Mass Relationships in Chemical Reactions

Chapter 3

Atomic Mass: Average Atomic Mass

- The mass of an atom depends on the number of electrons, protons, and neutrons it contains.

Micro World → Macro World
atoms & molecules → grams

Atomic mass is the mass of an atom in atomic mass units (amu)

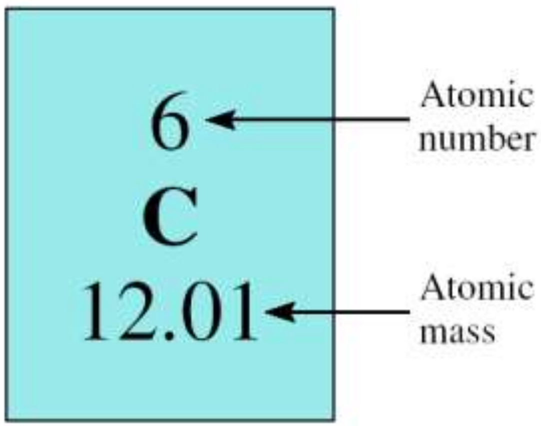
> By definition:
> 1 atom ^{12}C "weighs" 12 amu

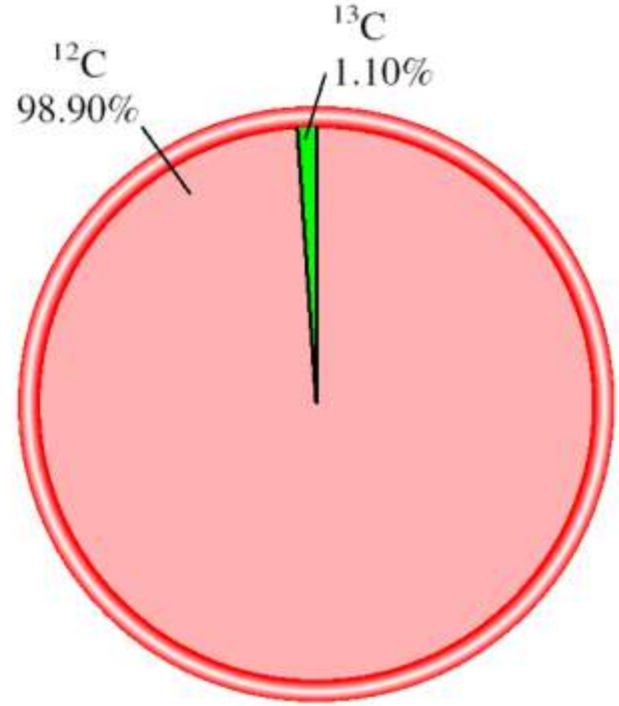
On this scale

^{1}H = 1.008 amu

^{16}O = 16.00 amu

The ***average atomic mass*** is the weighted average of all of the naturally occurring isotopes of the element.





$$\begin{aligned}\text{average atomic mass of natural carbon} &= (0.9890)(12.00000\ \text{amu}) + (0.0110)(13.00335\ \text{amu}) \\ &= 12.01\ \text{amu}\end{aligned}$$

Naturally occurring lithium is:

7.42% ^{6}Li (6.015 amu)

92.58% ^{7}Li (7.016 amu)

Average atomic mass of lithium:

$$\frac{7.42 \text{ x } 6.015 + 92.58 \text{ x } 7.016}{100} = 6.941 \text{ amu}$$

Example 3.1

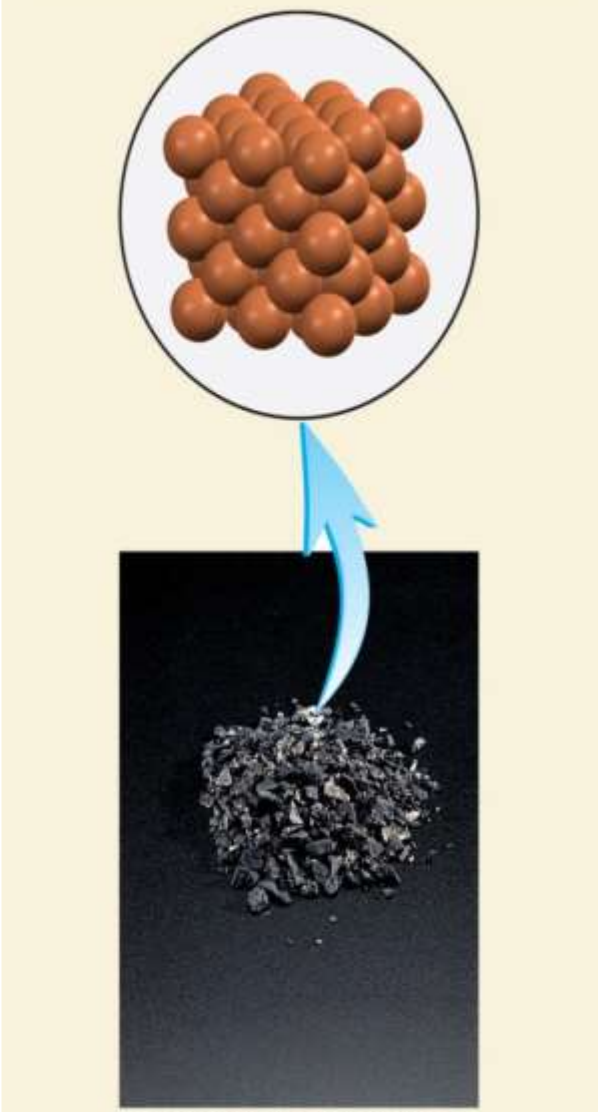
Copper, a metal known since ancient times, is used in electrical cables and pennies, among other things.

The atomic masses of its two stable isotopes,

$^{63}_{29}\text{Cu}\,(69.09\text{ percent})$ and $^{65}_{29}\text{Cu}\,(30.91\text{ percent})$,

are 62.93 amu and 64.9278 amu, respectively.

Calculate the average atomic mass of copper. The relative abundances are given in parentheses.



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Example 3.1

Solution

First the percent's are converted to fractions:

69.09 percent to $69.09/100$ or 0.6909

30.91 percent to $30.91/100$ or 0.3091

We find the contribution to the average atomic mass for each isotope, then add the contributions together to obtain the average atomic mass.

$$(0.6909)\,(62.93\text{ amu}) + (0.3091)\,(64.9278\text{ amu}) = 63.55\text{ "amu"}$$

10 Ne Neon 20.18	
10	Atomic number
20.18	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 Hydrogen 1.008																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.88	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.59	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3
55 Cs Cesium 132.9	56 Ba Barium 137.3	57 La Lanthanum 138.9	72 Hf Hafnium 178.5	73 Ta Tantalum 180.9	74 W Tungsten 183.9	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium (210)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (257)	105 Db Dubnium (260)	106 Sg Seaborgium (263)	107 Bh Bohrium (262)	108 Hs Hassium (265)	109 Mt Meitnerium (266)	110 Ds Darmstadtium (269)	111 Rg Roentgenium (272)	112	113	114	115	116	(117)	118

58 Ce Cerium 140.1	59 Pr Praseodymium 140.9	60 Nd Neodymium 144.2	61 Pm Promethium (147)	62 Sm Samarium 150.4	63 Eu Europium 152.0	64 Gd Gadolinium 157.3	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0
90 Th Thorium 232.0	91 Pa Protactinium (231)	92 U Uranium 238.0	93 Np Neptunium (237)	94 Pu Plutonium (242)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (249)	99 Es Einsteinium (254)	100 Fm Fermium (253)	101 Md Mendelevium (256)	102 No Nobelium (254)	103 Lr Lawrencium (257)

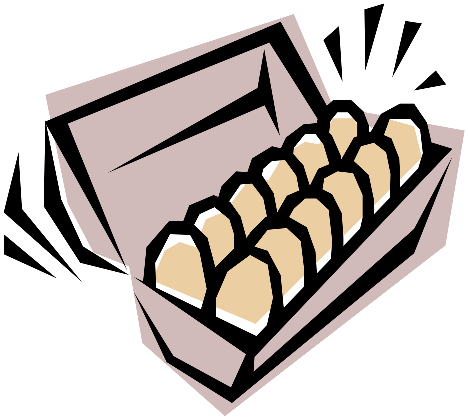
Metals

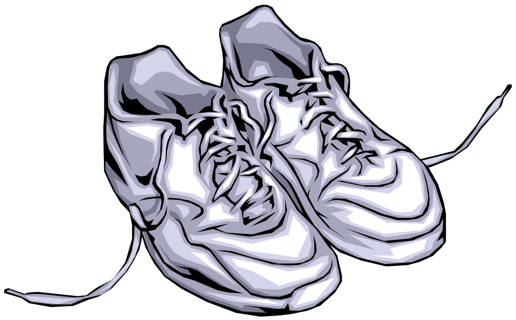
Metalloids

Nonmetals

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

Dozen = 12





Pair = 2

The ***mole (mol)*** is the amount of a substance that contains as many elementary entities as there are atoms in exactly 12.00 grams of ^{12}C

$$1 \text{ mol} = N_A = 6.0221367 \times 10^{23}$$

This number is expressed by Avogadro's number (N_A)

eggs
shoes
marbles
atoms

Molar mass is the mass of 1 mole of eggs / shoes / marbles / atoms in grams

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

1 ^{12}C atom = 12.00 amu

atomic unit

1 amu = 1.66 x 10^{-24}g

1 Oxygen atom = 16.00 amu

1 O_2 molecule = 2(16.00 amu) = 32.00 amu

1 mole ^{12}C atoms = 12.00 g ^{12}C

1 mole lithium atoms = 6.941 g of Li

For any element

atomic mass (amu) = molar mass (grams)

The Mole

- Number of atoms in exactly 12 grams of ^{12}C atoms

How many atoms in 1 mole of ^{12}C ?

