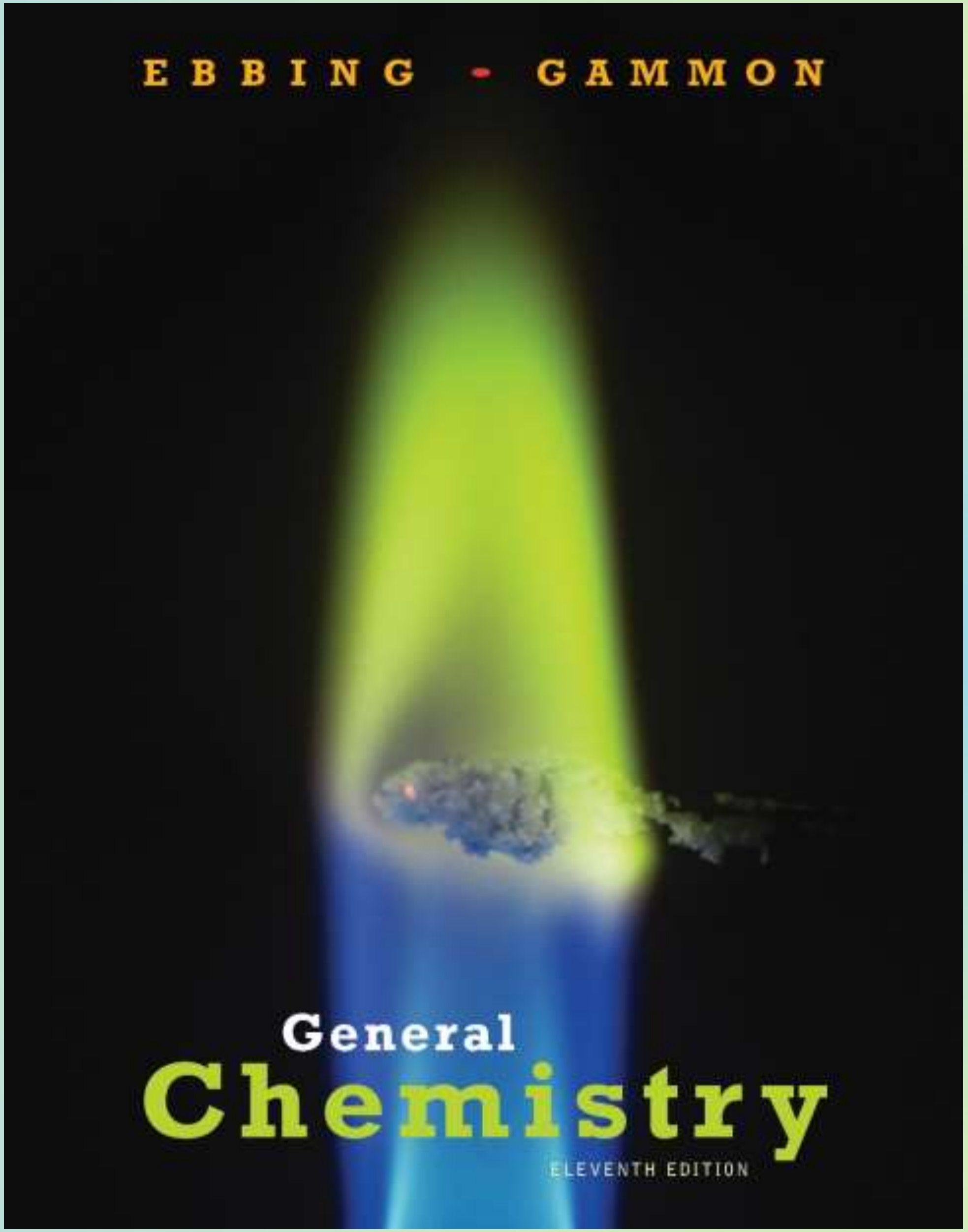


اللهم صلّ وسلّم على نبينا محمد وعلى آله وصحبه أجمعين

Calculations with Chemical Formulas and Equations

DISPLAY PROPERTY/TREND

Chemical Group Block

Atomic Number | 17 | 35.45 | Atomic Mass, u

Cl | Symbol

Chlorine | Name

Halogen | Chemical Group Block

Plot Atomic Mass

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 1.0080 H Hydrogen Nonmetal																	2 4.002... He Helium Noble Gas
2	3 7.0 Li Lithium Alkali Metal	4 9.0121... Be Beryllium Alkaline Eart...											5 10.81 B Boron Metalloid	6 12.011 C Carbon Nonmetal	7 14.007 N Nitrogen Nonmetal	8 15.999 O Oxygen Nonmetal	9 18.99... F Fluorine Halogen	10 20.180 Ne Neon Noble Gas
3	11 22.9... Na Sodium Alkali Metal	12 24.3... Mg Magnesium Alkaline Eart...											13 26.9... Al Aluminum Post-Transiti...	14 28.0... Si Silicon Metalloid	15 30.9... P Phosphorus Nonmetal	16 32.07 S Sulfur Nonmetal	17 35.45 Cl Chlorine Halogen	18 39.9 Ar Argon Noble Gas
4	19 39.0... K Potassium Alkali Metal	20 40.08 Ca Calcium Alkaline Eart...	21 44.9... Sc Scandium Transition Me...	22 47.8... Ti Titanium Transition Me...	23 50.9... V Vanadium Transition Me...	24 51.9... Cr Chromium Transition Me...	25 54.9... Mn Manganese Transition Me...	26 55.84 Fe Iron Transition Me...	27 58.9... Co Cobalt Transition Me...	28 58.6... Ni Nickel Transition Me...	29 63.55 Cu Copper Transition Me...	30 65.4 Zn Zinc Transition Me...	31 69.723 Ga Gallium Post-Transiti...	32 72.63 Ge Germanium Metalloid	33 74.9... As Arsenic Metalloid	34 78.97 Se Selenium Nonmetal	35 79.90 Br Bromine Halogen	36 83.80 Kr Krypton Noble Gas
5	37 85.4... Rb Rubidium Alkali Metal	38 87.62 Sr Strontium Alkaline Eart...	39 88.9... Y Yttrium Transition Me...	40 91.22 Zr Zirconium Transition Me...	41 92.9... Nb Niobium Transition Me...	42 95.95 Mo Molybden... Transition Me...	43 96.9... Tc Technetium Transition Me...	44 101.1 Ru Ruthenium Transition Me...	45 102.... Rh Rhodium Transition Me...	46 106.... Pd Palladium Transition Me...	47 107.... Ag Silver Transition Me...	48 112.41 Cd Cadmium Transition Me...	49 114.... In Indium Post-Transiti...	50 118.71 Sn Tin Post-Transiti...	51 121.7... Sb Antimony Metalloid	52 127.6 Te Tellurium Metalloid	53 126.... I Iodine Halogen	54 131.29 Xe Xenon Noble Gas
6	55 132.... Cs Cesium Alkali Metal	56 137.33 Ba Barium Alkaline Eart...		72 178.49 Hf Hafnium Transition Me...	73 180.... Ta Tantalum Transition Me...	74 183.... W Tungsten Transition Me...	75 186.... Re Rhenium Transition Me...	76 190.2 Os Osmium Transition Me...	77 192.22 Ir Iridium Transition Me...	78 195.... Pt Platinum Transition Me...	79 196.... Au Gold Transition Me...	80 200.... Hg Mercury Transition Me...	81 204.... Tl Thallium Post-Transiti...	82 207 Pb Lead Post-Transiti...	83 208.... Bi Bismuth Post-Transiti...	84 208.... Po Polonium Metalloid	85 209.... At Astatine Halogen	86 222.... Rn Radon Noble Gas
7	87 223.... Fr Francium Alkali Metal	88 226.... Ra Radium Alkaline Eart...		104 26... Rf Rutherford... Transition Me...	105 26... Db Dubnium Transition Me...	106 26... Sg Seaborgium Transition Me...	107 27... Bh Bohrium Transition Me...	108 26... Hs Hassium Transition Me...	109 27... Mt Meitnerium Transition Me...	110 282... Ds Darmstadt... Transition Me...	111 282... Rg Roentgeni... Transition Me...	112 286... Cn Coperniciu... Transition Me...	113 28... Nh Nihonium Post-Transiti...	114 29... Fl Flerovium Post-Transiti...	115 290... Mc Moscovium Post-Transiti...	116 29... Lv Livermorium Post-Transiti...	117 294... Ts Tennessine Halogen	118 29... Og Oganesson Noble Gas
				57 138.... La Lanthanum Lanthanide	58 140.... Ce Cerium Lanthanide	59 140.... Pr Praseody... Lanthanide	60 144.... Nd Neodymium Lanthanide	61 144.... Pm Promethium Lanthanide	62 150.4 Sm Samarium Lanthanide	63 151.... Eu Europium Lanthanide	64 157.2 Gd Gadolinium Lanthanide	65 158.... Tb Terbium Lanthanide	66 162.... Dy Dysprosium Lanthanide	67 164.... Ho Holmium Lanthanide	68 167.... Er Erbium Lanthanide	69 168.... Tm Thulium Lanthanide	70 173.05 Yb Ytterbium Lanthanide	71 174.9... Lu Lutetium Lanthanide
				89 227.... Ac Actinium Actinide	90 232.... Th Thorium Actinide	91 231.... Pa Protactini... Actinide	92 238.... U Uranium Actinide	93 237.... Np Neptunium Actinide	94 244.... Pu Plutonium Actinide	95 243.... Am Americium Actinide	96 247.... Cm Curium Actinide	97 247.... Bk Berkelium Actinide	98 251.... Cf Californium Actinide	99 252.... Es Einsteinium Actinide	100 25... Fm Fermium Actinide	101 25... Md Mendelevi... Actinide	102 25... No Nobelium Actinide	103 26... Lr Lawrencium Actinide

3.1 Molecular Weight and Formula Weight

Ionic compound

molecular weight (MW) of a substance is: *the sum of the atomic weights of all the atoms in a molecule of the substance*.

