3-1

Chemical Stoichiometry

$$Sn + O_2 \rightarrow SnO_2$$

Stoichiometry - The study of quantities of materials consumed and produced in chemical reactions.

AKA mole ratio

3-2

The Mole

The number equal to the number of carbon atoms in exactly 12 grams of pure ^{12}C.

1 mole of anything = 6.022×10^{23} units of that thing

Avogadro's number = 6.022×10^{23}

3-3

Atomic Masses and Atomic Weight

As discussed in chapter two, elements occur in nature as mixtures of isotopes:

Carbon = 98.89% ^{12}C
1.11% ^{13}C
<0.01% ^{14}C

Carbon atomic weight = 12.011 amu

3-4

Molar Mass

- molar mass (molecular weight) is the mass, in grams, of one mole of the compound.
- Sometimes called the formula weight (F.W.)

3-5

Molar Mass

Use *formula weight* to determine *molar mass*

CH_4

1 C = 12.011 amu = 12.011 amu
1 H = 1.0079 amu = 4.032 amu
formula weight = 16.043 amu

molar mass = 16.043 g/mol

3-6

Percent Composition

- Mass percent of an element:

$$\text{mass \%} = \frac{\text{mass of element in compound}}{\text{total mass of compound}} \times 100\%$$

3-7

For Iron(III) Oxide, (Fe_2O_3)

$$\text{mass \% Fe} = \frac{55.847 \times 2}{(55.847 \times 2) + (15.9994 \times 3)} \times 100\%$$

$$= \frac{111.694}{159.6922} \times 100\% = 69.94\%$$

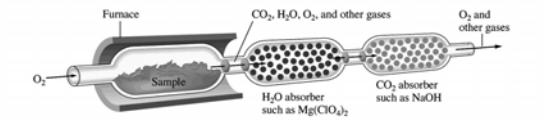
$$\text{mass \% O} = \frac{15.9994 \times 3}{(55.847 \times 2) + (15.9994 \times 3)} \times 100\%$$

$$= \frac{47.9982}{159.6922} \times 100\% = 30.05\%$$

3-8

Determining the Formula of a Compound

Figure 3.5 Zumdahl and Zumdahl, 6th Ed.



- If other elements contained, must have way of measuring them quantitatively

3-9

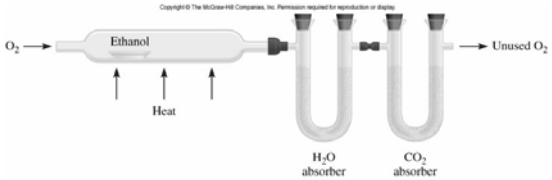
Reiteration of Empirical and Molecular Formulas

- molecular formula = (empirical formula)$_n$
  [n = integer]
- e.g., molecular formula = C_6H_6 = $(CH)_6$
- empirical formula = CH

3-10

Determining the Formula of a Compound

Figure 3.6 Chang, 9th Ed.

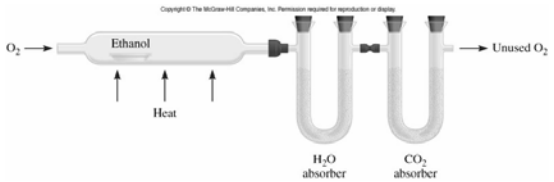


- example:
- empirical formula CH_5N
- If molar mass is 62, what is molecular formula?

3-10

Determining the Formula of a Compound

Figure 3.6 Chang, 9th Ed.



- example:
- empirical formula CH_5N
- If molar mass is 62, what is molecular formula?

3-11

Sample Test Question (STQ)

Empirical Formula Determination

What is the empirical formula for a compound that contains 70.1% silver, 6.90% carbon and 23.0% oxygen?

1. Base calculation on 100 grams of compound.
2. Determine moles of each element in 100 grams of compound.
3. Divide each value of moles by the smallest of the values.
4. Multiply each number by an integer to obtain all whole numbers.

