

# Lecture 1 - 5<sup>th</sup> September 2024

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

## 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  $\frac{\text{Btu}}{\text{lbf}^{\circ}\text{F}}$  expressed in units of Jg<sup>-1</sup>K<sup>-1</sup>?  
↳ given: 1J =  $9.4782 \cdot 10^{-4}$  Btu, 1kg = 2.20462 lbf, 1K = 1.8°F

$$7.2 \frac{\text{Btu}}{\text{lbf}^{\circ}\text{F}} \cdot \frac{1\text{J}}{9.4782 \cdot 10^{-4} \text{Btu}} \cdot \frac{2.20462 \text{lbf}}{1\text{kg}} \cdot \frac{1\text{kg}}{1000\text{g}} \cdot \frac{1.8^{\circ}\text{F}}{1\text{K}}$$

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

Example: you are interested in calculating the annual reduction in CO<sub>2</sub> emissions by replacing your car with an all-electric vehicle. Assume that you drive an annual distance 14000 km. Your car has a fuel efficiency 28 miles per gallon. The electric car has an average energy consumption of 157 Wh/km.

↳ Data for CO<sub>2</sub> emission: 25 g CO<sub>2</sub>/kWh, 2.3 kg CO<sub>2</sub>/L

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

$$\cdot \frac{1000 \text{ L}}{264.172 \text{ gal}} \cdot \frac{2.3 \text{ kg CO}_2}{1 \text{ L}} = 2705 \frac{\text{kg CO}_2}{\text{year}}$$

$$\text{Electric: } 14000 \frac{\text{km}}{\text{year}} \cdot \frac{157 \text{ Wh}}{1 \text{ km}} \cdot \frac{25 \text{ g CO}_2}{1 \text{ kWh}} \cdot \frac{1 \text{ kWh}}{1000 \text{ Wh}} \cdot \frac{1 \text{ kg}}{1000 \text{ g}}$$

$$= 55 \frac{\text{kg CO}_2}{\text{year}}$$

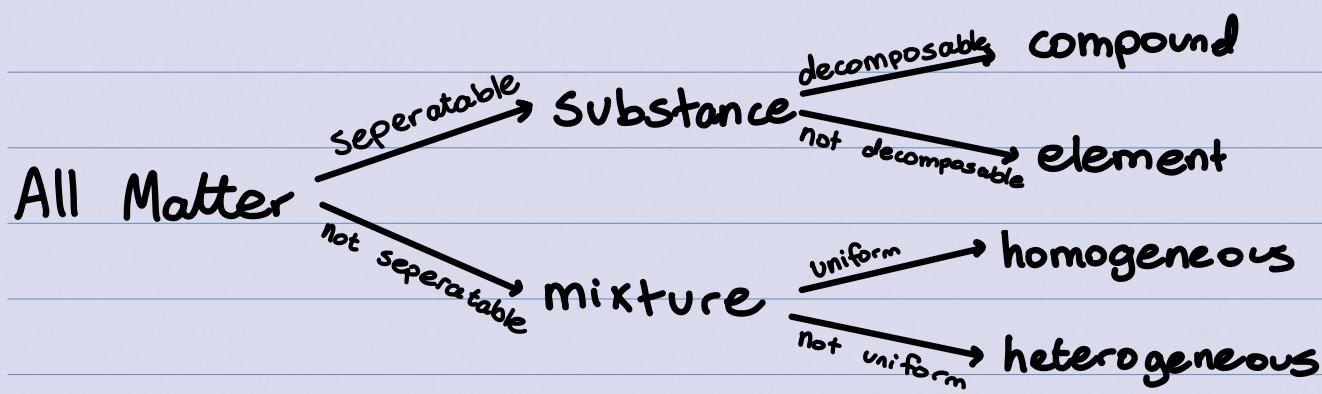
$$\text{reductions: gas - electric} = 2705 - 55 = 2650 \frac{\text{kg CO}_2}{\text{year}}$$