$^{A}_{Z}E$

Example: H_2O, is 18.0 amu

(2 x 1.0 amu +16.0 amu) =18 amu = *Molar Mass H2O = 18 g/mol*

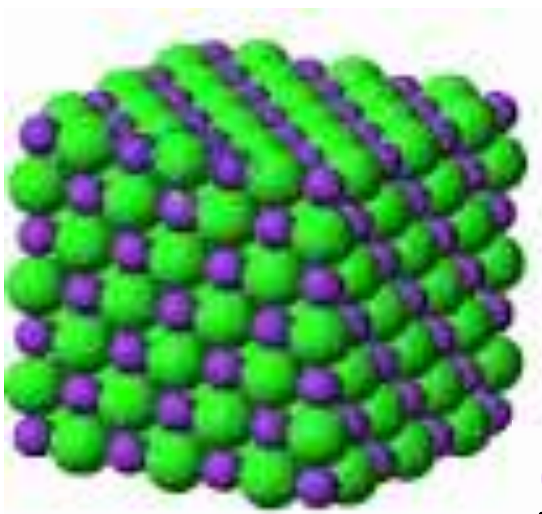
Amount of substance

Molar Mass = *Atomic weight* / mass

formula weight (FW) of a substance is: *the sum of the atomic weights of all atoms in a* ***formula unit*** *of the compound*, whether molecular or not

Example: NaCl, has a formula weight of 58.44 amu

(22.99 amu + 35.45 amu) = 58.44 amu = *Molar Mass Nacl = 58.44 g/mol*



Molecular weight = formula weight

Molar mass = Molecular weight

The Mass of one mole

$$\boxed{\text{Molecular Mass} = \text{formula Mass}}$$

$$\boxed{\text{Molar Mass } (g\backslash mol) = \text{molecular wight } (amu) = \text{formula wight } (amu)}$$

Example 3.1 Calculating the Formula Weight from a Formula

Calculate the formula weight of each of the following to three significant figures, using a table of atomic weights (AW):

a. chloroform, $CHCl_3$ b. $\boxed{\text{iron(III) sulfate, } Fe_2(SO_4)_3}$.

A.

$1 \times$ AM of C $=$ 12.0 amu

$1 \times$ AM of H $=$ 1.0 amu

$3 \times$ AM of Cl $= 3 \times 35.45$ amu $=$ 106.4 amu

FM of $CHCl_3$ $=$ 119.4 amu → Molecular weight ⇒ 119 amu

Molar Mass → Periodic table

B.

Molar Mass → Periodic table

$2 \times$ AM of Fe $= 2 \times 55.8$ amu $=$ 111.6 amu

$3 \times$ AM of S $= 3 \times 32.1$ amu $=$ 96.3 amu

$3 \times 4 \times$ AM of O $= 12 \times 16.00$ amu $=$ 192.0 amu

FM of $Fe_2(SO_4)_3$ $=$ 399.9 amu → Formula wight ⇒ 4.00×10^2 amu

"Atomic weight" → Periodic table

بنقرب الرقم لـ (2)
[decimal places]
ممكن يطلع في أعل ←
Roundung ...

Exercise 3.1 Calculate the formula weights of the following compounds, using a table of atomic weights. Give the answers to three significant figures. a. nitrogen dioxide, NO_2; b. glucose, $C_6H_{12}O_6$; c. sodium hydroxide, NaOH; d. magnesium hydroxide, $Mg(OH)_2$.

See Problems 3.27 and 3.28.

Ionic compound

A.) NO_2 = N → (1 × 14.007) +
O → (2 × 15.999)

46.005 = 46.0 (amu)

B.) $C_6H_{12}O_6$ = C → (6 × 12.001) +
H → (12 × 1) +
O → (6 × 15.999)

180 (amu)

C.) NaOH = Na → (1 × 22.989) +
O → (1 × 15.999) +
H → (1 × 1.008)

39.988 → 40.0 (amu)

D.) $Mg(OH)_2$ = Mg → (1 × 24.3) +
O → (2 × 15.999) +
H → (2 × 1)

58.298 → 58.3 (amu)

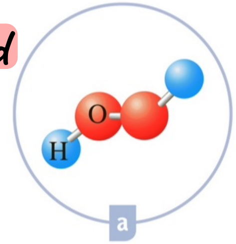
Example 3.2 Calculating the Formula Weight from Molecular Models

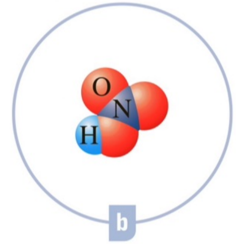
Gaining Mastery Toolbox

Critical Concept 3.2
A molecular model depicts the number and type of each atom present in the molecular compound. Molecular models can be used to determine the chemical formula of a molecular compound.

For the following two compounds, write the molecular formula and calculate the formula weight to four significant figures:

Ionic compound





(continued)

A. H_2O_2 = H → $(2 \times 1{,}008)$
O → (2×15.999) +

34.014 (amu)

Molecular formula

hydrogen peroxide

B. HNO_3 = H → (1×1.008) +
N → (1×14.007) +
O → (3×15.999)

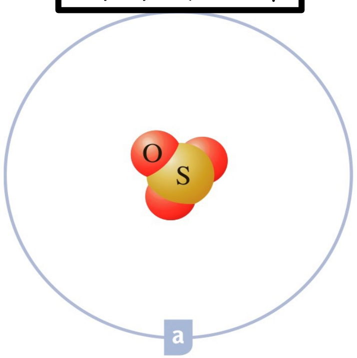
63.012 (amu)

Molecular formula

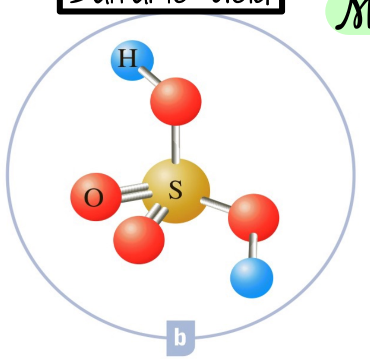
nitric acid

Exercise 3.2 For the following compounds, write the molecular formula and calculate the formula weight to three significant figures.

Sulfur trioxide



Sulfuric acid



A. SO_3 = S → (1×32.07)
O → (3×15.999) +

80.067 = 80.1 (amu)

Molecular formula

B. H_2SO_4 = H → (2×1) +
S → (1×32.07)
O → (4×15.999)

98.1 (amu)

3.2 The Mole Concept

The Mole (mol): A unit to count numbers of particles

A mole (symbol mol) is defined as *the quantity of a given substance that contains as many molecules or formula units as the number of atoms in exactly 12 g of carbon-12*

1 mole of Na2CO3 contains 6.02* 10 (23) C $\boxed{\text{Ions}}$

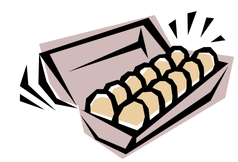
Pair = 2

$$1 \text{ mol} = N_A = 6.0221415 \times 10^{23} \quad = \text{Avogadro's number}$$

Sodium carbonate

$\boxed{1 \text{ mole}}$ of Na_2CO_3 contains 6.02×10^{23} Na_2CO_3 $\boxed{\text{units}}$

Dozen = 12



1 mole of Na_2CO_3 contains $\boxed{2x}$ 6.02×10^{23} Na^+ $\boxed{\text{ions}}$

1 mole of Na_2CO_3 contains 6.02×10^{23} CO_3^{2-} $\boxed{\text{ions}}$

molar mass: of a substance is *the mass of one mole of the substance.* ((1 mole of carbon = 12g/c))

C has a molar mass of exactly 12 g/mol, = Atomic Mass (amu)

$\boxed{C_2H_5OH}$ has a molar mass of exactly 46.1 g/mol Ethanol

And has a molecular weight of exactly 46.1 amu

For any element

atomic mass (amu) = molar mass (grams)

I Think You're OVERREACTING ...

➢ Mole Calculations

3.4 Example

(Q) A chemist determines from the amounts of elements that 0.0654 mol ZnI_2 can form. How many grams of zinc iodide is this? molar mass of ZnI_2 is 319 g/mol

Number of moles (mol) = $\frac{\text{Mass(g)}}{\text{Molar Mass}}$

Number of moles = mass(g) / molar mass

$0.0654 \text{ (mol)} = \frac{\text{Mass (g)}}{319 \text{ (g/mol)}} \Rightarrow 20.9 \text{ g}$

(Q) In a preparation rxn., 45.6 g of lead(II) chromate is obtained as a precipitate. How many moles of $PbCrO_4$ is this? molar mass of $PbCrO_4$ = 323 g/mol

Example 3.5

Number of mole (mol) = $\frac{45.6 \text{ (g)}}{323 \text{ (g/mol)}} \Rightarrow .1412 \text{ mol}$

Very important

$NA = 6.022 * 10 (23)$

(M)

Mass (g)$_{(A)}$ $\xrightarrow{\times \frac{1}{Mm_{(A)}}}$ Mole (mol)$_{(A)}$ $\xrightarrow{\times N_A}$ Number of atoms

Mole (mol)$_{(A)}$ $\xrightarrow{\times Mm_{(A)}}$ Mass (g)$_{(A)}$

Number of atoms $\xrightarrow{\times \frac{1}{N_A}}$ Mole (mol)$_{(A)}$

(N)

How many atoms?

How many molecule?

Mole (mol)$_{(A)}$ $\xrightarrow{\times \text{Ratio}}$ Mole (mol)$_{(B)}$

(N)

Mole (mol)$_{(B)}$ $\xrightarrow{\times Mm_{(B)}}$ Mass (g)$_B$

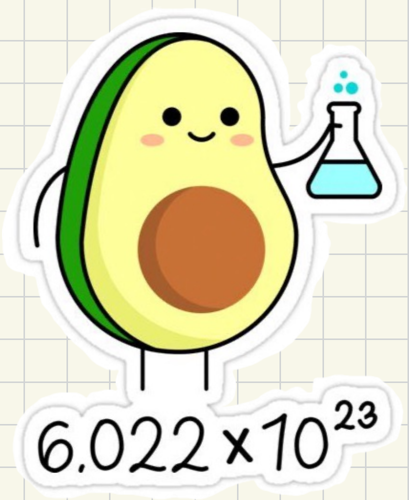
Mass (g)$_B$ $\xrightarrow{\times \frac{1}{Mm_{(B)}}}$ Mole (mol)$_{(B)}$

(M)

Desire quantity

$$\text{Ratio} = \frac{\text{number of mole B}}{\text{number of mole A}}$$

Given quantity



Exercise 3.4 Hydrogen peroxide, H_2O_2, is a colorless liquid. A concentrated solution of it is used as a source of oxygen for rocket propellant fuels. Dilute aqueous solutions are used as a bleach. Analysis of a solution shows that it contains 0.909 mol H_2O_2 in 1.00 L of solution. What is the mass of hydrogen peroxide in this volume of solution?

See Problems 3.37, 3.38, 3.39, and 3.40.

