**Limiting-Excess- Theoretical Yield- % Yield**

1. You must always first balance a reaction, make sure the number of moles of each element is present in both the reactant and product side.

2 C2H6 + 7 O2 🡪 4 CO2 + 6 H2O

2. Given starting amounts of both reactants, you must find the limiting and theoretical yield. The theoretical yield is the maximum amount of product you can make if all of the limiting reactant is used.

150 grams of C2H6 and 320 grams of oxygen gas are given for problem above.

To Find limiting- identify which product you are going to use to compare, in our case H2O

**C2H6**

150 g C2H6 X (1mole C2H6/ 30 g) this steps changes grams of reactant to moles

X(6 moles of H2O/ 2moles of C2H6) this step converts moles of given to moles of what you want

X(18g of H2O/1 mole of H2O) this step converts moles of what you want to grams

**So if you use all of the 150 g C2H6 you make 16200/60= 270 g of H2O**

**O2**

320 g O2 X (1mole O2/ 32 g of O2)

X ( 6 moles H2O/ 7 moles of O2)

X ( 18g H2O/ 1mole H2O= 34560/224= 150 g H2O

**So if you use all 320 g of O2 you make 150 g of H2O**

**As you can see O2 is my limiting because it restricts the amount of water that I would make. 150 g of H2O is the theoretical yield of water. The maximum amount I can make.**

3. To find the excess amount you compare the limiting reactant to the excess one.

320 g of O2 X (1mole O2/ 32g)

X (2 moles C2H6/ 7 moles of O2)

X (30 g of C2H6/ 1 mole C2H6) = 19200/224= 86 g of C2H6 were used

**Excess = 150g (amount started)- 86 g (amount used) = 64g using sig figs rules of subtraction= 60g C2H6 are in excess**

4. % yield = (actual yield/ theoretical yield) X 100

So if the reaction had a% yield of 81.5% to find actual yield

81.5= (A/150) X100 A= 120g H2O

**How to find Moles?**

Remember you **must** use moles when using a balanced equation. How to find moles?

1. Given, mass (grams) you must divide by the molar mass of the compound: n= mass/MM

Example: 12.5 grams of CO2 🡪 moles of CO2

12.5 g x(1 mole CO2/ Molar mass =44)= 12.5/44= .284 moles CO2

1. Given Pressure, Volume and Temperature. You can use: PV=nRT; this is for gases.

P= kPa V= dm3 = Liters T= Kelvin= C+273 R= 8.31

Example: Find moles of a gas given a pressure of 125 kPa, volume of 150 cm3 at 35C

Change cm3 to dm3 by dividing by a 1000, remember cm3 = milliliters

Change Celsius to Kelvin; K= C + 273

Also remember that if they give you Pa you divide by 1000 to make it kPa

Given m3 you divide by 1000 to make it dm3

PV=nRT

125 x.150 = n x8.31 x 308

n= .0073 moles

1. Given a solution; Molarity = moles of solute/ dm# of solution; c=n/dm3

Find moles of a solution of NaCl that consists of 150 cm3 of 0.200 mol/dm3 of NaCl

So using c= n/dm3

n= .150 x .200= .030 moles of NaCl

1. density = mass/ volume

D= grams/dm3

And Grams = moles X Molar Mass

So; D= (moles x Molar Mass)/ dm3

Find the Molar mass of a gas that has a density of 0.81g cm -3 and you have 2.1 moles of it in a 250 cm3 volume.

0.81 x 250= 2.1 X MM

MM= 96 g/mol

1. Remember Avogadro’s Law: If Pressure and Temperature are constant then moles are proportionally related to volume

PV= nRT Avogadro’s Law V=kn in which k= RT/P a constant value

If this is the case, then you can find Limiting and theoretical Volumes by using Volume as moles

Example:

2 C2H6(g) + 7 O2 (g) 🡪 4 CO2 (g) + 6 H2O (g)

Given 200. cm3 of C2H6 and150. cm3 of O2 at constant temperature and pressure Find volume of CO2

200. cm3 of C2H6  x (4 CO2/ 2 C2H6) = 400. cm3 of CO2

150. cm3 of O2 x ( 4 CO2/ 7 O2) = 85.7 cm3 of CO2

So your Theoretical Yield of CO2 is 85.7 cm3

**Finding Empirical and Molecular Formulas**

1. Given mass r % of each element you find Empirical Formula by first dividing each amount of each element by its atomic mass turning them into moles and the dividing all by the smallest moles.

Example:

Carbon = 50.70%, Hydrogen = 9.86 % and Nitrogen =39.44% and the molecular mass is 213 g/mol

Step 1 Turn into moles

C= 50.70/12= 4.225

H= 9.86/1= 9.86

N= 39.44/14= 2.817

Step 2 Divide by smallest mole

C= 4.225/2.817= 1.5

H= 9.86 / 2.817= 3.5

N= 2.817/2.917= 1

Notice that two of the element send in .5 moles so we must multiply by 2 to make them whole numbers. Had any ended in .333 I must multiply by 3

So Empirical Formula is C3H7N2

Now notice they gave me a molecular mass of 213 g-mol-1

So I find molar mass of Empirical Formula:

3 C + 7H + 2N= 36+ 7+ 28= 71

Molecular mass/ Empirical mass 🡪 213/71 = 3

So Molecular Formula is

3(Empirical Formula) 🡪 C9H21N6

**Empirical Formula given Combustion Data**

A sample of 3.50g of a compound that consists of Carbon, Hydrogen and Oxygen is burned in excess oxygen gas to produce 6.71g of CO2 and 4.14 g of H2O. Fine Empirical formula of the compound.