## Temperature Conversions:

- T(°F) = 1.8 T(°C) + 32
- T(°C) =  $\frac{T(\text{°F}) - 32}{1.8}$
- 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



Example: Classify the following:

- a) a piece of metal: can't determine
- b) distilled water: pure substance ↳ bc when separated, both are still  $H_2O$ !
- c) distilled water w/ ice: pure substance
- d) oil on water: heterogeneous
- e) coffee w/ cream and sugar: heterogeneous

I
H
1.00794

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 occurring 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.

$$\hookrightarrow \text{mass of } {}^{63}\text{Cu} : 29 + 34 = 63\text{u}$$

$$\text{mass of } {}^{65}\text{Cu} : 29 + 36 = 65\text{u}.$$

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

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

• 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  $\text{Na}^+$  and  $\text{Cl}^-$  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:  $\text{CH}_4$ ,  $\text{C}_6\text{H}_{12}\text{O}_6$
- 
- Two or more atoms, joined by covalent bonds, may form a charged polyatomic ion.  
↳ Eg: ammonium ( $\text{NH}_4^+$ ), carbonate ( $\text{CO}_3^{2-}$ ), etc.
  - Molecular formula: shows the actual number of atoms in a molecule.
  - Empirical 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 empirical 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

$$\therefore \text{weight of } CH_2O = 12.011 + 2.016 + 15.999 \\ = 30.026 \text{ u.}$$

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

$$150.13 \div 30.026 = 5.$$

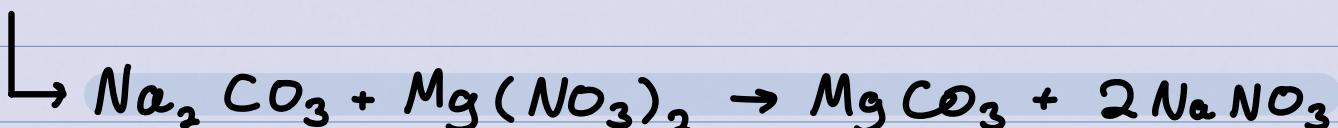
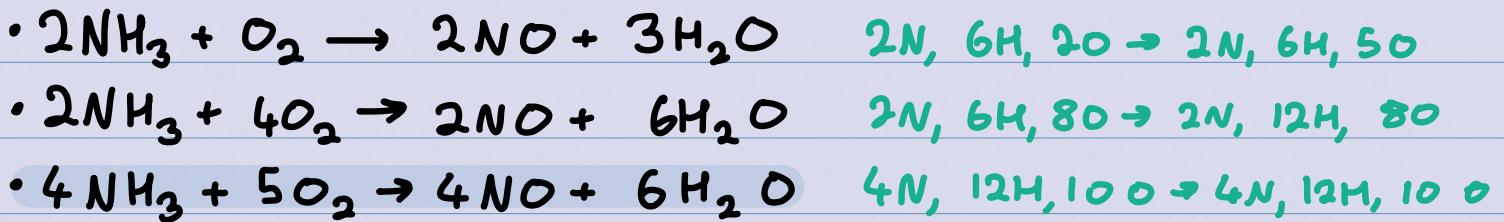
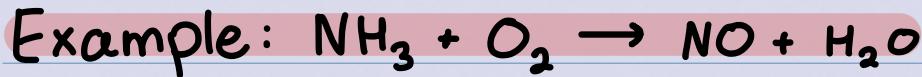
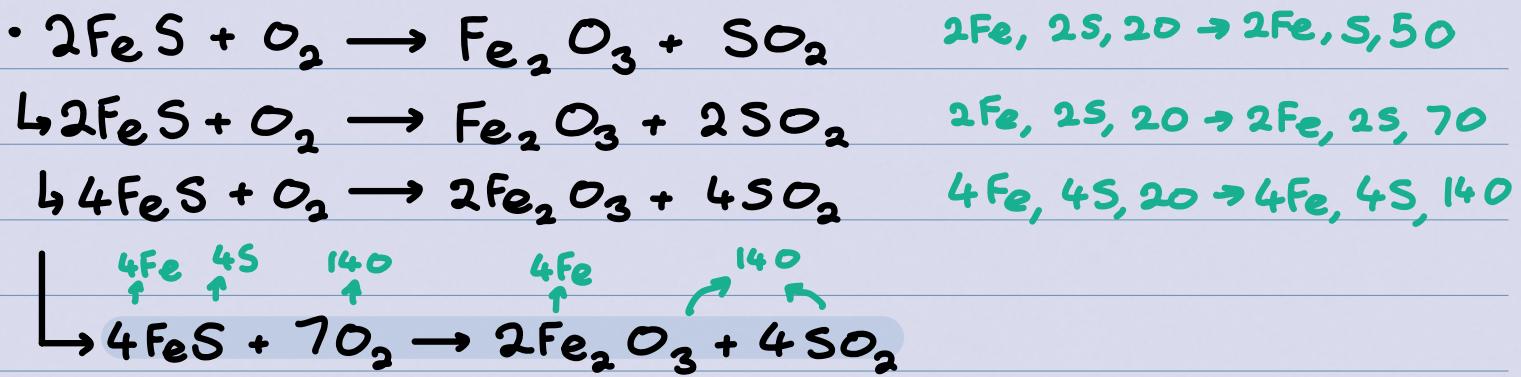
$$\therefore \text{compound is } 5CH_2O = C_5H_{10}O_5.$$

