

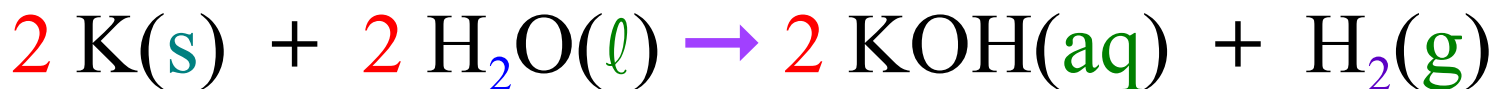
Chapter 3

Chemical Equations & Reaction Stoichiometry

I) Chemical Equations

Symbolic representation of
a chemical reaction

potassium + water \rightarrow potassium hydroxide + hydrogen



Coefficients

indicate the number of atoms, molecules
or formula units of each substance
involved in the rxn.

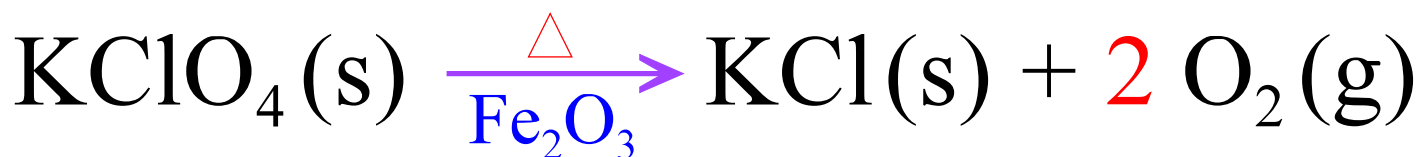
2 atoms of potassium react w.
2 molecules of water to produce
2 f.u. of potassium hydroxide &
1 molecule of hydrogen

(s) ≡ solid

(l) ≡ liquid

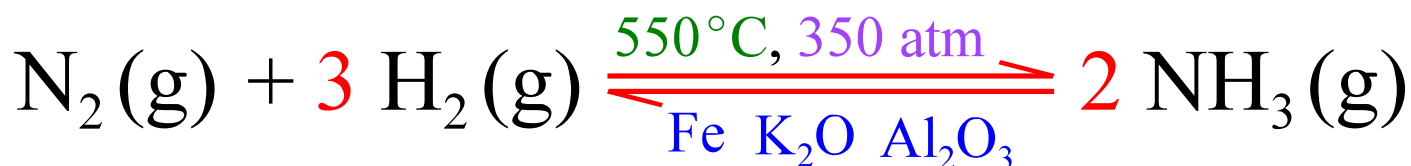
(g) ≡ gas

(aq) ≡ aqueous soln
(in H₂O)



Δ ≡ heat

Fe_2O_3 ≡ catalyst (makes rx happen faster)



II) Balancing Chemical Eqn.

Law of Conservation of Mass

- mass neither created nor destroyed
- just rearrangement of atoms

atoms of each element in reactants \equiv # atoms of each element in products

Balance atoms \Rightarrow balance mass

A) General Method for Balancing Equations

Requirements

1. Correct chemical formulas must be used for all reactants and products.
2. The number of atoms of each element in the reactants must equal the number of atoms of each element in the products.
3. Any charge on the left must equal any charge on the right.
4. Only the smallest whole number coefficients are acceptable.

Guidelines

1. Disregarding hydrogen, oxygen, and polyatomic ions, find the molecule containing the largest number of atoms of a single element. Balance that element first.
2. Balance polyatomic ions as a unit, if they remain unchanged.
3. Balance hydrogen and oxygen last. If either appears in elemental form, it is balanced last.
4. Check to see that all atoms are balanced and the smallest whole number coefficients are used.

B) Ex 1 : Solid potassium nitrate decomposes when heated to produce solid potassium nitrite & oxygen gas

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C) Ex 2 : Write the balanced eqn.
for the production of acetylene, C_2H_2 ,
from calcium carbide, CaC_2 , & water.



D) Ex 3 : Write the balanced eqn.
for the **combustion** of acetylene.

E) Ex 4 : Calcium oxide reacts w.
phosphoric acid to produce calcium
phosphate and water.

