Because learning changes everything.®

Chapter 3

Mass Relationships in Chemical Reactions

NASA

Atomic mass

Micro World
atoms & molecules → Macro World
grams

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

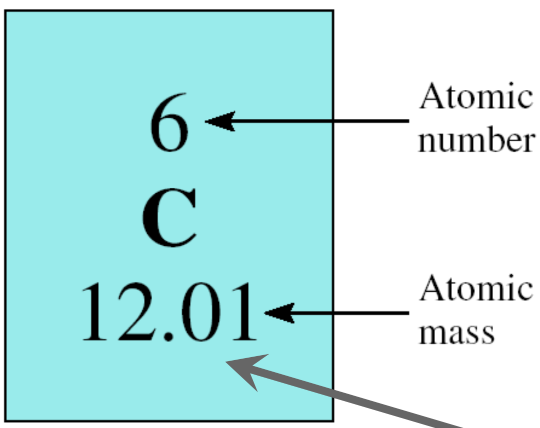
amu definition: the mass exactly equal to 1/12 the mass of one ^{12}C atom

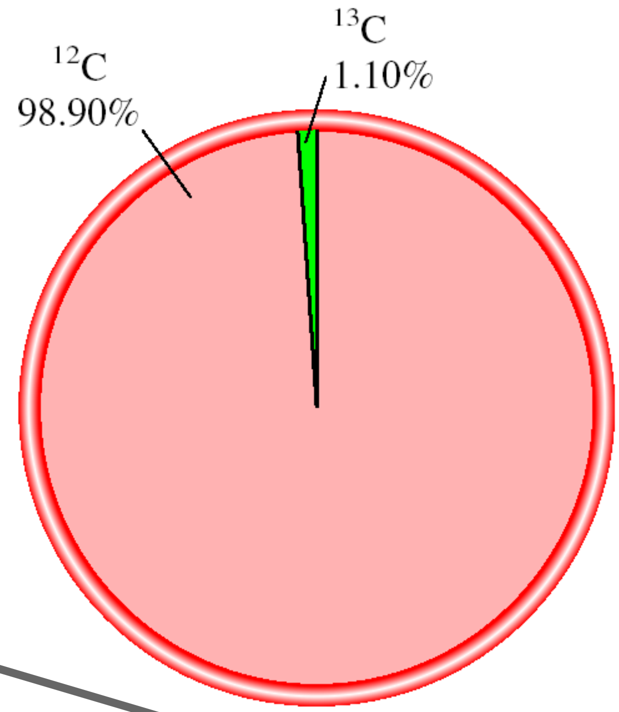
^{12}C = 6p, 6n = 12.00 amu $m_{e^-} = 0$

Experiment shows one atom ^{1}H = 8.400% of ^{12}C atom thus; mass of one atom ^{1}H = 1.008 amu
(12.00 x .08400)

^{16}O = 16.00 amu, ^{26}Fe = 55.85 amu

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





^{13}C = 13.00335 amu

average atomic mass of C =

(0.9890 x 12.00000 amu) + (0.0110 x 13.00335) = 12.01 amu

Naturally occurring lithium is:

7.42% ^{6}Li (6.015 amu)

92.58% ^{7}Li (7.016 amu)

Average atomic mass of lithium:

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

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, $^{63}_{29}Cu$ (69.09 percent) and $^{65}_{29}Cu$ (30.91 percent), are 62.93 amu and 64.9278 amu, respectively. Calculate the average atomic mass of copper. The relative abundances are given in parentheses.

Strategy Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of an isotope by its fractional abundance (not percent) will give the contribution to the average atomic mass of that particular isotope.

Solution First the percents are converted to fractions: 69.09 percent to 69.09/100 or 0.6909 and 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.

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

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	9 8B	10	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

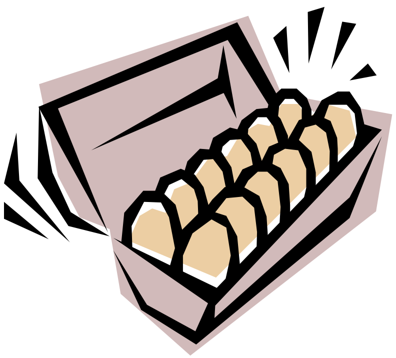
Metalloids

Nonmetals

The Mole

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

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

Avogadro's number (N_A)