- 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 conservation principle)

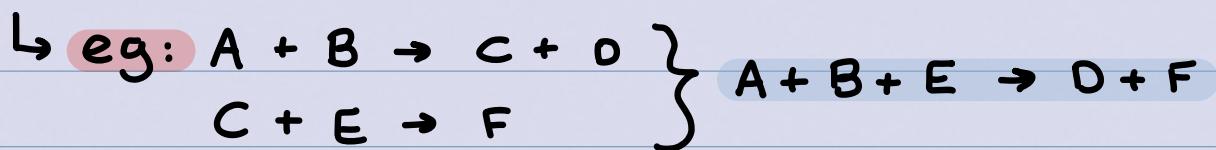
## Balancing a chemical equation:

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

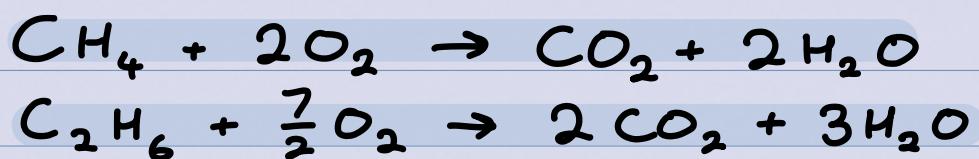


The mechanism equation: describes the steps required to go from reactants to products at a molecular level.

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



Example: consider a natural gas composed of a mixture of methane ( $\text{CH}_4$ ) and ethane ( $\text{C}_2\text{H}_6$ ) combusted in oxygen to form carbon dioxide and water. Write balanced chemical equations to describe what occurs.



↳ since these are simultaneous 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  $\text{O}_2$  molecules =  $6.022 \cdot 10^{23}$   $\text{O}_2$  molecules  
1 mole of apples =  $6.022 \cdot 10^{23}$  apples

↳ 1 molecule of  $\text{CH}_4$  contains 1 C atom and 4 H atoms, while 1 mol of  $\text{CH}_4$  contains 1 mol of C and 4 mols of H.