III) Simple Patterns of Chemical Reactivity

Classify reactions by **general**
type and **predict** products

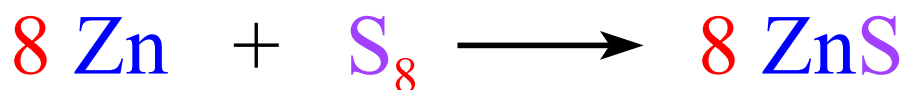
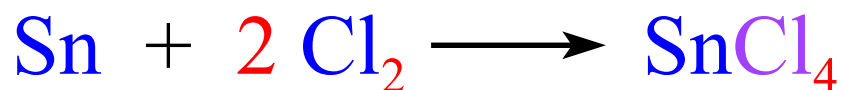
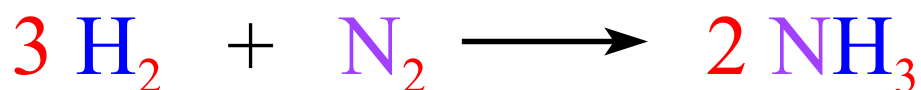
Note:

Elements in the **same**
group tend to undergo
similar reactions

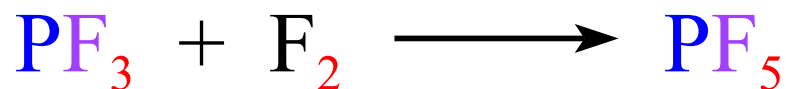
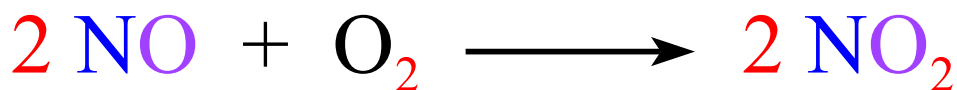
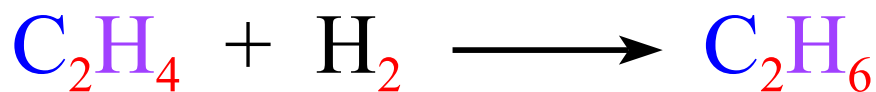
A) Combination Reactions

2 or more reactants combine to give 1 product.

1) Direct combination of Elements



2) Combination of Cmpd. & Element

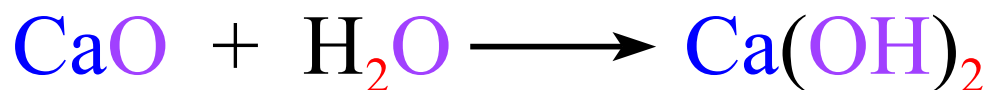
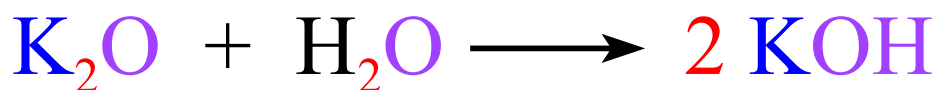


3) Rx. of oxides w. water

a) Metal Oxides

Basic oxides

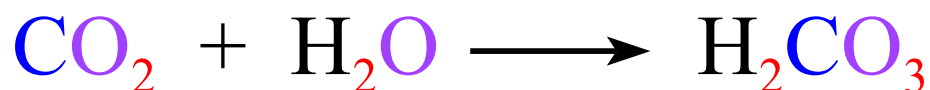
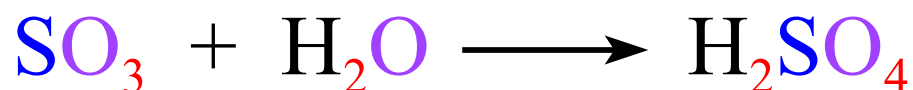
- produce basic metal hydroxides



b) Nonmetal oxides

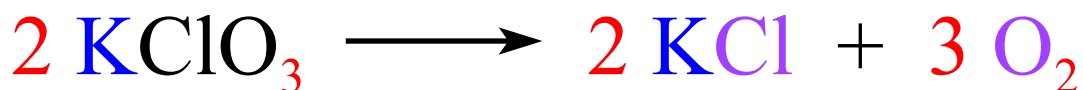
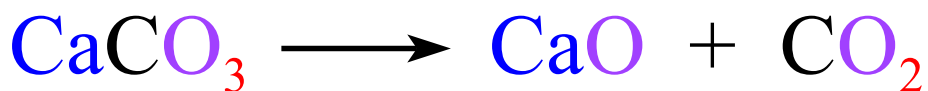
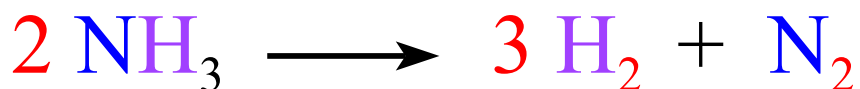
Acidic oxides