* H_2O_2 = 0.909 mol in 1.00 L (N =)
* Mass (H_2O_2) → $\boxed{N = \frac{M}{M_m}}$ → 30.9 (g) H_2O_2

$$M_{m\,(H_2O_2)} = H \rightarrow (2 \times 1,008) + O \rightarrow (2 \times 15.99)$$

M_m = 33.998 g/mol

Exercise 3.5 Nitric acid, HNO_3, is a colorless, corrosive liquid used in the manufacture of nitrogen fertilizers and explosives. In an experiment to develop new explosives for mining operations, a sample containing 28.5 g of nitric acid was poured into a beaker. How many moles of HNO_3 are there in this sample of nitric acid?

See Problems 3.41 and 3.42.

* HNO_3 → 28.5 (g) (M =)
* How many moles ?? → $\boxed{N = \frac{M}{M_m}}$ →

0.452 (mol)

$$M_{m\,(HNO_3)} = H \rightarrow (1 \times 1.008) + N \rightarrow (1 \times 14.007) + O_3 \rightarrow (3 \times 15.999)$$

M_m = 63.004 g/mol

(Q) How many molecules are there in a 3.46-g sample of hydrogen chloride, HCl?

1. * Molar mass =
$(1 \times 1) + (1 \times 35.453)$
$\boxed{36.453 \text{ g/mol}}$

2. * $N = \frac{M}{Mm} \Rightarrow \frac{3.46 (g)}{36.453 (g/mol)} =$

$.09491674 \rightarrow \boxed{.095 \text{ Mol}}$

3. $\boxed{1 \text{ mol} \rightarrow 6.022 \times 10^{23}}$

Number of Moles = ? .095

$.095 \times 6.022 \times 10^{23} =$

$\boxed{0.572 \times 10^{23} \text{ molecule}}$

Number of moles (mol) = $\frac{\text{Mass(g)}}{\text{Molar Mass}}$

(Q) How many S atoms are there in 16.3 g of S?

1. Molar mass = 32.065 g/mol

2. * $N = \frac{M}{Mm} \Rightarrow \frac{16.3 (g)}{32.065 (g/mol)} =$

$\boxed{.058 \text{ mol}}$

3. $\boxed{1 \text{ mol} \rightarrow 6.022 \times 10^{23}}$

Number of Moles = ?

$0.508 \times 6.022 \times 10^{23} =$

$\boxed{3.061 \times 10^{23} \text{ atom}}$

Exercise 3.6 Hydrogen cyanide, HCN, is a volatile, colorless liquid with the odor of certain fruit pits (such as peach and cherry pits). The compound is highly poisonous. How many molecules are there in 56 mg HCN, the average toxic dose?

See Problems 3.45, 3.46, 3.47, and 3.48.

* (HCN) How many Molecule ? → 56 mg HCN

10^{-3}

$$\text{Mass (g)} \times \frac{1\ \text{mol}}{\text{Mm (HCN) g}} \times NA =$$

$$10^{-3} \times 56\ g \times \frac{1}{27.018\ g} \times 6.022 \times 10^{23} = \boxed{1.248 \times 10^{21}\ \text{atom}}$$

Mm (HCN) = H → (1 x 1)

C → (1 x 12.011) +

N → (1 x 14.007) +

27.018 g/mol

Example 3.3 Calculating the Mass of an Atom or Molecule

a. What is the mass in grams of one chlorine atom, Cl?

b. What is the mass in grams of one HCl molecule?

A.) $\boxed{N = \frac{M}{M_m}}$ $= 1 \text{ atom (Cl)} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atom}} \times \underbrace{(35.45)\ g/mol}_{\text{Mm CL}} = 5.9 \times 10^{-23} (g)$

B.) $\boxed{N = \frac{M}{M_m}}$ $= 1 \text{ molecule} \times \frac{1 \text{ mol}}{6.022 \times 10^{23}} \times \underbrace{(36.45)\ g/mol}_{\text{Mm (HCL)}} = 6.052 \times 10^{-23} (g)$

Mm (HCL) $= (1 \times 1) + (1 \times 35.45) = 36.45$ g/mol

Exercise 3.3 a. What is the mass in grams of a calcium atom, Ca? b. What is the mass in grams of an ethanol molecule, C_2H_5OH?

See Problems 3.33, 3.34, 3.35, and 3.36.

A. $Ca = 1 \text{ atom } Ca \times \dfrac{1}{\underbrace{6.022 \times 10^{23}}_{NA}} \times \underbrace{40.08}_{Mm\ el} = \boxed{6.66 \times 10^{-23} (g)}$

B. $C_2H_5OH = 1 \text{ molecule} \times \dfrac{1}{\underbrace{6.022 \times 10^{23}}_{NA}} \times \overbrace{(46.021)}^{Mm\ C_2H_5OH} = \boxed{7.65 \times 10^{-23} (g)}$

Mm (C_2H_5OH) →

$C = (2 \times 12.011)$

$H = (6 \times 1)$ +

$O = (1 \times 15.999)$ +

46.021 (g/moL)

(Q) How much, in grams, do 8.85×10^{24} atoms of zinc weigh?

A. 3.49×10^{49} g

B. 961 g

C. 4.45 g

D. 5.33×10^{47} g

E. 1.47 g

$$8.85 \times 10^{24}\,\cancel{\text{atoms}} \times \left(\frac{1\,\cancel{\text{mol}}}{6.022 \times 10^{23}\,\cancel{\text{atoms}}}\right) \times \left(\frac{65.41\text{ g Zn}}{1\,\cancel{\text{mol}}}\right)$$

= 961 g Zn

(Q) How many hydrogen atoms are present in 25.6 g of urea $[(NH_2)_2CO]$? molar mass of urea = 60.06 g/mol. بدي أطلع مادة من مادة *

grams of urea ⟶ moles of urea ⟶ moles of H ⟶ atoms of H

Desire quantity

$$25.6\ \cancel{g\ (NH_2)_2CO} \times \frac{1\ \cancel{mol\ (NH_2)_2CO}}{60.06\ \cancel{g\ (NH_2)_2CO}} \times \boxed{\frac{4\ \cancel{mol\ H}}{1\ \cancel{mol\ (NH_2)_2CO}}} \times \frac{6.022 \times 10^{23}\ H\ atoms}{1\ \cancel{mol\ H}}$$

Given quantity

$$= \boxed{1.03 \times 10^{24}\ H\ atoms}$$

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

A. 2.53 mol Ca

B. 0.432 mol Ca

C. 3.00 mol Ca

D. 7.59 mol Ca

E. 0.843 mol Ca

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

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

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

= 7.59 mol Ca

(Q) A sample of $\boxed{\text{sodium carbonate, } Na_2CO_3}$, is found to contain 10.8 moles of sodium. How many moles of oxygen atoms (O) are present in the sample?

A. 10.8 mol O

B. 7.20 mol O

C. 5.40 mol O

D. 32.4 mol O

E. 16.2 mol O

10.8 moles of Na = ? mol O

2 mol Na $\Leftrightarrow$ 3 mol O

$$10.8\ \cancel{\text{mol Na}} \boxed{\left(\frac{3\ \text{mol O}}{2\ \cancel{\text{mol Na}}}\right)}$$

= 16.2 mol O

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

A. 59.6 g

B. 29.8 g

C. 89.4 g

D. 134 g

E. 52.4 g

$$\boxed{\textbf{mass O} \rightarrow \textbf{mol O} \rightarrow \textbf{mol Fe} \rightarrow \textbf{mass Fe}}$$

$$25.6 \text{ g O} \rightarrow ? \text{ g Fe}$$

$$3 \text{ mol O} \Leftrightarrow 2 \text{ mol Fe}$$

$$25.6\,\cancel{\text{g O}} \times \left(\frac{1\,\cancel{\text{mol O}}}{16.0\,\cancel{\text{g O}}}\right) \times \boxed{\left(\frac{2\,\cancel{\text{mol Fe}}}{3\,\cancel{\text{mol O}}}\right)} \times \left(\frac{55.845\text{ g Fe}}{1\,\cancel{\text{mol Fe}}}\right)$$

= 59.6 g Fe

(Q) Silver is often found in nature as the ore, argentite (Ag_2S). How many grams of pure silver can be obtained from a 836 g rock of argentite?

A. 7.75 g

B. 728 g

C. 364 g

D. 775 g

E. 418 g

$$\boxed{\textbf{mass } Ag_2S \rightarrow \textbf{mol } Ag_2S \rightarrow \textbf{mol Ag} \rightarrow \textbf{mass Ag}}$$

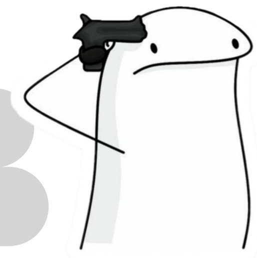
$$836 \text{ g } Ag_2S \rightarrow ? \text{ g Ag}$$

$$1 \text{ mol } Ag_2S \Leftrightarrow 2 \text{ mol Ag}$$

$$836 \text{ g } \cancel{Ag_2S} \times \left(\frac{1 \text{ mol } \cancel{Ag_2S}}{247.8 \text{ g } \cancel{Ag_2S}}\right) \times \boxed{\left(\frac{2 \cancel{\text{mol Ag}}}{1 \text{ mol } \cancel{Ag_2S}}\right)} \times \left(\frac{107.9 \text{ g Ag}}{1 \cancel{\text{mol Ag}}}\right)$$

= 728 g Ag

➢ Percentage Composition 3.3



$$\%\text{ by mass of element} = \frac{\text{mass of element}}{\text{mass of sample}} \times 100\%$$

Example: A sample of a liquid with a mass of 8.657 g was decomposed into its elements and gave 5.217 g of carbon, 0.9620 g of hydrogen, and 2.478 g of oxygen. What is the percentage composition of this compound?

$$\left(\frac{\text{g C}}{\text{g total}}\right) = 100\% = \frac{5.217\text{ g C}}{8.657\text{ g}} \times 100\% = 60.26\%\text{ C}$$

+

$$\left(\frac{\text{g H}}{\text{g total}}\right) = 100\% = \frac{0.9620\text{ g H}}{8.657\text{ g}} \times 100\% = 11.11\%\text{ H}$$

+

$$\left(\frac{\text{g O}}{\text{g total}}\right) = 100\% = \frac{2.478\text{ g O}}{8.657\text{ g}} \times 100\% = 28.62\%\text{ O}$$

Sum of percentages: 100 %

(Q) A sample was analyzed and found to contain 0.1417 g nitrogen and 0.4045 g oxygen.

What is the percentage composition of this compound?