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

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

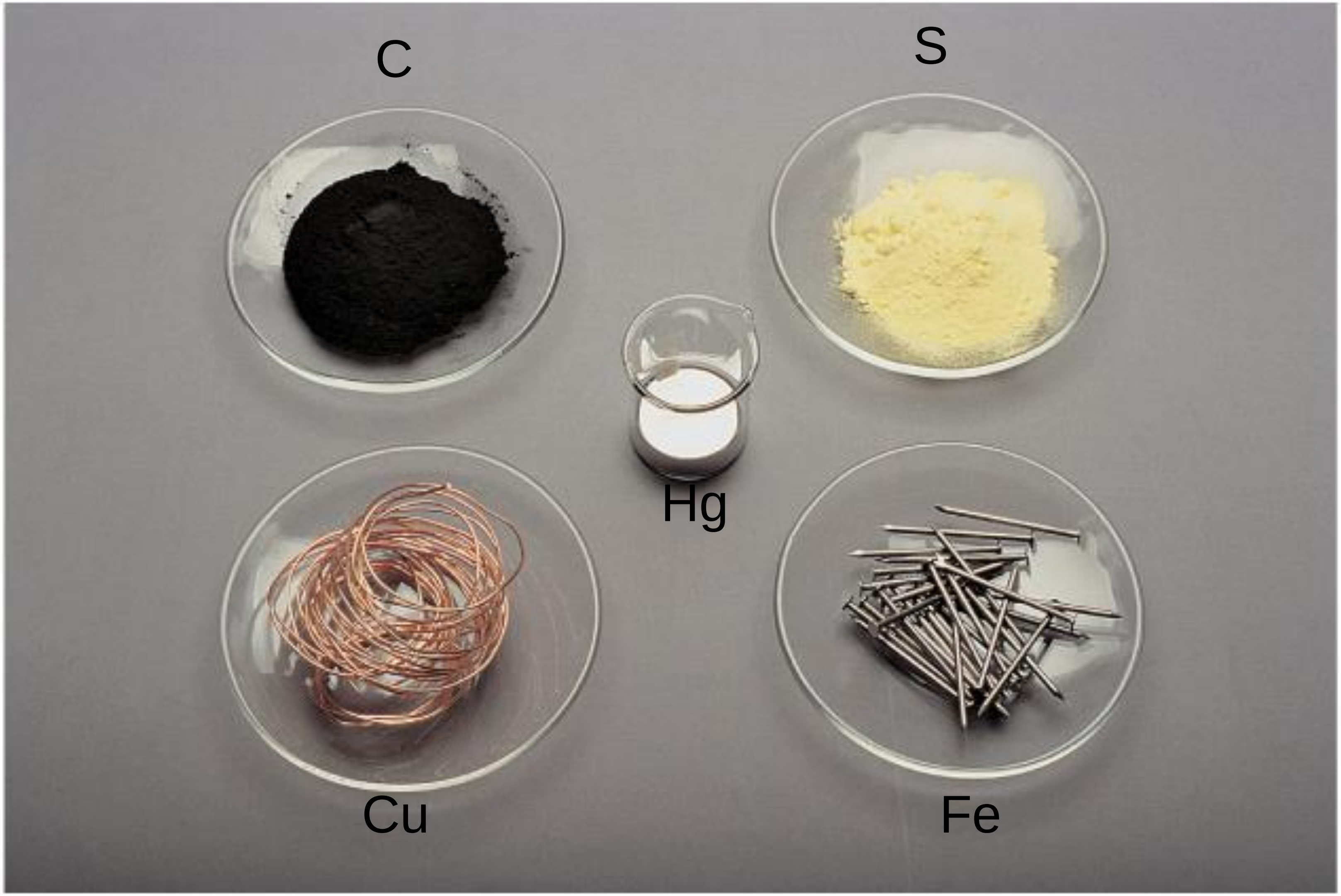
1 ^{12}C atom = 12.00 amu

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)

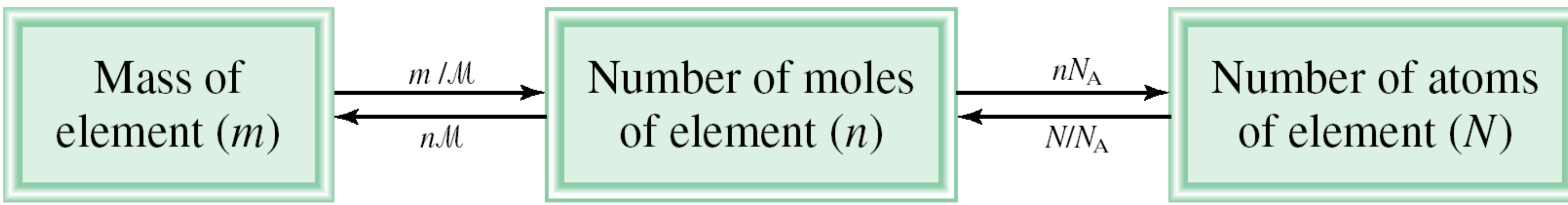
One Mole of:



Conversion Between Mass and Atoms

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

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



$\mathcal{M}$ = molar mass in (**g/mol**)

N_A = Avogadro's number

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

$$1 \text{ mol K} = 39.10 \text{ g K}$$

$$1 \text{ mol K} = 6.022 \times 10^{23} \text{ atoms K}$$

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

$$\mathbf{8.49 \times 10^{21} \text{ atoms K}}$$

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

$$\frac{1 \text{ mol He}}{4.003 \text{ g He}} \quad \text{and} \quad \frac{4.003 \text{ g He}}{1 \text{ mol He}}$$

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

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

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?

