Exam 1 is Tuesday October 1st

Make sure to have your ID with you.

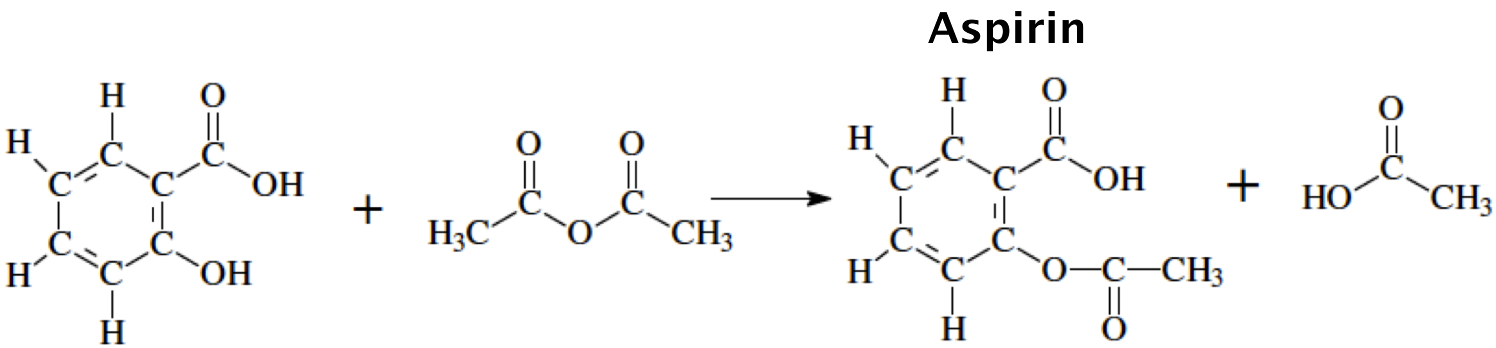
You will be assigned an exam seat.

If you have any seating requests (left handed, aisle, etc.) email me by
Wednesday September 25th.

Take note of the make-up policy in the syllabus.

Requests for re-grading must be brought to my attention within 48 hours of the exam being handed back in class.

Chapter 3



10.0 g + excess ⟶ ? g

Stoichiometry

Average Atomic Mass

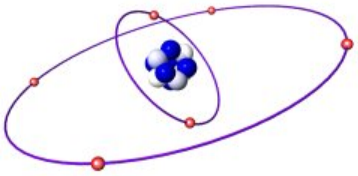
Atomic Mass: Mass of an atom in atomic mass units

1 amu = 1/12 of the mass of 1 C-12 atom

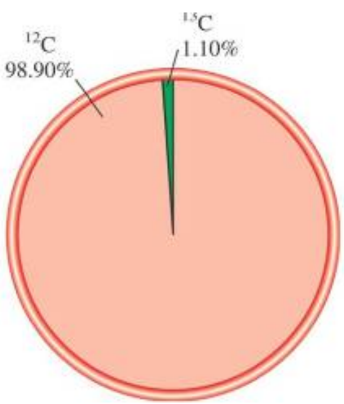
→ The mass of a ^{12}C atom = 12 amu

= 1.661×10^{-24} g

= mass 1 proton or 1 neutron



Naturally occurring carbon is a mixture of isotopes



$^{12}_{6}$C 98.90% 6 protons 6 neutrons 12.000 amu

$^{13}_{6}$C 1.100% 6 protons 7 neutrons 13.003 amu

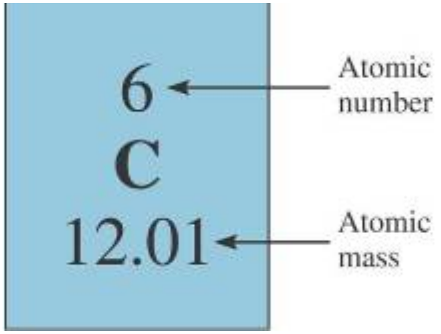
$^{14}_{6}$C ~ 10^{-12} 6 protons 8 neutrons 14.003 amu

(C-14 is unstable)

Atomic mass of naturally occurring carbon:

(0.9890 x 12.000 amu)+ (0.0110 x 13.003 amu)

= Atomic Mass of C = 12.01 amu



The Mole – like a dozen but a lot more!