- produce acids



B) Decomposition Reactions

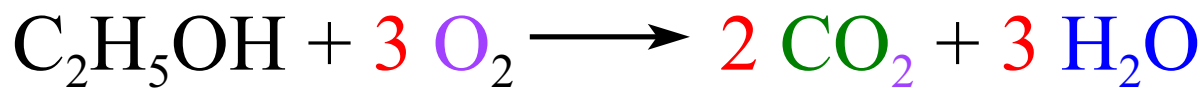
Single compd. breaks down into
2 or more simpler substances



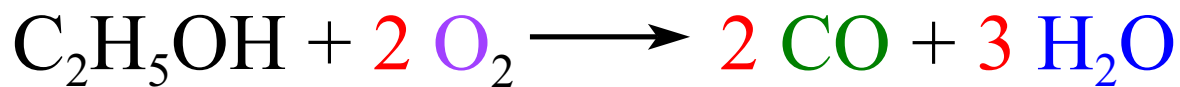
C) Combustion

Reaction w. O_2

1) Complete Combustion



2) Incomplete Combustion



IV) Molecular & Formula Weights

A) Molecular Weights

Sum of the **atomic** weights of the **atoms** in the **molecule**

1) Ex 1: Find the **M.W.** of the ethyl alcohol (ethanol), C_2H_6O

B) Formula Weights

Used for ionic substances

- consists of **formula units**,
NOT molecules

Formula Wt.

Sum of the **atomic weights** of the **atoms** as given in the **formula**

- 1) Ex 1: What is the **F.W.** of
 $(\text{NH}_4)_2\text{CO}_3$?

C) Percent Composition

$$\text{mass \% of an element} = \frac{\text{mass of element}}{\text{Total mass}} \times 100\%$$

- 1) Ex 1: Determine the mass % comp. of copper (II) nitrate.

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V) Avogadro's Number & the Mole

How do you weigh out the same number of items?

If we weigh quantities in ratios of weights of individual items, we obtain equal numbers of items.

A) The Mole

A.W. of H
1.008 amu

A.W. of C
12.011 amu

1 atom of H
1.008 amu

1 atom of C
12.011 amu

10 atoms of H
10.08 amu

10 atoms of C
120.11 amu

X atoms of H
1.008 g

Y atoms of C
12.011 g

$$X = Y = 6.022 \times 10^{23} \text{ atoms}$$

Avogadro's Number, N_A

The unit for a **very** large number of **particles** is

Mole

1 mole = **6.02** x **10²³** **particles**

1 mole C = **6.02** x **10²³** **C atoms** = **12.011g C**

1) **Molar Mass**

Mass in **grams** numerically equal to **A.W.**, **M.W.**, or **F.W.**

a) A given **AW** tells you:

1) **avg. mass** of a **single atom**;
amu

2) mass of a **mole** of **atoms**;
grams / mole

2) Apply to Molecules & f.u.

$$\begin{aligned} 1 \text{ mol C}_2\text{H}_6\text{O} &= 6.02 \times 10^{23} \text{ molecules C}_2\text{H}_6\text{O} \\ &= 46.08 \text{ g C}_2\text{H}_6\text{O} \end{aligned}$$

$$\begin{aligned} 1 \text{ mol (NH}_4\text{)}_2\text{CO}_3 &= 6.02 \times 10^{23} \text{ (NH}_4\text{)}_2\text{CO}_3 \text{ f.u.} \\ &= 96.11 \text{ g (NH}_4\text{)}_2\text{CO}_3 \end{aligned}$$

Note: 1 mol of $(\text{NH}_4)_2\text{CO}_3$ contains:

2 x (1 mol) NH_4^+ ions

2 x (6.02×10^{23}) NH_4^+ ions

8 mol H atoms

B) Calculations

1) Ex 1: How many **moles** of He are
in 40.0 **g** of He?