Solution The conversion factor needed to convert between moles and grams is the molar mass. In the periodic table (see inside front cover) we see the molar mass of Zn is 65.39 g. This can be expressed as

$$1 \text{ mol Zn} = 65.39 \text{ g Zn}$$

From this equality, we can write two conversion factors

$$\frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \quad \text{and} \quad \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}}$$

The conversion factor on the right is the correct one. Moles will cancel, leaving unit of grams for the answer. The number of grams of Zn is

$$0.356 \cancel{\text{mol Zn}} \times \frac{65.39 \text{ g Zn}}{1 \cancel{\text{mol Zn}}} = \boxed{23.3 \text{ g Zn}}$$

Thus, there are 23.3 g of Zn in 0.356 mole of Zn.

EXAMPLE 3.4

Sulfur (S) is a nonmetallic element that is present in coal. When coal is burned, sulfur is converted to sulfur dioxide and eventually to sulfuric acid that gives rise to the acid rain phenomenon. How many atoms are in 16.3 g of S?

Solution We need two conversions: first from grams to moles and then from moles to number of particles (atoms). The first step is similar to Example 3.2. Because

$$1 \text{ mol S} = 32.07 \text{ g S}$$

the conversion factor is

$$\frac{1 \text{ mol S}}{32.07 \text{ g S}}$$

Avogadro's number is the key to the second step. We have

$$1 \text{ mol} = 6.022 \times 10^{23} \text{ particles (atoms)}$$

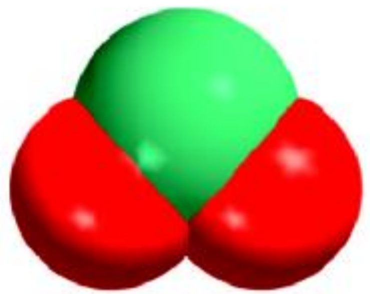
and the conversion factors are

$$\frac{6.022 \times 10^{23} \text{ S atoms}}{1 \text{ mol S}} \quad \text{and} \quad \frac{1 \text{ mol S}}{6.022 \times 10^{23} \text{ S atoms}}$$

$$16.3 \cancel{\text{g S}} \times \frac{1 \cancel{\text{mol S}}}{32.07 \cancel{\text{g S}}} \times \frac{6.022 \times 10^{23} \text{ S atoms}}{1 \cancel{\text{mol S}}} = \boxed{3.06 \times 10^{23} \text{ S atoms}}$$

Molecular Mass

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

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?

Solution The conversion factor needed to convert between grams and moles is the molar mass. First we need to calculate the molar mass of CH_4, following the procedure in Example 3.5:

$$\begin{aligned} \text{molar mass of } CH_4 &= 12.01 \text{ g} + 4(1.008 \text{ g}) \\ &= 16.04 \text{ g} \end{aligned}$$

Because

$$1 \text{ mol } CH_4 = 16.04 \text{ g } CH_4$$

the conversion factor we need should have grams in the denominator so that the unit g will cancel, leaving the unit mol in the numerator:

$$\frac{1 \text{ mol } CH_4}{16.04 \text{ g } CH_4}$$

We now write

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

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

How many H atoms are in 72.5 g of C_3H_8O?

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

$$1 \text{ mol } C_3H_8O \text{ molecules} = 8 \text{ mol H atoms}$$

$$1 \text{ mol H} = 6.022 \times 10^{23} \text{ atoms H}$$

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

$$\mathbf{5.82 \times 10^{24} \text{ atoms H}}$$

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.

Solution To calculate the number of H atoms, we first must convert grams of urea to moles of urea using the molar mass of urea. This part is similar to Example 3.2. The molecular formula of urea shows there are four moles of H atoms in one mole of urea molecule, so the mole ratio is 4:1. Finally, knowing the number of moles of H atoms, we can calculate the number of H atoms using Avogadro's number. We need two conversion factors: molar mass and Avogadro's number. We can combine these conversions

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

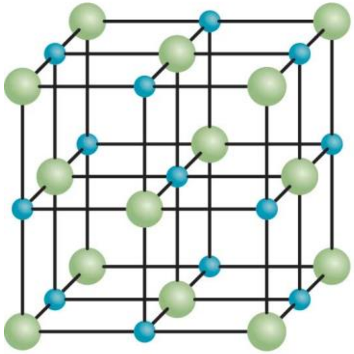
into one step:

$$25.6\ \cancel{\text{g }(\text{NH}_2)_2\text{CO}} \times \frac{1\ \cancel{\text{mol }(\text{NH}_2)_2\text{CO}}}{60.06\ \cancel{\text{g }(\text{NH}_2)_2\text{CO}}} \times \frac{4\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol }(\text{NH}_2)_2\text{CO}}} \times \frac{6.022 \times 10^{23}\ \text{H atoms}}{1\ \cancel{\text{mol H}}}$$