– Based on experimental evidence

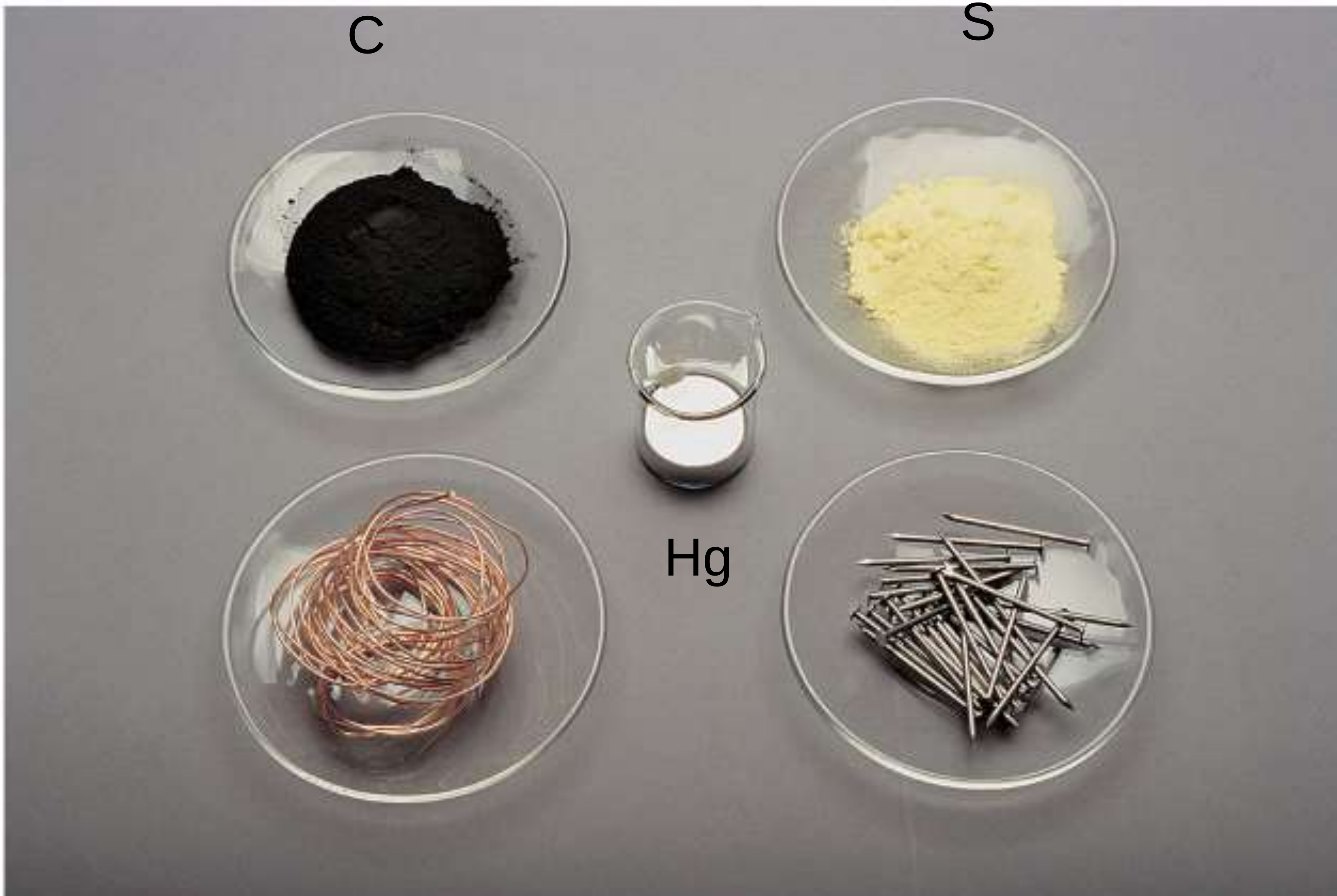
1 mole of ^{12}C = 6.022×10^{23} atoms = 12.011 g

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

One Mole of:



Cu

Fe

Moles of Compounds

Atoms

- **Atomic Mass**
  - Mass of atom (from periodic table)

- **1 mole of atoms = gram atomic mass**
  = 6.022×10^{23} atoms

Molecules

- **Molecular Mass**
  - Sum of atomic masses of all atoms in compound's formula

1 mole of molecule X = gram molecular mass of X

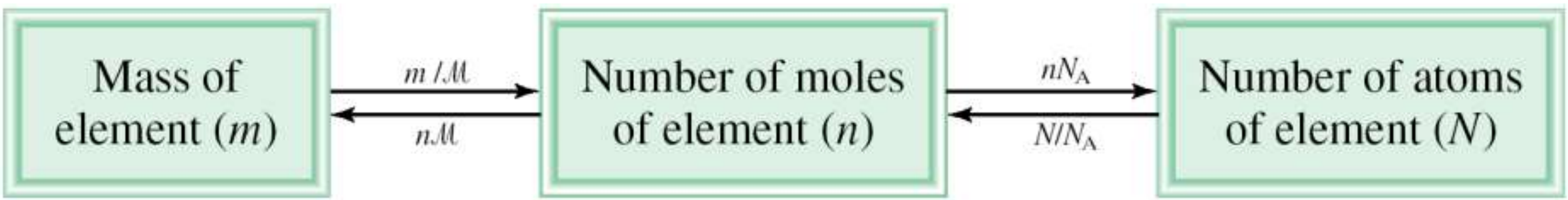
= 6.022×10^{23} molecules

12.00 g ^{12}C atoms $\longrightarrow$ 6.022 x 10^{23} ^{12}C atoms

Mass of g ^{12}C atom $\longleftarrow$ 1 ^{12}C atom

$$\frac{1\ {}^{12}\text{C}\ \cancel{\text{atom}}}{12.00\ \text{amu}} \times \frac{12.00\ \text{g}}{6.022 \times 10^{23}\ {}^{12}\text{C}\ \cancel{\text{atoms}}} = \frac{1.66 \times 10^{-24}\ \text{g}}{1\ \text{amu}}$$

1 amu = 1.66 x 10^{-24} g or 1 g = 6.022 x 10^{23} amu



M = molar mass in g/mol

N_A = Avogadro’s number

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

1 mol K = 39.10 g K

1 mol K = 6.022×10^{23} atoms K

$$0.551 \cancel{\text{g K}} \times \frac{\cancel{\text{1 mol K}}}{39.10 \cancel{\text{g K}}} \times \frac{6.022 \times 10^{23} \text{ atoms K}}{\cancel{\text{1 mol K}}} =$$

8.49×10^{21} atoms K

Example 3.2

Helium (He) is a valuable gas used in industry, low-temperature research, deep-sea diving tanks, and balloons.

How many moles of He atoms are in 6.46 g of He?

Solution

The conversion factor needed to convert between grams and moles is the molar mass. In the periodic table (see inside front cover) we see that the molar mass of He is 4.003 g. This can be expressed as

$$1 \text{ mol He} = 4.003 \text{ g He}$$

From this equality, we can write two conversion factors

A scientific research helium balloon.

The conversion factor on the left is the correct one.
Grams will cancel, leaving the unit mol for the answer, that is,

$$6.46\ \cancel{\text{g He}} \times \frac{1\ \text{mol He}}{4.003\ \cancel{\text{g He}}} = 1.61\ \text{mol He}$$

Thus, there are 1.61 moles of He atoms in 6.46 g of He.

Example 3.3

Zinc (Zn) is a silvery metal that is used in making brass (with copper) and in plating iron to prevent corrosion.

How many grams of Zn are in 0.356 mole of Zn?

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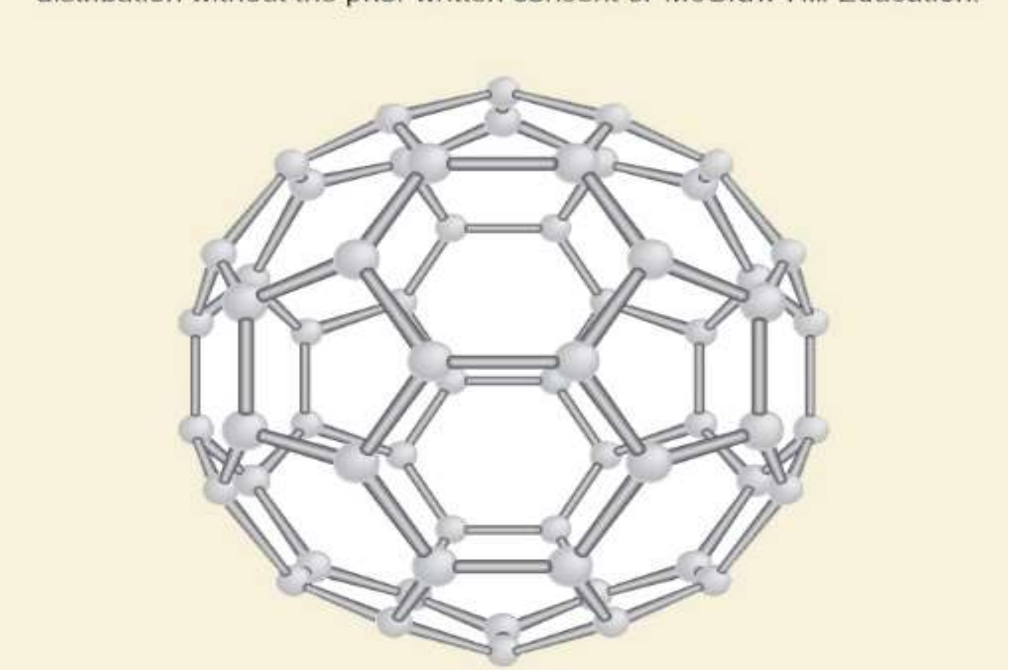
Zinc

Example 3.4

The C_{60} molecule is called buckminsterfullerene because its shape resembles the geodesic domes designed by the visionary architect R. Buckminster Fuller.

What is the mass (in grams) of one C_{60} molecule?



Example 3.4

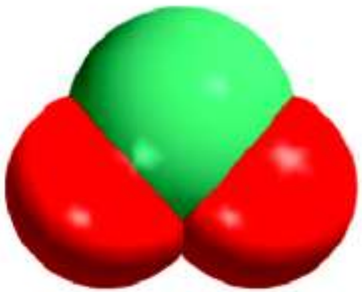
Solution

Because one C_{60} molecule contains 60 C atoms, and 1 mole of C contains $6.022 \times 10^{23}\ C$ atoms and has a mass of 12.011 g, we can calculate the mass of one C_{60} molecule as follows:

$$1\ \cancel{C_{60}}\ \cancel{\text{molecule}} \times \frac{60\ \cancel{C}\ \cancel{\text{atoms}}}{1\ \cancel{C_{60}}\ \cancel{\text{molecule}}} \times \frac{1\ \cancel{\text{mol}}\ \cancel{C}}{6.022 \times 10^{23}\ \cancel{C}\ \cancel{\text{atoms}}}$$

$$\times \frac{12.01\ \text{g}\ C}{1\ \cancel{\text{mol}}\ \cancel{C}} = 1.197 \times 10^{-21}\ \text{g}$$

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 (amu) = molar mass (grams)

1 molecule SO_2 = 64.07 amu

1 mole SO_2 = 64.07 g SO_2

How many H atoms are in 72.5 g of C_3H_8O ?

1 mol $C_3H_8O = (3 \times 12) + (8 \times 1) + 16 = 60$ g C_3H_8O

1 mol C_3H_8O molecules = 8 mol H atoms

1 mol H = 6.022×10^{23} atoms H

$$\cancel{72.5\ \text{g}\ C_3H_8O} \times \frac{1\ \cancel{\text{mol}\ C_3H_8O}}{60\ \cancel{\text{g}\ C_3H_8O}} \times \frac{8\ \cancel{\text{mol H atoms}}}{1\ \cancel{\text{mol}\ C_3H_8O}} \times \frac{6.022 \times \cancel{10^{23}}\ \text{H atoms}}{1\ \text{mol H atoms}} =$$

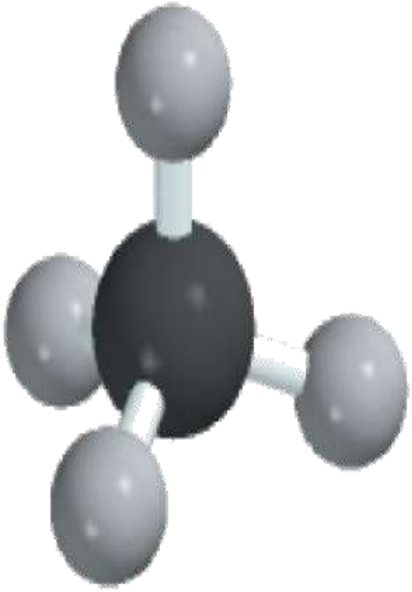
5.82×10^{24} atoms H

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
2 P	2 x 30.97
8 O	+ 8 x 16.00
	310.18 amu

Example 3.6



Methane (CH_4) is the principal component of natural gas.

How many moles of CH_4 are present in 6.07 g of CH_4?

molar mass of $CH_4 = 12.01\,g + 4(1.008\,g) = 16.04\,g$

$$6.07\,\cancel{g}\ \cancel{CH_4} \times \frac{1\,mol\,CH_4}{16.04\,\cancel{g}\ \cancel{CH_4}} = 0.378\,mol\,CH_4$$

Thus, there is 0.378 mole of CH_4 in 6.07 g of CH_4.

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Example 3.7

How many hydrogen atoms are present in 25.6 g of urea [$(NH_2)_2CO$], which is used as a fertilizer, in animal feed, and in the manufacture of polymers?

The molar mass of urea is 60.06 g.

