Lecture 1 - 5th September 2024

density: $\rho = \frac{m}{V}$ → water density: 997 kg/m^3

Significant Digits

- all nonzero digits are significant!
- for quantities less than one, any zero preceding the first non-zero digit is not significant.
  - ↳ eg: 0.00237 has 3 sig figs.
- in quantities greater than one, any zero following the last non-zero digit is not significant, unless there is a decimal point.
  - ↳ eg: 200 has only one sig fig, while 200.00 has five.

- when adding or subtracting, retain digits to the magnitude of the value with the least significant digits.

- when multiplying or dividing, keep the same number of digits as the value with the least significant digits.

Stoichiometry:

Example: what is $7.2 \frac{Btu}{lbm\,°F}$ expressed in units of $J g^{-1} K^{-1}$?

↳ given: $1J = 9.4782 \cdot 10^{-4}$ Btu, $1kg = 2.20462$ lbm, $1K = 1.8°F$

$$\hookrightarrow 7.2 \frac{Btu}{lbm\,°F} \cdot \frac{1\,J}{9.4782 \cdot 10^{-4}\,Btu} \cdot \frac{2.20462\,lbm}{1\,kg} \cdot \frac{1\,kg}{1000\,g} \cdot \frac{1.8\,°F}{1\,K}$$

$= 30.1 \frac{J}{gK}$

Example: you are interested in calculating the annual reduction in CO_2 emissions by replacing your car with an all-electric vehicle. Assume that you drive an annual distance 14000 km. Your car has a fuel effeciency 28 miles per gallon. The electric car has an average energy consumption of 157 Wh/km.

↳ Data for CO_2 emission: 25 g CO_2/kWh, 2.3 kg CO_2/L

↳ Gas: $14000 \frac{km}{year} \cdot \frac{6.21371 \cdot 10^{-4} mi}{1 m} \cdot \frac{1000 m}{1 km} \cdot \frac{1 gal}{28 mi}$

$\cdot \frac{1000 L}{264.172 gal} \cdot \frac{2.3 kg\, CO_2}{1 L} = 2705 \frac{kg\, CO_2}{year}$

Electric: $14000 \frac{km}{year} \cdot \frac{157 Wh}{1 km} \cdot \frac{25 g\, CO_2}{1 kWh} \cdot \frac{1 kWh}{1000 Wh} \cdot \frac{1 kg}{1000 g}$

$= 55 \frac{kg\, CO_2}{year}$

reductions: gas - electric = $2705 - 55 = 2650 \frac{kg\, CO_2}{year}$

Temperature Conversions:

- $T(°F) = 1.8\, T(°C) + 32$
- $T(°C) = \frac{T(°F) - 32}{1.8}$
- $T(K) = T(°C) + 273.15$
- $T(°R) = 1.8\, T(K)$
- $T(°R) = T(°F) + 491.67$

all given on exams!

Classification of Matter

- All Matter
  - $\xrightarrow{\text{seperatable}}$ substance
    - $\xrightarrow{\text{decomposable}}$ compound
    - $\xrightarrow{\text{not decomposable}}$ element
  - $\xrightarrow{\text{not seperatable}}$ mixture
    - $\xrightarrow{\text{uniform}}$ homogeneous
    - $\xrightarrow{\text{not uniform}}$ heterogeneous

Example: classify the following:

a) a piece of metal: can't determine

b) distilled water: pure substance

c) distilled water w/ ice: pure substance ↳ bc when seperated, both are still H_2O!

d) oil on water: heterogeneous

e) coffe w/ cream and sugar: heterogenous

1
H
1.00794

1 → atomic number: # of protons

1.00794 → relative atomic mass (u)

↳ # protons + # electrons = relative atomic mass!

Example: the mass number and the atomic number of an atom that contains 15 electrons, 15 protons, and 16 neutrons are, respectively:

15 protons means that the atomic number is 15.
the mass is # protons + # electrons = 15 + 16 = 31.