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

Formula Mass

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



	1 Na	22.99 amu
NaCl	1 Cl	<u>+35.45 amu</u>
	NaCl	58.44 amu

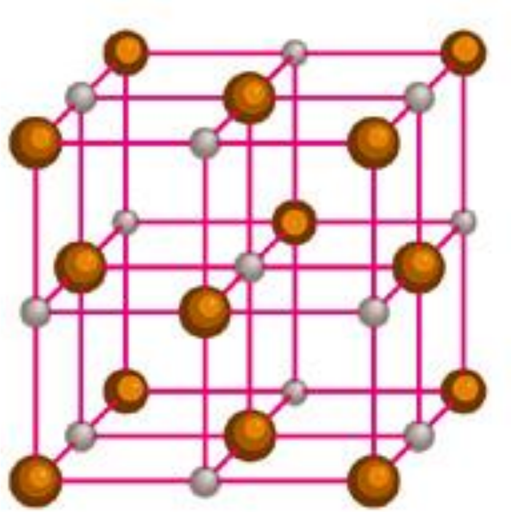
For any ionic compound

formula mass (amu) = molar mass (grams)

1 formula unit NaCl = 58.44 amu

1 mole NaCl = 58.44 g NaCl

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



Cl

	22.99 amu
1Cl	+ 35.45 amu
NaCl	58.44 amu

$$\boxed{\begin{array}{c}\text{For any ionic compound} \\ \text{formula mass (amu) = molar mass (grams)}\end{array}}$$

1 formula unit NaCl = 58.44 amu

1 mole NaCl = 58.44 g NaCl

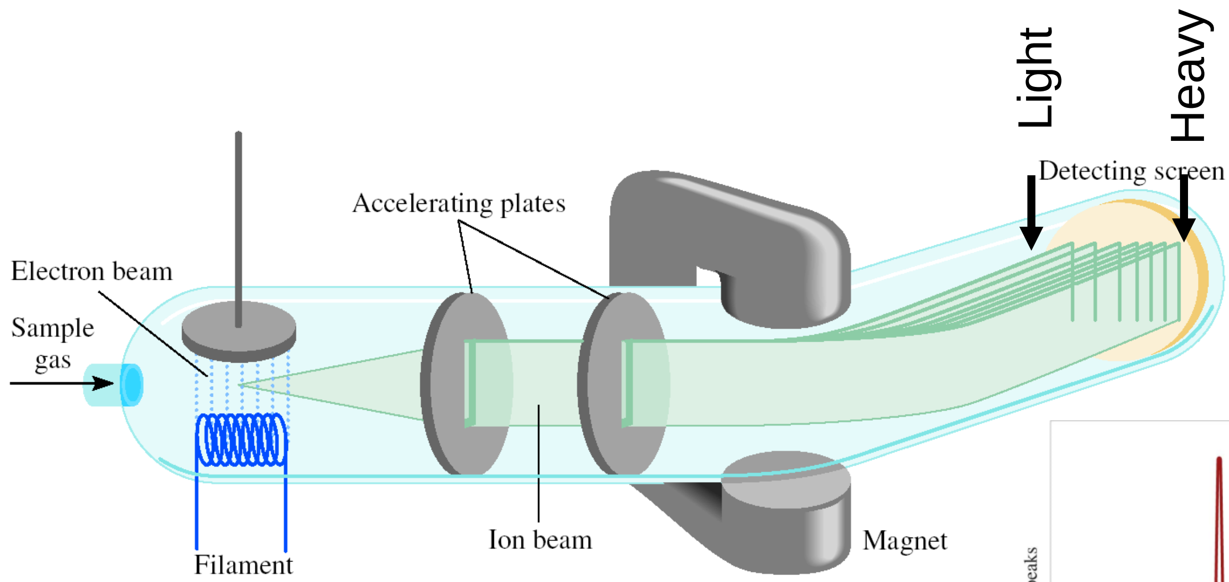
What is the formula mass of $Ca_3(PO_4)_2$?

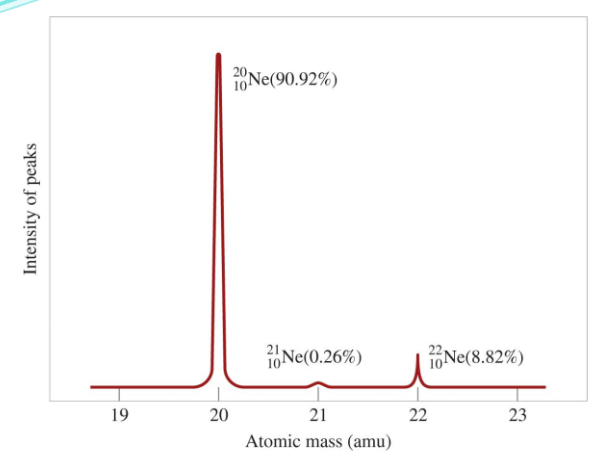
1 formula unit of $Ca_3(PO_4)_2$

3 Ca		3 x 40.08
2 P		2 x 30.97
8 O	+	8 x 16.00
		310.18 amu

Mass Spectrometer

The most direct and most accurate method for determining atomic and molecular masses is mass spectrometry