1 mol He = **6.02** x 10^{**23**} He **atoms** = **4.00 g** He

2) Ex 2: How many **grams** of $(\text{NH}_4)_2\text{S}$ are required to obtain 0.50 **mol** of NH_4^+ ?

a) Need FW

$$\underline{\quad} \text{ N x } 14.01 \text{ amu} =$$

$$\underline{\quad} \text{ H x } 1.01 \text{ amu} =$$

$$\underline{\quad} \text{ S x } 32.06 \text{ amu} = \underline{\hspace{10em}}$$

$$1 \text{ mol } (\text{NH}_4)_2\text{S} =$$

$$1 \text{ mol } (\text{NH}_4)_2\text{S} =$$

3) Ex 3: Typically smog contains about 0.040 g CO per m³ of air. How many molecules of CO are in a m³ of air?

MW of CO =

$$1 \text{ mol} \left\{ \begin{array}{l} \text{atoms} \\ \text{molecules} \\ \text{f.u.} \end{array} \right\} = \left\{ \begin{array}{l} \text{AW (g)} \\ \text{MW (g)} \\ \text{FW (g)} \end{array} \right\} = 6.02 \times 10^{23} \left\{ \begin{array}{l} \text{atoms} \\ \text{molecules} \\ \text{f.u.} \end{array} \right\}$$

VI) Empirical & Molecular Formulas

A) Molecular Formula

Actual numbers and kinds of atoms in a molecule

C_6H_6 Benzene

C_2H_5OH Ethanol

B) Empirical Formula

Relative number of atoms of each kind in a molecule

- smallest whole-number ratio of atoms

C_1H_1 Benzene or acetylene

Subscripts in a molecular formula are always some integer multiple of subscripts in empirical formula

C) Procedure for Determining E.F.

1) Express composition in **grams**.

If % comp. given,

assume 100 g sample

2) Determine **# moles** of **each** element

3) **Divide** by **smallest # moles** to obtain **mole ratio** - this is also the **atom ratio**

4) **If needed: Multiply** by **simplest factor** to get **whole** numbers

5) Write the formula

6) Ex 1: A 10.45 g sample of Bi combines w. oxygen to produce 11.65 g of a bismuth oxide.

Determine the E.F. of the oxide.

a) Step 1: determine mass of oxygen

b) Step 2: Convert to moles

c) Step 3: Determine mole ratio
Divide by smallest # moles

d) Step 4: Multiply by factor to get
whole numbers

Bi:

O:

e) Step 5: Write formula

D) Molecular Formula Determination

Molecular formula is **always** some **integer multiple** of the E.F.

	<u>EF</u>	<u>MF</u>
Benzene	CH	C ₆ H ₆
Acetylene	CH	C ₂ H ₂

$$\text{MF} = (\text{CH})_n$$

n = multiplying factor

$$n = \frac{\text{MW}}{\text{EFW}}$$

Find **MW** experimentally

<u>Benzene</u>	<u>Acetylene</u>
$n = \frac{78.1 \text{ amu}}{13.0 \text{ amu}} = 6$	$n = \frac{26.0 \text{ amu}}{13.0 \text{ amu}} = 2$

1) Ex : Analysis of an unknown compd.
gave 39.72% C, 1.67% H, 58.61% Cl.
The MW was found to be 181.4 amu.
Determine the molecular formula.

a) Determine Emp. Formula

C:

H:

Cl:

C:

H:

Cl:

$$\text{E.F.} = \text{C H Cl}$$

$$\text{E.F.W.} =$$

b) Determine Molecular Formula

$$n = \frac{\text{MW}}{\text{EFW}} = \frac{181.4 \text{ amu}}{\quad} =$$

$$\text{MF} =$$

E) Combustion Analysis

“Burn” a compd. containing C, H & O and use the quantities of the products, CO₂ & H₂O, to determine the amt's. of C, H & O in **original** compd.

mass CO₂ \Rightarrow mass C

mass H₂O \Rightarrow mass H

1) Ex : An unknown compd. contains only C, H & O. Complete combustion of a 0.1000g sample produced:

0.1910 g CO_2 0.1172 g H_2O

a) Step 1: find mass of each element

$$? \text{ g C} = 0.1910 \text{ g CO}_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2} = 0.05212 \text{ g C}$$

