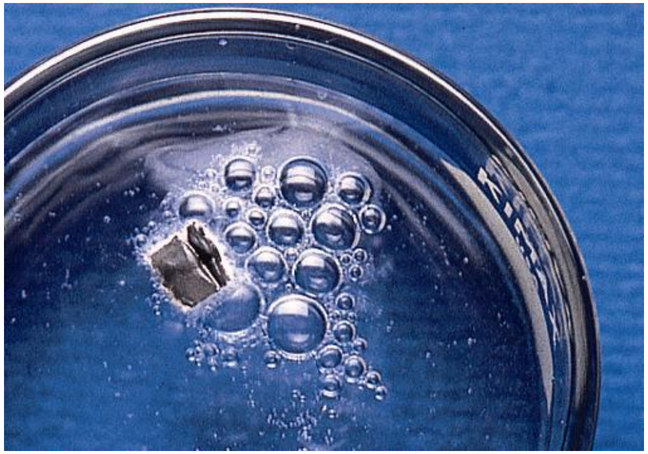
Mass Relationships in Chemical Reactions

Chapter 3



Micro World
atoms & molecules

→

Macro World
grams

Atomic number: number of protons which equal the number of electrons in neutral atom

Atomic mass is the mass of an atom in atomic mass units (amu) (1amu=1.6 x10^{-24}g)

amu: *define as mass exactly equal to 1/12 of the mass of Carbon-12 ……*

By definition:
1 atom ^{12}C "weighs" 12 amu

On this scale

^{1}H = 1.00794 amu

^{16}O = 15.9994 amu

Ex sulfer-36 has mass of 35.967 amu, which is around 3 times the mass of C-12 [35.967/12=2.99]

- When express the mass in amu, mass of atom is approximately equal the number of protons and neutrons.

Average atomic mass: is the weighted average of all the naturally occurring isotopes

Ex: Natural lithium is:7.42% ^{6}Li (6.015 amu) and 92.58% ^{7}Li (7.016 amu)

Average atomic mass of lithium=

$$[(7.42/100) \times 6.015] + [(92.58/100) \times 7.016] = 6.941 \text{ amu}$$

24 Cr 52.00	
24	Atomic number
Cr	
52.00	Atomic mass

Average atomic mass (6.941)

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 **H** 1.008																	2 **He** 4.003
3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.88	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
55 **Cs** 132.9	56 **Ba** 137.3	57 **La** 138.9	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.9	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (210)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (257)	105 **Ha** (260)	106 **Sg** (263)	107 **Ns** (262)	108 **Hs** (265)	109 **Mt** (266)	110	111	112						

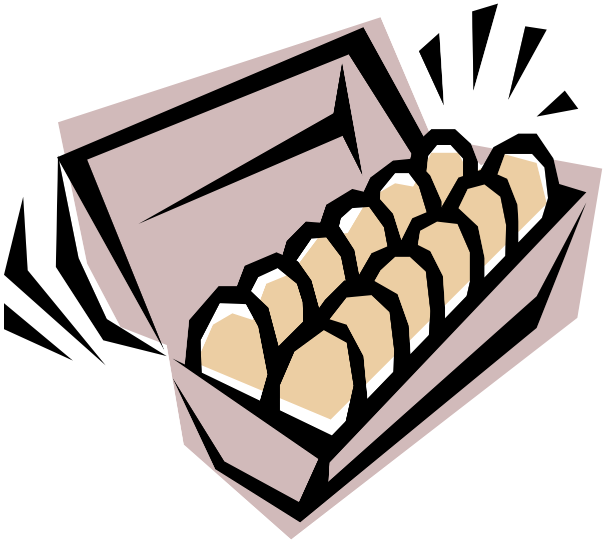
58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (147)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.0	71 **Lu** 175.0
90 **Th** 232.0	91 **Pa** (231)	92 **U** 238.0	93 **Np** (237)	94 **Pu** (242)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (249)	99 **Es** (254)	100 **Fm** (253)	101 **Md** (256)	102 **No** (254)	103 **Lr** (257)

Metals

Metalloids

Nonmetals

Dozen = 12



Pair = 2

Mole = 6.022×10^{23} units

<u>The mole</u>* *(mol) is a unit to account the number of particles(atoms, molecules,…)

- Number of atoms in exactly 12 grams of ^{12}C = **6.022×10^{23} atoms** **(**experimentally**)**

1 mole of ^{12}C =N_A= 6.022×10^{23} atoms = 12.011 g

Avogadro's number = N_A

– Number of atoms, molecules or particles in one mole

1 mole of X = 6.022×10^{23} units of X

- 1 mole Xe = 6.022×10^{23} Xe atoms
- 1 mole NO_2 = 6.022×10^{23} NO_2 molecules

Molar mass: ($\mathcal{M}$), defined as the mass (in grams or kilograms) of 1 mole of units (such as atoms or molecules) of a substance