∴, 31 and 15.

- Atoms are most stable when their outer shell is completely empty or completely full.
  ↳ octet rule!

- Elements combine to form compounds by atomic bonding.

- Bonding involves electrons that are either shared (covalent bond) or transferred (ionic bond)

- Energy is absorbed or released when electrons are transferred or shared.

Example: copper has two major naturally occuring isotopes, which have 34 and 36 neutrons. Using the periodic table, determine the percentage abundance of copper atoms having 34 neutrons.

- From the periodic table, the atomic number of copper is 29 and the atomic mass is 63.546 amu.

↳ mass of ^{63}Cu : $29 + 34 = 63u$
mass of ^{65}Cu : $29 + 36 = 65u$.

$$63.546u = x \cdot 63u + (1-x) \cdot 65u$$

$\therefore, x = 0.727 = 72.7\%$.

- Electronegativity: a measure of how strongly an atom competes for electrons in bonds formed with other atoms.
  ↳ increases rightwards in the periodic table!

Ionic Compounds: one atom donates one or more electron to another atom.
- Usually a metal combined with a non-metal.
- Dissociate into ions when dissolved in water.
  ↳ Eg: NaCl, ordered crystal of Na^+ and Cl^- ions.

Covalent Compounds: one or more atoms share one or more electrons.
- Large, organic molecules are usually covalent.
- Do not dissociate into ions when dissolved in water.
  ↳ Eg: CH_4, $C_6H_{12}O_6$

- Two or more atoms, joined by covalent bonds, may form a charged polyatomic ion.
  ↳ Eg: ammonium (NH_4^+), carbonate (CO_3^{2-}), etc.

- Molecular formula: shows the actual number of atoms in a molecule.

- Emperical formula: gives the relative number of atoms in a molecule.

- The molecular formula can be obtained from the empirical formula if the molecular mass of the compound is known.

Example: C_6H_{12} is the molecular formula, while CH_2 is the emperical formula of hexene.

Example: A compound has the empirical formula CH_2O and a molecular weight of 150.13 u. What is the molecular formula?

↳ weight of C : 12.011
weight of H_2 : $1.008 \cdot 2 = 2.016$
weight of O : 15.999

$\therefore$, weight of $CH_2O = 12.011 + 2.016 + 15.999$
$= 30.026$ u.

we know that the compound should have a molecular mass of 150.13u, so how many times does one CH_2O go into it?

$150.13 \div 30.026 = 5.$

$\therefore$, compound is $5CH_2O = C_5H_{10}O_5$.

- During a chemical reaction, atoms rearrange through the breaking and formation of chemical bonds.

- Chemical equations describe changes that occur in chemical reactions.
  ↳ equations must be balanced (atom conversion principle)

Balancing a chemical equation:

- Done by including stoichiometric coefficients
- The stoichiometric chemical equation describes only the net changes taking place in a reaction.

Example: $FeS + O_2 \longrightarrow Fe_2O_3 + SO_2$

- $2FeS + O_2 \longrightarrow Fe_2O_3 + SO_2$ 2Fe, 2S, 2O → 2Fe, S, 5O
  ↳ $2FeS + O_2 \longrightarrow Fe_2O_3 + 2SO_2$ 2Fe, 2S, 2O → 2Fe, 2S, 7O
  ↳ $4FeS + O_2 \longrightarrow 2Fe_2O_3 + 4SO_2$ 4Fe, 4S, 2O → 4Fe, 4S, 14O
  ↳ $\underset{4Fe\ 4S}{4FeS} + \underset{14O}{7O_2} \longrightarrow \underset{4Fe\ \ 14O}{2Fe_2O_3 + 4SO_2}$

Example: $NH_3 + O_2 \longrightarrow NO + H_2O$

- $2NH_3 + O_2 \longrightarrow 2NO + 3H_2O$ 2N, 6H, 2O → 2N, 6H, 5O
- $2NH_3 + 4O_2 \rightarrow 2NO + 6H_2O$ 2N, 6H, 8O → 2N, 12H, 8O
- $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$ 4N, 12H, 10O → 4N, 12H, 10O

Example: $Na_2CO_3 + Mg(NO_3)_2 \rightarrow MgCO_3 + NaNO_3$

↳ $Na_2CO_3 + Mg(NO_3)_2 \rightarrow MgCO_3 + 2NaNO_3$

The mechanism equation: describes the steps required to go from reactants to products at a molecular level.