Mass fraction of C in CO_2

$$? \text{ g H} = 0.1172 \text{ g H}_2\text{O} \times \frac{2(1.008) \text{ g H}}{18.02 \text{ g H}_2\text{O}} = 0.01311 \text{ g H}$$

$$\begin{aligned} \text{mass O} &= 0.1000 \text{ g} - 0.05212 \text{ g} - 0.01311 \text{ g} \\ &= 0.03477 \text{ g O} \end{aligned}$$

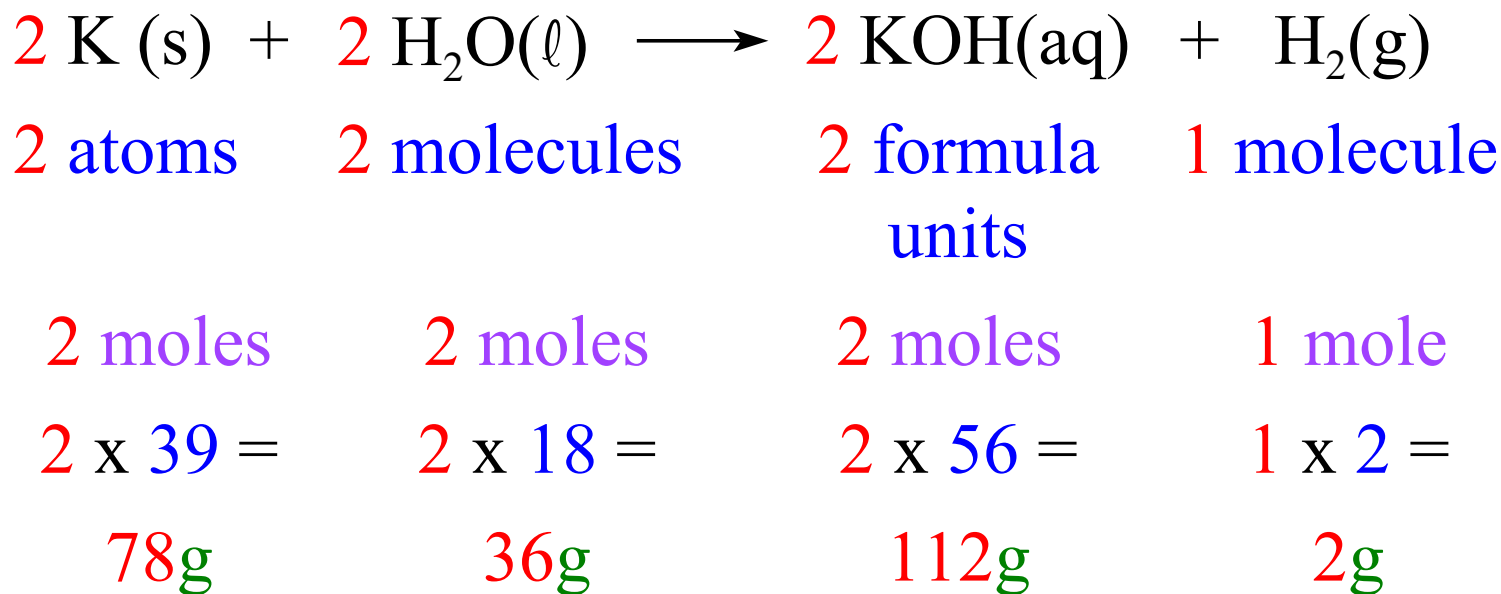
b) Step 2: find mol each element

VII) Stoichiometry & the Balanced Eqn.

Determination of **quantities** of **reactants** & **products** involved in chem. rx's.

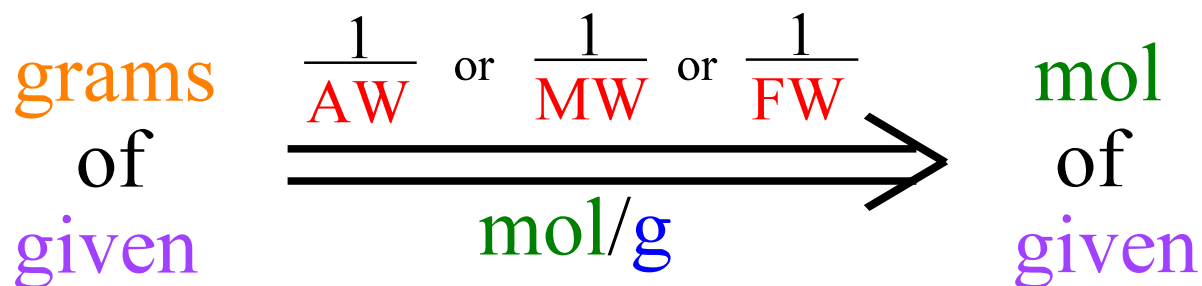
Use **balanced chem. eqn.**

- tells you **not only** the **reactants** & **products** but **also how much** of each is involved in the chem. rx.