1 mole ^{12}C atoms = 12.00 g = 6.022 x 10^{23} atoms

1 mole lithium atoms = 6.941 g of Li

For any element

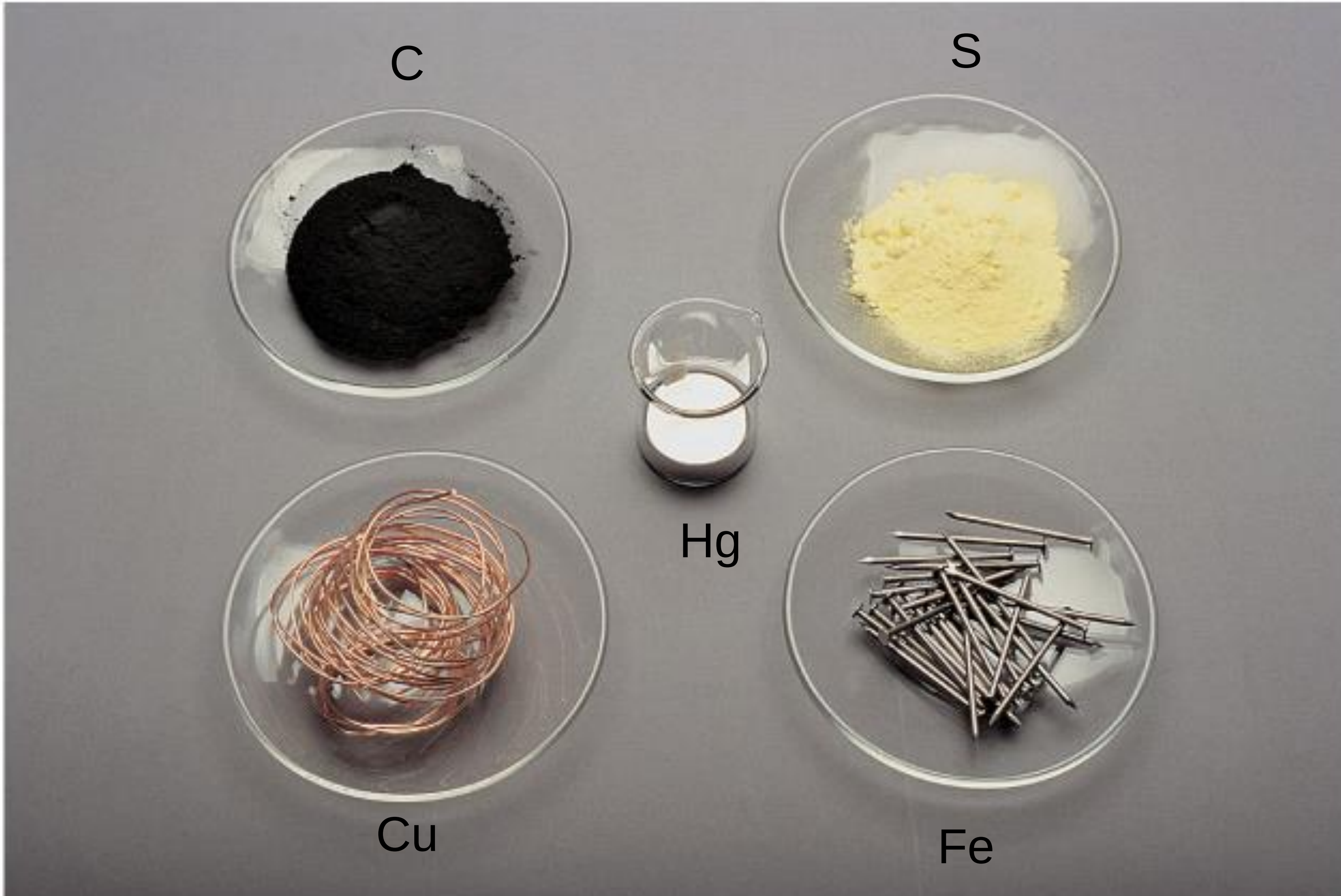
atomic mass (amu) = molar mass (grams/mol)

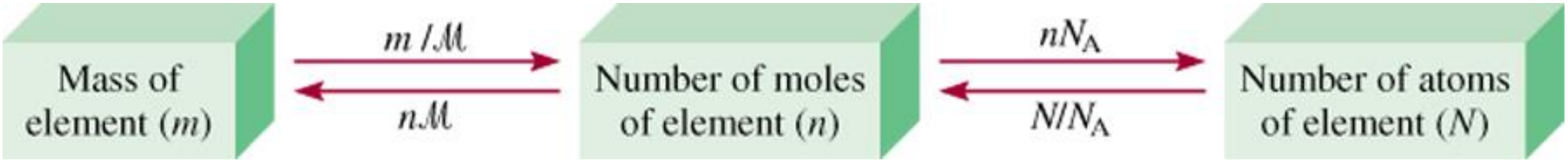
from periodic table

Atomic mass of O=16 amu

Molar mass of O = 16g/mol

One Mole of:





$\mathcal{M}$ = molar mass in g/mol   , n = mole

N_A = Avogadro's number

Two main rules

1- mole=mass/molar mass

$$\boldsymbol{n = \frac{m}{\mu}}$$

2- number of atoms (or molecules)= moles x Avogadro's #

$$N = n.N_A$$

Learning Check: Using Molar Mass

Ex. How many moles of iron (Fe) are in 15.34 g Fe?

- What do we know?

1 mol Fe = 55.85 g Fe

- What do we want to determine?