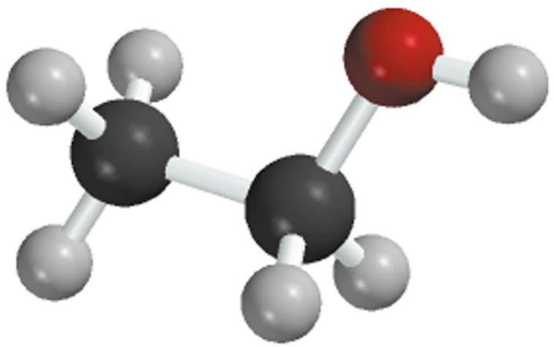


Mass Spectrum of Ne

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%**

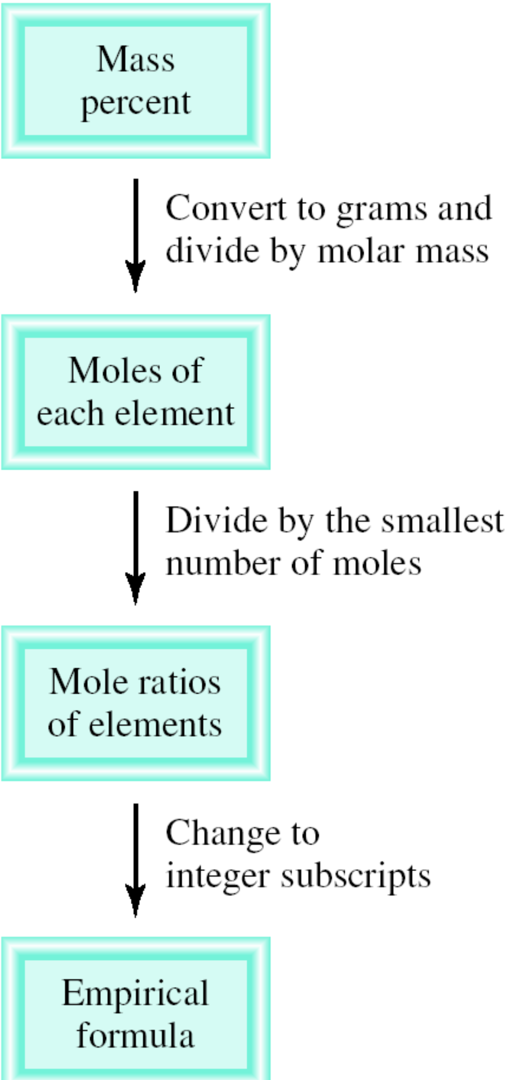
Examples

What is the mass of H ,Cl in 10 g HCl?

What is % composition of the elements C in CH_3COOH?

What is % composition of the elements in 25.00 g H_2SO_4 if $m_H = 0.5142$ g and $m_O = 16.3239$ g and $m_S = 8.1619$ g ?

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

Mass percent

↓ Convert to grams and divide by molar mass

Moles of each element

↓ Divide by the smallest number of moles

Mole ratios of elements

↓ Change to integer subscripts

Empirical formula

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

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

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

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

$$KMnO_4$$

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

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

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

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

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

$$1.33 \times 1 = 1.33$$
$$1.33 \times 2 = 2.66$$
$$1.33 \times 3 = 3.99 \approx 4$$

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.

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.