Mole

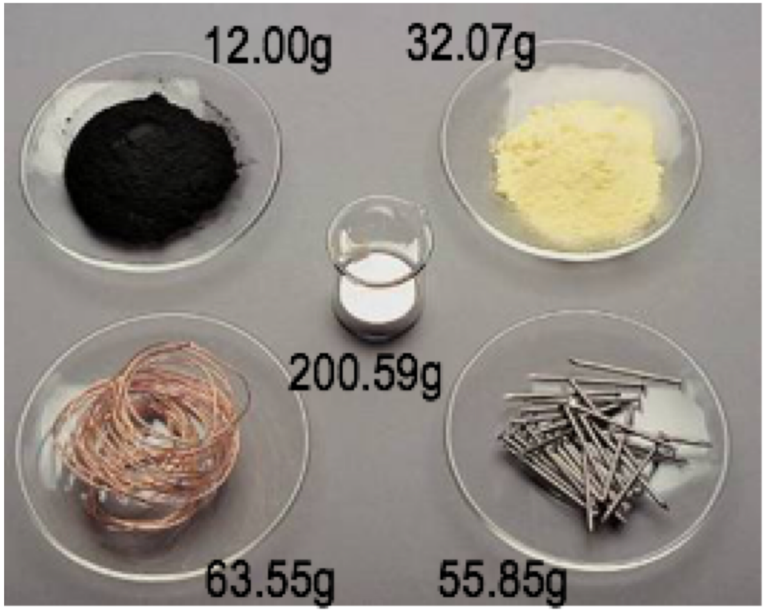
of atoms in 12.00g of C-12

Avogadro's number (N_a)

- # particles in 1 mole
- $N_a = 6.022 \times 10^{23}$ particles/mol
- Determined experimentally

Similar to the word "dozen"

- Makes numbers more manageable



1 Mole of each substance

For most chemicals, a mole is an amount that can be measured in a lab (using a balance, etc.)

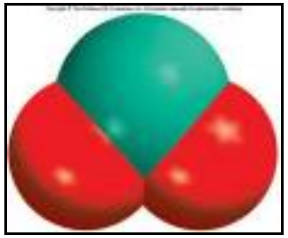
(Atoms are too small to measure on a balance)

$$\text{C-12:} \quad \frac{12.00g}{1mole} x \frac{1mole}{6.022x10^{23} atoms} = 1.993x10^{-23} g/1atom$$

Molar Mass

The mass of one mole of a substance

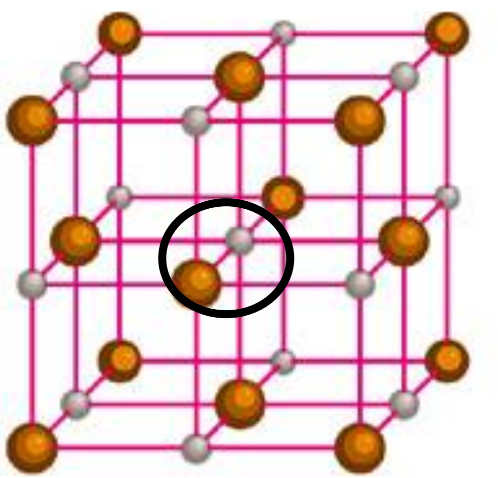
- Equal to AMU, but in units of g/mol



SO_2 - molecule

To calculate for a compound:

- Find atomic mass of each element
  → located on Periodic table (often below symbol)
- Multiply atomic mass of element by subscript, then add all elements together.
- Molecular mass: mass of molecule
  → include every atom
- Formula mass: mass of ions in a salt
  → use smallest ratio



NaCl – ionic compound

Examples:

1 mol Na = 22.99 g/mol

1 mol SO_2 = 64.07g/mol

1 mole NaCl = 58.44g/mol

Calculating Molar Mass

1.) NaCl

2.) SO_2

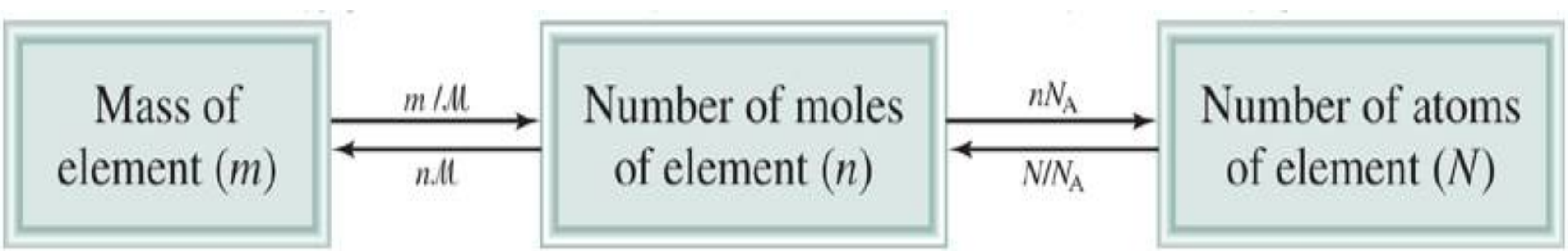
3.) $Pb(NO_3)_2$

Mole-based Calculations (Mass/Mole/Particle Conversions)

Molar Mass (M): grams/mol – from Periodic Table!

Avogadro's Number N_a: 6.022×10^{23} particles/mol

$$10.0g_C \text{ x } \frac{1mol_C}{12.0g_C} = 0.833mol_C \text{ x } \frac{6.02 \times 10^{23}{}_{atoms}}{1mol_C} = 5.01 \times 10^{23} \text{ atoms}$$



$$10.0g_C = \frac{12.0g_C}{1mol_C} \text{ x } 0.833mol_C = \frac{1mol_C}{6.02 \times 10^{23}{}_{atoms}} \text{ x } 5.01 \times 10^{23} \text{ atoms}$$

moles → mass

What is the mass, in grams, of 0.557 mol K_2O? (52.5 g)

mass → # moles

How many moles are there in 25.64 g of K_2O? (0.2722 mol)

moles → # particles

How many molecules are in 2.6 moles of CO_2? (1.6 x 10^{24} molecules)

moles → # particles

How many ions are in 2.6 moles of NaCl? (3.1 x 10^{24} ions)

particles → # moles (→ grams!)

If you have 2.5 x 10^{22} atoms of gold, how many moles do you have? How many grams do you have? (0.042mol, 8.2g)

Combined!

How many atoms are there in 2.578 g of SO_2 (MM = 64.065 g/mol)?

Mass → Moles → Molecules → Atoms

A: 7.270×10^{22} atoms

Percent Composition of Compounds by Mass (Mass % Compostion)

- General idea for percentages is "part / total"
- For mass %: mass of each element in the compound divided by the total mass of the compound
- Units should be the same for both values (usually g)

To Determine the Mass % of a Compound:

- Assume 1 mole of compound.
  - This will make subscripts = # moles of each element
- Calculate molar mass of compound.
- Calculate mass of each element based on subscripts.
- For each element, divide mass by molar mass of compound

Mass % Compostion

3 pieces Pepperoni (Pe) – 10. g per piece
2 pieces Cheese (Ch) – 9.0 g per piece
5 pieces Veggie (Ve) – 12 g per piece

$Pe_3Ch_2Ve_5$

Total: 10 slices, 108 g

Percent by slice:

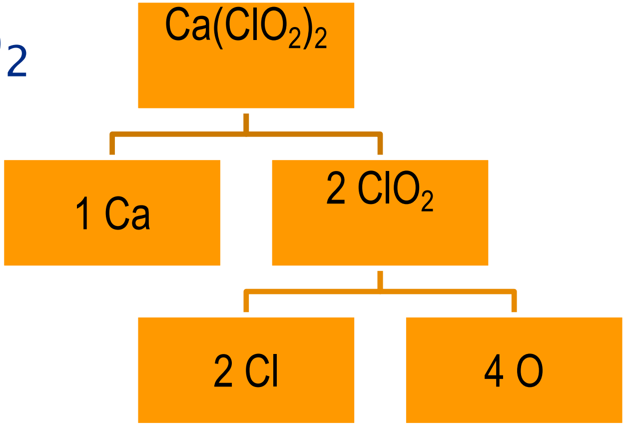
Pe: (3/10)*100 = 30%
Ch: (2/10)*100 = 20%
Ve: (5/10)*100 = 50%

Percent by mass:

Pe: (30./108)*100 = 27%
Ch: (18/108)*100 = 17%
Ve: (60./108)*100 = 56%

Mass % Composition of Calcium Chlorite, $Ca(ClO_2)_2$

Step 1: Find the molar mass of $Ca(ClO_2)_2$

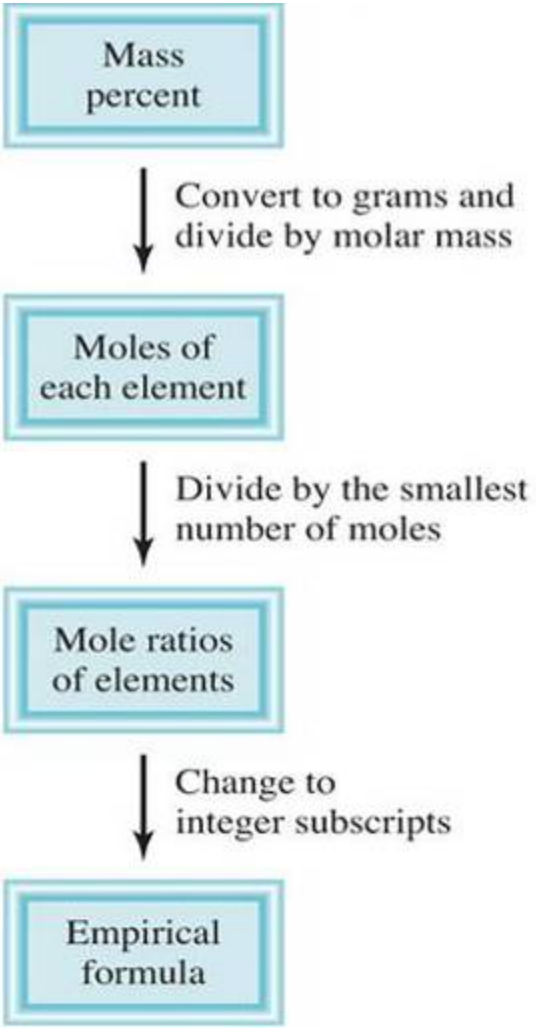


Step 2: Divide each elemental mass by the molar mass of $Ca(ClO_2)_2$ (Total should equal approximately 100%)

Empirical Formulas from Mass % Composition

What is the empirical formula for a compound with a mass composition of 2.2% H, 26.7% C, and 71.1% O?

Assume 100g, then can change % of each element to grams:



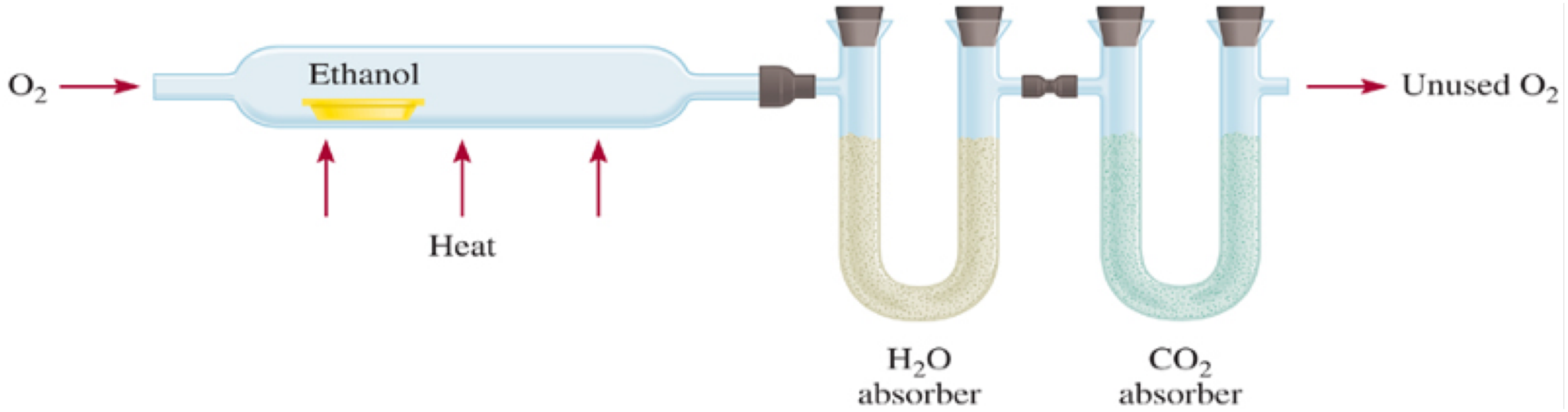
Convert grams of each element to moles

Divide by smallest # of moles

Use integers for subscripts

$H_1C_1O_2$ = HCO_2 (empirical formula)