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

↳ 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!!!

$$\begin{aligned} \text{C}_7\text{H}_{14}\text{O}_2: \quad & 7(12.011) + 14(1.0079) + 2(15.999) \\ & = 130.186 \text{ g/mol} \end{aligned}$$

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

Example: isopentyl acetate ( $\text{C}_7\text{H}_{14}\text{O}_2$ ) is the compound responsible for the scent of bananas. Interestingly, bees release about 1ug 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?

p from previous example:

molar mass of  $\text{C}_7\text{H}_{14}\text{O}_2 = 130.186 \text{ g/mol}$

$$\hookrightarrow 130.186 \cdot \frac{g}{mol} \cdot \frac{1 \text{ mol}}{6.022 \cdot 10^{23} \text{ molecules}} \cdot \frac{1 \mu\text{g}}{10^{-6} \text{ g}} \\ = 2.1618 \cdot 10^{-16} \frac{\mu\text{g}}{\text{molecule}} \cdot$$

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

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

$\hookrightarrow$  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.

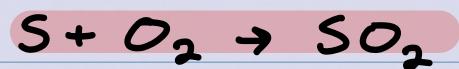
$\therefore 3.2 \cdot 10^{16}$  Carbon atoms.

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?

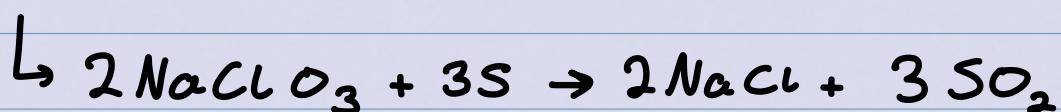
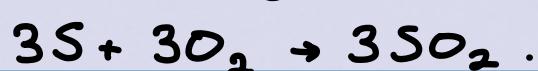
$$5 \text{ L blood} \times \frac{1000 \text{ mL}}{1 \text{ L blood}} \times \frac{15.5 \text{ g hemoglobin}}{100 \text{ mL blood}} \times \frac{1 \text{ mol hemoglobin}}{64500 \text{ g hemoglobin}} \\ \times \frac{4 \text{ mol Fe}}{1 \text{ mol hemoglobin}} \times \frac{55.845 \text{ g Fe}}{1 \text{ mol Fe}} = 2.7 \text{ g Fe}$$

Example: how many mols of  $NaClO_3$  are needed to produce 6 mol of  $SO_2$  in the following two-step

reaction:



Since  $\text{O}_2$  is produced in the first reaction and used in the second, and since we aren't told that the equations happen simultaneously, we know that it's sequential.



to get 6 mols of  $\text{SO}_2$ , we simply multiply both sides of the reaction by 2:

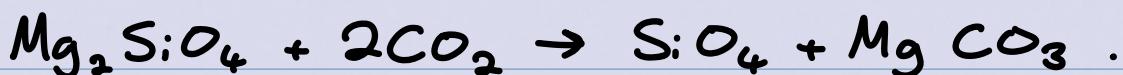


$\therefore$ , 4 mols of  $\text{NaClO}_3$  are needed to make 6 mols of  $\text{SO}_2$ .

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  $\text{Mg}_2\text{SiO}_4$ . The dissolved carbon dioxide will react with basalt to form the stable mineral,  $\text{MgCO}_3$ , according to the following unbalanced equation:  $\text{Mg}_2\text{SiO}_4 + \text{CO}_2 \rightarrow \text{SiO}_2 + \text{MgCO}_3$ .

a) What mass of  $\text{CO}_2$ , in g, can be stored in one tonne of  $\text{Mg}_2\text{SiO}_4$ ?