We can combine these conversions

$$\text{grams of urea} \rightarrow \text{moles of urea} \rightarrow \text{moles of H} \rightarrow \text{atoms of H}$$

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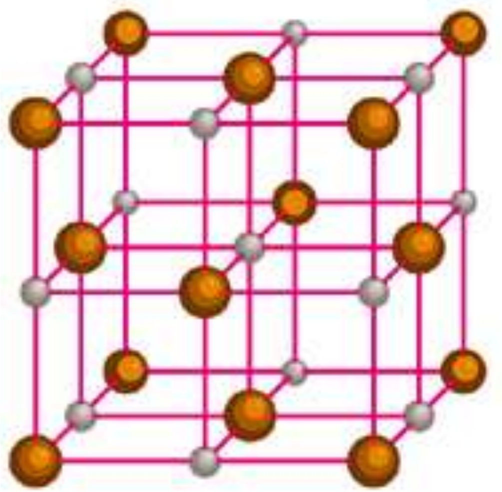
$$= 25.6\,\text{g}\,(NH_2)_2\,CO \times \frac{1\,\text{mol}\,(NH_2)_2\,CO}{60.06\,\text{g}\,(NH_2)_2\,CO} \times \frac{4\,\text{mol}\,H}{1\,\text{mol}\,(NH_2)_2\,CO} \times \frac{6.022\times10^{23}\ H\ \text{atoms}}{1\,\text{mol}\,H}$$

urea

$$= 25.6\,\cancel{\text{g}}\,(\cancel{NH_2})_{\cancel{2}}\,\cancel{CO} \times \frac{1\,\text{mol}\,(\cancel{NH_2})_{\cancel{2}}\,\cancel{CO}}{60.06\,\cancel{\text{g}}\,(\cancel{NH_2})_{\cancel{2}}\,\cancel{CO}} \times \frac{4\,\cancel{\text{mol}}\,\cancel{H}}{1\,\text{mol}\,(\cancel{NH_2})_{\cancel{2}}\,\cancel{CO}} \times \frac{6.022\times10^{23}\ H\ \text{atoms}}{1\,\cancel{\text{mol}}\,\cancel{H}}$$

$$= 1.03\times10^{24}\ H\ \text{atoms}$$

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 (grams)

1 formula unit NaCl = 58.44 amu

1 mole NaCl = 58.44 g NaCl

Formula Weights (FW)

- A formula weight is the sum of the atomic weights for the atoms in a chemical formula.
- So, the formula weight of calcium chloride, $CaCl_2$, would be

$$\begin{array}{rl} \text{Ca:} & 1(40.1\ \text{amu*}) \\ \underline{+\ \text{Cl:}} & \underline{2(35.5\ \text{amu})} \\ & 111.1\ \text{amu} \end{array}$$

- Formula weights are generally reported for ionic compounds.

*atomic mass unit

Molecular Weight (MW)

- A molecular weight is the sum of the atomic weights of the atoms in a molecule.

Formula and Molecular Weights

The **formula weight of a substance is the sum of the atomic weights of the atoms in the** chemical formula of the substance. Using atomic weights, we find, for example, that the formula weight of sulfuric acid (H_2SO_4) is 98.1 amu:*

$$\begin{aligned} \text{FW of } H_2SO_4 &= 2(\text{AW of H}) + (\text{AW of S}) + 4(\text{AW of O}) \\ &= 2(1.0 \text{ amu}) + 32.1 \text{ amu} + 4(16.0 \text{ amu}) \\ &= 98.1 \text{ amu} \end{aligned}$$

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**

Ex. If we need 0.168 mole $Ca_3(PO_4)_2$ for an experiment, how many grams do we need to weigh out?

- Calculate MM of $Ca_3(PO_4)_2$

$3 \times$ mass Ca $= 3 \times 40.08$ g $= 120.24$ g

$2 \times$ mass P $= 2 \times 30.97$ g $= 61.94$ g

$8 \times$ mass O $= 8 \times 16.00$ g $= 128.00$ g

1 mole $Ca_3(PO_4)_2 = 310.18$ g $Ca_3(PO_4)_2$

- What do we want to determine?

0.168 g $Ca_3(PO_4)_2$ = ? Mol Fe

Start End

$$0.160\,\text{mol}\,Ca_3(PO_4)_2 \times \left(\frac{310.18\,\text{g}\,Ca_3(PO_4)_2}{1\,\text{mol}\,Ca_3(PO_4)_2}\right)$$ **= 52.11 g $Ca_3(PO_4)_2$**

Examples

1- Calculate how many atoms there are in 0.200 moles of copper.

The number of atoms in one mole of Cu is equal to the Avogadro number = 6.02×10^{23} .

Number of atoms in 0.200 moles of Cu

$$= (0.200 \text{ mol}) \times (6.02 \times 10^{23} \text{ mol}^{-1}) = 1.20 \times 10^{23} .$$

2- Calculate how molecules of H_2O there are in 12.10 moles of water.

Number of water molecules = $(12.10 \text{ mol}) \times (6.02 \times 10^{23})$

$$= 7.287 \times 10^{24}$$

3- Calculate the mass, in grams, of 0.433 mol of $Ca(NO_3)_2$.

Mass = o.433 mol x 164.1 g/mol = 71.1 g.

4- Calculate the number of moles of glucose ($C_6H_{12}O_6$) in 5.380 g of $C_6H_{12}O_6$.

$$\text{Moles of } C_6H_{12}O_6 = \frac{5.380 \text{ g}}{180.0 \text{ gmol}^{-1}} = 0.02989 \text{ mol.}$$

5- How many glucose molecules are in 5.23 g of $C_6H_{12}O_6$? How many oxygen atoms are in this sample?

$$\text{Molecules of } C_6H_{12}O_6 = \frac{5.23 \text{ g}}{180.0 \text{ gmol}^{-1}} \text{ x } (6.02\text{x}10^{23})$$

$$= 1.75 \text{ x } 10^{22} \text{ molecules}$$

$$\text{Atoms of O} = 1.75 \text{ x } 10^{22} \text{ x } 6 = 1.05 \text{ x } 10^{23}$$

What is the mass, in grams, of one molecule of octane, C_8H_{18}?

Molecules octane $\longrightarrow$ mol octane $\longrightarrow$ g octane

1. Calculate molar mass of octane

Mass C = 8×12.01 g = 96.08 g

Mass H = 18×1.008 g = 18.14 g

1 mol octane = 114.22 g octane

2. Convert 1 molecule of octane to grams

$$\left(\frac{114.22\ \text{g octane}}{1\ \cancel{\text{mol octane}}}\right) \times \left(\frac{1\ \cancel{\text{mol octane}}}{6.022 \times 10^{23}\ \text{molecules octane}}\right)$$

= **1.897×10^{-22} g octane**

- Calculate the number of formula units of Na_2CO_3 in 1.29 moles of Na_2CO_3.

$$1.29\ \cancel{\text{mol}\ Na_2CO_3}\left(\frac{6.0223\times10^{23}\ \text{formula units}\ Na_2CO_3}{1\ \cancel{\text{mol}\ Na_2CO_3}}\right)$$

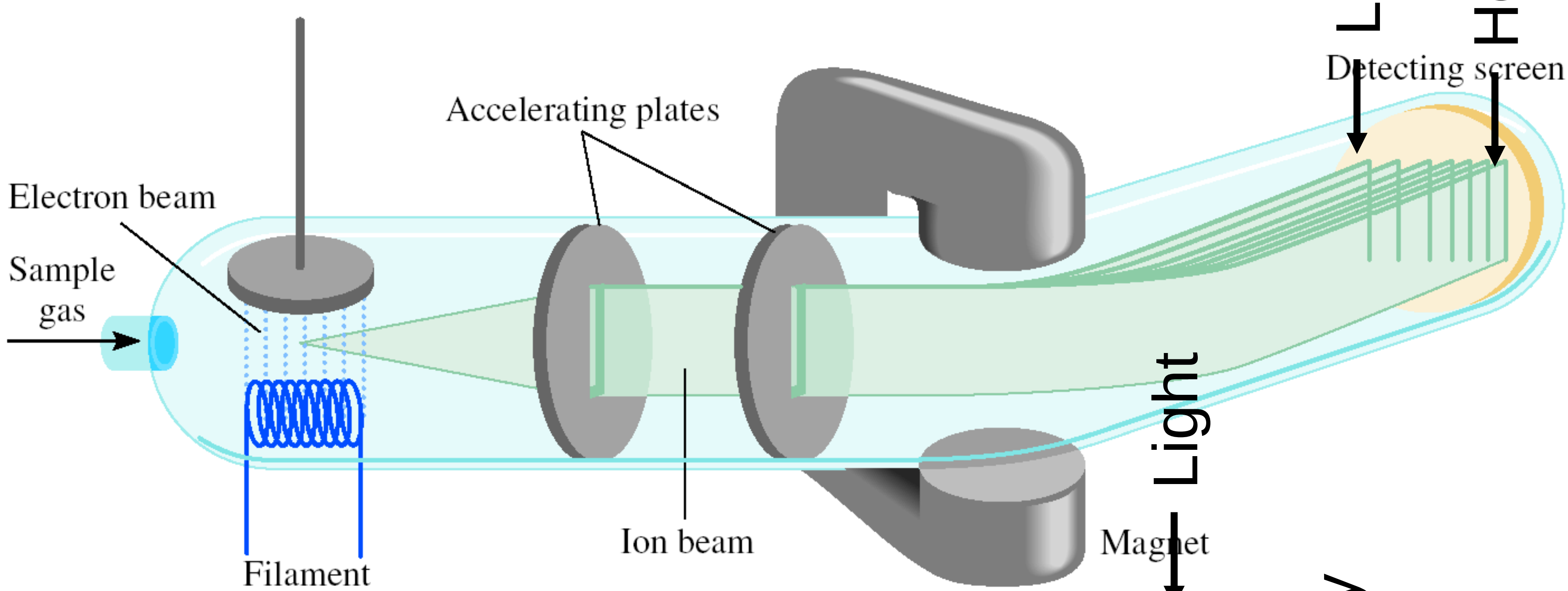
$$= \mathbf{7.77\times10^{23}\ \text{particles}\ Na_2CO_3}$$

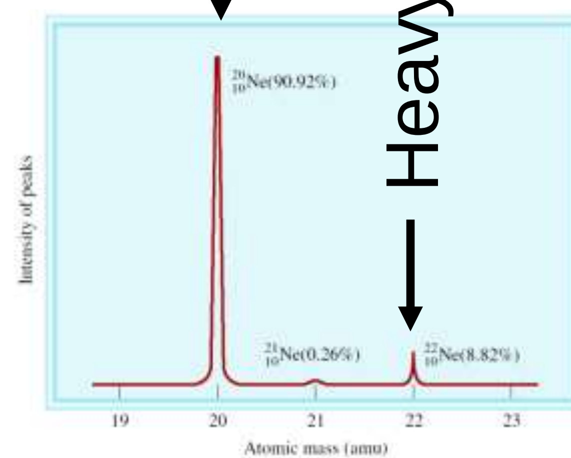
- How many moles of Na_2CO_3 are there in 1.15×10^5 formula units of Na_2CO_3 ?

