Stoichiometry!

Atomic Mass

- Atoms are so small, it is difficult to discuss how much they weigh in grams
- Use atomic mass units.
- an atomic mass unit (amu) is one twelfth the mass of a carbon-12 atom
- This gives us a basis for comparison
- The decimal numbers on the table are atomic masses in amu

They are not whole numbers

- Because they are based on averages of atoms and of isotopes.
- can figure out the average atomic mass from the mass of the isotopes and their relative abundance.
- add up the percent as decimals times the masses of the isotopes.

Examples

- There are two isotopes of carbon ^{12}C with a mass of 12.00000 amu(98.892%), and ^{13}C with a mass of 13.00335 amu (1.108%)
- There are two isotopes of nitrogen , one with an atomic mass of 14.0031 amu and one with a mass of 15.0001 amu. What is the percent abundance of each?

The Mole

- The mole is a number
- a very large number, but still, just a number
- 6.022×10^{23} of anything is a mole
- a large dozen
- The number of atoms in exactly 12 grams of carbon-12

The Mole

- Makes the numbers on the table the mass of the average atom
- Average atomic mass
- Just atomic mass

Molar mass

- mass of 1 mole of a substance
- often called molecular weight.
- To determine the molar mass of an element, look on the table.
- To determine the molar mass of a compound, add up the molar masses of the elements that make it up.

Find the molar mass of

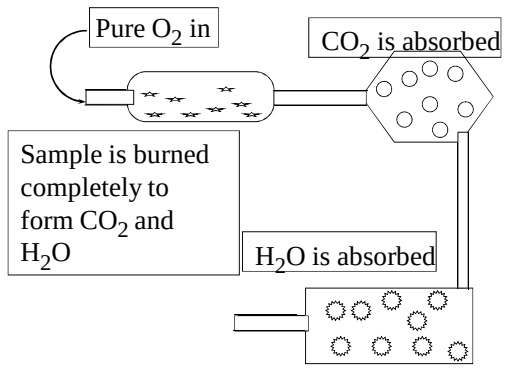
- CH_4
- Mg_3P_2
- $Ca(NO_3)_2$
- $Al_2(Cr_2O_7)_3$
- $CaSO_4 \cdot 2H_2O$

Percent Composition

- Percent of each element a compound is composed of.
- Find the mass of each element, divide by the total mass, multiply by a 100.
- Easiest if you use a mole of the compound.
- find the percent composition of CH_4
- $Al_2(Cr_2O_7)_3$
- $CaSO_4 \cdot 2H_2O$

Working backwards

- From percent composition, you can determine the empirical formula.
- Empirical Formula the lowest ratio of atoms in a molecule
- Based on mole ratios
- A sample is 59.53% C, 5.38%H, 10.68%N, and 24.40%O what is its empirical formula.



- A 0.2000 gram sample of a compound (vitamin C) composed of only C, H, and O is burned completely with excess O_2 . 0.2998 g of CO_2 and 0.0819 g of H_2O are produced. What is the empirical formula?

Empirical To Molecular Formulas

- Empirical is lowest ratio
- Molecular is actual molecule
- Need Molar mass
- Ratio of empirical to molar mass will tell you the molecular formula
- Must be a whole number because...

Example

- A compound is made of only sulfur and nitrogen. It is 69.6% S by mass. Its molar mass is 184 g/mol. What is its formula?

Chemical Equations

- Are sentences.
- Describe what happens in a chemical reaction.
- Reactants → Products
- Equations should be balanced
- Have the same number of each kind of atoms on both sides because ...