First, we must balance the equation:



$$1 \text{ tonne Mg}_2\text{SiO}_4 \times \frac{1000 \text{ kg}}{1 \text{ tonne}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol Mg}_2\text{SiO}_4}{140.69 \text{ g}} \times \frac{2 \text{ mol CO}_2}{1 \text{ mol Mg}_2\text{SiO}_4}$$

$$\times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 6.26 \cdot 10^5 \text{ g CO}_2.$$

b) What mass of  $\text{Mg CO}_3$ , in g, is produced?

$$1 \text{ tonne} \times \frac{1000 \text{ kg}}{1 \text{ tonne}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol Mg}_2\text{SiO}_4}{140.69 \text{ g}} \times \frac{2 \text{ mol Mg CO}_3}{1 \text{ mol Mg}_2\text{SiO}_4} \times \frac{84.3139 \text{ g Mg CO}_3}{1 \text{ mol Mg CO}_3}$$

$$= 1.2 \cdot 10^6 \text{ g Mg CO}_3.$$

c) the average annual  $\text{CO}_2$  emission per person is 17.7 tonne. What mass of  $\text{Mg}_2\text{SiO}_4$ , in tonne, is required to store Canada's annual  $\text{CO}_2$  emission? Assume that the population of Canada is  $\sim 40$  million.

$$17.7 \frac{\text{tonne}}{\text{person}} \times \frac{40,000,000 \text{ person}}{\text{population}} \times \frac{1 \text{ tonne Mg}_2\text{SiO}_4}{6.26 \cdot 10^5 \text{ g CO}_2} \times \frac{10^6 \text{ g}}{1 \text{ tonne}}$$

$$= 1.13 \cdot 10^9 \frac{\text{tonne Mg}_2\text{SiO}_4}{\text{population}}.$$

Example: A natural gas is composed of 70 mol-% methane ( $\text{CH}_4$ ) and 30 mol-% ethane ( $\text{C}_2\text{H}_6$ ).

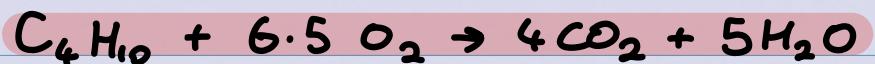
a) what is the average molar mass of the gas mixture?

$$1 \text{ mol} \times \frac{70 \text{ mol } \text{CH}_4}{100 \text{ mol}} \times \frac{(12.011) + 4(1.008) \text{ g } \text{CH}_4}{1 \text{ mol } \text{CH}_4} = 11.23 \text{ g of } \text{CH}_4$$

$$1 \text{ mol} \times \frac{30 \text{ mol } \text{C}_2\text{H}_6}{100 \text{ mol}} \times \frac{2(12.011) + 6(1.008)}{1 \text{ mol } \text{C}_2\text{H}_6} = 9.02 \text{ g of } \text{C}_2\text{H}_6$$

$$\therefore \text{total} = 11.23 + 9.02 = 20.3 \text{ g!}$$

Example: how many grams of  $\text{CO}_2$  are produced in the complete combustion of 406g of a bottled gas that consists of 72.7%  $\text{C}_3\text{H}_8$  and 27.3%  $\text{C}_4\text{H}_{10}$ ?



$$\text{mass of } \text{C}_3\text{H}_8: 406 \text{ g} \times \frac{72.7}{100} = 295.162 \text{ g } \text{C}_3\text{H}_8$$

$$\text{mass of } \text{C}_4\text{H}_{10}: 406 \text{ g} - 295.162 \text{ g} = 110.838 \text{ g } \text{C}_4\text{H}_{10}$$

$$\text{C}_3\text{H}_8: 3(12.011) + 8(1.008) = 44.09 \text{ g/mol}$$

$$\text{C}_4\text{H}_{10}: 4(12.011) + 10(1.008) = 58.12 \text{ g/mol}$$

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

$$n_{\text{C}_3\text{H}_8} : \frac{295.16 \text{ g}}{44.09 \text{ g/mol}} = 6.695 \text{ mol } \text{C}_3\text{H}_8$$

$$n_{\text{C}_4\text{H}_{10}} : \frac{110.838 \text{ g}}{58.12 \text{ g/mol}} = 1.907 \text{ mol } \text{C}_4\text{H}_{10}$$

from the reaction equations, we see that:

1 mol  $C_3H_8 \sim 3$  mol  $CO_2$ , 1 mol  $C_4H_{10} \sim 4$  mol  $CO_2$ .

$$\hookrightarrow \text{mol } CO_2 \text{ from } C_3H_8 = 3(6.695) = 20.085 \text{ mol } CO_2$$

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

$$\therefore \text{total mol } CO_2 = 20.085 + 7.628 = 27.713 \text{ mol } CO_2.$$

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

$$= 1219.68 \text{ g } CO_2!$$

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.