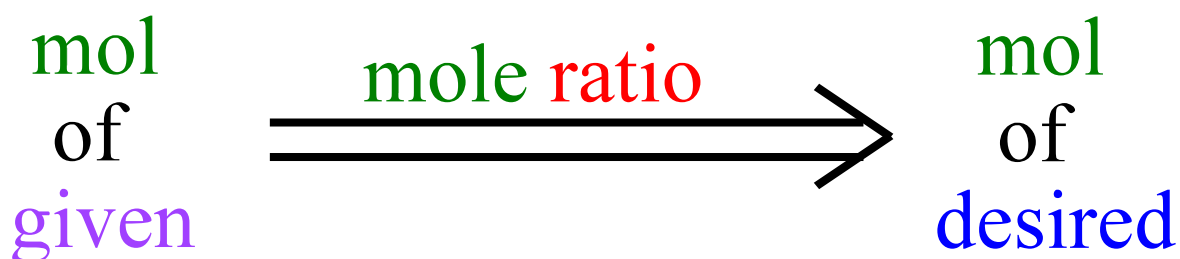
A) Procedure

1) Calc. Number of **moles** of a **given** substance



2) Determine **moles** of **desired** subst.

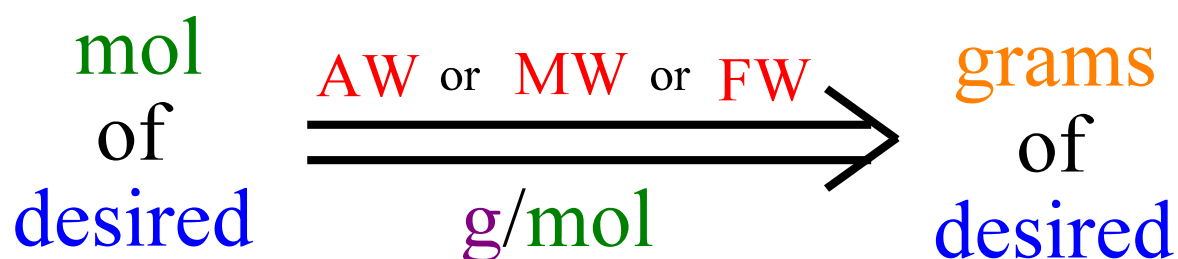
- use **coeff.** in the **bal. eqn.**
- **convert** **moles** of **given** substance **to** **moles** of **desired** substance