Total sample mass = 0.1417 g + 0.4045 g = 0.5462 g

% Composition of N

$$\left(\frac{\text{g N}}{\text{g total}}\right)\times 100\% = \left(\frac{0.1417\ \text{g N}}{0.5462\ \text{g}}\right)\times 100\% \ \mathbf{= 25.94\%\ N}$$

100 − 25.94% = 74.06%

% Composition of O

$$\left(\frac{\text{g O}}{\text{g total}}\right)\times 100\% = \left(\frac{0.4045\ \text{g O}}{0.5462\ \text{g}}\right)\times 100\% \ \mathbf{= 74.06\%\ O}$$

(Q) a.Calculate the mass percentages of the elements in formaldehyde (CH_2O) molar mass = 30g/mol

ممكن ما يعطيني إياه

Example 3.7

Molar Mass

$$\% \text{ C} = \frac{12.0 \text{ g}}{30.0 \text{ g}} \times 100\% = \mathbf{40.0\%}$$

$$\% \text{ H} = \frac{2 \times 1.01 \text{ g}}{30.0 \text{ g}} \times 100\% = \mathbf{6.73\%}$$

% **O** = 16/30 x100% = 53.3 %

$$\% \text{ O} = 100\% - (40.0\% + 6.73\%) = \mathbf{53.3\%}$$

Example 3.8

b. How many grams of carbon are there in 83.5 g of CH_2O?

CH_2O is 40.0% C, so the mass of carbon in 83.5 g CH_2O is:
83.5 g X 0.400 = **33.4 g**

1g CH_2O = 40% C
83.5 g CH_2O = ??

$$\% \text{ Mass of element (A)} = \frac{\text{mass(A)}}{\text{mass sample}}$$

Percentage composition : *If the masses of the sample and The element are given:

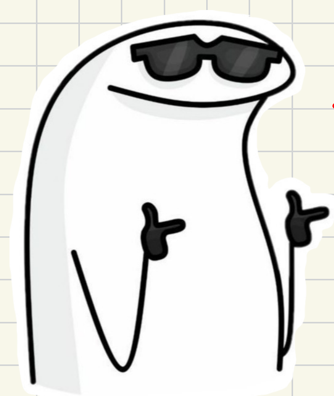
$$\% \text{ Mass of element } (A) = \frac{\text{mass}(A)}{\text{mass sample}} \times \% 100$$

(Number of moles for the element from the formula)

$$\boxed{N = \frac{M}{Mm}} = \frac{Mm\,(A) * N\,(A)}{Mm \text{ sample} * N \text{ sample}} \times \% 100$$

always = 1

Very important



* If the masses of the sample and The element arent given we use the molar mass :

Exercise 3.7 Ammonium nitrate, NH_4NO_3, which is prepared from nitric acid, is used as a nitrogen fertilizer. Calculate the mass percentages of the elements in ammonium nitrate (to three significant figures).

See Problems 3.57, 3.58, 3.59, and 3.60.

* NH_4NO_3 = Mm → N → (2 x 14.007)
H → (4 x 1) +
O → (3 x 15.999) +

80.011 g/mol

Exercise 3.8 How many grams of nitrogen, N, are there in a fertilizer containing 48.5 g of ammonium nitrate and no other nitrogen-containing compound? See Exercise 3.7 for the percentage composition of NH_4NO_3.

$$\text{percentage } N = \frac{\text{Mass (N)}}{\text{Mass (sample)}} =$$

$$.35 \times 48.5 = 17.9 \text{ (N)}$$

* $$N = \frac{Mm(N) \times N(N)}{Mm(NHNO_3)} = \frac{14.007 \times 2}{80.011} = 0.35 \times 100\% = \boxed{35.0\%}$$

* $$H = \frac{Mm(H) \times N(H)}{Mm(NHNO_3)} = \frac{1,008 \times 4}{80.011} = .0499 \times 100\% = \boxed{5.0\%}$$

* $$O = \frac{Mm(O) \times N(O)}{Mm(NHNO_3)} = \frac{3 \times 15.999}{80.011} = .599 \times 100\% \quad \boxed{6.0\%}$$

(Q) Calculate the mass percentages of the elements in H_3PO_4

molar mass = 97.99 g/mol

Phosphoric acid

$$\%H = \frac{3(1.008\ g)\ H}{97.99\ g\ H_3PO_4} \times 100\% = \boxed{3.086\%}$$

$$\%P = \frac{30.97\ g\ P}{97.99\ g\ H_3PO_4} \times 100\% = \boxed{31.61\%}$$

$$\%O = \frac{4(16.00\ g)\ O}{97.99\ g\ H_3PO_4} \times 100\% = \boxed{65.31\%}$$

Example 3.9 Calculating the Percentages of C and H by Combustion

$(C_2H_4O_2) = C_2H_3O_2^{-1}$

Gaining Mastery Toolbox

Critical Concept 3.9

The mass percentages of carbon and hydrogen can be calculated from the masses of CO_2 and H_2O. During the reaction, *all* of the carbon atoms in the sample end up in the CO_2 and *all* of the hydrogen atoms end up in the H_2O. Therefore, the mass of carbon in the CO_2 is equal to the mass of carbon in the sample, and the mass of hydrogen in the H_2O is equal to the mass of hydrogen in the sample. We *cannot* make the statement that the mass of oxygen in the H_2O and CO_2 is equal to the mass of oxygen in the sample because oxygen is both in the sample and is a reactant.

Solution Essentials:

- Elemental analysis
- Percentage composition
- Mass percentage
- Molar mass
- Dimensional analysis
- Rules for significant figures and rounding

Acetic acid contains only C, H, and O. A 4.24-mg sample of acetic acid is completely burned. It gives 6.21 mg of carbon dioxide and 2.54 mg of water. What is the mass percentage of each element in acetic acid?

Problem Strategy Note that in the products of the combustion, all of the carbon from the sample ends up in the CO_2, all of the hydrogen ends up in the H_2O, and the oxygen is in both compounds. Because of this, you should concentrate first on the carbon and hydrogen and worry about the oxygen last. If we can determine the mass of carbon and hydrogen in the original sample, we should then be able to determine the percentage of each of these elements present in the compound. Once we know the mass percentages of carbon and hydrogen in the original compound, the remaining mass percentage (100% total) must be due to oxygen. Let's start by determining the mass of carbon that was originally contained in the compound. You first convert the mass of CO_2 to moles of CO_2. Then you convert this to moles of C, noting that 1 mol C produces 1 mol CO_2. Finally, you convert to mass of C. Similarly, for hydrogen, you convert the mass of H_2O to mol H_2O, then to mol H, and finally to mass of H. (Remember that 1 mol H_2O produces 2 mol H.) Once you have the masses of C and H, you can calculate the mass percentages. Subtract from 100% to get % O.

Solution Following is the calculation of grams C:

$$6.21 \times 10^{-3}\ \cancel{g\ CO_2} \times \frac{1\ \cancel{mol\ CO_2}}{44.0\ \cancel{g\ CO_2}} \times \frac{1\ \cancel{mol\ C}}{1\ \cancel{mol\ CO_2}} \times \frac{12.0\ g\ C}{1\ \cancel{mol\ C}}$$

$$= 1.69 \times 10^{-3}\ g\ C\ (\text{or } 1.69\ mg\ C)$$

For hydrogen, you note that 1 mol H_2O yields 2 mol H, so you write

$$2.54 \times 10^{-3}\ \cancel{g\ H_2O} \times \frac{1\ \cancel{mol\ H_2O}}{18.0\ \cancel{g\ H_2O}} \times \frac{2\ \cancel{mol\ H}}{1\ \cancel{mol\ H_2O}} \times \frac{1.01\ g\ H}{1\ \cancel{mol\ H}}$$

$$= 2.85 \times 10^{-4}\ g\ H\ (\text{or } 0.285\ mg\ H)$$

You can now calculate the mass percentages of C and H in acetic acid.

$$\text{Mass\% C} = \frac{1.69\ mg}{4.24\ mg} \times 100\% = 39.9\%$$

$$\text{Mass\% H} = \frac{0.285\ mg}{4.24\ mg} \times 100\% = 6.72\%$$

You find the mass percentage of oxygen by subtracting the sum of these percentages from 100%:

$$\text{Mass\% O} = 100\% - (39.9\% + 6.72\%) = 53.4\%$$

$C_3H_4O_2 \longleftarrow$ ✓ $C_6H_8O_6 \rightarrow C_6H_7O_6^{-1}$??

Exercise 3.9 A 3.87-mg sample of ascorbic acid (vitamin C) gives 5.80 mg CO_2 and 1.58 mg H_2O on combustion. What is the percentage composition of this compound (the mass percentage of each element)? Ascorbic acid contains only C, H, and O.

See Problems 3.63 and 3.64.

I CHEM

* Sample = 3.87 mg
* CO_2 = 5.80 mg
* H_2O = 1.58 mg

* $CO_2 = (1 \times 12.011) + (2 \times 15.999) = 44.009$ g/mol

* $H_2O = (2 \times 1) + (1 \times 15.999) = 17.999$ g/mol

* Sample − (C + H) = (3.87 − 1.582 − .1755) = 2.1125 mg ⇒ Mass(O)

→ CO_2 → Mass (C) g

$$\text{Mass } CO_2 \; 5.80\,\cancel{mg} \times 10^{-3} \times \frac{1\ \cancel{mol\ (CO_2)}}{Mm\ 44.009\,\cancel{(g)}} \times \frac{1\ \cancel{mol}\ C}{1\ \cancel{mol}\ (CO_2)} \times \frac{12.001\ g}{1\ \cancel{mol}\ C}$$

⟹ 1.582 (mg) : C

⟹ H_2O → Mass (H) g

$$\text{Mass } H_2O \; 1.58\,\cancel{mg} \times \frac{1\ mol\ \cancel{H_2O}}{17.999\ \cancel{(g)}} \times \frac{2 \times 1\ \cancel{mol}\ (H)}{1\ mol\ \cancel{H_2O}} \times \frac{1\ (g)}{1\ \cancel{mol}\ H}$$