$$1.15\times10^{5}\ \cancel{\text{formula units}\ Na_2CO_3}\left(\frac{1\ \text{mol}\ Na_2CO_3}{6.0223\times10^{23}\ \cancel{\text{formula units}\ Na_2CO_3}}\right)$$

$$= \mathbf{1.91\times10^{-19}\ \text{mol}\ Na_2CO_3}$$

Mass Spectrometer



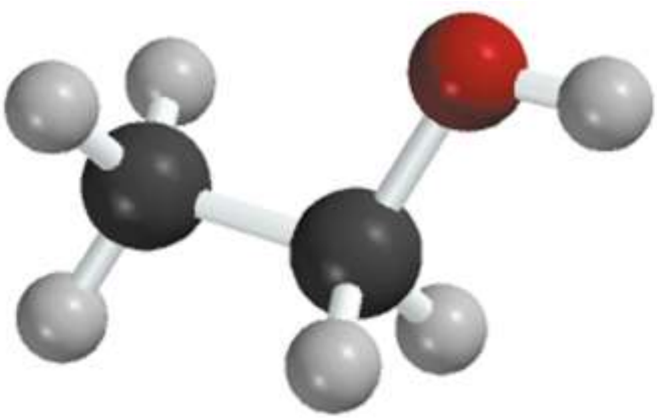


Mass Spectrum of Ne

Percent composition of an element in a compound =

$$\frac{n \text{ x molar mass of element}}{\text{molar mass of compound}} \text{ x } 100\%$$

n is the number of moles of the element in 1 mole of the compound



C_2H_6O

$$\%C = \frac{2 \text{ x } (12.01 \text{ g})}{46.07 \text{ g}} \text{ x } 100\% = 52.14\%$$

$$\%H = \frac{6 \text{ x } (1.008 \text{ g})}{46.07 \text{ g}} \text{ x } 100\% = 13.13\%$$

$$\%O = \frac{1 \text{ x } (16.00 \text{ g})}{46.07 \text{ g}} \text{ x } 100\% = 34.73\%$$

52.14% + 13.13% + 34.73% = 100.0%

- In phosphoric acid (H_3PO_4), calculate the percent composition in each elements

Molar mass of H_3PO_4 = 97.99 g

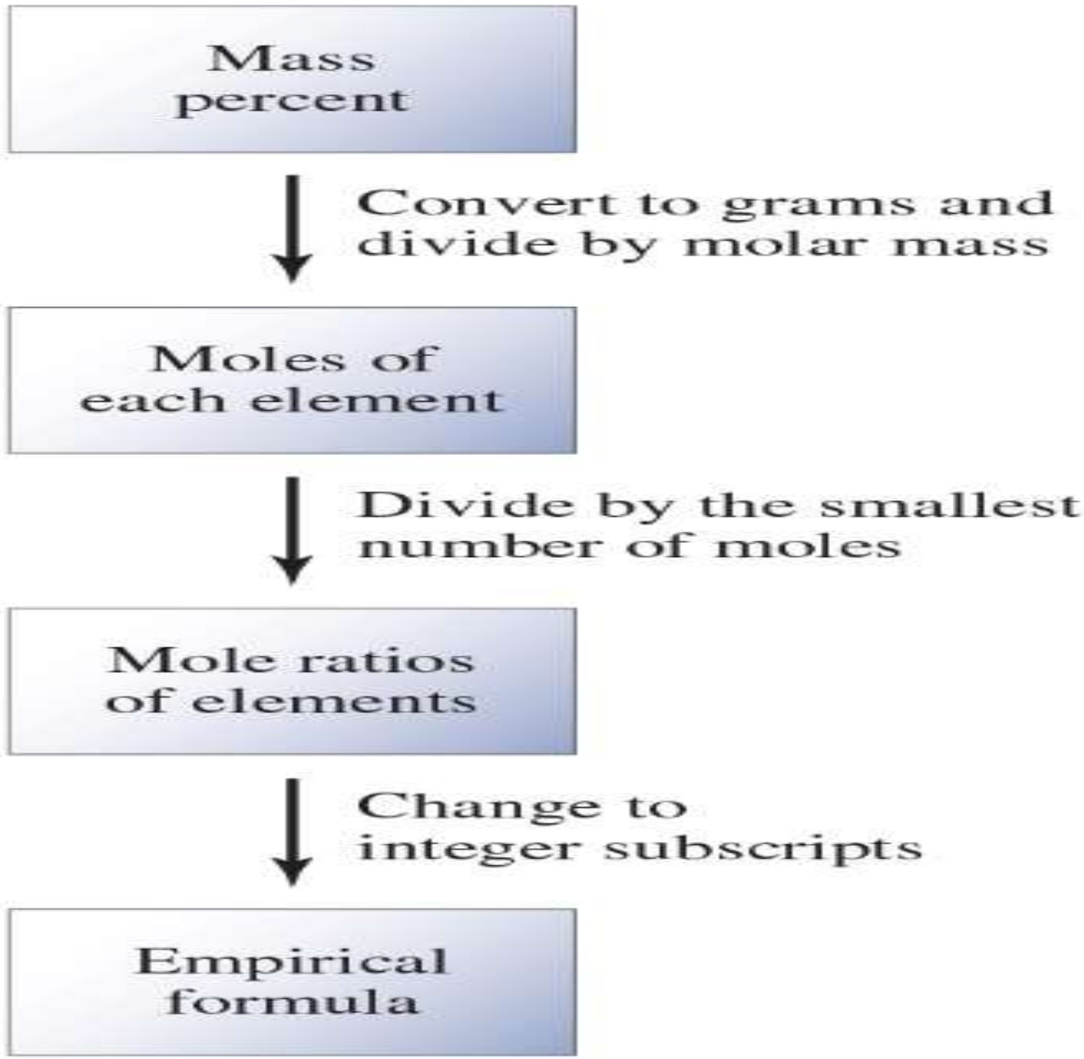
$$\%H = \frac{3 \times (1.008\ g)}{97.99\ g\ H_3PO_4} \times 100\% = 3.086\%$$

$$\%P = \frac{1 \times (30.97\ g)}{97.99\ g\ H_3PO_4} \times 100\% = 31.61\%$$

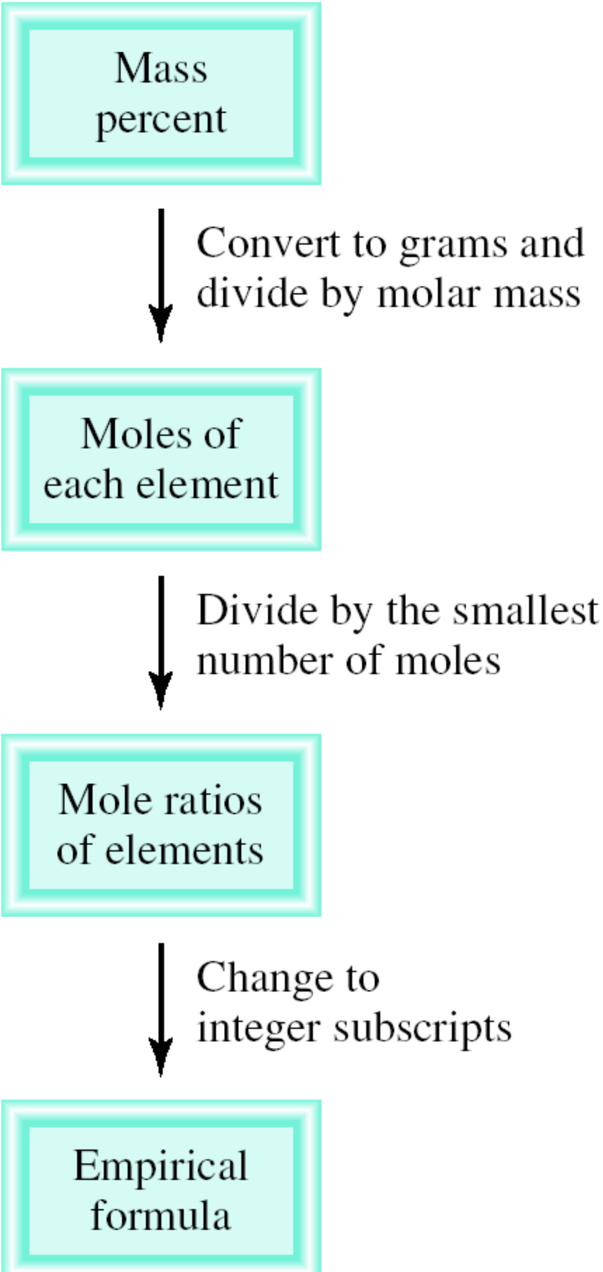
$$\%O = \frac{4 \times (16.00\ g)}{97.99\ g\ H_3PO_4} \times 100\% = 65.31\%$$

3.086% + 31.61% + 65.31% = 100.0%

Percent Composition and Empirical Formulas



Percent Composition and Empirical Formulas



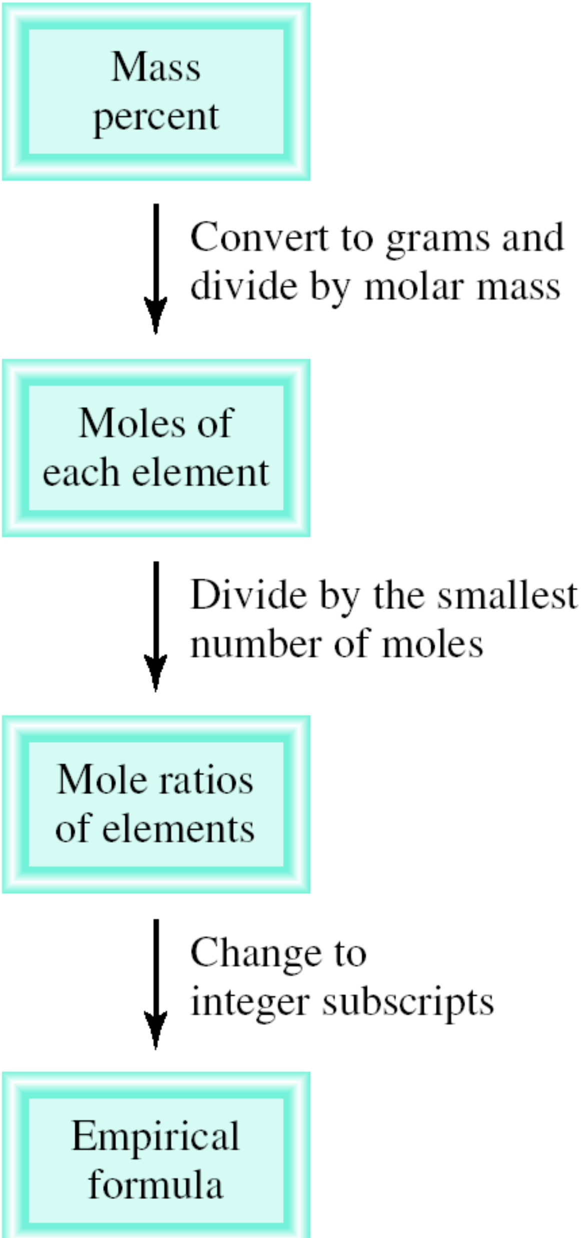
Determine the empirical formula of a compound that has the following percent composition by mass: K 24.75, Mn 34.77, O 40.51 percent.

$$n_K = 24.75\ \cancel{\text{g K}} \times \frac{1\ \text{mol K}}{39.10\ \cancel{\text{g K}}} = 0.6330\ \text{mol K}$$

$$n_{Mn} = 34.77\ \cancel{\text{g Mn}} \times \frac{1\ \text{mol Mn}}{54.94\ \cancel{\text{g Mn}}} = 0.6329\ \text{mol Mn}$$

$$n_O = 40.51\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 2.532\ \text{mol O}$$

Percent Composition and Empirical Formulas



$$n_K = 0.6330,\ n_{Mn} = 0.6329,\ n_O = 2.532$$

$$K : \frac{0.6330}{0.6329} \approx 1.0$$

$$Mn : \frac{0.6329}{0.6329} = 1.0$$

$$O : \frac{2.532}{0.6329} \approx 4.0$$

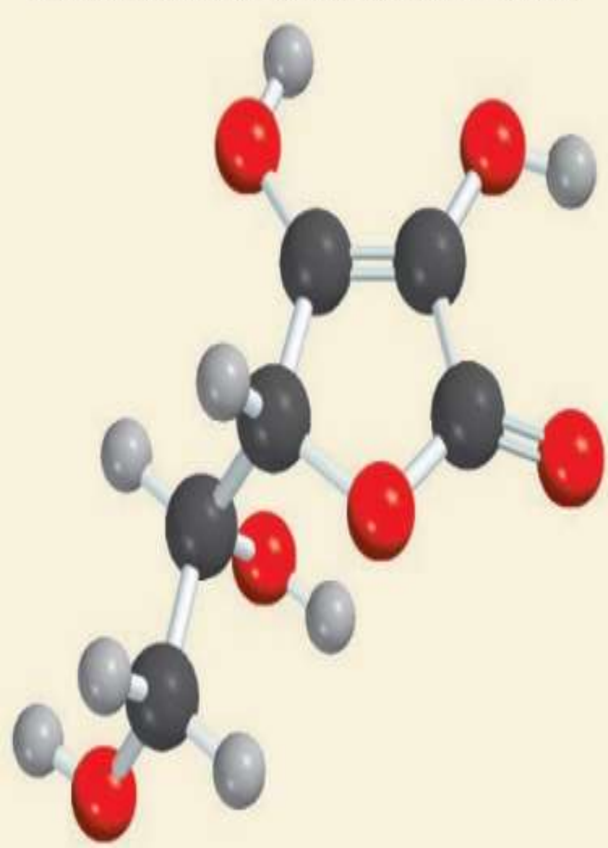
$$KMnO_4$$

Example 3.9

Ascorbic acid (vitamin C) cures scurvy.