Sequential Reactions: when the product of an initial chemical reactions becomes a reactant for further reactions.

↳ eg: $A + B \rightarrow C + D$
$C + E \rightarrow F$ } $A + B + E \rightarrow D + F$

Example: consider a natural gas composed of a mixture of methane (CH_4) and ethane (C_2H_6) combusted in oxygen to form carbon dioxide and water. Write balanced chemical equations to describe what occurs.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$
$$C_2H_6 + \frac{7}{2}O_2 \rightarrow 2CO_2 + 3H_2O$$

↳ since these are simultoneous reactions, we cannot add them.

- Never combine simultaneous reactions into a single equation!!

Avogadro's Constant: $6.022 \cdot 10^{23}$ units = 1 mole

↳ Example: 1 mole of O_2 molecules = $6.022 \cdot 10^{23}$ O_2 molecules
1 mole of apples = $6.022 \cdot 10^{23}$ apples

↳ 1 molecule of CH_4 contains 1 C atom and 4 H atoms, while 1 mol of CH_4 contains 1 mol of C and 4 mols of H.

- Molar Mass/Weight: the mass of one mole of a substance. ⟶ $M = \frac{m}{n}$

↳ Example: determine the molar mass of the following:

C: 12.011 g/mol
H: 1.0079 g/mol
O: 15.999 g/mol

these are just the mass given on the periodic table!!!

$C_7H_{14}O_2$: $7(12.011) + 14(1.0079) + 2(15.999)$
$= 130.186$ g/mol

- Since one mole of carbon is defined as 12g of ^{12}C, the mass given on the periodic table is the molar mass!

Example: isopentyl acetate ($C_7H_{14}O_2$) is the compound responsible for the scent of bananas. Interestingly, bees release about 1μg of this compound when they sting to attract other bees to join the attack.

a) how many molecules of isopentyl acetate are released in a typical bee sting?

from previous example!

molar mass of $C_7H_{14}O_2$ = 130.186 g/mol

$$\hookrightarrow 130.186 \frac{g}{mol} \cdot \frac{1 \text{ mol}}{6.022 \cdot 10^{23} \text{ molecules}} \cdot \frac{1 \,\mu g}{10^{-6} \, g}$$

$$= 2.1618 \cdot 10^{-16} \frac{\mu g}{\text{molecule}}.$$

$$\therefore, \frac{1}{2.1618 \cdot 10^{-16} \frac{\mu g}{\text{molecule}}} = 4.6 \cdot 10^{15} \frac{\text{molecules}}{\mu g}$$

b) how many atoms of carbon are present in this?

↳ since each molecule of $C_7H_{14}O_2$ contains seven carbon atoms, $4.6 \cdot 10^{15}$ molecules of $C_7H_{14}O_2$ contain $7(4.6 \cdot 10^{15})$ carbon atoms.

$\therefore$, $3.2 \cdot 10^{16}$ Carbon atoms.