→ 0.1755 (mg) : H

Mass of element

$$* \quad C\% = \frac{1.582\ mg}{3.87\ mg} = .4087 \times 100\%$$

Mass of the Sample

$$\boxed{40.9}$$

$$* \quad H\% = \frac{1.755\ mg}{3.87\ mg} = .453 \times 100\%$$

$$\boxed{45.4}$$

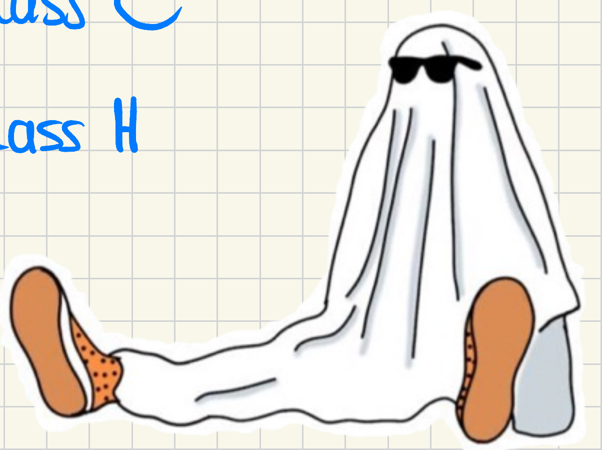
$$* \quad O\% = \frac{2.1125\ mg}{3.87\ mg} = .545 \times 100\%$$

$$\boxed{54.5}$$

* Note:

Mass CO_2 → Mol CO_2 → Mol C → Mass C

Mass H_2O → Mol H_2O → Mol H → Mass H



➢ Determining Empirical and Molecular Formulas 3.5

Empirical Formula :

– Simplest ratio of atoms of each element in compound

– Obtained from experimental analysis of compound

Molecular Formula :

– Exact composition of one molecule

– Exact whole number ratio of atoms of each element in molecule

glucose

Molecular formula : $\boxed{C_6H_{12}O_6}$

Empirical formula : $\boxed{CH_2O}$

1 : 2 : 1

➢ Three Ways to Calculate Empirical Formulas

1. **From Masses of Elements**

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

2. **From Percentage Composition**

   e.g., 43.64% P and 56.36% O

3. **From Combustion Data**

   - Given masses of combustion products

   e.g., 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

1. Empirical Formula from $\boxed{\text{Mass Data}}$

When a 0.1156 g sample of a compound was analyzed, it was found to contain $\boxed{0.04470\ \text{g}}$ of C, $\boxed{0.01875\ \text{g}}$ of H, and $\boxed{0.05215\ \text{g}}$ of N. Calculate the empirical formula of this compound.

Step 1: Calculate moles of each substance

$$\boxed{N = \frac{M}{Mm}}$$

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

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

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

Step 2: Select the smallest number of moles

- Smallest is 3.722×10^{-3} mole

	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

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

Empirical formula = CH_5N

Example 3.10

A compound of nitrogen and oxygen is analyzed, and a sample weighing 1.587 g is found to contain 0.483 g N and 1.104 g O. What is the empirical formula of the compound?

$$N = 0.483\,(g)\,N \times \frac{1}{14.007} = \frac{.03448\ mol}{.03448\ mol} = 1$$

$$O = 1.104\,(g)\,O \times \frac{1}{15.999} = \frac{.069004\ mol}{0.3448\ mol} = 2$$

$\Rightarrow NO_2$

Exercise 3.10 A sample of compound weighing 83.5 g contains 33.4 g of sulfur. The rest is oxygen. What is the empirical formula?

See Problems 3.65 and 3.66.

Sample = S + O

83.5 (g) = 33.4 (g) + O

O = 50.1 (g)

$$S = 33.4\,(g) \times \frac{1}{32.07} = \frac{1.0414\ mol}{1.0414\ mol} = 1$$

$$O = 50.1\,(g) \times \frac{1}{15.999} = \frac{3.1314\ mol}{1.0414\ mol} = 3$$

$\Rightarrow SO_3$

Smallest

2. Empirical Formula from $\boxed{\text{Percentage Composition}}$

Calculate the empirical formula of a compound whose percentage composition data is $\boxed{43.64\%\ P}$ and $\boxed{56.36\%\ O}$. If the molar mass is determined to be $\boxed{283.9\ g/mol}$, what is the molecular formula?

??
بدال ما أضرب بـ
ال molar Mass تبعت ال Compound

Mass of the sample

Step 1: Assume 100 g of compound

Periodic table ⇐ بجيب
الـ Mass إلـها = 100g أصلاً بجيبها ...

- 43.64 g P
- 56.36 g O

1 mol P = 30.97 g

1 mol O = 16.00 g

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

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

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 to get $\boxed{\text{integers}}$

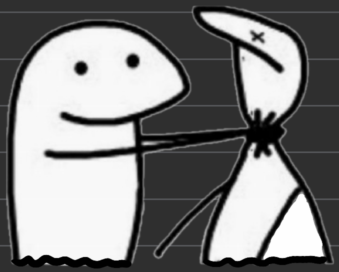
$1.000 \times 2 = 2$

$2.500 \times 2 = 5$

Empirical formula = P_2O_5

.

"Molecular formula" :



Mm P

$$P = 2 \times 30.9 = 61.8\ g/mol$$

Mm O

$$O = 5 \times 15.999 = 79.995\ g/mol$$

$+$

$$141.795\ g/mol$$ ← Molar Mass For the Empirical Formula.

Sample

$$n = \frac{283.9\ g/mol}{141.795\ g/mol} = 2$$

$$P_2O_5 \times 2 = \boxed{P_4O_{10}}$$

Empirecal formula ⟶ molecular formula

(Q) Ascorbic acid (vitamin C) is composed of $\boxed{40.92 \text{ percent}}$ 100% carbon (C), $\boxed{4.58 \text{ percent}}$ hydrogen (H), and $\boxed{54.50 \text{ percent}}$ oxygen (O) by mass. Determine its empirical formula.

Assume you have 100 g. 1.

2.

$$n_C = 40.92 \ \cancel{g\,C} \times \frac{1 \text{ mol C}}{12.01 \ \cancel{g\,C}} = 3.407 \text{ mol C}$$

$$n_H = 4.58 \ \cancel{g\,H} \times \frac{1 \text{ mol H}}{1.008 \ \cancel{g\,H}} = 4.54 \text{ mol H}$$

→formula $C_{3.407}H_{4.54}O_{3.406}$

$$n_O = 54.50 \ \cancel{g\,O} \times \frac{1 \text{ mol O}}{16.00 \ \cancel{g\,O}} = 3.406 \text{ mol O}$$

3.

$$C: \frac{3.407}{3.406} \approx 1 \qquad H: \frac{4.54}{3.406} = 1.33 \qquad O: \frac{3.406}{3.406} = 1$$

→formula $C_1H_{1.33}O_1$ X 3 → formula $\boxed{C_3H_4O_3}$

(($Na_2Cr_2O_7$))
+1 −2

EXAMPLE 3.11

Chromium forms compounds of various colors. (The word *chromium* comes from the Greek *khroma,* meaning "color"; see Figure 3.9.) Sodium dichromate is the most important commercial chromium compound, from which many other chromium compounds are manufactured. It is a bright orange, crystalline substance. An analysis of sodium dichromate gives the following mass percentages: 17.5% Na, 39.7% Cr, and 42.8% O. What is the empirical formula of this compound? (Sodium dichromate is ionic, so it has no molecular formula.)

$$Na = 17.5\% \times \underbrace{100}_{\text{Mass Sample}} = 17.5g \times \frac{1}{22.9} = \frac{.761 \text{ mol Na}}{.761 \text{ mol}} = 1 \times 2 \Rightarrow 2$$

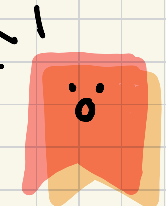
$$Cr = 39.7\% \times 100 = 39.7 g \times \frac{1}{51.9} = \frac{.763 \text{ mol Cr}}{.761 \text{ mol}} = 1 \times 2 \Rightarrow 2$$

$$O = 42.8\% \times 100 = 42.8g \times \frac{1}{15.999} = \frac{2.68 \text{ mol O}}{.761 \text{ mol}} = 3.5 \times 2 \Rightarrow 7$$

$\Rightarrow Na_2Cr_2O_7$

Exercise 3.11 Benzoic acid is a white, crystalline powder used as a food preservative. The compound contains 68.8% C, 5.0% H, and 26.2% O, by mass. What is its empirical formula?

See Problems 3.67, 3.68, 3.69, and 3.70.



Mass Sample

$$C = 68.8\% \times 100 = 68.8 \times \frac{1}{12.011} = \frac{5.728 \text{ mol}}{1.637} = 3.49 \times 2 = 6.99 = 7$$

$$H = 5.0\% \times 100 = 5.0 \times \frac{1}{1} = \frac{5.0 \text{ mol}}{1.637} = 3 \times 2 = 6$$

$$O = 26.2\% \times 100 = 26.2 \times \frac{1}{15.999} = \frac{1.637 \text{ mol}}{1.637} = 1 \times 2 = 2$$

$\Rightarrow C_7H_6O_2$

3. Empirical Formulas from Indirect Analysis: (Combustion analysis)

Combustion Analysis

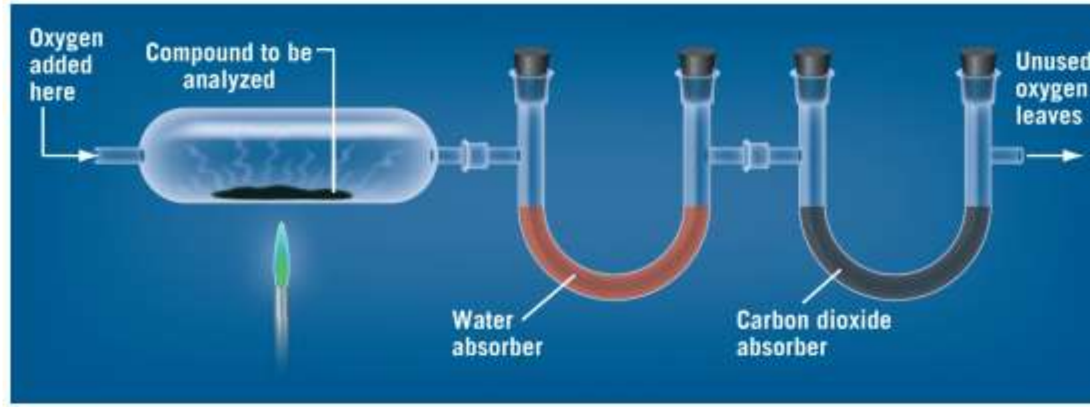
Compounds containing carbon, hydrogen, and oxygen, can be burned completely in pure oxygen gas.