Determination of Empirical Formulas by Elemental Analysis (Combustion)



- Burn measured amount of compound with excess O_2.
  - $C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O +$ Unused O_2
- Measure mass of products (must know what they are)
- Use mass of products to determine moles & mass of each element present
  - CO_2 and H_2O contain all C and H atoms
  - Determine amount of oxygen by difference
- If know molar mass can determine molecular formula

Determining Empirical Formula from Experiment

A 0.595g sample of a CHO compound burns in O_2 to produce 1.188g CO_2 and 0.486g H_2O. What is the empirical formula?

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O + \text{Unused } O_2$$

MM CO_2 = 44.01 g/mol MM H_2O = 18.016 g/mol

Determine Moles & Mass of C from CO_2

Determine Moles & Mass of H from H_2O

Determine Mass & Moles of O from what is left

Divide by smallest # moles to get formula: C_2H_4O

What If You Don't get Whole Numbers?

Results from Empirical Formula Calculation:

C = 1.5

O = 1

H = 3

Formula = $C_3H_6O_2$

Chemical Reactions & Chemical Equations:
Chemical Equations

- Shorthand description of a chemical reaction
  - Like a recipe!

3 
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\+ 2c 

→ 

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- Symbols & formulas represent elements & compounds

$$H_2(g) + 2C(s) + Cl_2(g) \rightarrow C_2H_2Cl_2(g)$$

Chemical Equations

$$H_2(g) + 2C(s) + Cl_2(g) \rightarrow C_2H_2Cl_2(g)$$

Reactants: Starting substances on left: H_2, C, Cl_2

Products: Substances formed on right: $C_2H_2Cl_2$

Values in front of symbols: **Stoichiometric coefficients**

Coefficients = # moles of that substance

→ If there is no #, the coefficient is 1

+ sign: Think of it as "and"; not mathematical adding!

Arrow (produces, yields) – change from products to reactants

→ Shows the direction of reaction ($\rightarrow$, $\leftarrow$, $\leftrightarrow$, $\rightleftarrows$)

(g), (s), (l), (aq): chemical phase: gas, solid, liquid, aqueous

Rules & Hints For Balancing Chemical Equations

Cannot make something out of nothing!

- **ONLY COEFFICIENTS CAN BE CHANGED!!! $H_2O \neq H_2O_2$**

- If an element(s) is present in just 1 compound on each side of the equation, balance that element(s) first.

- Balance free elements last. (O_2, C, H_2, etc.)

- Fractions can be cleared at any time by multiplying all coefficients by a common multiplier (often denominator).

$$2\,[C_4H_{10} + 13/2\ O_2 \rightarrow 4\ CO_2 + 5\ H_2O] \rightarrow 2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$$

- Groupings of atoms (such as in polyatomic ions) may remain unchanged. In such cases, you can balance these groupings as a unit.

Balancing Chemical Equations

Starting – Unbalanced (no coefficients):

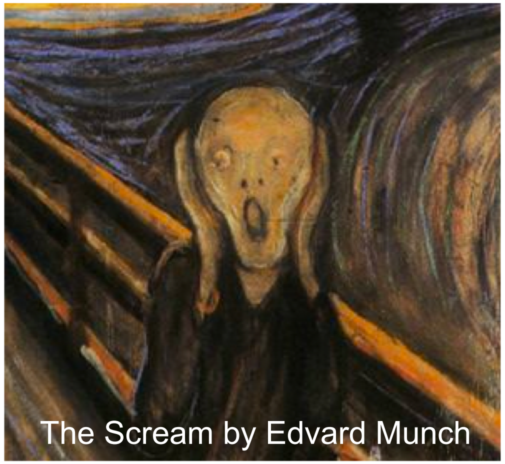
1.) $H_2 + O_2 \rightarrow H_2O$

2.) $C_2H_5O_2 + O_2 \rightarrow CO_2 + H_2O$

Balanced: 1.) $\underline{2}H_2 + O_2 \rightarrow \underline{2}H_2O$

2.) $\underline{4}C_2H_5O_2 + \underline{9}O_2 \rightarrow \underline{8}CO_2 + \underline{10}H_2O$

Amounts of Reactants and Products: Stoichiometry



Calculations based on chemical reactions

How much do you need to make what you want?

