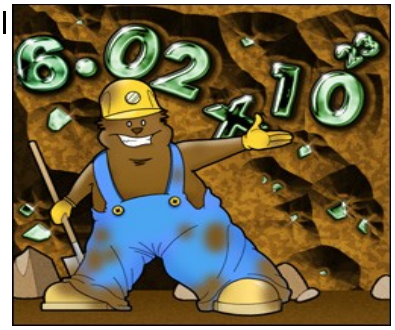
Stoichiometry **-**Study of the mathematics of chemical reactions



Calculation of Atomic Masses

How many isotopes of the element exist and how much does each contribute?

Ex. chlorine has an atomic weight of 35.45 amu because it is composed of 76% of Cl-35 (34.96 amu) and 24% Cl-37 (36.97 amu). How?

atomic mass unit (amu) – Mass of 1 atom; precise definition is that it is one twelfth of the mass of an unbound atom of carbon-12 (12C) at rest and in its ground state.

Mass Spectrometer is the best instrument for measuring the masses of atoms and ions; It is used because it is much easier and more precise to *compare* masses of atoms and molecules (determine *relative* masses) than to measure their *absolute* masses.

MOLE MOLE MOLE MOLE MOLE MOLE MOLE MOLE

Relationships of the Mole

Remember **1 mol = ? g...**

The average mass of 1 atom of a substance expressed in amu is the same number as the mass of 1 mol of a substance expressed in grams. So, **1 g = 6.02 x 10^{23} amu**

ex. In 5.56 mg there are how many molecules of Fr-12 (freon). $C\,Cl_2F_2$

Answer: 2.77 x $10^{19 \text{molecules}}$

Molar Mass = 1 mol of substance

Molecular wt Formula Wt

$CHCl_3$ (chloroform) aluminum carbonate (all ionic compounds)

Molar Volume = 1 mol = 22.4 L (@ STP)

How many L are in 3.0 molecules of CO_2?

Mass percent OR % composition from formulas:

gfm/total gfm x 100

ex. potassium ferricyanide - $K_3Fe(CN)_6$

Answer: 329.27 g/mol

Determining the Formula of a Compound

- *Empirical Rhyme*
  Percent to mass
  Mass to mole
  Divide by small
  Times till whole

N – 87. 4%   H -12.6%

NH_2

- molecular formula – Divide molar mass by the empirical formula mass; it will give the # of empirical f.u. in the actual molecule.

ex. mass of the empirical formula – NH_2 (16.026g)

mass of compound – 32.05 g   SO, units of empir. formula in compound – 32.05g/16.026 = 2

N_2H_4 (hydrazine)

Stoichiometry:Using the mole ratio of substances to solve problems

Start off with -   A BALANCED EQUATION

butane   + oxygen -----> carbon dioxide + water

C_4H_{10} + $13/2O_2$ -----> $4CO_2$ +   5 H_2O

Rules for determining amount of substance reacting or formed in chemical equations:

1. Balance the equation.
2. Decide what quantity of each substance is given.
3. Decide what quantity you want to calculate. Represent it with a question mark directly under the substance (one line down if a mole quantity, two lines down if a laboratory quantity).
4. Complete the diagram with arrows by going from the known laboratory quantity to moles, then to moles of unknown quantity, and finally to laboratory quantity of the unknown. (Step(s) can be omitted if either or both the given quantity and the unknown quantity are moles.)

5. Calculate each quantity following the steps in the diagram.

How many liters (at STP) of NH_3 are needed to react with 80.0 g of O_2?

1) $4\,NH_3 + 5\,O_2 \rightarrow 4\,NO + 6\,H_2O$

$$\begin{array}{ccc} \text{moles} & \xleftarrow{\text{(step 2)}} & \text{moles} \\ \downarrow \text{(step 3)} & & \uparrow \text{(step 1)} \\ ?\ \text{L} & & 80.0\ \text{g} \end{array}$$

2) moles ← moles

Answer: 44.8 liters NH_3 (STP)

ex. $KClO_3 \longrightarrow KCl + O_2$ (steps to solving problems - p. 109)

If a 100.00g sample of potassium chlorate decomposes, how many g of oxygen will be given off?

<u>Limiting Reagents</u>: - All reactants in a chemical rxn may not be used.