Only carbon dioxide and water are produced

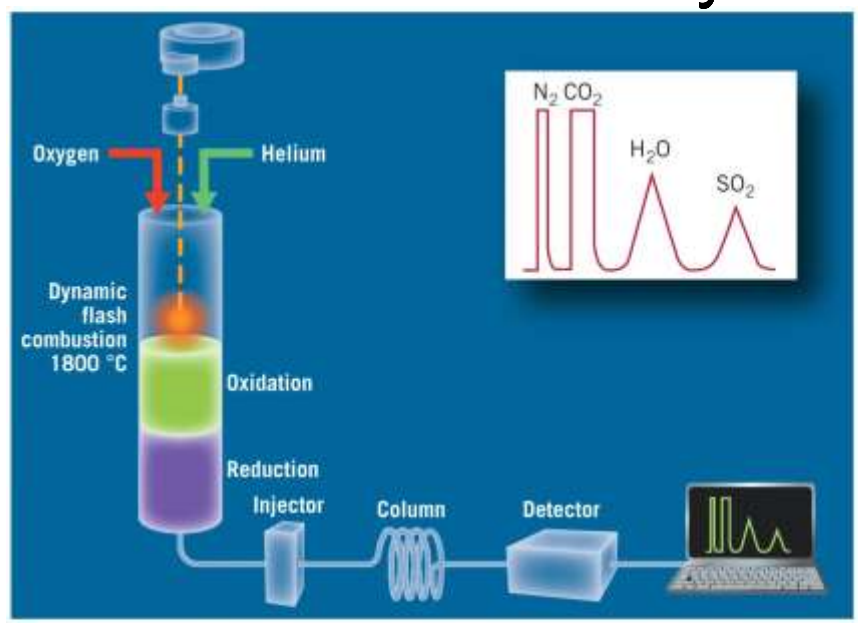
e.g., Combustion of methanol (CH_3OH)

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

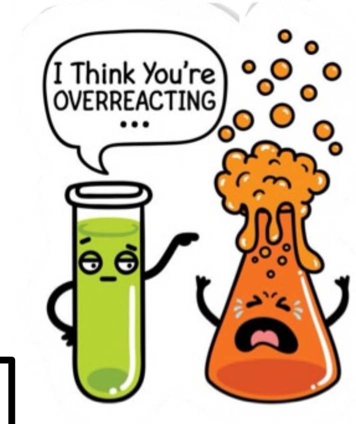
Classic



Modern CHN analysis



- Carbon dioxide and water are separated and weighed separately
  - All C ends up as CO_2
  - All H ends up as H_2O
  - بدي أطلع مادة من مادة
  - Mass of C can be derived from amount of CO_2
    - $\boxed{\text{mass } CO_2 \rightarrow \text{mol } CO_2 \rightarrow \text{mol } C \rightarrow \text{mass } C}$
  - Mass of H can be derived from amount of H_2O
    - $\boxed{\text{mass } H_2O \rightarrow \text{mol } H_2O \rightarrow \text{mol } H \rightarrow \text{mass } H}$
  - **Mass of oxygen is obtained by difference**
    - $\boxed{\textbf{mass O = mass sample – (mass C + mass H)}}$



(Q) 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.

1. Calculate mass of C from mass of CO_2.

mass CO_2 → mole CO_2 → mole C → mass C

$$7.406\,\text{g}\,CO_2\left(\frac{1\ \text{mol}\,CO_2}{44.01\,\text{g}\,CO_2}\right)\left(\frac{1\,\text{mol}\,C}{1\,\text{mol}\,CO_2}\right)\left(\frac{12.011\,\text{g}\,C}{1\ \text{mol}\,C}\right) \quad \mathbf{= 2.021\ g\ C}$$

2. Calculate mass of H from mass of H_2O.

mass H_2O → mol H_2O → mol H → mass H

$$4.512\,\text{g}\,H_2O\left(\frac{1\,\text{mol}\,H_2O}{18.015\,\text{g}\,H_2O}\right)\left(\frac{2\,\text{mol}\,H}{1\,\text{mol}\,H_2O}\right)\left(\frac{1.008\,\text{g}\,H}{1\,\text{mol}\,H}\right) \quad \mathbf{= 0.5049\ g\ H}$$

3. Calculate mass of O from difference.

$$5.217\ \text{g sample} - (2.021\ \text{g C} + 0.5049\ \text{g H}) \quad \mathbf{= 2.691\ g\ O}$$

4. Calculate mol of each element

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

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

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

- Preliminary empirical formula

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

5. Calculate mol ratio of each element

$$C_{\frac{0.1683}{0.1682}}H_{\frac{0.5009}{0.1682}}O_{\frac{0.1682}{0.1682}} = C_{1.00}H_{2.97}O_{1.00}$$

Empirical Formula = CH_3O

(Q) The combustion of a $\boxed{\text{13.660 g sample}}$ of a compound of C, H, and S in pure oxygen gave $\boxed{19.352\text{ g }CO_2}$ and $\boxed{11.882\text{ g of }H_2O}$. Calculate the empirical formula of the compound.

A. $C_4H_{12}S$
B. CH_3S
C. C_2H_6S
D. $C_2H_6S_3$
E. CH_3S_2

(1) mass CO_2 → mole CO_2 → mole C → mass C
(2) mass H_2O → mole H_2O → mole H → mass H
(3) Calculate mass of S from difference

$$19.352\text{ g }CO_2\left(\frac{1\text{ mol }CO_2}{44.01\text{ g }CO_2}\right)\left(\frac{1\text{ mol C}}{1\text{ mol }CO_2}\right)\left(\frac{12.011\text{ g C}}{1\text{ mol C}}\right) = 5.281\text{ g C}$$

$$11.882\text{ g }H_2O\left(\frac{1\text{ mol }H_2O}{18.015\text{ g }H_2O}\right)\left(\frac{2\text{ mol H}}{1\text{ mol }H_2O}\right)\left(\frac{1.008\text{ g H}}{1\text{ mol H}}\right) = 1.330\text{ g H}$$

(4)
13.66 g sample – 5.281 g C – 1.330 g H = 7.049 g S

5.

$$\text{mol C} = \frac{\text{g C}}{\text{MM C}} = \frac{5.281\ \cancel{\text{g}}}{12.011\ \cancel{\text{g}}/\text{mol}} = 0.4497 \text{ mol C}$$

$$\text{mol H} = \frac{\text{g H}}{\text{MM H}} = \frac{1.330\ \cancel{\text{g}}}{1.008\ \cancel{\text{g}}/\text{mol}} = 1.319 \text{ mol H}$$

$$\text{mol S} = \frac{\text{g S}}{\text{MM S}} = \frac{7.049\ \cancel{\text{g}}}{32.065\ \cancel{\text{g}}/\text{mol}} = 0.2198 \text{ mol S}$$

- Preliminary empirical formula
  - $C_{0.4497}H_{1.319}S_{0.2198}$

$$C_{\frac{0.4497}{0.2198}}H_{\frac{1.319}{0.2198}}O_{\frac{0.2198}{0.2198}} = C_{2.03}H_{6.00}S_{1.00}$$

Empirical Formula = C_2H_6S

➢ Determining Molecular Formula from empirical formula

-In some cases molecular and empirical formulas are the same

-When they are different, the subscripts of molecular formula are integer multiples of those in empirical formula

– If empirical formula is A_xB_y

– Molecular formula will be $A_{(n \times x)}B_{(n \times y)}$

(Q)The empirical formula of hydrazine is NH_2, and its molecular mass is 32.0. What is its molecular formula?

Atomic masses: N = 14.007; H = 1.008

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

n = (32.0/16.02) = 2

$(NH_2) \times 2 = N_2H_4$

In Example 3.9, we found the percentage composition of acetic acid to be 39.9% C, 6.72% H, and 53.4% O. Determine the empirical formula. The molecular weight of acetic acid was determined by experiment to be 60.0 amu. What is its molecular formula?

Example 3.12

Sample mass

* $C = 39.9\% \times 100g = 39.9_{(g)} \times \frac{1}{12.011} = \frac{3.32 \text{ mol}}{3.32 \text{ mol}} = 1$
* $H = 6.72\% \times 100g = 6.72_{(g)} \times \frac{1}{1.008} = \frac{6.72 \text{ mol}}{3.32 \text{ mol}} = 2 \rightarrow CH_2O$
* $O = 53.4\% \times 100g = 53.4_{(g)} \times \frac{1}{15.999} = \frac{3.33 \text{ mol}}{3.32 \text{ mol}} = 1 \Rightarrow C_2H_4O_2$

Molecular Formula =

$(1 \times 12.011) + (2 \times 1) + (1 \times 15.999) = 30.01 \Rightarrow$ * $n = \frac{60.0 \text{ g/mol}}{30.0 \text{ g/mol}} = 2$

Exercise 3.12 The percentage composition of acetaldehyde is 54.5% C, 9.2% H, and 36.3% O, and its molecular weight is 44 amu. Obtain the molecular formula of acetaldehyde.

See Problems 3.73, 3.74, 3.75, and 3.76.

Sample mass

* $H = 9.2\% \times 100 = 9.2 \times \frac{1}{1} = \frac{9.2 \text{ mol}}{2.268} = 4$
* $O = 36.3\%\ 100 = 36.3 \times \frac{1}{15.999} = 2.268 \text{ mol} = 1$
* $C = 54.5\%\ 100 = 54.5 \times \frac{1}{12.011} = \frac{4.537 \text{ mol}}{2.268} = 2$

C_2H_4O — Empirical and molecular formula

* $n = \frac{44 \text{ g/mol}}{44.012} = 1$

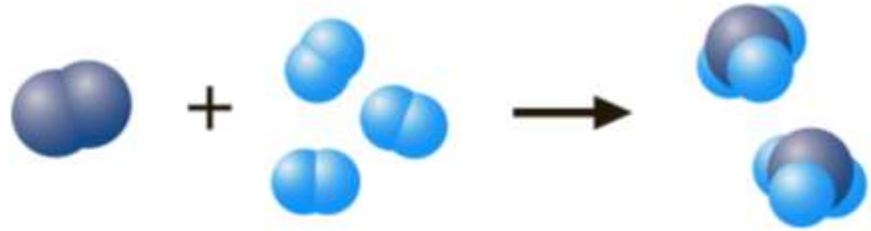
C_2H_4O

Molecular Formula $= (2 \times 12.011) + (4 \times 1) + (1 \times 15.999) = 44.012$

Stoichiometry:Quantitative Relations in Chemical Reactions

Molar Interpretation of a Chemical Equation

Ratio



$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

N_2	+	$3H_2$	$\longrightarrow$	$2NH_3$	
1 molecule N_2	+	3 molecules H_2	$\longrightarrow$	2 molecules NH_3	(molecular interpretation)
1 mol N_2	+	3 mol H_2	$\longrightarrow$	2 mol NH_3	(molar interpretation)
28.0 g N_2	+	3×2.02 g H_2	$\longrightarrow$	2×17.0 g NH_3	(mass interpretation)

$(2 \times 14.007) = 28.0\,g$

$\times$! number of molecule

Exercise 3.13 In an industrial process, $\boxed{\text{hydrogen chloride, HCl,}}$ is prepared by burning hydrogen gas, H_2, in an atmosphere of chlorine, Cl_2. Write the chemical equation for the reaction. Below the equation, give the molecular, molar, and mass interpretations.