15.34 g Fe = ? Mol Fe

Start

End

- Set up ratio so that what you want is on top & what you start with is on the bottom

$$15.34\ \cancel{\text{g Fe}} \times \left(\frac{1\ \text{mol Fe}}{55.85\ \cancel{\text{g Fe}}} \right)$$ **= 0.2747 mole Fe**

Or using direct way

$$n = \frac{m}{\mu} = \frac{15.34}{55.85} = \mathbf{0.2747\ mole\ Fe}$$

Ex: How many potassium atoms are in 0.551 g of potassium (K) ?

1 mol of K = 39.10 g of K

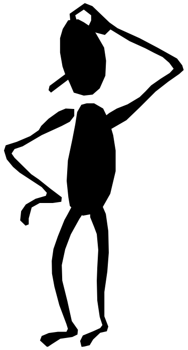
$$N = n.N_A = \left(\frac{0.551}{39.10}\right).(6.022\ x10^{23}) = 8.49 \text{ x } 10^{21} \text{ atoms of K}$$

Ex: calculate the mass of one atom of Na (Na=23g/mol)

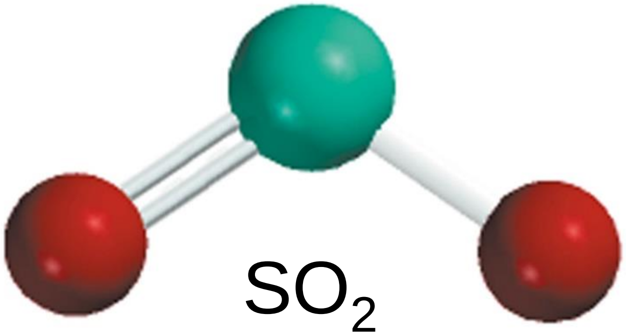
$$\boldsymbol{N = n.N_A}$$

$$1 = \left(\frac{m}{23}\right).(6.022x10^{23})$$

$$m = 3.82x10^{-23}\text{g}$$



Molecular mass (or molecular weight) is the sum of the atomic masses (in amu) in a molecule.



SO_2

1S	32.07 amu
2O	+ 2 x 16.00 amu
SO_2	64.07 amu

For any molecule

molecular mass in amu = molar mass in grams

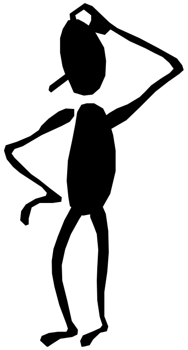
1 molecule of SO_2 weighs 64.07 amu

1 mole of SO_2 weighs 64.07 g

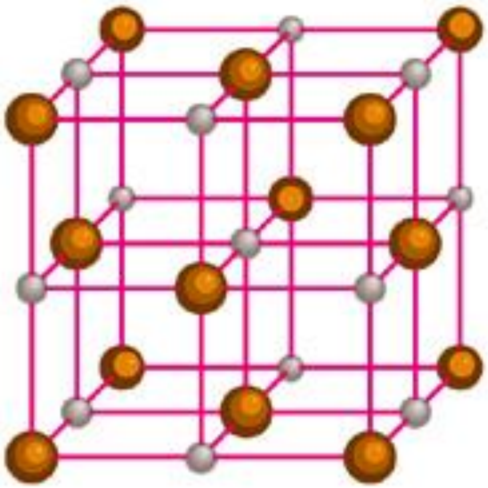
Ex How many H atoms are in 72.5 g of C_3H_8O ?

$$N = (n)_{molecule} . N_A . (\# \ of \ H \ per \ molecule)$$

$$N = \left(\frac{72.5}{60.09}\right) . (6.022x10^{23}) . (8) =$$ 5.82×10^{24} H atoms



Formula mass is the sum of the atomic masses (in amu) in a formula unit of an ionic compound.



NaCl

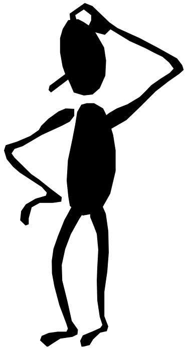
1Na		22.99 amu
1Cl	+	35.45 amu
NaCl		58.44 amu

For any ionic compound

formula mass (amu) = molar mass (gram/mol)

1 formula unit of NaCl = 58.44 amu

1 mole of NaCl = 58.44 g of NaCl



Do You Understand Formula Mass?

What is the formula mass of $Ca_3(PO_4)_2$?

1 formula unit of $Ca_3(PO_4)_2$

3 Ca		3 x 40.08 g/mol
2 P		2 x 30.97 g/mol
8 O	+	8 x 16.00 g/mol
		310.18 g/mol

Units of <u>grams per mole</u> are the most practical for chemical calculations!