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

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

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

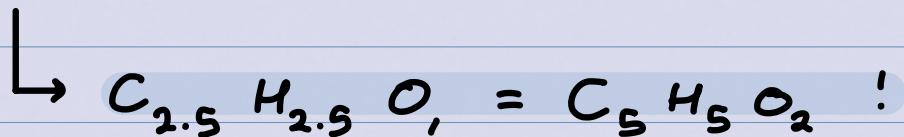
$$0.524 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.528 \text{ g H}$$

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

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

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

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



**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.

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

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



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  $n = C \cdot V$ .

$$\hookrightarrow n = 0.8 \times 0.75 = 0.6.$$

$$\therefore 0.6 \text{ mol } N_2H_4$$

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

$$\hookrightarrow C = \frac{n}{V} \rightarrow C = \frac{0.6}{1.25} = 0.48 \text{ M}.$$

• 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

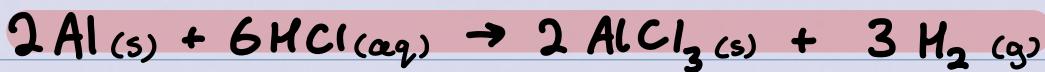
◦ Molality =  $\frac{\text{amount of solute (moles)}}{\text{amount of solvent (kg)}}$

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

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

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.



What mass of H<sub>2</sub> is produced in the reaction?

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

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

$$44.06 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{3 \text{ mol H}_2}{6 \text{ mol HCl}} \times \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} = 1.22 \text{ g H}_2$$

Example: a mixture contains only CuCl<sub>2</sub> and FeCl<sub>3</sub>.

A 0.7391 g sample of the mixture is completely dissolved in water and then treated with AgNO<sub>3</sub>(aq). The following reactions occur:



If it takes 86.91 mL of 0.1463 M AgNO<sub>3</sub> 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<sub>CuCl<sub>2</sub></sub> + m<sub>FeCl<sub>3</sub></sub> = 0.7391 g.

amount of AgNO<sub>3</sub> that reacts:  $86.91 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.1463 \text{ mol AgNO}_3}{1 \text{ L}}$

$$= 0.01271 \text{ mol } \text{AgNO}_3.$$

$$\therefore 0.01271 = n_{\text{AgNO}_3 \text{ in 1}} + n_{\text{AgNO}_3 \text{ in 2}}$$

We see that 2 mol  $\text{AgNO}_3 \sim 1$  mol  $\text{CuCl}_2$   
3 mol  $\text{AgNO}_3 \sim 1$  mol  $\text{FeCl}_3$

let  $x$  be the moles of  $\text{CuCl}_2$  and  $y$  be the moles of  $\text{FeCl}_3$ .

$$\hookrightarrow 0.01271 = 2x + 3y$$

$$0.7391 = M_{\text{CuCl}_2} + M_{\text{FeCl}_3} = n_{\text{CuCl}_2} M_{\text{CuCl}_2} + n_{\text{FeCl}_3} M_{\text{FeCl}_3}$$
$$\downarrow$$

$$0.7391 = 134.452x + 162.204y.$$

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

$$\therefore 0.0019719 \text{ mol } \text{CuCl}_2.$$

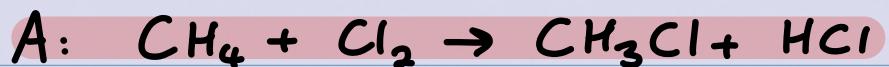
$$0.0019719 \text{ mol } \text{CuCl}_2 \times \frac{1 \text{ mol Cu}}{1 \text{ mol } \text{CuCl}_2} \times \frac{63.546 \text{ g Cu}}{1 \text{ mol Cu}} = 0.125 \text{ g Cu.}$$

$$\hookrightarrow \% \text{ Cu} = \frac{0.125 \text{ g Cu}}{0.7391 \text{ g mixture}} = 16.95\% \text{ Cu by mass!}$$