See Problems 3.77 and 3.78.

$$H_2 + Cl_2 \rightarrow 2HCl$$ 8 Chemical equation

1 molecules H_2 + 1 molecules Cl_2 → 2 molecules HCl 8 Molecular interpretation

1 mole H_2 + 1 mole Cl_2 → 2 mole HCl 8 Molar interpretation

1 x 2.02 (g) H_2 + 1 x 70.9 (g) Cl_2 → 2 x 36.46 (g) HCl 8 Mass interpretation

3.7 Amounts of Substances in a Chemical Reaction

Example 3.13

mass

g → من (g) لمادة أخرى = لمادة معينة

Relating the Quantity of Reactant to Quantity of Product

In the following reaction:

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

How many grams of Fe(s) can be produced from 1.00 kg Fe_2O_3?

Molar masses are: Fe = 55.8 g/mol and Fe_2O_3 = 160 g/mol

Solution The calculation is as follows:

$$\boxed{1.00 \times 10^3\ \cancel{\text{g Fe}_2\text{O}_3} \times \frac{1\ \cancel{\text{mol Fe}_2\text{O}_3}}{160\ \cancel{\text{g Fe}_2\text{O}_3}} \times \frac{2\ \cancel{\text{mol Fe}}}{1\ \cancel{\text{mol Fe}_2\text{O}_3}} \times \frac{55.8\ \text{g Fe}}{1\ \cancel{\text{mol Fe}}} = \mathbf{698\ g\ Fe}}$$

* The equation must be balanced *

Exercise 3.14 Sodium is a soft, reactive metal that instantly reacts with water to give hydrogen gas and a solution of sodium hydroxide, NaOH. How many grams of sodium metal are needed to give 7.81 g of hydrogen by this reaction? (Remember to write the balanced equation first.)

See Problems 3.83, 3.84, 3.85, and 3.86.

$$2Na + 2H_2O \longrightarrow H_2 + 2NaOH$$ ← Balanced equation

Mass Na = ?

$$7.81\,\cancel{g}\,(\cancel{H}) \times \frac{1\ \text{mol}\ \cancel{H}}{\cancel{2} \times (1.01)\,\cancel{(g)H}} \times \frac{\cancel{2}\ \text{mol}\ \cancel{Na}}{1\ \text{mol}\ \cancel{H}} \times \frac{22.9\ (g)}{1\ \text{mol}\ \cancel{Na}}$$

$$7.81 \times 22.9 = 178\ (g)\ Na$$

Example 3.14

Relating the Quantities of Two Reactants (or Two Products)

In the following reaction:

$4HCl(aq) + MnO_2(s) \rightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$

How many grams of HCl react with 5.00 g of MnO_2, according to this equation?

$$5.00\ \text{g}\ \cancel{\text{MnO}_2} \times \frac{1\ \cancel{\text{mol MnO}_2}}{86.9\ \cancel{\text{g MnO}_2}} \times \frac{4\ \cancel{\text{mol HCl}}}{1\ \cancel{\text{mol MnO}_2}} \times \frac{36.5\ \text{g HCl}}{1\ \cancel{\text{mol HCl}}} = \mathbf{8.40\ g\ HCl}$$

Exercise 3.16 oxygen can be prepared by heating mercury(II) oxide, HgO. Mercury metal is the other product. If 6.47 g of oxygen is collected, how many grams of mercury metal are also produced? $2HgO \rightarrow 2Hg + O_2$

3.SF

$$6.47\ \text{g O}_2 \times \frac{1\ \text{mol O}_2}{32.00\ \text{g O}_2} \times \frac{2\ \text{mol Hg}}{1\ \text{mol O}_2} \times \frac{200.59\ \text{g Hg}}{1\ \text{mol Hg}} = 81.\underline{1}1 = 81.1\ \text{g Hg}$$

Exercise 3.15 Sphalerite is a $\boxed{\text{zinc sulfide (ZnS)}}$ mineral and an important commercial source of zinc metal. The first step in the processing of the ore consists of heating the sulfide with oxygen to give zinc oxide, ZnO, and sulfur dioxide, SO_2. How many kilograms of oxygen gas combine with 5.00×10^3 g of zinc sulfide in this reaction? (You must first write the balanced chemical equation.)

See Problems 3.87 and 3.88.

number of moles

$$2\,ZnS + 3\,O_2 \longrightarrow 2\,ZnO + 2\,SO_2 \rightarrow \text{Balanced equation}$$

مابدخل بحساب ال (Mm)

Mm

Mass O_2 = ?

$$5.00 \times 10^3 (g)\ ZnS \times \frac{1\ \cancel{\text{mole ZnS}}}{97.47\ \cancel{(g)}} \times \frac{3\ \text{moles } O_2}{2\ \text{moles ZnS}} \times 31.998\ O_2$$

$2.46\ (g) \rightarrow O_2$

How many grams of Al_2O_3 are produced when 41.5 g Al react?

$$2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(l)$$

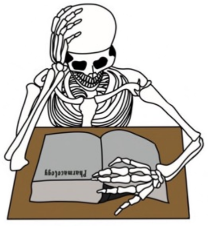
A. 78.4 g

B. 157 g

C. 314 g

D. 22.0 g

E. 11.0 g



$$\boxed{41.5\ \cancel{\text{g Al}}\left(\frac{1\ \cancel{\text{mol Al}}}{26.98\ \cancel{\text{g Al}}}\right)\left(\frac{1\ \cancel{\text{mol Al}_2\text{O}_3}}{2\ \cancel{\text{mol Al}}}\right)\left(\frac{101.96\ \text{g Al}_2\text{O}_3}{1\ \cancel{\text{mol Al}_2\text{O}_3}}\right)}$$

$$= 78.4\ \text{g Al}_2\text{O}_3$$

How many grams of sodium dichromate are required to produce 24.7 g iron(III) chloride from the following reaction?

$$14HCl + Na_2Cr_2O_7 + 6FeCl_2 \rightarrow 2CrCl_3 + 7H_2O + 6FeCl_3 + 2NaCl$$

A. 6.64 g $Na_2Cr_2O_7$

B. 0.152 g $Na_2Cr_2O_7$

C. 8.51 g $Na_2Cr_2O_7$

D. 39.9 g $Na_2Cr_2O_7$

E. 8.04 g $Na_2Cr_2O_7$

$$\boxed{\begin{gathered}24.7\ \cancel{\text{g FeCl}_3} \times \left(\frac{1\ \cancel{\text{mol FeCl}_3}}{162.2\ \cancel{\text{g FeCl}_3}}\right) \times \left(\frac{1\ \cancel{\text{mol Na}_2\text{Cr}_2\text{O}_7}}{6\ \cancel{\text{mol FeCl}_3}}\right) \times \left(\frac{262.0\ \text{g Na}_2\text{Cr}_2\text{O}_7}{1\ \cancel{\text{mol Na}_2\text{Cr}_2\text{O}_7}}\right) \\ = 6.64\ \text{g Na}_2\text{Cr}_2\text{O}_7\end{gathered}}$$

3.8 Limiting Reactant; Theoretical and Percentage Yields

Example 3.15 Calculating with a Limiting Reactant (Involving Moles)

Reactant A = B there is no limiting reactant

Zinc metal reacts with hydrochloric acid by the following reaction:

$1Zn(s) + 2HCl(aq) \rightarrow 1ZnCl_2(aq) + 1H_2(g)$

If 0.30 mol Zn is added to a solution containing 0.52 mol HCl, how many moles of H_2 are produced? Coefficient

Zn → 0.30 mol / 1 → .30

HCL → 0.52 mol / 2 → .26 L

$.52 \text{ mol } \cancel{HCL} \times \frac{1 \text{ mol } H_2}{2 \text{ mol } \cancel{HCl}} = \boxed{.26 \text{ mol } H_2}$

limiting

L

KO_2 : .25 / 4 → $\boxed{.0625 \text{ mol}}$

H_2O : .15 / 2 → .075 mol

محاوير

$.25 \text{ mol } \cancel{KO_2} \times \frac{3 \text{ mol } O_2}{4 \text{ mol } \cancel{KO_2}} = \boxed{.1875 \text{ mol } O_2}$

limiting

3.91 Potassium superoxide, KO_2, is used in rebreathing gas masks to generate oxygen.

$4KO_2(s) + 2H_2O(l) \rightarrow 4KOH(s) + 3O_2(g)$

If a reaction vessel contains 0.25 mol KO_2 and 0.15 mol H_2O, what is the limiting reactant? How many moles of oxygen can be produced?

Coefficient :

Explanation

$$1\,Bread + \underline{\underline{3}}\ falafel \longrightarrow\longrightarrow 1\,sandwich$$

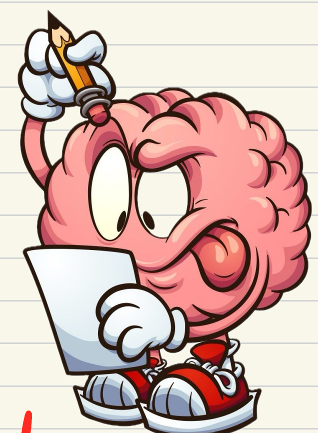
$$\underbrace{Bread + 3\ falafel}_{\text{Reactant}} \qquad \underbrace{sandwich}_{\text{Product}}$$

5 bread
12 falafel } How many sandwiches can be produced?

1. Step 1: $\frac{5}{1} = 5$ → Bread is excess

Coefficient

$\frac{12}{3} = 4$ → Falafel is the limiting



2. Step 2: $12\ \text{Falafel} \times \frac{1\ \text{Sandwich}}{3\ \text{falafil}} = 4\ \text{Sandwich}$

(The maximum mass) Masses NOT Moles ← Theorelical yield دائمًا ب *

Exercise 3.17 Aluminum chloride, $AlCl_3$, is used as a catalyst in various industrial reactions. It is prepared from hydrogen chloride gas and aluminum metal shavings.

