- Divide each number of moles by the smallest number

4. Find Whole Number Ratio

- Multiply to eliminate all decimals (or fractions)

Example # 1

- A 43.2 g lump of charcoal (carbon) burns with oxygen and produces 159.0 g of a new compound. What is the empirical formula of this compound?
- Step 1:
- Mass of Carbon = 43.2 g
- Mass of Oxygen = 159.0 - 43.2 g
- = 115.8 g

2. Find moles of each element

(divide by molar mass)

- $\frac{43.2 \text{ g C}}{} \Big| \frac{1 \text{ mol C}}{12.0 \text{ g C}}$ = 3.60 mol C
- $\frac{115.8 \text{ g O}}{} \Big| \frac{1 \text{ mol O}}{16.0 \text{ g O}}$ = 7.24 mol O

3. Divide to find mole ratio

(always divide by smallest # of moles)

- $\frac{43.2 \text{ g C}}{} \Big| \frac{1 \text{ mol C}}{12.0 \text{ g C}} = \frac{3.60 \text{ mol C}}{\mathbf{\textit{3.60}}}$ = 1.00 mol C
- $\frac{115.8 \text{ g O}}{} \Big| \frac{1 \text{ mol O}}{16.0 \text{ g O}} = \frac{7.24 \text{ mol O}}{\mathbf{\textit{3.60}}}$ = 2.01 mol O

Step 4 unnecessary for this question!

- 2 moles of oxygen for each 1 mole of carbon
- Empirical Formula is . . .
- CO_2

Example #2

- Analysis of a compound showed that it contained 70.0 % iron and 30.0 % oxygen. What is the formula of this compound?
- Assume we have 100 g

Multiply to find Whole Numbers

Element	mass in 100.0 g	mol	mol ratio	simplest ratio
Fe	70.0	$\frac{70.0}{55.8}$ = 1.25	$\frac{1.25}{1.25}$ = 1.00	2
O	30.0	$\frac{30.0}{16.0}$ = 1.88	$\frac{1.88}{1.25}$ = 1.50	3

Empirical Formula is . . . Fe_2O_3

Example #3: Find Empirical Forumula if a compound has 20.2 % Al, 79.8 % Cl

Element	mass in 100 g	mol	mol ratio	simplest ratio
Al	20.2 g			
Cl	79.8 g			

Example #3: Find Empirical Forumula if a compound has 20.2 % Al, 79.8 % Cl

Element	mass in 100 g	mol	mol ratio	simplest ratio
Al	20.2 g	$\frac{20.2}{27.0}$ =.748	$\frac{.748}{.748}$ = 1.00	1
Cl	79.8 g	$\frac{79.8}{35.5}$ =2.25	$\frac{2.25}{.748}$ = 3.01	3

Empirical Formula is $AlCl_3$

HOMEWORK:
Percent Composition & Empirical Formulas Worksheet

http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_03&folder=percent_composition

6. Molecular Formula

Actual Formula of a Compound

Molecular Formula

- Experimental data can only give the simplest whole number ratios for a formula
- Some compounds may have the same empirical formula, but have different molecular formulas
- E.g. NO_2 and N_2O_4
- C_2H_2 and C_6H_6

- Molecular formula is always a whole number multiple of the empirical formula
- To find the multiplier, divide the molar mass by the empirical mass (mass of empirical formula)

Example #1

- Many hydrocarbons (compounds containing only hydrogen and carbon) have the same empirical formula.
- Two compounds both have an empirical formula of CH_2. The molar mass of compound A is 28.0 g/mol. The molar mass of compound B is 70.0 g.mol What is the molecular formula of each?

Compound A

- Empirical mass = 12.0 + 2(1.0) = 14.0 g/mol

$$\frac{\text{Molar Mass A}}{\text{Empirical Mass}} = \frac{28.0\,\text{g/mol}}{14.0\,\text{g/mol}} = 2\ \ (\textit{multiplier})$$

- Molecular formula of A = (CH_2) x 2 = C_2H_4

Compound B

- Empirical mass = 12.0 + 2(1.0) = 14.0 g/mol
-
- Multiplier = $\frac{\text{Molar Mass B}}{\text{Empirical Mass}} = \frac{70.0\,\text{g/mol}}{14.0\,\text{g/mol}} = 5$

- Molecular formula of B = (CH_2) x 5 = C_5H_{10}

Example # 2

- The empirical formula for a compound is NO_2. If 0.405 moles of the substance has a mass of 37.2 g. What is the molecular formula?
- Empirical mass = 14.0 + 2(16.0) = 46.0 g/mol
- Molar mass = $\frac{37.2\text{ g}}{0.405\text{ mol}} = 91.9\text{ g/mol}$

Find Multiplier

$$\frac{\text{Molar Mass}}{\text{Empirical Mass}} = \frac{91.9\,\text{g/mol}}{46.0\,\text{g/mol}} = 1.998 \approx 2$$

Formula is (NO_2) X 2 or N_2O_4

HOMEWORK:
Empirical & Molecular
Formula Worksheet