It is composed of 40.92 percent carbon (C), 4.58 percent hydrogen (H), and 54.50 percent oxygen (O) by mass.

Determine its empirical formula.



Solution

If we have 100 g of ascorbic acid, then each percentage can be converted directly to grams. In this sample, there will be 40.92 g of C, 4.58 g of H, and 54.50 g of O.

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

$$n_H = 40.92\ \cancel{\text{g}}\ \cancel{H} \times \frac{1\ \text{mol}\ H}{1.008\ \cancel{\text{g}}\ \cancel{H}} = 4.54\ \text{mol}\ H$$

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

Thus, we arrive at the formula $C_{3.407}H_{4.54}O_{3.406}$,

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

This can be done by a trial-and-error procedure:

$$1.33 \times 1 = 1.33$$

$$1.33 \times 2 = 2.66$$

$$1.33 \times 3 = 3.99 < 4$$

Because 1.33 × 3 gives us an integer (4), we multiply all the subscripts by 3 and obtain $C_3H_4O_3$ as the empirical formula for ascorbic acid.

Combust 11.5 g ethanol

Collect 22.0 g CO_2 and 13.5 g H_2O

g CO_2 $\longrightarrow$ mol CO_2 $\longrightarrow$ mol C $\longrightarrow$ g C   6.0 g C = 0.5 mol C

g H_2O $\longrightarrow$ mol H_2O $\longrightarrow$ mol H $\longrightarrow$ g H   1.5 g H = 1.5 mol H

g of O = g of sample – (g of C + g of H)   4.0 g O = 0.25 mol O

Empirical formula $C_{0.5}H_{1.5}O_{0.25}$

Divide by smallest subscript (0.25)

Empirical formula C_2H_6O

Experimental Determination of Empirical Formulas

$$\text{mass of C} = 22.0\ \text{g}\ \cancel{CO_2} \times \frac{1\ \cancel{\text{mol}\ CO_2}}{44.01\ \cancel{\text{g}\ CO_2}} \times \frac{1\ \cancel{\text{mol C}}}{1\ \cancel{\text{mol}\ CO_2}} \times \frac{12.01\ \text{g C}}{1\ \cancel{\text{mol C}}}$$
$$= 6.00\ \text{g C}$$

$$\text{mass of H} = 13.5\ \text{g}\ \cancel{H_2O} \times \frac{1\ \cancel{\text{mol}\ H_2O}}{18.02\ \cancel{\text{g}\ H_2O}} \times \frac{2\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol}\ H_2O}} \times \frac{1.008\ \text{g H}}{1\ \cancel{\text{mol H}}}$$
$$= 1.51\ \text{g H}$$

Thus, 11.5 g of ethanol contains 6.00 g of carbon and 1.51 g of hydrogen. The remainder must be oxygen, whose mass is

$$\begin{aligned}\text{mass of O} &= \text{mass of sample} - (\text{mass of C} + \text{mass of H})\\ &= 11.5\ \text{g} - (6.00\ \text{g} + 1.51\ \text{g})\\ &= 4.0\ \text{g}\end{aligned}$$

The number of moles of each element present in 11.5 g of ethanol is

$$\text{moles of C} = 6.00\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 0.500\ \text{mol C}$$

$$\text{moles of H} = 1.51\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}} = 1.50\ \text{mol H}$$

$$\text{moles of O} = 4.0\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.25\ \text{mol O}$$

Divide by smallest subscript (0.25) Empirical formula C_2H_6O

Molecular Formulas

Molecular weight of the compound should be known

$$X = \frac{\mathcal{M}_{\text{actual}}}{\mathcal{M}_{\text{empirical}}}$$

Multiply the empirical formula by the integer x

A compound has empirical formula $C_6H_{10}S_2O$ but its molecular weight is 324 g/mol !

$C_{12}H_{20}S_4O_2$

Calculate the number of grams of Al in 371 g of Al_2O_3?

196.5 g

Example 3.10

Chalcopyrite ($CuFeS_2$) is a principal mineral of copper.

Calculate the number of kilograms of Cu in 3.71×10^3 kg of chalcopyrite.

©The Natural History Museum/Alamy Stock Photo

Chalcopyrite

Solution

The molar masses of Cu and $CuFeS_2$ are 63.55 g and 183.5 g, respectively. The mass percent of Cu is therefore

$$\%Cu = \frac{\text{molar mass of Cu}}{\text{molar mass of CuFeS}_2} \times 100\% = \frac{63.55\text{ g}}{183.5\text{ g}} \times 100\% = 34.63$$

$$\text{mass of Cu in CuFeS}_2 = 0.3463 \times (3.71 \times 10^3\text{ kg}) = 1.28 \times 10^3\text{ kg}$$

Determining Empirical Formula by Combustion Analysis

*Combustion of 0.255 g of isopropyl alcohol produces 0.561 g of CO_2 and 0.306 g of H_2O. Determine the empirical formula of isopropyl alcohol.

$$\text{Grams of C} = \frac{0.561 \text{ g } CO_2}{44 \text{ gmol}^{-1}} \times 1 \times 12 \text{ g mol-1} = 0.153 \text{ g}$$

$$\text{Grams of H} = \frac{0.306 \text{ g } H_2O}{18 \text{ gmol}^{-1}} \times 2 \times 1 \text{ g mol-1} = 0.0343 \text{ g}$$

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

= 0.255 g – (0.153 g + 0.0343 g) = 0.068 g

$$\text{Moles of C} = \frac{0.153 \text{ g C}}{12 \text{ gmol}^{-1}} = 0.0128 \text{ mol}$$

$$\text{Moles of H} = \frac{0.0343 \text{ g H}}{1 \text{ gmol}^{-1}} = 0.0343 \text{ mol}$$

$$\text{Moles of O} = \frac{0.068 \text{ g O}}{16 \text{ gmol}^{-1}} = 0.0043 \text{ mol}$$

Calculate the mole ratio by dividing by the smallest number of moles:

C : H : O

2.98 : 7.91 : 1.00

C_3H_8O

Example 3.11

A sample of a compound contains 30.46 percent nitrogen and 69.54 percent oxygen by mass.

In a separate experiment, the molar mass of the compound is found to be between 90 g and 95 g.

Determine the molecular formula and the accurate molar mass of the compound.

Let *n* represent the number of moles of each element so that

$$n_N = 30.46\ \cancel{\text{g}}\ \cancel{\text{N}} \times \frac{1\ \text{mol N}}{14.01\ \cancel{\text{g}}\ \cancel{\text{N}}} = 2.174\ \text{mol N} \qquad n_O = 69.54\ \cancel{\text{g}}\ \cancel{\text{O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g}}\ \cancel{\text{O}}} = 4.346\ \text{mol O}$$

The molar mass of the empirical formula NO_2 is

empirical molar mass = 14.01 g + 2(16.00 g) = 46.01 g

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{90\ \text{g}}{46.01\ \text{g}} \approx 2$$

molecular formula is $(NO_2)_2$ or N_2O_4

Experimental Determination of Empirical Formulas

“*empirical*” means “based only on observation and measurement.”

Ex: *The empirical formula of ethanol is determined from analysis of the compound in terms of its component elements.*

No knowledge of how the atoms are linked together in the compound is required.

Question 1

Determine the number of moles of aluminum in 0.2154 kg of Al.

A) 1.297×10^{23} mol
B) 5.811×10^{3} mol
C) 7.984 mol
D) 0.1253 mol
E) 7.984×10^{-3} mol

Question 2

How many phosphorus atoms are there in 2.57 g of P?

A) 4.79×10^{25}
B) 1.55×10^{24}
C) 5.00×10^{22}
D) 8.30×10^{-2}
E) 2.57

Question 3

Determine the mass percent of iron in $Fe_4[Fe(CN)_6]_3$.

A) 44% Fe
B) 26% Fe
C) 33% Fe
D) 58% Fe
E) None of the above.

Question 4

How many oxygen atoms are present in 5.2 g of O_2?

A) 5.4×10^{-25} atoms
B) 9.8×10^{22} atoms
C) 2.0×10^{23} atoms
D) 3.1×10^{24} atoms
E) 6.3×10^{24} atoms

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***

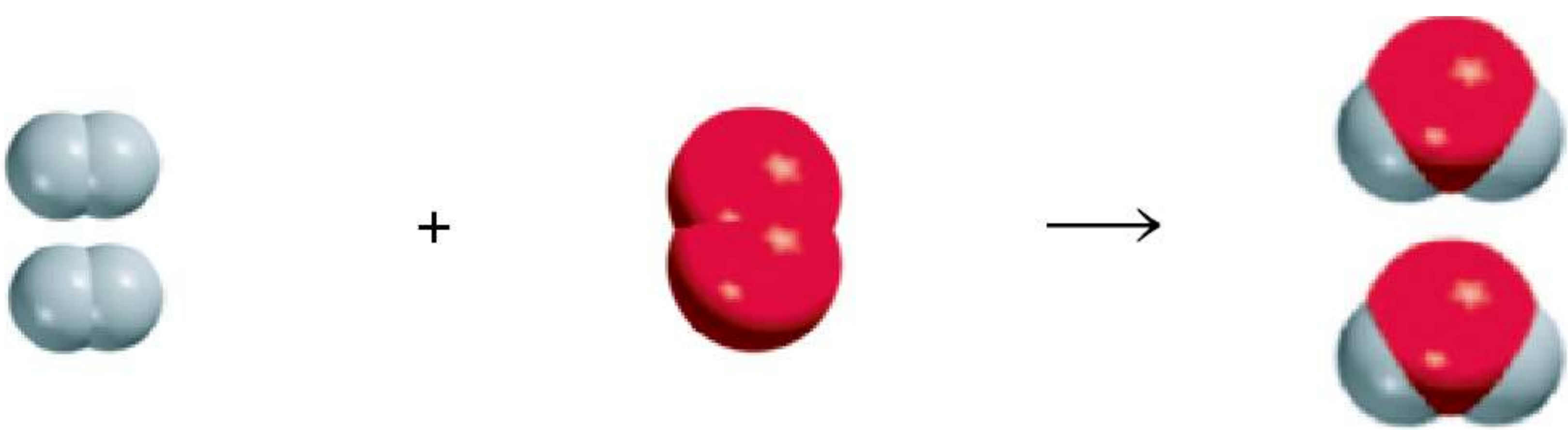
A ***chemical equation*** uses chemical symbols to show what happens during a chemical reaction

reactants $\longrightarrow$ products

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

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



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

NOT

2 grams Mg + 1 gram O_2 makes 2 g MgO

Balancing Chemical Equations

1. Write the **correct** formula(s) for the reactants on the left side and the **correct** formula(s) for the product(s) on the right side of the equation.