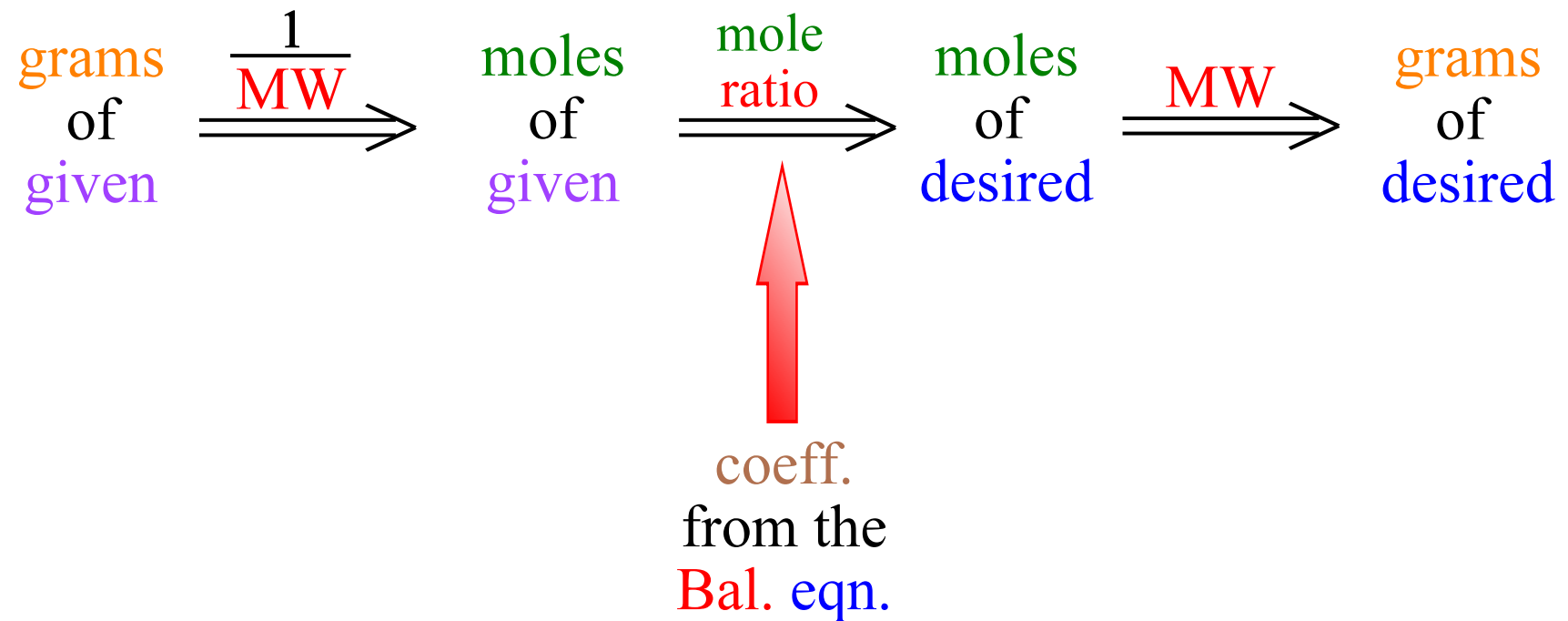
$$\text{mole ratio} = \frac{\text{moles of desired substance}}{\text{moles of given substance}}$$

$$\text{mole ratio} = \frac{\text{coeff. of desired substance}}{\text{coeff. of given substance}}$$

3) Convert moles of desired substance to grams



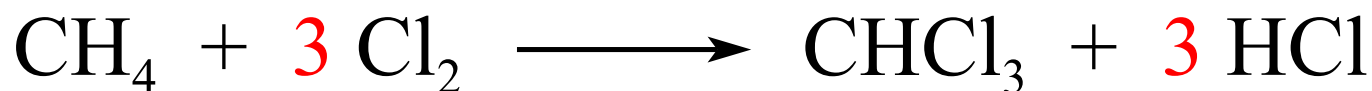
4) Summary



VIII) Solving Stoichiometry Problems

A) Ex 1: mole \longrightarrow mole
given desired

How many moles of chloroform, CHCl_3 ,
would be produced by the reaction
1.3 mol of chlorine?



B) Ex 2: mole \longrightarrow mole \longrightarrow grams
 given desired desired

How many grams of chlorine are req.
to produce 0.67 mol of CHCl_3 ?

C) Ex 3:

grams → moles → moles → grams
given given desired desired

Hydrazine reacts w. hydrogen peroxide according to the following eqn.,



How many grams of H_2O_2 are required to react w. $1.0 \times 10^3 \text{ g}$ of N_2H_4 ?

IX) Limiting Reactant

Limiting reactant (reagent):

Reactant which is **entirely** consumed when a rxn. goes to **completion**

- **limits** the amount of product formed

other reactants are called excess reactants (reagents)

A) Ex 1: For the following rx., if a rxn. mixture is prepared using 1.00 kg of each reactant, how much HCN could be produced?



Determine the limiting reactant:

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X) Theoretical & Percent Yields

Theoretical Yield: amt. product which can be obtained from a given amt. of reactant if **ALL** of it **reacts completely** according to the given eqn.

Actual Yield : amt. product actually obtained

- usually < theoretical yield

Percent Yield : how close actual yield is to theoretical yield

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

A) Ex 1 : If only 410 g HCN is produced in the lab from the rxn. mixture in the previous ex., what is the % yield?

$$\% \text{ yield} = \frac{\text{g HCN}}{\text{g HCN}} \times 100 \%$$