Example: the hemoglobin content of blood is about 15.5 g/100 mL. The molar mass of hemoglobin is about 64500 g/mol. There are 4 iron atoms in a hemoglobin molecule. Approximately how many grams of iron are present in the 5 L of blood of a typical adult?

$$5 \text{ L } \cancel{\text{blood}} \times \frac{1000 \text{ mL}}{1 \cancel{\text{L blood}}} \times \frac{15.5 \cancel{\text{ g hemoglobin}}}{100 \cancel{\text{ mL blood}}} \times \frac{1 \cancel{\text{ mol hemoglobin}}}{64500 \cancel{\text{ g hemoglobin}}}$$

$$\times \frac{4 \cancel{\text{ mol Fe}}}{1 \cancel{\text{ mol hemoglobin}}} \times \frac{55.845 \text{ g Fe}}{1 \cancel{\text{ mol Fe}}} = 2.7 \text{g Fe}$$

Example: how many mols of $NaClO_3$ are needed to produce 6 mol of SO_2 in the following two-step reaction: $2NaClO_3 \rightarrow 2NaCl + 3O_2$

$S + O_2 \rightarrow SO_2$

Since O_2 is produced in the first reaction and used in the second, and since we aren't told that the equations happen simutaneously, we know that it's sequential.

↳ $2NaClO_3 \rightarrow 2NaCl + 3O_2$

$3S + 3O_2 \rightarrow 3SO_2$.

↳ $2NaClO_3 + 3S \rightarrow 2NaCl + 3SO_2$

to get 6 mols of SO_2, we simply multiply both sides of the reaction by 2:

$4NaClO_3 + 6S \rightarrow 4NaCl + 6SO_2$.

∴, 4 mols of $NaClO_3$ are needed to make 6 mols of SO_2.

Example: an approach for carbon capture is to dissolve carbon dioxide from power plant exhaust in water and then inject that solution into sub-surface basalt rocks that contain the mineral Mg_2SiO_4. The disssolved carbon dioxide will react with basalt to form the stable mineral, $MgCO_3$, according to the following unbalanced equation: $Mg_2SiO_4 + CO_2 \rightarrow SiO_2 + MgCO_3$.

a) what mass of CO_2, in g, can be stored in one tonne of Mg_2SiO_4?

First, we must balance the equation:

$$Mg_2SiO_4 + 2CO_2 \rightarrow SiO_4 + MgCO_3 .$$

$$1 \text{ tonne } Mg_2SiO_4 \times \frac{1000 \text{ kg}}{1 \text{ tonne}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol } Mg_2SiO_4}{140.69 \text{ g}} \times \frac{2 \text{ mol } CO_2}{1 \text{ mol } Mg_2SiO_4}$$

$$\times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 6.26 \cdot 10^5 \text{ g } CO_2.$$

b) What mass of $MgCO_3$, in g, is produced?

$$1 \text{ tonne} \times \frac{1000 \text{ kg}}{1 \text{ tonne}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol } Mg_2SiO_4}{140.69 \text{ g}} \times \frac{2 \text{ mol } MgCO_3}{1 \text{ mol } Mg_2SiO_4} \times \frac{84.3139 \text{ g } MgCO_3}{1 \text{ mol } MgCO_3}$$

$$= 1.2 \cdot 10^6 \text{ g } MgCO_3 .$$

c) the average annual CO_2 emission per person is 17.7 tonne. What mass of Mg_2SiO_4, in tonne, is required to store Canada's annual CO_2 emission? Assume that the population of Canada is ~40 million.

$$17.7 \frac{\text{tonne}}{\text{person}} \times \frac{40{,}000{,}000 \text{ person}}{\text{population}} \times \frac{1 \text{ tonne } MgSiO_4}{6.26 \cdot 10^5 \text{ g } CO_2} \times \frac{10^6 \text{ g}}{1 \text{ tonne}}$$

$$= 1.13 \cdot 10^9 \frac{\text{tonne } Mg_2SiO_4}{\text{population}} .$$

Example: A natural gas is composed of 70 mol-% methane (CH_4) and 30 mol-% ethane (C_2H_6).

a) what is the average molar mass of the gas mixture?

$$1\ mol \times \frac{70\ mol\ CH_4}{100\ mol} \times \frac{(12.011)+4(1.008)\ g\ CH_4}{1\ mol\ CH_4} = 11.23\ g\ of\ CH_4$$

$$1\ mol \times \frac{30\ mol\ C_2H_6}{100\ mol} \times \frac{2(12.011)+6(1.008)}{1\ mol\ C_2H_6} = 9.02\ g\ of\ C_2H_6$$

$\therefore$ total $= 11.23 + 9.02 = 20.3$ g!