Ethane reacts with oxygen to form carbon dioxide and water

$$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$$

2. Change the numbers in front of the formulas (***coefficients***) to make the number of atoms of each element the same on both sides of the equation. Do not change the subscripts.

$2C_2H_6$ **NOT** C_4H_{12}

Balancing Chemical Equations

3. Start by balancing those elements that appear in only one reactant and one product.

$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$ start with C or H but not O

2 carbon on left

1 carbon on right

multiply CO_2 by 2

$C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$

6 hydrogen on left

2 hydrogen on right

multiply H_2O by 3

$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$

Balancing Chemical Equations

4. Balance those elements that appear in two or more reactants or products.

$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$ multiply O_2 by $\frac{7}{2}$

2 oxygen on left

4 oxygen (2x2) + 3 oxygen (3x1) = 7 oxygen on right

$C_2H_6 + \frac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$

remove fraction
multiply both sides by 2

$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$

Balancing Chemical Equations

5. Check to make sure that you have the same number of each type of atom on both sides of the equation.

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

4 C (2 x 2)	4 C
12 H (2 x 6)	12 H (6 x 2)
14 O (7 x 2)	14 O (4 x 2 + 6)

Reactants	Products
4 C	4 C
12 H	12 H
14 O	14 O

Examples of balancing chemical equation

$Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$ (unbalanced)

$O_2 + NO \longrightarrow NO_2$ (unbalanced)

$$2HCl + Zn \longrightarrow ZnCl_2 + H_2$$

$$C_3H_8 + 5\,O_2 \longrightarrow 3CO_2 + 4H_2O$$

$$Zn + 2\,HNO_3 \longrightarrow Zn(NO_3)_2 + H_2$$

PRACTICE EXERCISE

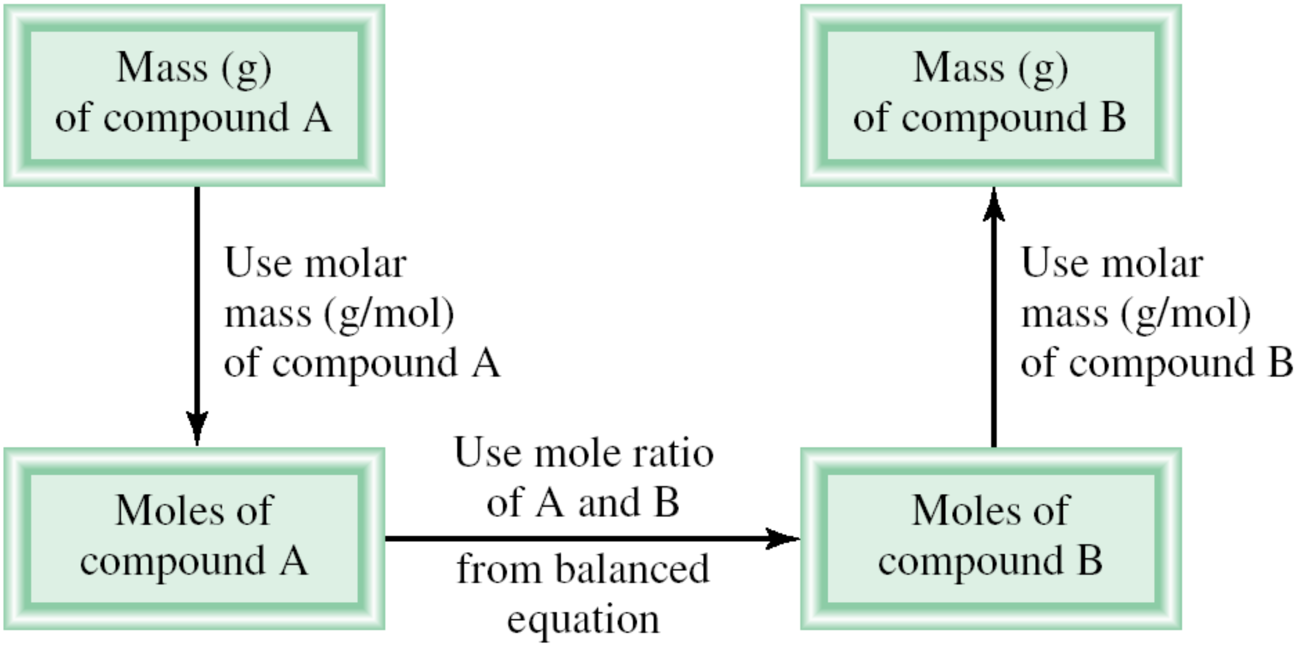
Balance these equations by providing the missing coefficients:

$\_\_Al(s) + \_\_HCl(aq) \longrightarrow \_\_AlCl_3(aq) + \_\_\ H_2\ (g)$

$\_\_C_2H_4(g) + \_\_O_2(g) \longrightarrow \_\_CO_2(g) + \_\_\ H_2O(g)$

$\_\_Fe(s) + \_\_O_2(g) \longrightarrow \_\_Fe_2O_3(s)$

Amounts of Reactants and Products



1. Write 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 desired units

Methanol burns in air according to the equation

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

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

grams CH_3OH $\longrightarrow$ moles CH_3OH $\longrightarrow$ moles H_2O $\longrightarrow$ grams H_2O

molar mass CH_3OH | coefficients chemical equation | molar mass H_2O

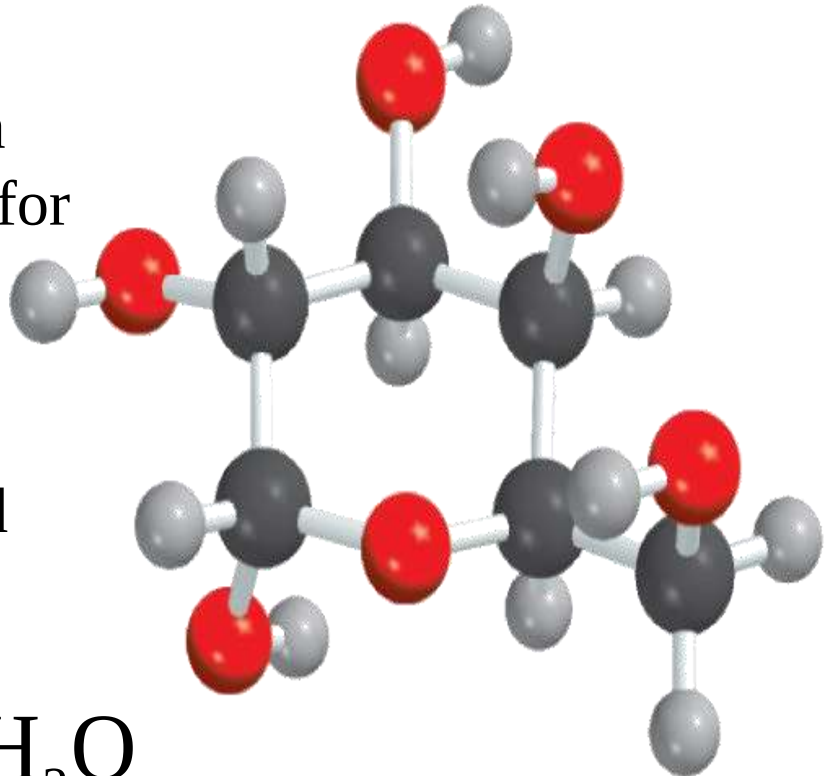
$$209\ \cancel{\text{g } CH_3OH} \times \frac{1\ \cancel{\text{mol } CH_3OH}}{32.0\ \cancel{\text{g } CH_3OH}} \times \frac{4\ \cancel{\text{mol } H_2O}}{2\ \cancel{\text{mol } CH_3OH}} \times \frac{18.0\text{ g } H_2O}{1\ \cancel{\text{mol } H_2O}} =$$

235 g H_2O

Example 3.13

The food we eat is degraded, or broken down, in our bodies to provide energy for growth and function. A general overall equation for this very complex process represents the degradation of glucose ($C_6H_{12}O_6$) to carbon dioxide (CO_2) and water (H_2O):

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$



$C_6H_{12}O_6$

If 856 g of $C_6H_{12}O_6$ is consumed by a person over a certain period, what is the mass of CO_2 produced?

Example 3.13

Solution

We follow the preceding steps and Figure 3.8.

Step 1: The balanced equation is given in the problem.

Step 2: To convert grams of $C_6H_{12}O_6$ to moles of $C_6H_{12}O_6$, we write

$$856\ \cancel{g\ C_6H_{12}O_6} \times \frac{1\ \text{mol}\ C_6H_{12}O_6}{180.2\ \cancel{g\ C_6H_{12}O_6}} = 4.750\ \text{mol}\ C_6H_{12}O_6$$

Step 3: From the mole ratio, we see that

$$1\ \text{mol}\ C_6H_{12}O_6 \simeq 6\ \text{mol}\ CO_2$$

Therefore, the number of moles of CO_2 formed is

$$4.750\ \cancel{\text{mol}\ C_6H_{12}O_6} \times \frac{6\ \text{mol}\ CO_2}{1\ \cancel{\text{mol}\ C_6H_{12}O_6}} = 28.50\ \text{mol}\ CO_2$$

Example 3.13

Step 4: Finally, the number of grams of CO_2 formed is given by

$$28.50 \text{ mol } CO_2 \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 1.25 \times 10^3 \text{ g } CO_2$$

After some practice, we can combine the conversion steps

$$\text{grams of } C_6H_{12}O_6 \rightarrow \text{moles of } C_6H_{12}O_6 \rightarrow \text{moles of } CO_2 \rightarrow \text{grams of } CO_2$$

into one equation:

$$\text{mass of } CO_2 = 856 \cancel{\text{ g } C_6H_{12}O_6} \times \frac{1 \cancel{\text{ mol } C_6H_{12}O_6}}{180.2 \cancel{\text{ g } C_6H_{12}O_6}} \times \frac{6 \cancel{\text{ mol } CO_2}}{1 \cancel{\text{ mol } C_6H_{12}O_6}} \times \frac{44.01 \text{ g } CO_2}{1 \cancel{\text{ mol } CO_2}}$$

$$= 1.25 \times 10^3 \text{ g } CO_2$$

Example 3.14

All alkali metals react with water to produce hydrogen gas and the corresponding alkali metal hydroxide.

A typical reaction is that between lithium and water:

$$2\text{Li}\ (s) + 2\text{H}_2\text{O}\ (l) \rightarrow 2\text{LiOH}\ (aq) + \text{H}_2\ (g)$$

©McGraw-Hill Education/Ken Karp

Lithium reacting with water to produce hydrogen gas.

How many grams of Li are needed to produce 9.89 g of H_2?