LR - limiting reagent ER - excess reagent

To determine which reactant is the LR, we calculate the amount of product expected from each rxt. AND the rxt that gives the smallest yield limits the amount of product formed (LR)

(study guide – p. 65 – look at ex. 3.9A)

- Check to be sure you have a balanced equation
- Convert the amount of reagent one that was given into the amount of product that you could form if that reagent was completely consumed.
- Convert the amount of reagent two that was given into the amount of product that you could form if that reagent was completely consumed.
- <u>Determine which of the answers in step 2 and 3 produced the LEAST amount of product and that will be your limiting reagent. Also the LEAST amount of product will be your theoretical yield.</u>

Example: 6 grams of H_2 and 160 grams of O_2. What is the limiting reagent and what is the theoretical yield in grams? If 20 grams of water is actually formed, what would be the Percent Yield?

$H_2 + O_2 \text{----->} H_2O$

1. check for balance $2H_2 + O_2 \text{---->} 2H_2O$

2. *Convert grams of given to moles of given H2*

6 grams H_2 X 1 mole H_2 / 2 grams H_2 = 3 moles H_2

3. Determine the moles of H_2O formed if all 3 moles of H_2 was consumed: We know from the balanced equation (coefficients) that: 2 moles H_2 produces 1 mole H_2O or 2 moles H_2 = 2 mole H_2O so ...

3 moles H_2 X 2 mole H_2O / 2 moles H_2 = 3 moles H_2O produced

4. Convert the grams of O_2 given to moles of O_2

160 grams O_2 X 1 mole O_2 / 32 grams O_2 = 5 moles O_2

5. *Determine the moles of H_2O formed if all 5 moles of O_2 was consumed:*

We know from the balanced equation that for every 1 mole of O_2 that is consumed then 2 moles H_2O is produced or 1 mole O_2 = 2 moles H_2O so 5 moles of O_2 X 2 moles H_2O / 1 mole O_2 = 10 moles H_2O

6. Compare the answers in step 3 and 5 and the one that produced the LEAST amount of water is the limiting reagent:

H_2 produced 3 moles water and O_2 produced 10 moles H_2O so...

H_2 is the limiting reagent and 3 moles of H_2O is the theoretical yield.

7. Convert the moles of H_2O theoretical yield to grams H_2O

1 mole H_2O = 18 grams

3 moles H_2O X 18 grams H_2O / 1 mole H_2O = 54 grams H_2O

8. Determine the Percent Yield - The Actual Yield was given as 20 grams

Percent Yield = (Actual Yield / Theoretical Yield) 100 = (20 grams / 54 grams) 100 = 37%

MOLARITY: Measures concentration of a solution

moles of solute/L of solution

ex. Calculate the molarity of a solution prepared by dissolving 1.56 g of HCl (g)in enough water to make 26.8 mL of solution.

<u>Stoichiometry Problem Involving Molarity:</u>

What is the molar concentration of $AgNO_3$ in a solution if titration of 25.00 mL of the solution with .300 M NaCl requires 37.05 mL of the NaCl solution to reach the endpt?

$AgNO_3$ + NaCl -----→

$$0.03705 \text{ L NaCl} \times \frac{0.300 \text{ mol NaCl}}{1 \text{ L NaCl}} \times \frac{1 \text{ mol } AgNO_3}{1 \text{ mol NaCl}} = 0.1111 \text{ mol } AgNO_3$$

M = mol/L

$$\frac{0.01111 \text{ mol } AgNO_3}{0.02500\text{L } AgNO_3} = 0.445 \text{ M } AgNO_3$$

How many milliliters of 0.5 *M* hydrochloric acid is required to react with an excess of zinc metal to produce 98.5 L of hydrogen gas @ STP. [Assume 100% yield]

Step 1) Write a balanced chemical equation

$$Zn(s) \quad + \quad 2\,HCl(aq) \quad \rightarrow \quad ZnCl_2(aq) \quad + \quad H_2\,(g)$$

Step 2) Add given information to equation

$$Zn(s) \quad + \quad 2\,HCl(aq) \quad \rightarrow \quad ZnCl_2(aq) \quad + \quad H_2\,(g)$$

"excess" x mL of 0.5 *M* 98.5 L @ STP

Step 3) Solve for number of moles of HCl needed

$$\text{x mol HCl} = 98.5\text{ L } H_2\left(\frac{1\text{ mol } H_2}{22.4\text{ L } H_2}\right)\left(\frac{2\text{ mol HCl}}{1\text{ mol } H_2}\right) = 8.8\text{ mol HCl}$$

Step 3) Solve for volume of HCl

$$M = \frac{\text{mol}}{\text{L}} \Rightarrow 0.5\,M = \frac{8.8\text{ mol HCl}}{\text{x L}}$$

$$\text{x} = 17.6\text{ L}\left(\frac{1000\text{ mL}}{1\text{ L}}\right) \Rightarrow 17{,}600\text{ mL of } 0.5\,M\text{ HCl}$$