Stoichiometry:
Mole Ratios in Chemical Reactions

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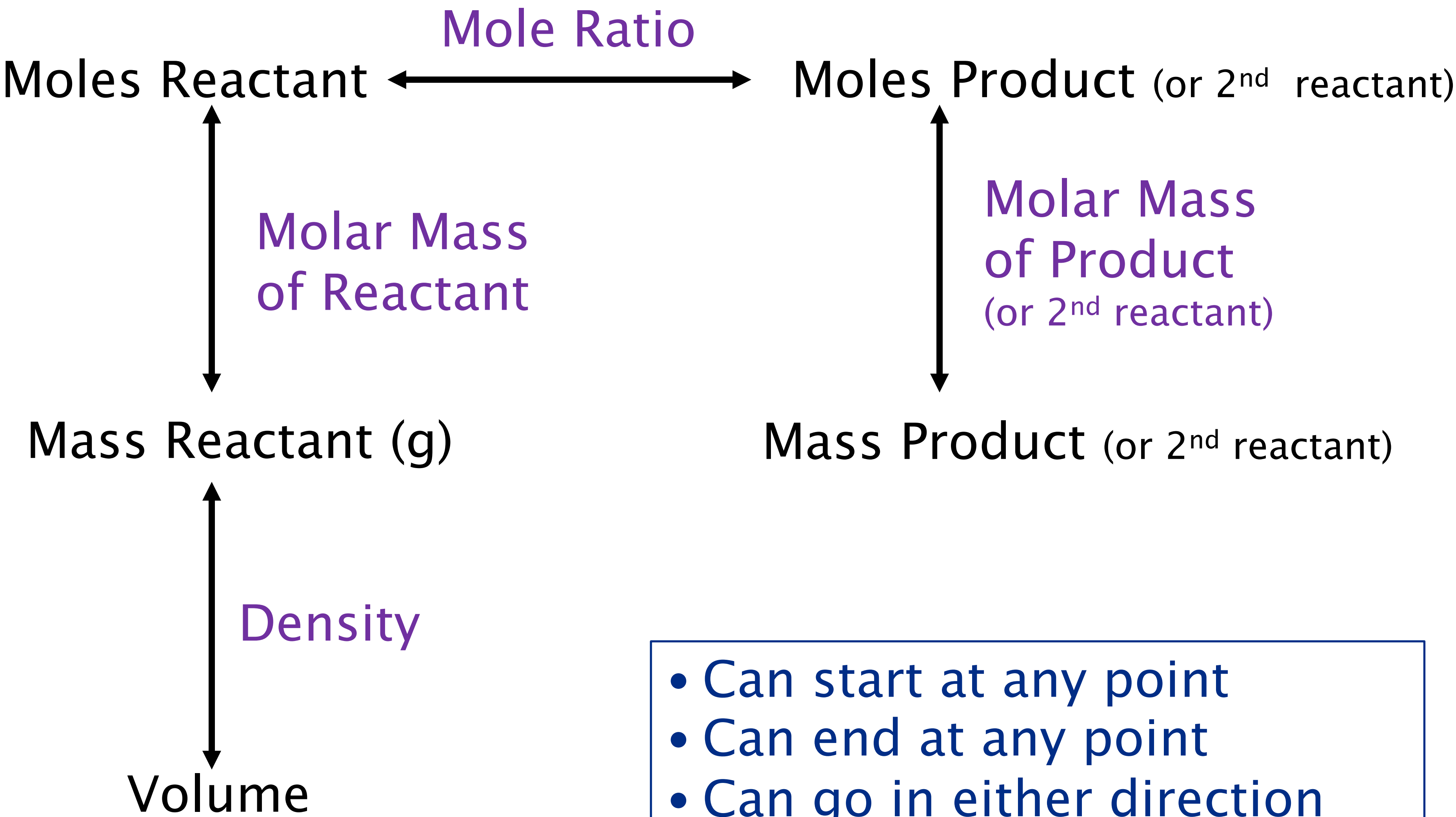
3 eggs and 2 cups of flour react to make one cake
ratio: 3:2:1

$$2C(s) + 1Cl_2(g) + 2H_2(g) \Rightarrow 1C_2H_4Cl_2(g)$$

2 moles of graphite (carbon), 1 mole of chlorine gas, and 2 moles of hydrogen gas react to form 1 mole of dichloroethane