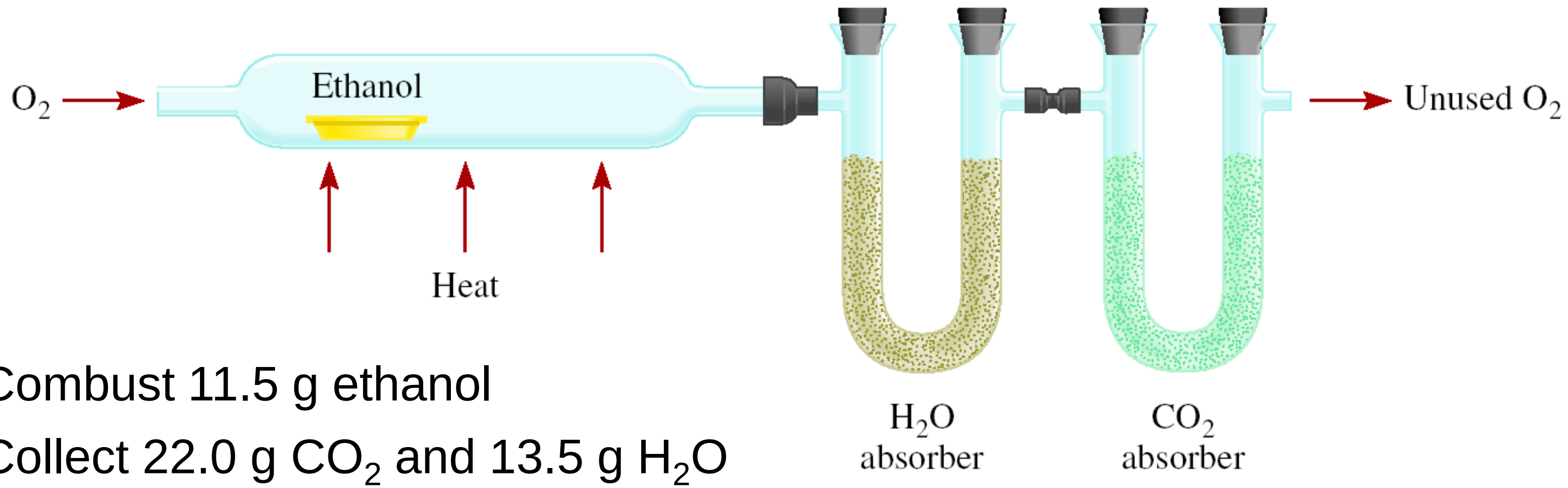
Strategy Chalcopyrite is composed of Cu, Fe, and S. The mass due to Cu is based on its percentage by mass in the compound. How do we calculate mass percent of an element?

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

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

To calculate the mass of Cu in a 3.71×10^3 kg sample of $CuFeS_2$, we need to convert the percentage to a fraction (that is, convert 34.63 percent to 34.63/100, or 0.3463) and write

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



Combust 11.5 g ethanol

Collect 22.0 g CO_2 and 13.5 g H_2O

g CO_2 ⟶ mol CO_2 ⟶ mol C ⟶ g C   6.0 g C = 0.5 mol C

g H_2O ⟶ mol H_2O ⟶ mol H ⟶ 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

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 3.11

A sample of a compound contains 1.52 g of nitrogen (N) and 3.47 g of oxygen (O). The molar mass of this compound is between 90 g and 95 g. Determine the molecular formula and the accurate molar mass of the compound.

Solution We are given grams of N and O. Use molar mass as a conversion factor to convert grams to moles of each element. Let n represent the number of moles of each element. We write

$$n_N = 1.52 \cancel{\text{g N}} \times \frac{1 \text{ mol N}}{14.01 \cancel{\text{g N}}} = 0.108 \text{ mol N}$$

$$n_O = 3.47 \cancel{\text{g O}} \times \frac{1 \text{ mol O}}{16.00 \cancel{\text{g O}}} = 0.217 \text{ mol O}$$

Thus, we arrive at the formula $N_{0.108}O_{0.217}$, which gives the identity and the ratios of atoms present. However, chemical formulas are written with whole numbers. Try to convert to whole numbers by dividing the subscripts by the smaller subscript (0.108). After rounding off, we obtain NO_2 as the empirical formula.

$$\text{empirical molar mass} = 14.01 \text{ g} + 2(16.00 \text{ g}) = 46.01 \text{ g}$$

Next, we determine the ratio between the molar mass and the empirical molar mass

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

The molar mass is twice the empirical molar mass. This means that there are two NO_2 units in each molecule of the compound, and the molecular formula is $(NO_2)_2$ or N_2O_4.

The actual molar mass of the compound is two times the empirical molar mass, that is, 2(46.01 g) or 92.02 g, which is between 90 g and 95 g.

Chemical Reaction

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

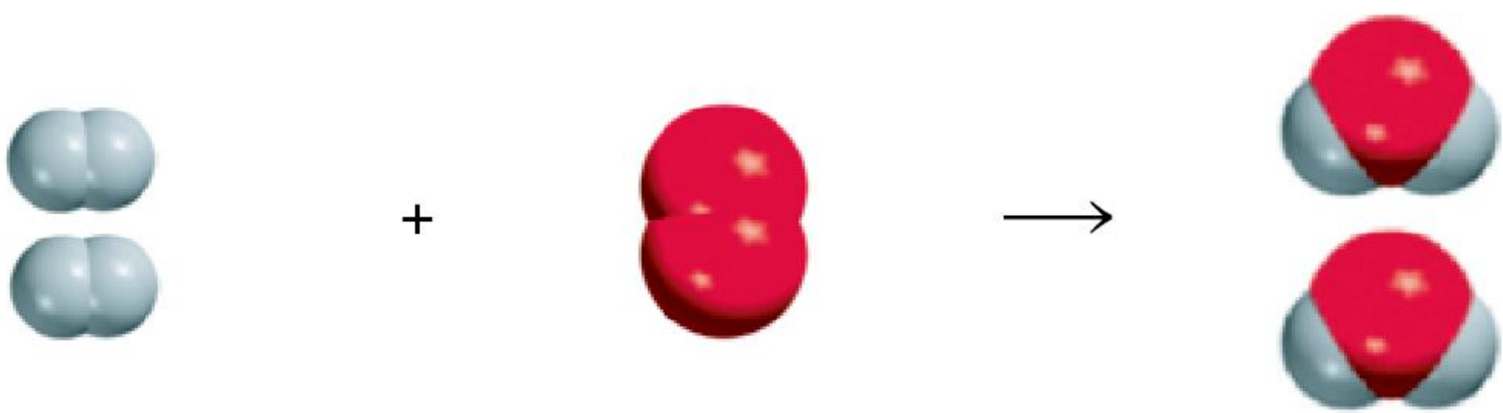
A ***chemical equation*** uses chemical symbols to show what happens during a chemical reaction