Don't Round off Until Your Final Answer!
This is an actual example from Chem I

A compound of mercury and iodine was analyzed and found to contain 2.65 g of mercury and 3.35 g of iodine. Determine the formula of the compound.

$2.65\text{g Hg} \times \frac{1\text{ mol}}{200.6} = 0.01\text{ mol}$

A student wrote: $3.35\text{g I} \times \frac{1\text{ mol}}{127.0\text{ g}} = 0.03\text{ mol}$, yielding a formula of HgI_3.

However, $2.65\text{g Hg} \times \frac{1\text{ mol}}{200.6\text{ g}} = 0.01321\text{ mol}$, and furthermore

$3.35\text{g I} \times \frac{1\text{ mol}}{127.0\text{ g}} = 0.02640\text{ mol}$

$\frac{0.02640\text{ mol}}{0.01321\text{ mol}} = 1.998$

so the correct formula is HgI_2

3-12

Reading Chemical Equations

Reactant(s) $\rightarrow$ Product(s)

Name compounds by the rules we learned in chapter 2

3-13

Chemical Equations

Chemical change involves a reorganization of the atoms in one or more substances.

3-14

Chemical Equations: *Coefficients*

$2\ K + Cl_2 \rightarrow 2\ KCl$

2 K denotes how many react.

The 2 relates to everything which follows in the compound.

3-15

Chemical Equations: Subscripts and Superscripts

$2\ K + Cl_2 \rightarrow 2\ KCl$

Cl_2 denotes 2 atoms of chlorine in a chlorine molecule

Subscripts refer only to the element immediately preceding

$MgO + H_2O \rightarrow Mg^{2+} + 2\ OH^-$

Superscripts will denote the charge on an ion

3-16

Chemical Equations: State

$2\ K_{(s)} + Cl_2\ (g) \rightarrow 2KCl\ (s)$

(s)(l)(g)(aq) denote the *state* of the molecules

aq = aqueous = dissolved in *water*

3-17

Chemical Equations Must Be Balanced

There must be an equal number of atoms of each element on both sides of the equation.

$2\ K(s) + Cl_2(g) \rightarrow KCl(s)$ is not balanced

$2\ K(s) + Cl_2(g) \rightarrow 2\ KCl(s)$ is balanced

3-18

Balancing Equations by Inspection

Don't change formulas of molecules or subscripts.

Don't forget that if you use a coefficient, all of the atoms in the molecule are increased.

Consider the substance with the most <u>atoms</u> first

3-19

Chemical Reactions

$CH_4 + O_2 \rightarrow CO_2 + H_2O$

$Fe + O_2 \rightarrow Fe_2O_3$

$C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$

$XeF_2 + H_2O \rightarrow Xe + HF + H_3PO_4$

$SiCl_4 + H_2O \rightarrow H_4SiO_4 + HCl$

3-20

Calculating Masses of Reactants and Products

1. Balance the equation.
2. Convert mass to moles.
3. Set up mole ratios.

Stoichiometry is an exact number

4. Use mole ratios to calculate moles of desired substituent.
5. Convert moles to grams, if necessary.

3-21

Sample Test Questions

Using the following balanced chemical reaction, calculate the theoretical yield of carbon dioxide if you burn 1.25 g C_4H_{10} in excess oxygen.

$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$

3-22

Limiting Reactant

The limiting reactant is the reactant that is consumed first, limiting the amount of product(s) formed.

AKA limiting reagent

3-23

Sample Test Question

Using the following equation, identify the limiting reagent if you burn 1.25 g C_4H_{10} in 16.9 g of O_2.

$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$

3-24

Percent Yield

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

3-25

Sample Test Question

Calculate the percent yield if you begin with 1.25 g C_4H_{10} in 16.9 g of O_2, and end up with 3.32 g of CO_2.

$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$

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Chemical Reactions and Energy Changes

All chemical reactions involve a change in energy: heat, light, electricity, *etc*.

Endothermic reaction- take up heat from surroundings (cold to the touch).

Reactants are lower energy state than products.

Exothermic reaction- produce heat (hot to the touch). Products are lower energy state than reactants.