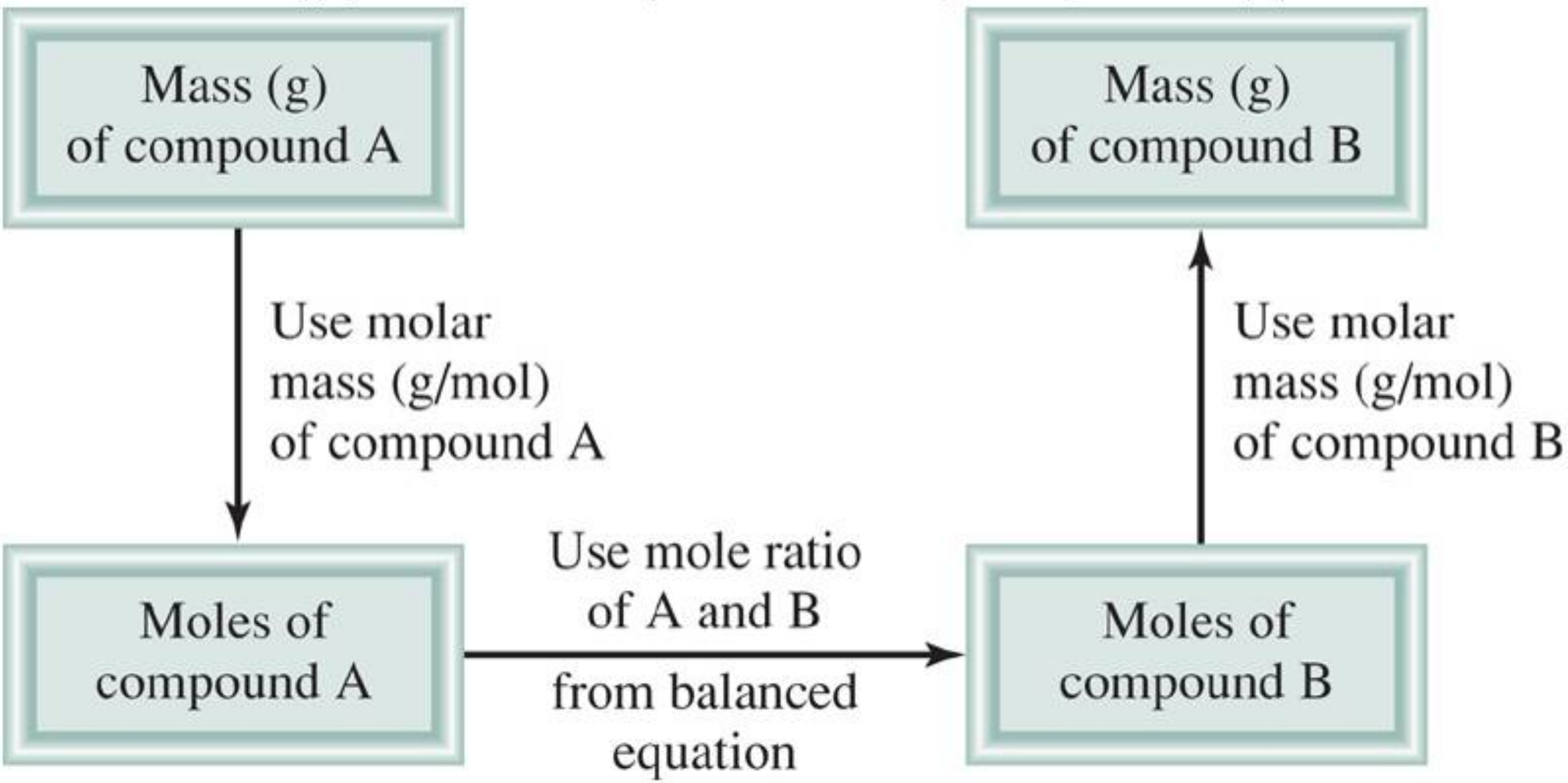
Mole ratio: 2:1:2:1

Stoichiometry Flow Chart



- Can start at any point
- Can end at any point
- Can go in either direction
- Can have additional steps