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

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

multiply CO_2 by 2

6 hydrogen on left

2 hydrogen on right

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

multiply H_2O by 3

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

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

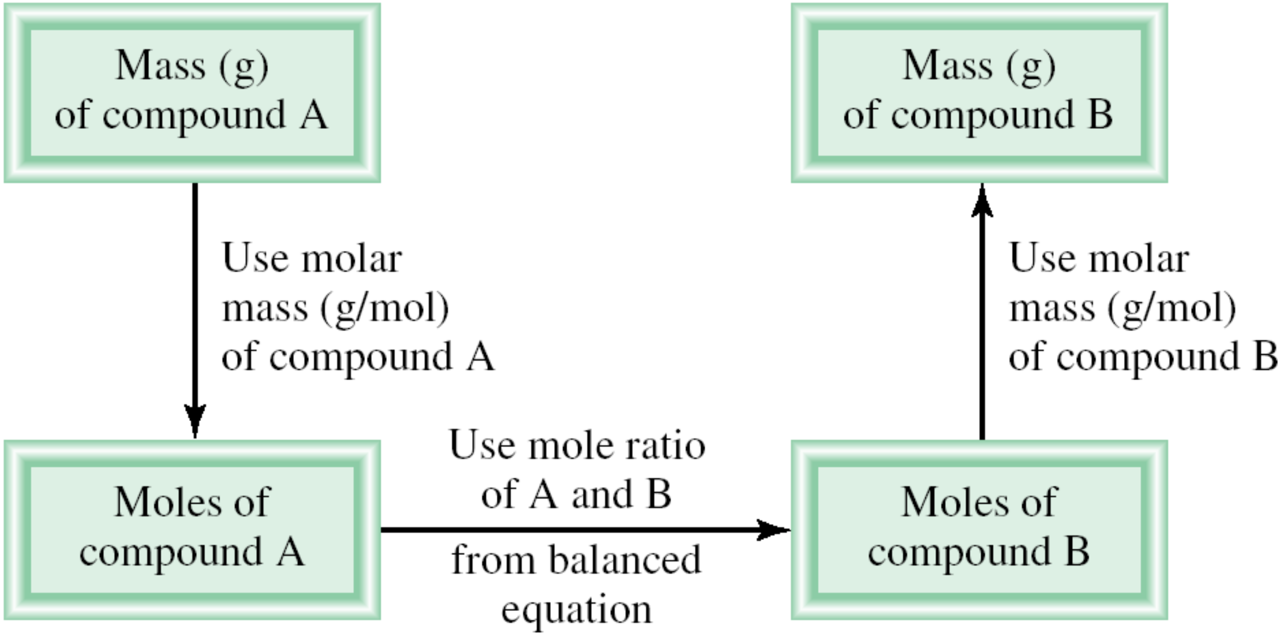
Balancing Chemical Equations

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

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

Reactants	Products
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)

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

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

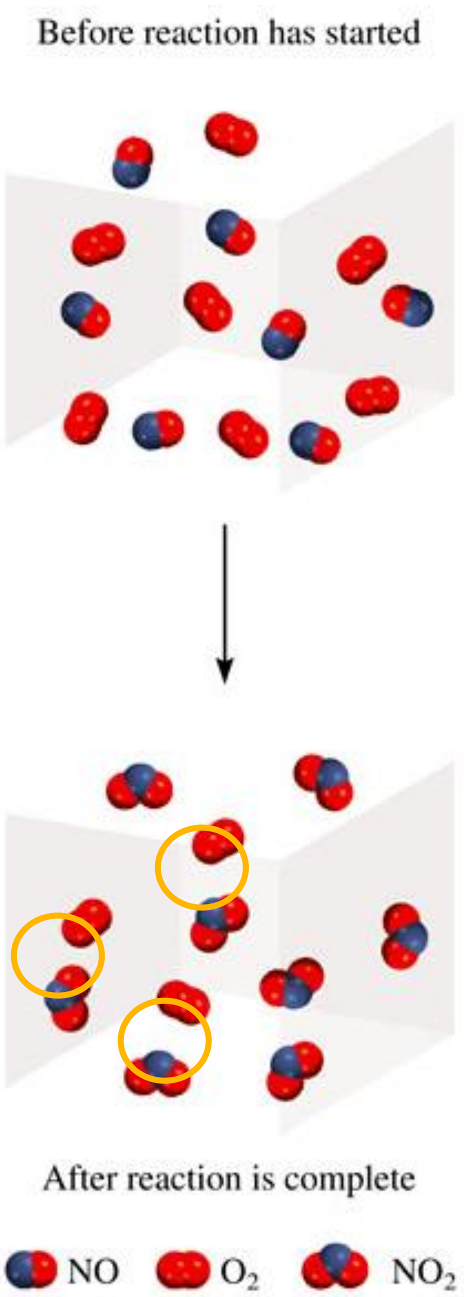
If 209 g of methanol are used up in the combustion, what mass of water is produced?