Solution The conversion steps are

$$\text{grams of H}_2 \rightarrow \text{moles of H}_2 \rightarrow \text{moles of Li} \rightarrow \text{grams of Li}$$

Combining these steps into one equation, we write

$$9.89\ \text{g H}_2 \times \frac{1\ \text{mol H}_2}{2.016\ \text{g}\ H_2} \times \frac{2\ \text{mol Li}}{1\ \text{mol H}_2} \times \frac{6.941\ \text{g Li}}{1\ \text{mol Li}} = 68.1\ \text{g Li}$$

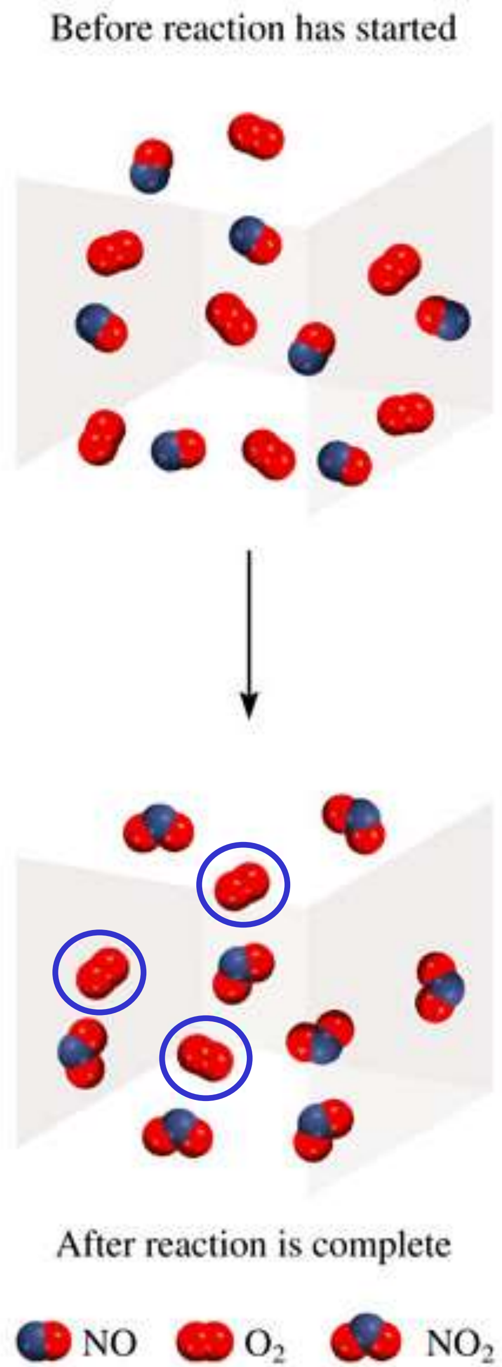
Limiting Reagent:

Reactant used up first in the reaction.

$$2NO + O_2 \longrightarrow 2NO_2$$

NO is the limiting reagent

O_2 is the excess reagent



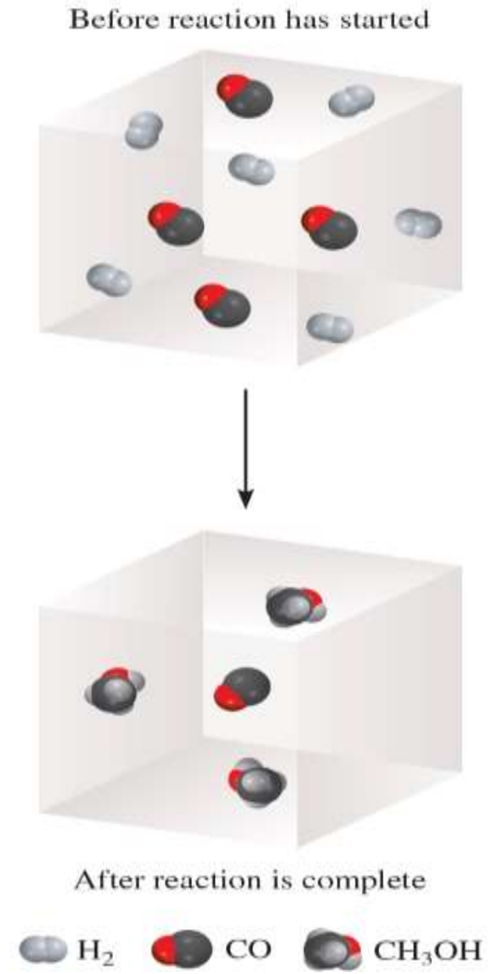
Limiting Reagent:

Reactant used up first in the reaction.

$CO + 2H_2 \rightarrow CH_3OH$

H_2 is the limiting reagent

CO is the excess reagent

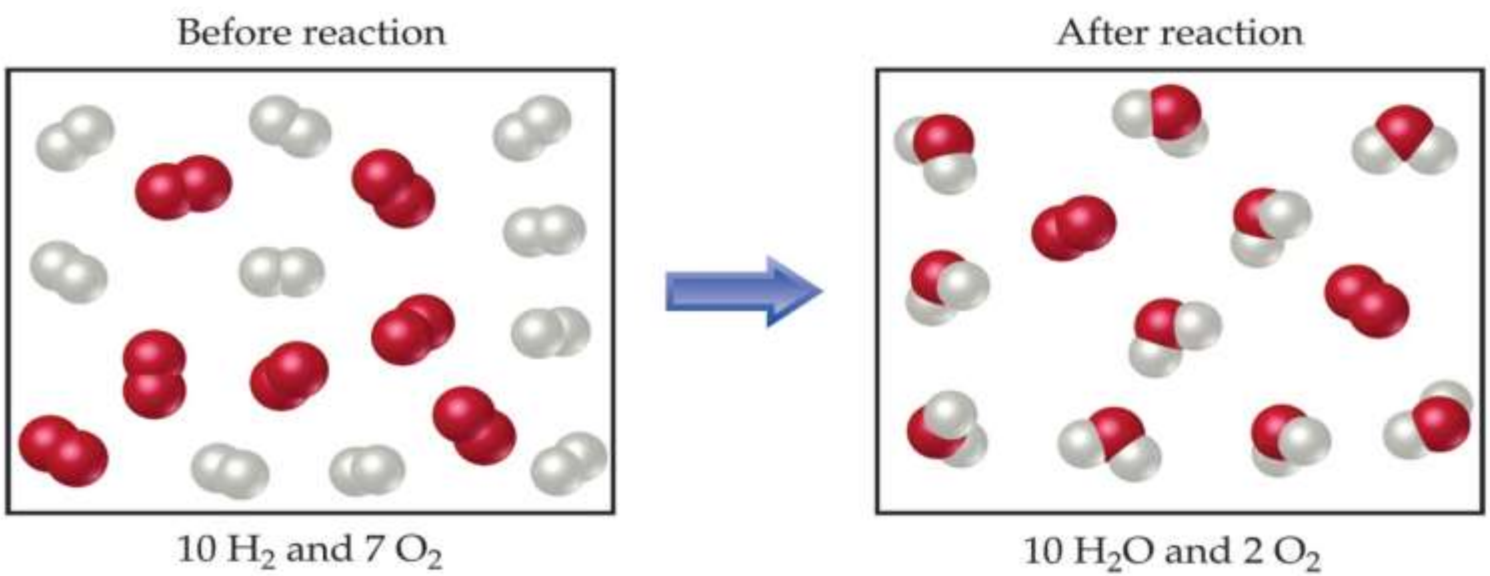


Limiting Reactants

- The reactant that is completely consumed in a reaction is called the **limiting reactant** (limiting reagent).
  - In other words, it's the reactant that run out first (in this case, the H_2).
  - In example below, O_2 would be the **excess reactant** (excess reagent).

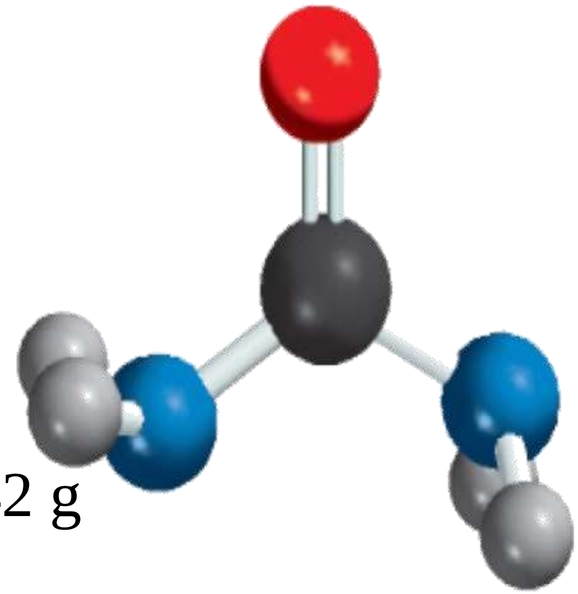
Because 2 mol $H_2 \simeq$ 1 mol O_2, the number of moles of O_2 needed to react with all the H_2 is

$$\text{Moles } O_2 = (10\ \cancel{\text{mol } H_2})\left(\frac{1 \text{ mol } O_2}{2\ \cancel{\text{mol } H_2}}\right) = 5 \text{ mol } O_2$$



Example 3.15

Urea $\left[(NH_2)_2CO\right]$ is prepared by reacting ammonia with carbon dioxide:



$$2NH_3(g) + CO_2(g) \rightarrow (NH_2)_2CO\ (aq) + H_2O\ (l)$$

In one process, 637.2 g of NH_3 are treated with 1142 g of CO_2.

$(NH_2)_2CO$

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

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

c) How much excess reagent (in grams) is left at the end of the reaction?

Example 3.15

Solution

We carry out two separate calculations. First, starting with 637.2 g of NH_3, we calculate the number of moles of $(NH_2)_2CO$ that could be produced if all the NH_3 reacted according to the following conversions:

$$\text{grams of } NH_3 \rightarrow \text{moles of } NH_3 \rightarrow \text{moles of } \left(NH_2\right)_2 CO$$

Combining these conversions in one step, we write

$$\text{moles of } \left(NH_2\right)_2 CO = 637.\cancel{\text{g}}\ \cancel{NH_3} \times \frac{1\ \cancel{\text{mol}}\ \cancel{NH_3}}{17.03\ \cancel{\text{g}}\ \cancel{NH_3}} \times \frac{1\ \text{mol}\ \left(NH_2\right)_2 CO}{2\ \cancel{\text{mol}}\ \cancel{NH_3}}$$

$$= 18.71\ \text{mol}\ \left(NH_2\right)_2 CO$$

Example 3.15

Second, for 1142 g of CO_2, the conversions are

$$\text{grams of } CO_2 \rightarrow \text{moles of } CO_2 \rightarrow \text{moles of } (NH_2)_2CO$$

The number of moles of $(NH_2)_2CO$ that could be produced if all the CO_2 reacted is

$$\text{moles of } (NH_2)_2CO = 1142 \cancel{\text{g}}\ \cancel{CO_2} \times \frac{1 \cancel{\text{mol}}\ \cancel{CO_2}}{44.01 \cancel{\text{g}}\ \cancel{CO_2}} \times \frac{1 \text{ mol } (NH_2)_2CO}{1 \cancel{\text{mol}}\ \cancel{CO_2}}$$

$$= 25.95 \text{ mol } (NH_2)_2CO$$

It follows, therefore, that NH_3 must be the limiting reagent because it produces a smaller amount of $(NH_2)_2CO$.

b-Solution

The molar mass of $(NH_2)_2CO$ is 60.06 g. We use this as a conversion factor to convert from moles of $(NH_2)_2CO$ to grams of $(NH_2)_2CO$:

Example 3.15

b-Solution

$$\text{mass of } (\text{NH}_2)_2\text{CO} = 18.71 \cancel{\text{mol}} \ \cancel{(\text{NH}_2)_2 \text{CO}} \times \frac{60.06 \ \cancel{g} \ \cancel{(\text{NH}_2)_2 \text{CO}}}{\cancel{1} \ \cancel{\text{mol}} \ \cancel{(\text{NH}_2)_2 \text{CO}}}$$

$$= 1124 \ g \ (\text{NH}_2)_2\text{CO}$$

C- Solution

Starting with 18.71 moles of $(NH_2)_2CO$, we can determine the mass of CO_2 that reacted using the mole ratio from the balanced equation and the molar mass of CO_2. The conversion steps are

$$\text{moles of } (\text{NH}_2)_2\text{CO} \rightarrow \text{moles of CO}_2 \rightarrow \text{grams of CO}_2$$