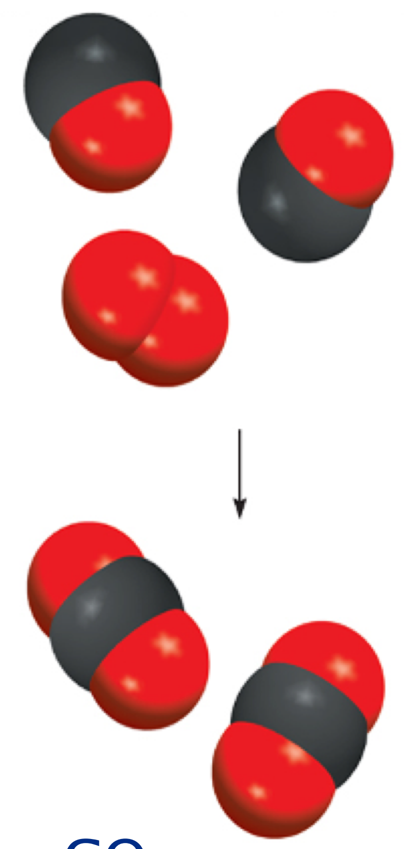
Stoichiometry Flow Chart 2



What is the mass of CO_2 produced when 10.7g of CO reacts with O_2 to form CO_2?

Write and balance the equation:

$$2CO\,(g) + 1O_2(g) \rightarrow 2CO_2(g)$$



Calculate moles of CO (28.0104 g/mol) in 10.7g of CO. (0.38195)

Calculate moles of CO_2 from mole ratio.

(0.38195 mol)

Calculate grams CO_2 (44.0098g/mol) from moles CO_2.

(16.8g)

Limiting Reagents & Reaction Yield:

- **Limiting Reagent:** Reactant that runs out first!
  - Determines how much product you can make
  - Find by calculating the moles of 1 product from each given amount of reactant
  - Limiting reagent is the reactant producing the smallest amount of product

- **Theoretical Yield:**
  - Max amount of product that you can make
  - Based on limiting reagent!
  - Generally reported in grams

3  + 2c 

→ 

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12	unlimited
unlimited	12
12	12
12	6

$$Sb_4O_{10} + 6\ H_2O \rightarrow 4\ H_3SbO_4$$

If you start with 3.0 moles Sb_4O_{10} and 8.0 moles of water, what is your limiting reagent?

169.87 g/mol 27.026 g/mol 133.886 g/mol

$$AgNO_3 + HCN \longrightarrow AgCN + HNO_3$$

If you make silver cyanide, which is used in electroplating, from 20.0 g of silver nitrate and 15.0 g of hydrogen cyanide gas, what is your limiting reagent? What is your theoretical yield of AgCN?

Step 1: Make sure equation is balanced.

Step 2: Moles of reactants – for limiting reagent, need both!

Why might you want to make either silver nitrate or hydrogen cyanide your limiting reagent?

Step 3: Cross the mole bridge. Limiting reagent produces smallest amount of product! (LR = $AgNO_3$)

Step 4: Use the limiting reagent to determine the mass of AgCN.

(15.8 g)

Yields of Chemical Reactions

Reactions rarely produce maximum product

a. Impure reactants
b. Incomplete reaction
c. All product not fully recovered
d. Side reactions may occur

Actual yield: Yield recovered during experiment

Theoretical yield: Yield calculated from limiting reagent

$$\text{Percent yield} = \left(\frac{\text{Actual yield}}{\text{Theoretical yield}}\right) \times 100$$

2.01588 g/mol   28.0134 g/mol   17.0305 g/mol

Step 1: $3\,H_2 + N_2 \longrightarrow 2\,NH_3$

If you start with 4.00 g of hydrogen gas and 22.00g of nitrogen gas, and make 18.5 g of ammonia, what is your percent yield?

Step 2:

Step 3:

Step 4:

Step 5:

82.1 % yield