<u>Ex</u> Calculate the mass in grams of $FeCl_3$ in 1.53×10^{23} formula units. (molar mass = 162.204 g/mol)

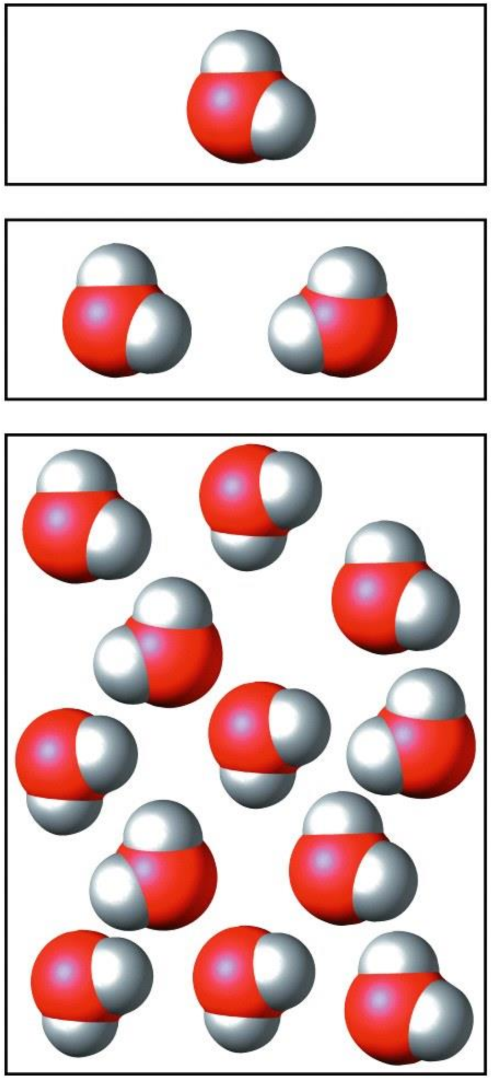
$$N = n.N_A$$
$$1.53\text{x}10^{23} = \left(\frac{m}{162}\right)\text{x}(6.022\text{x}10^{23})$$
$$m = 41.21g$$

<u>Ex</u> Calculate the number of formula units of Na_2CO_3 in 1.29 moles of Na_2CO_3

$$\mathrm{N} = \mathrm{n.N_A}$$
$$N = (1.29)\text{x}(6.022\text{x}10^{23}) = 7.77\times10^{23} \text{ particles } Na_2CO_3$$

Mole-to-Mole Conversion Factors

- Can use chemical formula to relate amou of each atom to amount of compound
- In H_2O there are 3 relationships:
  - 2 mol H $\Leftrightarrow$ 1 mol H_2O
  - 1 mol O $\Leftrightarrow$ 1 mol H_2O
  - 2 mol H $\Leftrightarrow$ 1 mol O
- Can also use these on atomic scale
  - 2 atom H $\Leftrightarrow$ 1 molecule H_2O
  - 1 atom O $\Leftrightarrow$ 1 molecule H_2O



Ex Calculate the number of **moles** of calcium in 2.53 moles of $Ca_3(PO_4)_2$

2.53 moles of $Ca_3(PO_4)_2$ = ? mol Ca

$$3 \text{ mol Ca} \Leftrightarrow 1 \text{ mol } Ca_3(PO_4)_2$$

$$2.53 \cancel{\text{mol } Ca_3(PO_4)_2} \left(\frac{3 \text{ mol Ca}}{1 \cancel{\text{mol } Ca_3(PO_4)_2}} \right)$$

= 7.59 mol Ca

Ex How many g of iron (Fe) are required to use up all of 25.6 g of oxygen atoms (O) to form Fe_2O_3?

- **mass O → mol O → mol Fe → mass Fe**

25.6 g O → ? g Fe

3 mol O ⇔ 2 mol Fe

$- n_O = \frac{25.6}{16} = 1.6\, mol$ O

3 mol O ⇔ 2 mol Fe

1.6 ⇔ ?? mol Fe

$- n_{Fe}$=(1.6 x 2)/3

=1.06

- mass of Fe = 1.06 x 55.85 =59.2 g of Fe

Determining Empirical & Molecular Formulas

- When making or isolating new compounds one must characterize them to determine structure &

Molecular Formula (MF)

- – Exact composition of one molecule
- – Exact whole # ratio of atoms of each element in molecule(**Ex** :**MF of glucose is $C_6H_{12}O_6$)**

Empirical Formula (EF)

- – Simplest ratio of atoms of each element in compound
- – Obtained from experimental analysis of compound

Ex **EF of glucose is CH_2O**

Ex what is the EF of pentane (C_5H_{12})

EF is C_5H_{12} same as molecular formula

Three Ways to Calculate Empirical Formulas

1. **From Masses of Elements**

   Ex. 2.448 g sample of which 1.771 g is Fe and 0.677 g is O.

2. **From Percentage Composition**

   Ex. 43.64 % P and 56.36 % O.

3. **From Combustion Data**

   Given masses of combustion products

   Ex. The combustion of a 5.217 g sample of a compound of C, H, and O in pure oxygen gave 7.406 g CO_2 and 4.512 g of H_2O.

Strategy for Determining Empirical Formulas

1. Determine mass in **g** of each element
2. Convert mass in **g** to **moles**
3. Divide all quantities by smallest number of moles to get smallest ratio of moles
4. Convert any non-integers into integer numbers.
   – Multiply by smallest number to make subscripts in step 3 integers

1. Empirical Formula from Mass Data

Ex: When a 0.1156 g sample of a compound was analyzed, it was found to contain 0.04470 g of C, 0.01875 g of H, and 0.05215 g of N. Calculate the empirical formula of this compound.