Example: how many grams of CO_2 are produced in the complete combustion of 406 g of a bottled gas that consists of 72.7 % C_3H_8 and 27.3% C_4H_{10}?

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

$$C_4H_{10} + 6.5\ O_2 \rightarrow 4CO_2 + 5H_2O$$

mass of C_3H_8: $406\ g \times \frac{72.7}{100} = 295.162\ g\ C_3H_8$

mass of C_4H_{10}: $406\ g - 295.162\ g = 110.838\ g\ C_4H_{10}$

C_3H_8: $3(12.011) + 8(1.008) = 44.09$ g/mol

C_4H_{10}: $4(12.011) + 10(1.008) = 58.12$ g/mol

we know $M = \frac{m}{n}$, so $n = \frac{m}{M}$:

$$n_{C_3H_8}: \frac{295.16\ g}{44.09\ g/mol} = 6.695\ mol\ C_3H_8$$

$$n_{C_4H_{10}}: \frac{110.838\ g}{58.12\ g/mol} = 1.907\ mol\ C_4H_{10}$$

from the reaction equations, we see that:

1 mol C_3H_8 ~ 3 mol CO_2, 1 mol C_4H_{10} ~ 4 mol CO_2.

↳ mol CO_2 from C_3H_8 = 3(6.695) = 20.085 mol CO_2

mol CO_2 from C_4H_{10} = 4(1.907) = 7.628 mol CO_2

∴ total mol CO_2 = 20.085 + 7.628 = 27.713 mol CO_2.

$$27.713 \text{ mol } CO_2 \times \frac{(12.011) + 2(16) \text{ g } CO_2}{1 \text{ mol } CO_2} = 27.713\ (44.011) \text{ mol } CO_2$$

$= 1219.68 \text{ g } CO_2$!

Example: a 10.2g sample of an organic compound containing carbon, hydrogen, and oxygen, was burned in excess oxygen. This yielded 23.1g CO_2 and 4.72g H_2O. Assuming complete combustion, what is the empirical formula of the compound?

- we know the C in the CO_2 and the H in the H_2O must be from the original compound, but the O is both from the original compound and the excess oxygen.

$$23.1\text{g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.009 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.525 \text{ mol C}$$

$$4.72\text{g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0148 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.524 \text{ mol H}$$

$$0.525 \text{ mol C} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 6.3\text{g C}$$

$0.524 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.528 \text{ g H}$

$m_o = m - m_c - m_H = 10.2 - 6.3 - 0.528 = 3.37 \text{ g O}$

$3.37 \text{ g O} \times \frac{1 \text{ mol O}}{16 \text{ g O}} = 0.211 \text{ mol O}$

divide all by the smallest # of mols (0.211):

$\frac{0.525}{0.211} \text{ mol C}, \quad \frac{0.524}{0.211} \text{ mol H}, \quad \frac{0.211}{0.211} \text{ mol O}$

↳ $C_{2.5}H_{2.5}O_1 = C_5H_5O_2$!

Solution: composed of a solvent and one or more solutes.

- the solvent is present in the largest amount.
- for aqueous solutions (aq), the solvent is water.
- the solute is present in smaller amounts and is dissolved in the solvent
- solutions are homogeneous (have uniform properties).

molarity: describes the concentration of a solution.

↳ $\text{Molarity} = \frac{\text{amount of solute (moles)}}{\text{volume of solution (litres)}} \rightarrow C = \frac{n}{V}$

Example: the fuel hydrazine can be produced by the reaction of solutions of sodium hypochlorite and ammonia:

$$NaClO_{(aq)} + 2NH_{3\,(aq)} \rightarrow N_2H_{4\,(aq)} + NaCl_{(aq)} + H_2O_{(l)}$$

a) 750 mL of a 0.8 M NaClO solution is mixed with excess ammonia. How many moles of hydrazine can be formed?

we know $c = \frac{n}{V}$, so $n = c \cdot V$.