- **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

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

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



Starting with 112g of  $\text{CH}_4$  in reaction A and an excess of  $\text{Cl}_2$ , how many grams of  $\text{CH}_2\text{Cl}_2$  are formed in reaction B?

$$112 \text{ g } \text{CH}_4 \times \frac{1 \text{ mol } \text{CH}_4}{16.04 \text{ g } \text{CH}_4} \times \frac{1 \text{ mol } \text{CH}_3\text{Cl}}{1 \text{ mol } \text{CH}_4} = 6.9825 \text{ mol } \text{CH}_3\text{Cl}$$

Since we only have a 84% yield (not 100%):

$$6.9825 \times 0.84 = 5.86 \text{ mol } \text{CH}_3\text{Cl}$$

$$5.86 \text{ mol } \text{CH}_3\text{Cl} \times \frac{1 \text{ mol } \text{CH}_2\text{Cl}_2}{1 \text{ mol } \text{CH}_3\text{Cl}} \times \frac{84.93 \text{ g } \text{CH}_2\text{Cl}_2}{1 \text{ mol } \text{CH}_2\text{Cl}_2} = 497.6898.$$

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

$$0.84 \times 497.6898 = 418 \text{ g } \text{CH}_2\text{Cl}_2$$

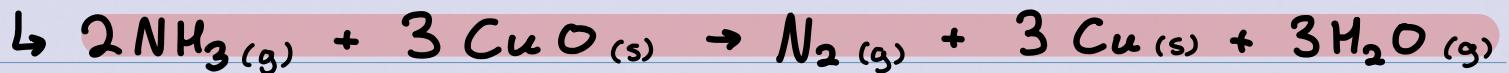
• **Limiting Reactant:** the reactant that is completely consumed.

↳ 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  $\text{NH}_3$  is reacted with 90.4 g of  $\text{CuO}$ , what is the limiting reactant?



Molar mass of  $\text{NH}_3$ :  $14.01 + 3(1.01) = 17.04 \text{ g/mol}$

molar mass of  $\text{CuO}$ :  $63.55 + 16 = 79.55 \text{ g/mol}$

$$n = \frac{m}{M} \rightarrow \frac{18.1 \text{ g}}{17.04 \text{ g/mol}} = 1.062 \text{ mol NH}_3$$

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

equation:  $2 \text{ mol NH}_3 \sim 3 \text{ mol CuO} \rightarrow \frac{2 \text{ mol NH}_3}{3 \text{ mol CuO}} = 0.667$

actual:  $1.067 \text{ mol NH}_3 \sim 1.137 \text{ mol CuO} \rightarrow \frac{1.067 \text{ mol NH}_3}{1.137 \text{ mol CuO}} = 0.937$

Since the actual mole ratio (0.937) is greater than the expected mole ratio (0.667), we have more  $\text{NH}_3$  than needed.

$\hookrightarrow$ ,  $\text{CuO}$  is the limiting reactant!

Determine the theoretical yield of  $\text{N}_2$  in grams.

$$90.4 \text{ g CuO} \times \frac{1 \text{ mol CuO}}{79.545 \text{ g CuO}} \times \frac{1 \text{ mol N}_2}{3 \text{ mol CuO}} \times \frac{28.014 \text{ g N}_2}{1 \text{ mol N}_2}$$

$$= 10.612 \text{ g N}_2$$

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

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