Step 1: Calculate moles of each substance

$$0.04470\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.011\ \cancel{\text{g C}}} = 3.722 \times 10^{-3}\ \text{mol C}$$

$$0.01875\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}} = 1.860 \times 10^{-2}\ \text{mol H}$$

$$0.05215\ \cancel{\text{g N}} \times \frac{1\ \text{mol N}}{14.0067\ \cancel{\text{g N}}} = 3.723 \times 10^{-3}\ \text{mol N}$$

1. Empirical Formula from Mass Data

Step 2: Select the smallest # of moles.

- The smallest is 3.722×10^{-3} mole

Step 3: Divide all # of moles by the smallest one

		Mole ratio	Integer ratio
• C =	$\dfrac{3.722\times10^{-3}\ \text{mol C}}{3.722\times10^{-3}\ \text{mol C}} =$	1.000	= 1
• H =	$\dfrac{1.860\times10^{-2}\ \text{mol H}}{3.722\times10^{-3}\ \text{mol C}} =$	4.997	= 5
• N =	$\dfrac{3.723\times10^{-3}\ \text{mol N}}{3.722\times10^{-3}\ \text{mol C}} =$	1.000	= 1

Empirical formula = CH_5N

1. Empirical Formula from Mass Data

Ex 2: One of the compounds of iron and oxygen, "black iron oxide," occurs naturally in the mineral magnetite. When a 2.448 g sample was analyzed it was found to have 1.771 g of Fe and 0.677 g of O. Calculate the empirical formula of this compound.

1. Calculate moles of each substance

$$1.771\ \cancel{\text{g Fe}} \times \frac{1\ \text{mol Fe}}{55.485\ \cancel{\text{g Fe}}} = 0.03171\ \text{mol Fe}$$

$$0.677\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.0423\ \text{mol O}$$

1. Empirical Formula from Mass Data

2. Divide both by smallest #mol to get smallest whole # ratio.

$$\frac{0.03171 \text{ mol Fe}}{0.03171 \text{ mol Fe}} = 1.000 \text{ Fe}$$

$$= Fe_{1.00}O_{1.33}$$

$$\frac{0.0423 \text{ mol O}}{0.03171 \text{ mol Fe}} = 1.33 \text{ O}$$

3-Multiply by smallest number to make subscripts in step 2 integers

$$Fe_{(1.00\times3)}O_{(1.33\times3)} = Fe_3O_{3.99}$$

Empirical Formula = Fe_3O_4

2. Empirical Formula from % Composition

Ex **:**Calculate the empirical formula of a compound whose % composition data is 43.64 % P and 56.36 % O.

Step 1: Assume 100 g of compound.

- 43.64 g P — 1 mol P = 30.97 g
- 56.36 g O — 1 mol O = 16.00 g

$$43.64\ \cancel{\text{g P}} \times \frac{1\ \text{mol P}}{30.97\ \cancel{\text{g P}}} = 1.409\ \text{mol P}$$

$$56.36\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 3.523\ \text{mol P}$$

2. Empirical Formula from % Composition

Step 2: Divide by smallest number of moles

$$\frac{1.409\ \text{mol P}}{1.409\ \text{mol P}} = 1.000$$

$$\frac{3.523\ \text{mol O}}{1.409\ \text{mol P}} = 2.500$$

Step 3: Multiple by **n** to get smallest integer ratio

Here **n = 2**

P: $1.00 \times 2 = 2$

O: $2.500 \times 2 = 5$

Empirical formula = P_2O_5

3. Empirical Formulas from Combustion Analysis:

Combustion Analysis

- Compounds containing carbon, hydrogen, & oxygen, can be burned completely in pure oxygen gas
  - Only carbon dioxide & water are produced

Ex. Combustion of ethanol (C_2H_5OH)

$$C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$$

Apparatus for determining the empirical formula of ethanol. The absorbers are substances that can retain water and carbon dioxide, respectively.

Combustion of ethanol

3. Empirical Formulas from Combustion Analysis:

- Carbon dioxide & water separated & weighed separately
  - All C ends up as CO_2
  - All H ends up as H_2O
  - Mass of **C** can be derived from amount of **CO_2**
    - mass $CO_2 \rightarrow$ mol $CO_2 \rightarrow$ mol C $\rightarrow$ mass C
  - Mass of **H** can be derived from amount of **H_2O**
    - mass $H_2O \rightarrow$ mol $H_2O \rightarrow$ mol H $\rightarrow$ mass H
  - Mass of oxygen is obtained by difference

    mass O = mass sample – (mass C + mass H)

Ex.The combustion of a 5.217 g sample of a compound of C, H, and O in pure oxygen gave 7.406 g CO_2 and 4.512 g of H_2O. Calculate the empirical formula of the compound.