↳ $n = 0.8 \times 0.75 = 0.6$.

∴ 0.6 mol N_2H_4

b) If the final volume of the resulting solution is 1.25 L, what will be the molarity of the hydrazine?

↳ $c = \frac{n}{V} \rightarrow c = \frac{0.6}{1.25} = 0.48$ M.

· Other ways to express concentration:

- Percentages
  - ↳ by mass: 25 % (w/w) = 25g solute per 100g solution
  - by volume: 25 % (v/v) = 25mL solute per 100mL solution
  - weight to volume: 25% (w/v) = 25g solute per 100mL solution

- molality = $\frac{\text{amount of solute (moles)}}{\text{amount of solvent (kg)}}$

- parts per million (ppm): $\frac{\text{g solute}}{\text{1,000,000 g solution}}$

- parts per billion (ppb): $\frac{\text{g solute}}{\text{1,000,000,000 g solution}}$

Example: an excess of aluminium foil is allowed to react with 225 mL of an aqueous solution of HCl that has a density of 1.088 g/mL and contains 18 % HCl by mass.

$$2\,Al_{(s)} + 6\,HCl_{(aq)} \rightarrow 2\,AlCl_{3\,(s)} + 3\,H_{2\,(g)}$$

What mass of H_2 is produced in the reaction?

mass = volume × density → mass = $0.225\ mL \times 1.088\ \frac{g}{mL}$ = 244.8 g.

18 % of this solution is HCl, so $0.18 \times 244.8\,g = 44.06\,g$ HCl

$$44.06\,g\ HCl \times \frac{1\ mol\ HCl}{36.46\,g\ HCl} \times \frac{3\ mol\ H_2}{6\ mol\ HCl} \times \frac{2.016\,g\ H_2}{1\ mol\ H_2} = 1.22\,g\ H_2$$

Example: a mixture contains only $CuCl_2$ and $FeCl_3$. A 0.7391 g sample of the mixture is completely dissolved in water and then treated with $AgNO_{3\,(aq)}$. The following reactions occur:

$$CuCl_{2\,(aq)} + 2\,AgNO_{3\,(aq)} \rightarrow 2\,AgCl_{(s)} + Cu(NO_3)_{2\,(aq)}$$
$$FeCl_{3\,(aq)} + 3\,AgNO_{3\,(aq)} \rightarrow 3\,AgCl_{(s)} + Fe(NO_3)_{3\,(aq)}$$

If it takes 86.91 mL of 0.1463 M $AgNO_3$ solution to precipitate all the chloride as AgCl, then what is the percentage by mass of copper in the original mixture?

we know that $m_{CuCl_2} + m_{FeCl_3} = 0.7391\,g$.

amount of $AgNO_3$ that reacts: $86.91\ mL \times \frac{1\,L}{1000\ mL} \times \frac{0.1463\ mol\ AgNO_3}{1\ L}$

$= 0.01271$ mol $AgNO_3$.

$\therefore, 0.01271 = n_{AgNO_3 \text{ in } 1} + n_{AgNO_3 \text{ in } 2}$

we see that 2 mol $AgNO_3$ ~ 1 mol $CuCl_2$
3 mol $AgNO_3$ ~ 1 mol $FeCl_3$

let x be the moles of $CuCl_2$ and y be the moles of $FeCl_3$.

↳ $0.01271 = 2x + 3y$

$0.7391 = m_{CuCl_2} + m_{FeCl_3} = n_{CuCl_2} M_{CuCl_2} + n_{FeCl_3} M_{FeCl_3}$

↓

$0.7391 = 134.452x + 162.204y$.