Combining these conversions in one step, we write

$$\text{mass of CO}_2 \text{ reacted} = 18.71 \cancel{\text{mol}} \ \cancel{(\text{NH}_2)_2 \text{CO}} \times \frac{1 \ \cancel{\text{mol}} \ \cancel{\text{CO}_2}}{1 \text{ mol } \cancel{(\text{NH}_2)_2 \text{CO}}} \times \frac{44.01 \ g \ \text{CO}_2}{1 \ \cancel{\text{mol}} \ \cancel{\text{CO}_2}}$$

$$= 823.4 \text{ g CO}_2$$

The amount of CO_2 remaining (in excess) is the difference between the initial amount (1142 g) and the amount reacted (823.4 g):

$$\text{mass of CO}_2 \text{ remaining} = 1142 \text{ g} - 823.4 \text{ g} = 319 \text{ g}$$

In one process, 124 g of Al are reacted with 601 g of Fe_2O_3

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

Calculate the mass of Al_2O_3 formed.

g Al $\longrightarrow$ mol Al $\longrightarrow$ mol Fe_2O_3 needed $\longrightarrow$ g Fe_2O_3 needed

OR

g Fe_2O_3 $\longrightarrow$ mol Fe_2O_3 $\longrightarrow$ mol Al needed $\longrightarrow$ g Al needed

$$\cancel{124\text{ g Al}} \times \frac{1\ \cancel{\text{mol Al}}}{27.0\ \cancel{\text{g Al}}} \times \frac{1\ \cancel{\text{mol } Fe_2O_3}}{2\ \cancel{\text{mol Al}}} \times \frac{160.\text{ g } Fe_2O_3}{1\ \cancel{\text{mol } Fe_2O_3}} = 367\text{ g } Fe_2O_3$$

Start with 124 g Al $\longrightarrow$ need 367 g Fe_2O_3

Have more Fe_2O_3 (601 g) so Al is limiting reagent

Use limiting reagent (Al) to calculate amount of product that can be formed.

$$\text{g Al} \longrightarrow \text{mol Al} \longrightarrow \text{mol } Al_2O_3 \longrightarrow \text{g } Al_2O_3$$

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

$$124 \cancel{\text{ g Al}} \times \frac{1 \cancel{\text{ mol Al}}}{27.0 \cancel{\text{ g Al}}} \times \frac{1 \cancel{\text{ mol } Al_2O_3}}{2 \cancel{\text{ mol Al}}} \times \frac{102.\text{ g } Al_2O_3}{1 \cancel{\text{ mol } Al_2O_3}} = 234 \text{ g } Al_2O_3$$

At this point, all the Al is consumed and Fe_2O_3 remains in excess.

Reaction Yield

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

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

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

Theoretical Yield

- The *theoretical yield* is the maximum amount of product that can be made.
- The amount of product actually obtained in a reaction is called the *actual yield*. The actual yield is almost always less than the theoretical yield.

Why?

- Part of the reactants may not react.
- Side reaction.
- Difficult recovery.

Percent Yield

- The percent yield of a reaction relates to the actual yield to the theoretical (calculated) yield.

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

Examples

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

If we start with 150 g of Fe_2O_3 as the limiting reactant, and found actual yield of Fe was 87.9 g, what is the percent yield?

$$\text{The percent yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

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

150 g Fe_2O_3

$\dfrac{150\ g}{159\ g\ mol^{-1}}$

0.943 mol Fe_2O_3 X $\dfrac{2\ mol\ Fe}{1\ mol\ Fe_2O_3}$

1.887 mol Fe

1.887 mol x 55.85 g mol^{-1}

105 g Fe

Theoretical yield = 105 g.

The percent yield = $\dfrac{87.9\ g}{105\ g}$ x 100 = 83.7 %

Calculate % yield if 803 g of CaO is produced?

$$5Ca + V_2O_5 \longrightarrow 5\,CaO + 2V$$

1.96×10^3 g $\quad 1.54 \times 10^3$ g $\quad$? g

48.90 mol $\quad$ 8.467 mol $\longrightarrow$ 42.34 mol

$$\frac{48.90 \text{ mol Ca}}{5 \text{ mol Ca}} \qquad \frac{8.467 \text{ mol } V_2O_5}{1 \text{ mol } V_2O_5}$$

2374.2 g CaO

9.78 $\quad$ 8.467

$$\textbf{\% Yield} = \frac{803 \text{ g}}{2374.2 \text{ g}} \times 100 = 33.8\%$$

Example 3.17

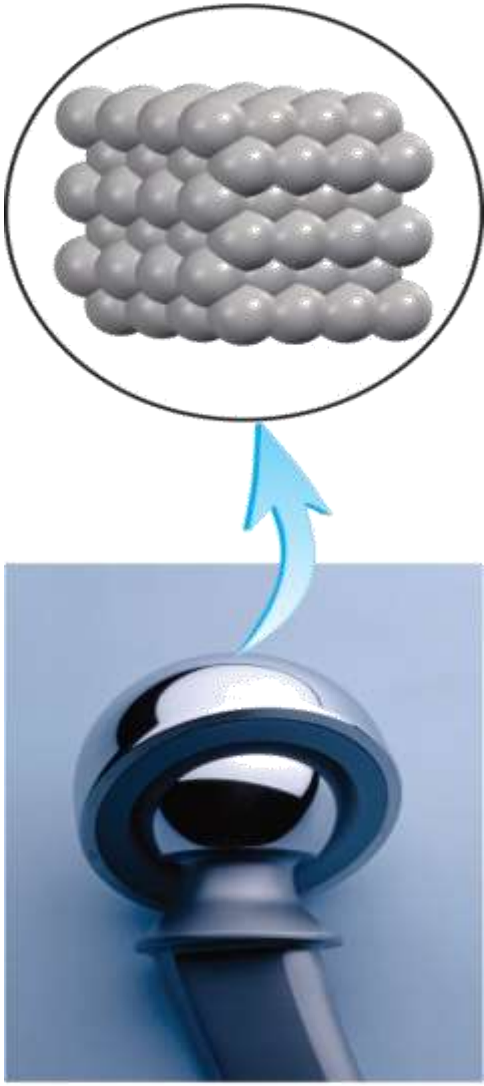
Titanium is a strong, lightweight, corrosion-resistant metal that is used in rockets, aircraft, jet engines, and bicycle frames. It is prepared by the reaction of titanium(IV) chloride with molten magnesium between 950° C and 1150° C:

$$TiCl_4(g) + 2Mg(l) \rightarrow Ti(s) + 2MgCl_2(l)$$

In a certain industrial operation 3.54×10^7 g of $TiCl_4$ are reacted with 1.13×10^7 g of Mg.

a) Calculate the theoretical yield of Ti in grams.

b) Calculate the percent yield if 7.91×10^6 g of Ti are actually obtained.



Example 3.17

Solution

Carry out two separate calculations to see which of the two reactants is the limiting reagent. First, starting with 3.54×10^7 g of $TiCl_4$, calculate the number of moles of Ti that could be produced if all the $TiCl_4$ reacted. The conversions are

$$\text{grams of TiCl}_4 \rightarrow \text{moles of TiCl}_4 \rightarrow \text{moles of Ti}$$

so that

$$\text{moles of Ti} = 3.54\times10_7\ \cancel{\text{g}}\ \cancel{\text{TiCl}_4} \times \frac{1\ \cancel{\text{mol}}\ \cancel{\text{TiCl}_4}}{189.7\ \cancel{\text{g}}\ \cancel{\text{TiCl}_4}} \times \frac{1\ \text{mol T}i}{1\ \cancel{\text{mol}}\ \cancel{\text{TiCl}_4}}$$

$$= 1.87\times10^5\ \text{mol Ti}$$

Example 3.17

Next, we calculate the number of moles of Ti formed from 1.13 × 10^7 g of Mg. The conversion steps are

$$\text{grams of Mg} \rightarrow \text{moles of Mg} \rightarrow \text{moles of Ti}$$

And we write

$$\text{moles of Ti} = 1.13 \times 10^7 \cancel{\text{g Mg}} \times \frac{1\ \cancel{\text{mol Mg}}}{24.31\ \cancel{\text{g Mg}}} \times \frac{1\ \text{mol Ti}}{2\ \cancel{\text{mol Mg}}}$$

$$= 2.32 \times 10^5\ \text{mol Ti}$$

Therefore, $TiCl_4$ is the limiting reagent because it produces a smaller amount of Ti.

The mass of Ti formed is

$$1.87 \times 10^5\ \cancel{\text{mol Ti}} \times \frac{47.88\ \text{g Ti}}{1\ \cancel{\text{mol Ti}}} = 8.95 \times 10^6\ \text{g Ti}$$

Example 3.17

Solution

The percent yield is given by

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

$$= \frac{7.91 \times 10^6\ \text{g}}{8.95 \times 10^6\ \text{g}} \times 100\%$$

$$= 88.4\%$$

Chemistry In Action: Chemical Fertilizers

Plants need: N, P, K, Ca, S, & Mg

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

$$NH_3\ (aq) + HNO_3\ (aq) \longrightarrow NH_4NO_3\ (aq)$$

fluorapatite

$$2Ca_5(PO_4)_3F\ (s) + 7H_2SO_4\ (aq) \longrightarrow$$

$$3Ca(H_2PO_4)_2\ (aq) + 7CaSO_4\ (aq) + 2HF\ (g)$$

The Oxidation Numbers of Elements in their Compounds

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 −1																	2 **He**
3 **Li** +1	4 **Be** +2											5 **B** +3	6 **C** +4 +2 −4	7 **N** +5 +4 +3 +2 +1 −3	8 **O** +2 $-\frac{1}{2}$ −1 −2	9 **F** −1	10 **Ne**
11 **Na** +1	12 **Mg** +2											13 **Al** +3	14 **Si** +4 −4	15 **P** +5 +3 −3	16 **S** +6 +4 +2 −2	17 **Cl** +7 +6 +5 +4 +3 +1 −1	18 **Ar**
19 **K** +1	20 **Ca** +2	21 **Sc** +3	22 **Ti** +4 +3 +2	23 **V** +5 +4 +3 +2	24 **Cr** +6 +5 +4 +3 +2	25 **Mn** +7 +6 +4 +3 +2	26 **Fe** +3 +2	27 **Co** +3 +2	28 **Ni** +2	29 **Cu** +2 +1	30 **Zn** +2	31 **Ga** +3	32 **Ge** +4 −4	33 **As** +5 +3 −3	34 **Se** +6 +4 −2	35 **Br** +5 +3 +1 −1	36 **Kr** +4 +2
37 **Rb** +1	38 **Sr** +2	39 **Y** +3	40 **Zr** +4	41 **Nb** +5 +4	42 **Mo** +6 +4 +3	43 **Tc** +7 +6 +4	44 **Ru** +8 +6 +4 +3	45 **Rh** +4 +3 +2	46 **Pd** +4 +2	47 **Ag** +1	48 **Cd** +2	49 **In** +3	50 **Sn** +4 +2	51 **Sb** +5 +3 −3	52 **Te** +6 +4 −2	53 **I** +7 +5 +1 −1	54 **Xe** +6 +4 +2
55 **Cs** +1	56 **Ba** +2	57 **La** +3	72 **Hf** +4	73 **Ta** +5	74 **W** +6 +4	75 **Re** +7 +6 +4	76 **Os** +8 +4	77 **Ir** +4 +3	78 **Pt** +4 +2	79 **Au** +3 +1	80 **Hg** +2 +1	81 **Tl** +3 +1	82 **Pb** +4 +2	83 **Bi** +5 +3	84 **Po** +2	85 **At** −1	86 **Rn**