	C	H	O	CO_2
Molar mass (g/mol)	12.011	1.008	15.999	44.01

1. Calculate mass of C from mass of CO_2.

mass CO_2 → mole CO_2 → mole C → mass C

$$7.406\ \cancel{\text{g CO}_2}\left(\frac{1\ \cancel{\text{mol CO}_2}}{44.01\ \cancel{\text{g CO}_2}}\right)\left(\frac{1\ \cancel{\text{mol C}}}{1\ \cancel{\text{mol CO}_2}}\right)\left(\frac{12.011\ \text{g C}}{1\ \cancel{\text{mol C}}}\right)$$

= 2.021 g C

2. Calculate mass of H from mass of H_2O.

mass $H_2O \rightarrow$ mol $H_2O \rightarrow$ mol H $\rightarrow$ mass H

$$4.512\ \cancel{g\ H_2O}\left(\frac{1\ \cancel{mol\ H_2O}}{18.015\ \cancel{g\ H_2O}}\right)\left(\frac{2\ \cancel{mol\ H}}{1\ \cancel{mol\ H_2O}}\right)\left(\frac{1.008\ g\ H}{1\ \cancel{mol\ H}}\right)$$

$$\mathbf{= 0.5049\ g\ H}$$

3. Calculate mass of O from difference.

Mass O= total mass-(C mass + H mass)

$= 5.217$ g– (2.021 g C + 0.5049 g H) **= 2.691 g O**

Or we can use the following rule:

$mass\ of\ element\ in\ sample = (\frac{mass\ of\ product\ contain\ this\ element}{it\ molar\ mass})$ x (#of element atoms in product) x (atomic mass of element)

$\therefore$ mass of C in CO_2: $\frac{7.406}{44.01}$ x 1 x 12.01= 2.02 mass of C

mass of H in H_2O: $\frac{4.512}{18}$ x 2 x 1.008 = 0.504 mass of H

	C	H	O
MM	12.011	1.008	15.999
g	2.021	0.5049	2.691

4. Calculate mol of each element

$$\text{mol C} = \frac{\text{g C}}{\text{MM C}} = \frac{2.021\ \text{g}}{12.011\ \text{g/mol}} = 0.1683\ \text{mol C}$$

$$\text{mol H} = \frac{\text{g H}}{\text{MM H}} = \frac{0.5049\ \text{g}}{1.008\ \text{g/mol}} = 0.5009\ \text{mol H}$$

$$\text{mol O} = \frac{\text{g O}}{\text{MM O}} = \frac{2.691\ \text{g}}{15.999\ \text{g/mol}} = 0.1682\ \text{mol O}$$

$$C_{\frac{0.1683}{0.1682}}H_{\frac{0.5009}{0.1682}}O_{\frac{0.1682}{0.1682}}$$

- Preliminary empirical formula
  - $C_{0.1683}H_{0.5009}O_{0.1682} = C_{1.00}H_{2.97}O_{1.00}$

5. Calculate mol ratio of each element

Empirical Formula = CH_3O

- – Since all values are close to integers, round to

Determining Molecular Formula

- Need molecular mass(molar mass) & empirical formula
- Calculate ratio of molecular mass to mass predicted by empirical formula & round to nearest integer

$$\textbf{Molecular formula} = \frac{\textit{molar mass of unknown}}{\textit{molar mass of EF}} \textbf{xEF}$$

$$\left(\frac{\textit{molar mass of unknown}}{\textit{molar mass of EF}}\right) = n$$

Ex. molar mass of Glucose is 180.16 g/mol

Empirical formula = CH_2O

Empirical formula molar mass = 30.03 g/mol, find the molecular formula for Glucose .

$$\text{Molecular formula} = \frac{180.16}{30.03} \times CH_2O = 6 \times CH_2O$$

Molecular formula is $C_6H_{12}O_6$

Ex The empirical formula of hydrazine is NH_2, and its molecular mass is 32.0. What is its molecular formula

Atomic Mass: N:14.007; H:1.008; O:15.999

Solution

Molar mass of NH_2 = (1×14.01) + (2×1.008) = 16.017g

$$\textbf{Molecular formula} = \frac{\textit{molar mass of unknown}}{\textit{molar mass of EF}} \text{ x EF}$$

$$= \frac{32}{16.017} \text{ x } NH_2$$

$$= 2 \text{ x } NH_2$$

Molecular formula : N_2H_4

Chemical reactions and chemical equations

A process in which one or more substances is changed into one or more new substances is a ***chemical reaction.***

A ***chemical equation*** uses chemical symbols to show what happens during a chemical reaction.

3 ways of representing the reaction of H_2 with O_2 to form H_2O

Two hydrogen molecules + One oxygen molecule $\longrightarrow$ Two water molecules