Balancing equations

$CH_4 + O_2 \rightarrow CO_2 + H_2O$

Reactants Products

1 C 1

4 H 2

2 O 3

Balancing equations

$CH_4 + O_2 \rightarrow CO_2 + 2\ H_2O$

Reactants Products

1 C 1

4 H ~~2~~ 4

2 O 3

Balancing equations

$CH_4 + O_2 \rightarrow CO_2 + 2\ H_2O$

Reactants Products

1 C 1

4 H ~~2~~ 4

2 O ~~3~~ 4

Balancing equations

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Reactants Products

1 C 1

4 H ~~2~~ 4

4 ~~2~~ O ~~3~~ 4

Abbreviations

- (*s*) , ↓ (for product)
- (*g*) , ↑ (for product)
- (*aq*)
- heat →
- Δ →
- catalyst →

Practice

- $Ca(OH)_2 + H_3PO_4 \rightarrow H_2O + Ca_3(PO_4)_2$
- $KClO_3(s) \rightarrow Cl_2(g) + O_2(g)$
- Solid iron(III) sulfide reacts with gaseous hydrogen chloride to form solid iron(III) chloride and dihydrogen monosulfide gas.
- $Fe_2O_3(s) + Al(s) \rightarrow Fe(s) + Al_2O_3(s)$

Meaning

- A balanced equation can be used to describe a reaction in molecules and atoms.
- Not grams.
- Chemical reactions happen molecules at a time
- or dozens of molecules at a time
- or moles of molecules.

Stoichiometry

- Given an amount of either starting material or product, determining the other quantities.
- use conversion factors from
  - molar mass (g - mole)
  - balanced equation (mole - mole)
- keep track

Examples

- One way of producing $O_2(g)$ involves the decomposition of potassium chlorate into potassium chloride and oxygen gas. A 25.5 g sample of Potassium chlorate is decomposed. How many moles of $O_2(g)$ are produced?
- How many grams of potassium chloride?
- How many grams of oxygen?

Examples

- A piece of aluminum foil 5.11 in x 3.23 in x 0.0381 in is dissolved in excess HCl(aq). How many grams of $H_2(g)$ are produced?
- How many grams of each reactant are needed to produce 15 grams of iron from the following reaction?

$$Fe_2O_3(s) + Al(s) \rightarrow Fe(s) + Al_2O_3(s)$$

Examples

- $K_2PtCl_4(aq) + NH_3(aq) \rightarrow Pt(NH_3)_2Cl_2\ (s) + KCl(aq)$
- what mass of $Pt(NH_3)_2Cl_2$ can be produced from 65 g of K_2PtCl_4 ?
- How much KCl will be produced?
- How much from 65 grams of NH_3?

Yield

How much you get from an chemical reaction

Limiting Reagent

- Reactant that determines the amount of product formed.
- The one you run out of first.
- Makes the least product.
- Book shows you a ratio method.
- It works.
- So does mine

Limiting reagent

- To determine the limiting reagent requires that you do two stoichiometry problems.
- Figure out how much product each reactant makes.
- The one that makes the least is the limiting reagent.

Example

- Ammonia is produced by the following reaction

$$N_2 + H_2 \rightarrow NH_3$$

What mass of ammonia can be produced from a mixture of 500. g N_2 and 100. g H_2 ?

- How much unreacted material remains?

Example

- A 2.00 g sample of ammonia is mixed with 4.00 g of oxygen. Which is the limiting reactant and how much excess reactant remains after the reaction has stopped?
- $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$

Excess Reagent

- The reactant you don't run out of.
- The amount of stuff you make is the yield.
- The theoretical yield is the amount you would make if everything went perfect.
- The actual yield is what you make in the lab.

Percent Yield

- $\%\ \text{yield} = \frac{\text{Actual}}{\text{Theoretical}} \times 100\%$
- $\%\ \text{yield} = \frac{\text{what you got}}{\text{what you could have got}} \times 100\%$

Examples

- Aluminum burns in bromine producing aluminum bromide. In a laboratory 6.0 g of aluminum reacts with excess bromine. 50.3 g of aluminum bromide are produced. What are the three types of yield.

Examples

- Years of experience have proven that the percent yield for the following reaction is 74.3%

  $Hg + Br_2 \rightarrow HgBr_2$

  If 10.0 g of Hg and 9.00 g of Br_2 are reacted, how much $HgBr_2$ will be produced?
- If the reaction did go to completion, how much excess reagent would be left?

Examples

- Commercial brass is an alloy of Cu and Zn. It reacts with HCl by the following reaction $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

  Cu does not react. When 0.5065 g of brass is reacted with excess HCl, 0.0985 g of $ZnCl_2$ are eventually isolated. What is the composition of the brass?