See Problems 3.91 and 3.92.

$$2Al(s) + 6HCl(g) \longrightarrow 2AlCl_3(s) + 3H_2(g)$$

Suppose a reaction vessel contains 0.15 mol Al and 0.35 mol HCl. How many moles of $AlCl_3$ can be prepared from this mixture?

$$2\,AL_{(s)} + 6HCL_{(g)} \rightarrow 2\,ALCl_{3\,(s)} + 3\,H_{2\,(g)}$$

$*$ $.15$ mol AL $\rightarrow \div 2 = ,075$ mol AL

$*$ $.35$ mol HCl $\rightarrow \div 6 = ,0583$ mol HCL

limiting

$$\Rightarrow .35 \text{ mol } \cancel{HCl} \times \underbrace{\frac{2 \text{ mol } ALCl_3}{6 \text{ mol } \cancel{HCl}}}_{\text{Ratio}} = .1166 \text{ mol } ALCL_3 = .12 \text{ mol}$$

2-Significant

Theoretical yield

Calculating with a Limiting Reactant (Involving Masses)

~~3.96~~ Hydrogen cyanide, HCN (g), is prepared from ammonia, air, and natural gas (CH_4) by the following process:

$$2NH_3(g) + 3O_2(g) + 2CH_4(g) \rightarrow 2HCN(g) + 6H_2O(g)$$ (l) hydrocyanic acid

If a reaction vessel contains 11.5 g NH_3, 12.0 g O_2, and 10.5 g CH_4, what is the maximum mass in grams of hydrogen cyanide that could be made, assuming the reaction goes to completion as written?

Masses → Moles

Mm

Coofecient from the balanced equation.

* $NH_3 = 11.5\,g \rightarrow \frac{1}{17.007} = .676\ mol \div 2 = .338\ mol$
* $O_2 = 12.0\,g \rightarrow \frac{1}{31.998} = .375\ mol \div 3 = .125\ mol$: limiting
* $CH_4 = 10.5 \rightarrow \frac{1}{16.001} = .656\ mol \div 2 = .328\ mol$

$$.375\ mol\ \cancel{O_2} \times \frac{2\ mol\ HCN}{3\ mol\ \cancel{O_2}} = .25\ mol\ HCN \times Mm\ 27.018 =$$

$$\boxed{6.754\ (g)\ HCN}$$

Theoretical hold

➢ Theoretical yield and percentage yield

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

3.97 Aspirin (acetylsalicylic acid) is prepared by heating salicylic acid, $C_7H_6O_3$, with acetic anhydride, $C_4H_6O_3$. The other product is acetic acid, $C_2H_4O_2$.

$$1C_7H_6O_3 + 1C_4H_6O_3 \rightarrow 1C_9H_8O_4 + 1C_2H_4O_2$$

What is the theoretical yield (in grams) of aspirin, $C_9H_8O_4$, when 2.00 g of salicylic acid is heated with 4.00 g of acetic anhydride? If the actual yield of aspirin is 1.86 g, what is the percentage yield?

3.97 8 * You Should balance the equation *

$$1C_7H_6O_3 + 1C_4H_6O_3 \longrightarrow 1C_9H_8O_4 + 1C_2H_4O_2$$

$C_7H_6O_3 = 2.00g$ $C_4H_6O_3 = 4.00g$?

Actual yield = 1.86g → percentage yield

Limiting Reactant

$$C_7H_6O_3 = 2.00g \times \frac{1}{138.004} = \frac{.01449\ mol}{1\ mol}$$

$$C_4H_6O_3 = 4.00g \times \frac{1}{102.001} = \frac{.0392\ mol}{1\ mol}$$

$$.01449 \times \frac{1\ mol\ C_9H_8O_4}{1\ mol\ C_7H_6O_3} =$$

×(Mm)

$$\#\ C_9H_8O_4 = .01449\ mol \times (180.05)$$

2.608g

(mole → Mass) = Theoretical yield

$$\text{percentage yield} = \frac{\text{actual yield (g)}}{\text{Theoretical yield (g)}} = \frac{1.86\ g}{2.608\ g}$$

→ .713 × 100%

Don't have a Unit → 71.3

(Q) When 6.40 g of CH_3OH was mixed with 10.2 g of O_2 and ignited, 6.12 g of CO_2 was obtained. What was the percentage yield of CO_2?

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

MM(g/mol) (32.04) (32.00) (44.01) (18.02)

A. 6.12%

B. 8.79%

C. 100%

D. 142%

E. 69.6%

$$6.40\text{ g }CH_3OH \times \frac{1\text{ mol }CH_3OH}{32.04\text{ g }CH_3OH} \times \frac{3\text{ mol }O_2}{2\text{ mol }CH_3OH} \times \frac{32.00\text{ g }O_2}{1\text{ mol }O_2}$$

= 9.59 g O_2 needed; CH_3OH limiting

$$6.40\text{ g }CH_3OH \times \frac{1\text{ mol }CH_3OH}{32.04\text{ g }CH_3OH} \times \frac{2\text{ mol }CO_2}{2\text{ mol }CH_3OH} \times \frac{44.01\text{ g }CO_2}{1\text{ mol }CO_2}$$

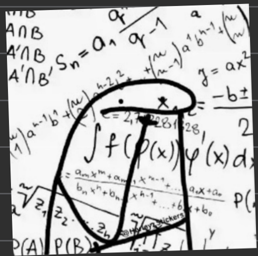
= 8.79 g CO_2 in theory

$$\frac{6.12\text{ g }CO_2\text{ actual}}{8.79\text{ g }CO_2\text{ theory}} \leftrightarrow 100\ \% = 69.6\%$$

Example = 3.16 :

$$2CH_3CHO + 1O_2 \rightarrow 2HC_2H_3O_2$$

CH_3CHO ↓ 20.0 g O_2 ↓ 10.0 g



A. $HC_2H_3O_2$ = ? (g)

B. the excess Reactant = ?

A. $$* \; CH_3CHO = 20.0\cancel{(g)} \times \frac{1\ \cancel{mol}}{44.1\cancel{(g)}} \times \frac{1\ mol\ \cancel{O_2}}{2\ mol\ CH_3\cancel{CHO}} \times \frac{32.(g)}{1\ mol\ \cancel{O_2}}$$

↪ Limiting

$$\Rightarrow 7.2\ (g)\ O_2 < (10.0\ g\ O_2)$$

عندي ١٠ داستهلكت
منه بس (7.2) وضايل
في ...excess

$$* \; 20.0\cancel{(g)} \times \frac{1\ mol\ \cancel{}}{44.1\ \cancel{(g)}} \times \frac{\cancel{2}\ HC_2H_3\cancel{O_2}}{\cancel{2}\ CH_3CH\cancel{O}} \times \frac{60.0\ (g)}{1\ \cancel{mol}}$$

→ 27.21 (g) $HC_2H_3O_2$ was produced.

B. The excess Reactant = O_2

(10 - 7.26) = 2.7 g O_2 (mass Remaining)

Exercise 3.18 In an experiment, 7.36 g of zinc was heated with 6.45 g of sulfur (Figure 3.16). Assume that these substances react according to the equation

$$8Zn + S_8 \longrightarrow 8ZnS$$

What amount of zinc sulfide was produced?

Answer = 11.0 g

Exercise 3.19 New industrial plants for acetic acid react liquid methanol with carbon monoxide in the presence of a catalyst.

$$CH_3OH(l) + CO(g) \longrightarrow HC_2H_3O_2(l)$$

In an experiment, 15.0 g of methanol and 10.0 g of carbon monoxide were placed in a reaction vessel. What is the theoretical yield of acetic acid? If the actual yield is 19.1 g, what is the percentage yield?

See Problems 3.97 and 3.98.

Answer = A. 21.4 g

B. 89.1 %

Learning Objectives	Important Terms
3.1 Molecular Weight and Formula Weight	
▪ Define the terms *molecular weight* and *formula weight* of a substance. ▪ Calculate the formula weight from a formula. Example 3.1 ▪ Calculate the formula weight from molecular models. Example 3.2	**molecular weight** **formula weight**
3.2 The Mole Concept	
▪ Define the quantity called the *mole*. ▪ Learn *Avogadro's number*. ▪ Understand how the *molar mass* is related to the formula weight of a substance. ▪ Calculate the mass of atoms and molecules. Example 3.3 ▪ Perform calculations using the mole. ▪ Convert from moles of substance to grams of substance. Example 3.4 ▪ Convert from grams of substance to moles of substance. Example 3.5 ▪ Calculate the number of molecules in a given mass of substance. Example 3.6	**mole (mol)** **Avogadro's number (N_A)** **molar mass**
3.3 Mass Percentages from the Formula	
▪ Define *mass percentage*. ▪ Calculate the percentage composition of the elements in a compound. Example 3.7 ▪ Calculate the mass of an element in a given mass of compound. Example 3.8	**percentage composition** **mass percentage**
3.4 Elemental Analysis: Percentage of Carbon, Hydrogen, and Oxygen	
▪ Describe how C, H, and O combustion analysis is performed. ▪ Calculate the percentage of C, H, and O from combustion data. Example 3.9	
3.5 Determining Formulas	
▪ Define *empirical formula*. ▪ Determine the empirical formula of a binary compound from the masses of its elements. Example 3.10 ▪ Determine the empirical formula from the percentage composition. Example 3.11 ▪ Understand the relationship between the molecular weight of a substance and its *empirical formula weight*. ▪ Determine the molecular formula from the percentage composition and molecular weight. Example 3.12	**empirical (simplest) formula**
3.6 Molar Interpretation of a Chemical Equation	
▪ Relate the coefficients in a balanced chemical equation to the number of molecules or moles (*molar interpretation*).	**stoichiometry**

3.7 Amounts of Substances in a Chemical Reaction

- Use the coefficients in a chemical reaction to perform calculations.
- Relate the quantities of reactant to the quantity of product. Example 3.13
- Relate the quantities of two reactants or two products. Example 3.14

3.8 Limiting Reactant; Theoretical and Percentage Yields

- Understand how a *limiting reactant* or *limiting reagent* determines the moles of product formed during a chemical reaction and how much *excess reactant* remains.
- Calculate with a limiting reactant involving moles. Example 3.15
- Calculate with a limiting reactant involving masses. Example 3.16
- Define and calculate the *theoretical yield* of chemical reactions.
- Determine the *percentage yield* of a chemical reaction.

limiting reactant (reagent)
theoretical yield
percentage yield

Key Equations

$$\text{Mass\% } A = \frac{\text{mass of } A \text{ in the whole}}{\text{mass of the whole}} \times 100\%$$

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$n = \frac{\text{molecular weight}}{\text{empirical formula weight}}$$

Done by: Joud Jaber

Thank You