$$\text{grams } CH_3OH \longrightarrow \text{moles } CH_3OH \longrightarrow \text{moles } H_2O \longrightarrow \text{grams } H_2O$$

molar mass CH_3OH | coefficients chemical equation | molar mass H_2O

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

235 g H_2O



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

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

$$124\ \cancel{\text{g Al}} \times \frac{1\ \cancel{\text{mol Al}}}{27.0\ \cancel{\text{g Al}}} \times \frac{1\ \cancel{\text{mol Fe}_2\text{O}_3}}{2\ \cancel{\text{mol Al}}} \times \frac{160.\ \text{g Fe}_2\text{O}_3}{1\ \cancel{\text{mol Fe}_2\text{O}_3}} = 367\ \text{g Fe}_2\text{O}_3$$

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

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

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

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

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

$$124\ \cancel{\text{g Al}} \times \frac{1\ \cancel{\text{mol Al}}}{26.98\ \cancel{\text{g Al}}} \times \frac{1\ \cancel{\text{mol } Al_2O_3}}{2\ \cancel{\text{mol Al}}} \times \frac{102.0\text{ g } Al_2O_3}{1\ \cancel{\text{mol } Al_2O_3}} = \mathbf{234.4\ g\ Al_2O_3}$$

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

Another method

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

124 g
4.596 mol

601 g
3.7633 mol

$$\frac{4.599 \text{ mol Al}}{2 \text{ mol Al}} \qquad \frac{3.763 \text{ mol Al}}{1 \text{ mol Al}}$$

2.300

3.763

Al is the **least** thus it is **the limiting reagent**

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

g Al $\longrightarrow$ mol Al $\longrightarrow$ mol Al_2O_3 $\longrightarrow$ g Al_2O_3

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

$$124\ \cancel{\text{g Al}} \times \frac{1\ \cancel{\text{mol Al}}}{26.98\ \cancel{\text{g Al}}} \times \frac{1\ \cancel{\text{mol } Al_2O_3}}{2\ \cancel{\text{mol Al}}} \times \frac{102.0\text{ g } Al_2O_3}{1\ \cancel{\text{mol } Al_2O_3}} = \mathbf{234.4\ g\ Al_2O_3}$$

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.

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

Calculate % yield if 803 g of CaO is produced?

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

1.96×10^3 g 1.54×10^3 g   ? g

48.90 mol   8.467 mol $\longrightarrow$ 42.34 mol

$\frac{48.90 \text{ mol Ca}}{5 \text{ mol Ca}}$ $\frac{8.467 \text{ mol } V_2O_5}{1 \text{ mol } V_2O_5}$ 2374.2 g CaO

9.78   8.467

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

EXAMPLE 3.16

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:

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

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.

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

$$\text{grams of } TiCl_4 \longrightarrow \text{moles of } TiCl_4 \longrightarrow \text{moles of Ti}$$

$$\text{moles of Ti} = 3.54 \times 10^7 \cancel{\text{g } TiCl_4} \times \frac{1 \cancel{\text{mol } TiCl_4}}{189.7 \cancel{\text{g } TiCl_4}} \times \frac{1 \text{ mol Ti}}{1 \cancel{\text{mol } TiCl_4}}$$
$$= 1.87 \times 10^5 \text{ mol Ti}$$

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

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

$$\text{moles of Ti} = 1.13 \times 10^7 \cancel{\text{g Mg}} \times \frac{1 \cancel{\text{mol Mg}}}{24.31 \cancel{\text{g Mg}}} \times \frac{1 \text{ mol Ti}}{2 \cancel{\text{mol Mg}}}$$
$$= 2.32 \times 10^5 \text{ mol Ti}$$

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

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

(Continued)

Solution The percent yield is given by

$$\begin{aligned}\%\text{yield} &= \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \\ &= \frac{7.91 \times 10^6 \text{ g}}{8.95 \times 10^6 \text{ g}} \times 100\% \\ &= \boxed{88.4\%}\end{aligned}$$