→ solve the system of equations: $x = 0.0019719$

$\therefore$ 0.0019719 mol $CuCl_2$.

$$0.0019719 \text{ mol } CuCl_2 \times \frac{1 \text{ mol } Cu}{1 \text{ mol } CuCl_2} \times \frac{63.546 \text{ g } Cu}{1 \text{ mol } Cu} = 0.125 \text{ g } Cu.$$

↳ $\% \, Cu = \dfrac{0.125 \text{ g } Cu}{0.7391 \text{ g mixture}} = 16.95\%$ Cu by mass!

- **Theoretical Yield:** the expected yield from given quantities of reactants.

- **Actual Yield:** the quantity that is actually produced.

- **Percentage Yield:** the ratio of the actual yield to the theoretical yield

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

Example: suppose that the reactions below occur consecutively, and each reaction has an 84% yield:

A: $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$
B: $CH_3Cl + Cl_2 \rightarrow CH_2Cl_2 + HCl$

Starting with 112g of CH_4 in reaction A and an excess of Cl_2, how many grams of CH_2Cl_2 are formed in reaction B?

$$112g\ CH_4 \times \frac{1\ mol\ CH_4}{16.04g\ CH_4} \times \frac{1\ mol\ CH_3Cl}{1\ mol\ CH_4} = 6.9825\ mol\ CH_3Cl$$

since we only have a 84% yield (not 100%):
$$6.9825 \times 0.84 = 5.86\ mol\ CH_3Cl$$

$$5.86\ mol\ CH_3Cl \times \frac{1\ mol\ CH_2Cl}{1\ mol\ CH_3Cl} \times \frac{84.93g\ CH_2Cl}{1\ mol\ CH_2Cl} = 497.6898.$$

however, we also only get 84% yield here as well:

$$0.84 \times 497.6898 = 418g\ CH_2Cl_2$$

- Limiting Reactant: the reactant that is completely consumed.
  ↳ all other reactants are called excess reactants.

- Percentage excess is the proportion by which an excess reactant exceeds the stoichiometric amount.

- In stoichiometric calculations, we must use the limiting reactant as the basis!

Example: if 18.1 g of NH_3 is reacted with 90.4 g of CuO, what is the limiting reactant?

↳ $2NH_{3\,(g)} + 3CuO_{(s)} \rightarrow N_{2\,(g)} + 3Cu_{(s)} + 3H_2O_{(g)}$

molar mass of NH_3: $14.01 + 3(1.01) = 17.04$ g/mol

molar mass of CuO: $63.55 + 16 = 79.55$ g/mol

$$n = \frac{m}{M} \rightarrow \frac{18.1\text{ g}}{17.04\text{ g/mol}} = 1.062 \text{ mol } NH_3$$

$$↳ \frac{90.4\text{ g}}{79.55\text{ g/mol}} = 1.137 \text{ mol } CuO$$

equation: $2 \text{ mol } NH_3 \sim 3 \text{ mol } CuO \rightarrow \dfrac{2 \text{ mol } NH_3}{3 \text{ mol } CuO} = 0.667$

actual: $1.067 \text{ mol } NH_3 \sim 1.137 \text{ mol } CuO \rightarrow \dfrac{1.067 \text{ mol } NH_3}{1.137 \text{ mol } CuO} = 0.937$

Since the actual mole ratio (0.937) is greater than the expected mole ratio (0.667), we have more NH_3 than needed.

↳ ∴, CuO is the limiting reactant!

Determine the theoretical yield of N_2 in grams.

$$90.4\text{ g } CuO \times \frac{1 \text{ mol } CuO}{79.545 \text{ g } CuO} \times \frac{1 \text{ mol } N_2}{3 \text{ mol } CuO} \times \frac{28.014 \text{ g } N_2}{1 \text{ mol } N_2}$$

$$= 10.612 \text{ g } N_2$$

- Atom Economy = $\frac{\text{mass of desired product(s)}}{\text{mass of all reactants}}$

↳ - calculated assuming a yield of 100%.
  - based on the stoichiometry of the chemical equation and does not consider the mass of excess reactants.