$$2H_2 + O_2 \longrightarrow 2H_2O$$

reactants $\longrightarrow$ products

How to "Read" Chemical Equations

$$2\,Mg + O_2 \longrightarrow 2\,MgO$$

2 atoms Mg + 1 molecule O_2 makes 2 formula units MgO

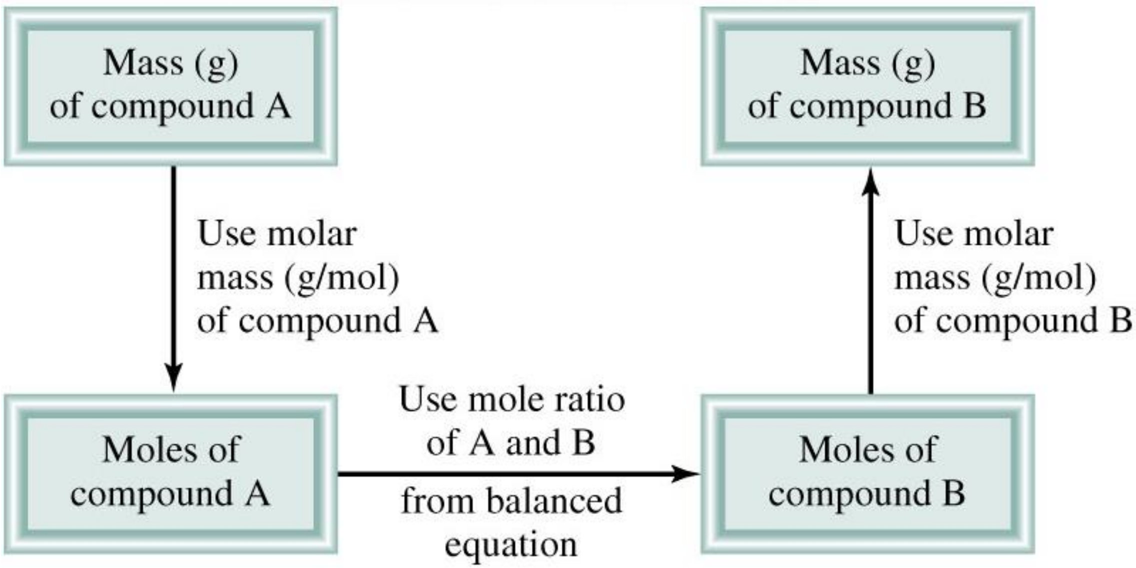
2 moles Mg + 1 mole O_2 makes 2 moles MgO

48.6 grams Mg + 32.0 grams O_2 makes 80.6 g MgO

IS NOT

2 grams Mg + 1 gram O_2 makes 2 g MgO

Stoichiometry Calculations: Amounts of Reactants and Products in chemical reaction



Use the **fabulous four steps**!

1. Write the **balanced chemical equation**.
2. Convert quantities of **known** substances into **moles**.
3. Use **coefficients** in balanced equation to calculate the number of **moles of the sought quantity**.
4. Convert moles of sought quantity into the **desired units**.

Using Balanced Equation to Determine Stoichiometry

Ex. What mass of O_2 will react with 96.1 g of propane (C_3H_8) gas, to form gaseous carbon dioxide & water?

Strategy

1. Write the balanced equation

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

2. Assemble the tools

96.1 g C_3H_8 → moles C_3H_8 → moles O_2 → g O_2

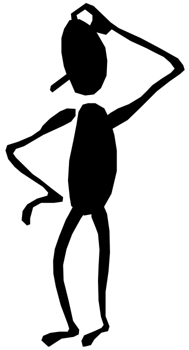
✓ $n_{C_3H_8} = \frac{96.1}{44.1} = 2.18$ mol C_3H_8

✓ 1mol C_3H_8 ⟶ 5 mol O_2

2.18 mol ⟶ ??? mol O_2

✓ $n_{O_2} = \frac{2.18 \text{ x } 5}{1} = 10.9$ mol O_2

✓ $mass_{O_2} = 10.9\ x\ 32 = 348.8\ g$



Ex: Methanol burns in air according to the equation

$$2\ CH_3OH + 3\ O_2 \longrightarrow 2\ CO_2 + 4\ H_2O$$

m=??

If 209 g of methanol are used up in the combustion, what mass of water is produced?

grams CH_3OH ⟶ moles CH_3OH ⟶ moles H_2O ⟶ grams H_2O

- ✓ Balanced equation
- ✓ $n_{CH_3OH} = \frac{209}{32} = 6.53\ mol$ CH_3OH
- ✓ n_{H2O} :

$2\ CH_3OH \longrightarrow 4H_2O$ (mol ratio)

6.53 mol ⟶ ???

- ✓ $n_{H_2O} = \frac{6.53x4}{2} = 13.06$
- ✓ $mass_{H_2O} = 13.06\ \text{x}\ 18 = 235.08\ g\ H_2O$

Limiting Reactant

- Reactant that is completely used up in the reaction
- Present in lower # of moles
- It determines the amount of product produced

Excess reactant

- Reactant that has some amount left over at end
- Present in higher # of moles

Four Steps to determine the limiting reagent

1. Balanced reaction: Done.
2. Find the mole of each reactant in the reaction
3. Divide the # of mole of each reactant on it coefficient
4. The one that give the smallest # in step 3 is the limiting reagent

Ex: For the following reaction $2NH_3 + CO_2 \longrightarrow (NH_2)_2CO + H_2O$

If we start with 637.2 g NH_3 and 1142g CO_2

a) which of the two reactants is the limiting reagent ?