Since compound was burned in excess O2 it means that all the carbon has turned into CO2 and all the Hydrogen has turned into water

6.71 g of CO2 x (1mole of CO2/ 44 g)= .153 moles of CO2 = .153 mole C= !.83 g C

4.14 g of H2O x ( 1 mole of H2O/ 18g)= .23 moles of H2O = .46 moles of H= .46g H

The mass of Oxygen is the found by subtractive masses of C and H from total amount.

3.50- (1.83+..46) =1.21g

Now turn them all into moles

C= 1.83/12= .153

H= 0.46/1= 0.46

O= 1.21/16= 0.756

Now divide by smallest mole

C= .153/.0756= 2

H= 0.46/0.756= 6

O= 0.0756/0.756= 1

So Empirical Formula is C2H6O

**Calculating the Formula of a Hydrate**

A hydrate is an ionic compound that its found in crystal form (with water)

Example CaF2 3H2O

This means that there are three moles of water for every mole of ionic anhydrate.

The formula is always compared to 1 mole of anhydrate.

Example

A student is given a sample of a hydrate with the Formula CaF2 XH2O

The mass of the sample is found to be 4.32 g

After heating the compound three times the mass is found to remain constant at 2.55g

Find the Formula of Hydrate

Mass of water= 4.32 – 2.55= 1.69

Mass of anhydrate (CaF2) = 2.55

Moles of water= 1.69/18= 0.094

Moles of anhydrate= 2.55/78= 0.033

X= Moles of water/ moles of anhydrate= 0.94/0.33= 3

So Formula of Hydrate is CaF2 3H2O

**Gas Laws**

Ideal Gas Law PV=nRT

Combined Gas Law (P1V1)/ (P2V2)= T1/T2 remember T in Kelvin

**Dilutions**

M1V1= M2V2

A student wants to make a 125 cm3 0.15 mol-dm-3 solution of NaOH from a bottle of 12.0 mol-dm-3 NaOH solution.

So

0.15 x 125= 12.0 x V

V= 1.6 cm3 of the 12.0 mol-dm-3 NaOH solution would need to be added to 123.4 cm3 of distilled water to make the solution they wanted. Notice that you can keep volumes in cm3 if you want answer in cm3 , but if you are asked for moles in each you must change to dm3

**Back Titrations**

A student has a sample of 20.0 g that includes CaCO3. Your job is to find mass of CaCO3 in the sample.

Since you know that CaCO3 reacts with HCl. You assume that the rest of sample wont react with HCl.

So first you add excess HCl to get all the CaCO3 to reacts.

CaCO3 +2 HCl 🡪 CaCl2 + H2O + CO2

1. Student added 500. cm3 of 0.12 mol-dm-3 of HCl to the 20.0g sample. The student then took a 25 cm3 sample of the solution and titrated it with 0.10 mol-dm=3  NaOH. It took 20. cm3 of NaOH solution to neutralize the acid in the 25 cm3 sample.

Two reactions are occurring

First the Calcium Carbonate reacts with excess acid

CaCO3 +2 HCl 🡪 CaCl2 + H2O + CO2

Second the neutralization of the acid sample.

NaOH + HCl 🡪 H2O + NaCl

A. Find moles of acid added to original sample.

Moles= c X dm3 so; .500 x 0.12=0.060 moles of HCl were added

B. Find moles of base added to 25. cm3 of acid smple

Moles= c X dm3 so; 0.020x 0.10= 0.0020 moles of NaOH was added

C. Find moles of acid in 25 cm3 sample.

Since Neutralization reaction shows relationship between NaOH and HCl is 1:1 then you had 0.0020 moles of HCl in the 25 cm3 sample.

D. Find moles of HCl in excess in the total solution.

0.0020 moles of HCl in a 25cm3; then in a 500. cm3 sample is

0.0020 x 500/25= 0.040 moles of HCl are in excess

E. How many moles of HCl reacted with CaCO3?

0.060-0.040= 0.020 moles of HCl reacted with CaCO3

F. Find mass of CaCO3 that was in the sample.

According to the reaction between CaCO3 and HCl. The ratio is 1 CaCO3 reacts with 2 HCl. So that means that;

0.020 moles HCl reacted x 1CaCO3 / 2 HCl= 0.010 moles of CaCO3

Mass of CaCO3 is 0.010 moles CaCO3 x 100g CaCO3/1 mole= 1.0 g CaCO3

G. Find % of CaCO3 in sample= 1.0/20.0 X100= 5.00%