1- Balanced reaction √

2- mol $NH_3 = \frac{637.2}{17.03} = 37.16$, and mol $CO_2 = \frac{1142}{44} = 25.9$

3- NH_3 : 37.16/2 =18.7 (Smallest) ∴ **NH_3 is the Limiting reagent**

and CO_2 : 25.9/1= 25.9 (CO_2 is excess)

b) Calculate the mass of product $(NH_2)_2CO$ formed

To find the mass formed of this product, we relate it to the limiting reagent (using mole ratio)

2 mol NH_3 $\longrightarrow$ 1 mol $(NH_2)_2CO$

37.16 mol $\longrightarrow$???? mol $(NH_2)_2CO$

✓ mol $(NH_2)_2CO = \frac{37.16 \times 1}{2} = 18.71$ mol

✓ mass $(NH_2)_2CO$ = 18.71 x 60.06 g/mol =1124 g (theoretical yield)

c) How many excess reagent (gram)is left at the end of reaction ?

The CO_2 excess left (left over)=(initial mass of CO_2 – reacted mass of CO_2)

- **First will find the Reacted mass of CO2 :**

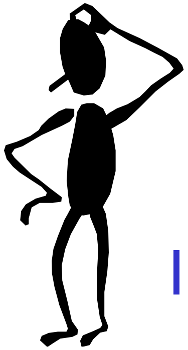
1 mol CO_2 $\longrightarrow$ 2mol NH_3

???? mol CO_2 $\longleftarrow$ 37.16 mol NH_3

✓ **mol CO_2 =37.16/2 =18.71**

✓ **reacted mass of CO_2 = 18.71 x 44= 823.4 g**

✓ **The mass left over of CO_2= initial mass - reacted mass= 1142-823.4= 319 g**



Do You Understand **Limiting Reactants**?

In a reaction, 124 g of Al are reacted with 601 g of Fe_2O_3.

$$2\,Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2\,Fe$$

Calculate the mass of Al_2O_3 formed in grams.

Also Limiting reagent can be determine based on the following statement
" the limiting reagent will yield the smaller amount of the product"

1. Balanced reaction: √
2. Moles of "given" reactants.

Moles of Al = 124 g / 26.9815 g/mol = 4.60 mol
Moles of Fe_2O_3 = 601 g / 159.6882 g/mol = 3.76 mol

3. Moles of "desired" product, Al_2O_3.

$$2\,Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2\,Fe$$

$$\text{Moles of } Al_2O_3 \text{ based on Al} = \frac{4.60 \text{ mol Al}}{1} \times \frac{1 \text{ mol } Al_2O_3}{2 \text{ mol Al}} = 2.30 \text{ mol } Al_2O_3$$

$$\text{Moles of } Al_2O_3 \text{ based on } Fe_2O_3 = \frac{3.76 \text{ mol } Fe_2O_3}{1} \times \frac{1 \text{ mol } Al_2O_3}{1 \text{ mol } Fe_2O_3} = 3.76 \text{ mole } Al_2O_3$$

Keep the smaller answer! Al is the limiting reactant.

4. Grams of Al_2O_3.

Grams of Al_2O_3 = 2.30 mol X 101.9612 g/mol = **235 g**

Reaction Yield

Theoretical Yield is the amount of product that would result if all the limiting reagent reacted.(calculated)

Actual Yield is the amount of product actually obtained from a reaction(experimentally).

-How much is obtained in mass units or in moles

- Actual yield usually is less than theoretical yield

Percent yield Relates the actual yield to the theoretical yield, It is calculated as:

$$\textbf{\% Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

<u>Ex</u>. If a cookie recipe predicts a yield of 36 cookies and yet only 24 are obtained, what is the % yield?

$$\text{percentage yield} = \left(\frac{24}{36}\right) \times 100 = 67\%$$

<u>Ex</u> When 18.1 g NH_3 and 90.4 g CuO are reacted, the theoretical yield is 72.2 g Cu. The actual yield is 58.3 g Cu. What is the percent yield?

$$2NH_3(g) + 3CuO(s) \rightarrow N_2(g) + 3Cu(s) + 3H_2O(g)$$

$$\%\text{ yield} = \frac{58.3\text{ g Cu}}{72.2\text{ g Cu}} \times 100\% \mathbf{= 80.7\%}$$

Ex A chemist set up a synthesis of solid phosphorus trichloride by mixing 12.0 g of solid phosphorus with 35.0 g chlorine gas and obtained 42.4 g of solid phosphorus trichloride. Calculate the percentage yield of this compound.

Analysis:

Write balanced equation

$$2P(s) + 3Cl_2(g) \longrightarrow 2PCl_3(s)$$

Determine Limiting Reagent → Determine Theoretical Yield → Calculate Percentage Yield

Solution:

$$2P(s) + 3Cl_2(g) \longrightarrow 2PCl_3(s)$$

1. Determine Limiting Reactant

mol of P= 12/30.97= 0.39/2=0.19

mol of Cl_2= 35/70.9= 0.49/3=0.16 (smallest) Cl_2 is limiting reagent

2. Determine Theoretical Yield **(mass of product** PCl_3**)**

3 Cl_2: 2 PCl_3

- ✓ mol PCl_3 =(0.49x2)/3= 0.33 mol PCl_3
- ✓ Mass PCl_3 = 0.33 x 137.32 = 44.9 g PCl_3 (theoretical yield)

3. Determine Percentage Yield

- Actual yield = 42.4 g

$$\text{percentage yield} = \left(\frac{42.4\ g\ PCl3}{44.9\ g\ PCl3}\right) x100=94.43\%$$

Stoichiometry Summary