The Mole

Atomic mass units and atoms are not convenient units to work with.
The concept of the mole was invented. This was the number of atoms of carbon-12 that were needed to make 12 g of carbon.
1 mole = 6.02×10^{23} things

One atom of C-12 has a mass of 12 amu.
One mole of C-12 has a mass of 12 g.
Grams we can use more easily.

Using the periodic table, we can find the molar mass of different atoms.
Molar mass is the mass of 1 mole.
The molar mass of sulfur is 32.06g mol^{-1}

Ex 1:

a) Find the number of moles in 20.0 g of sulfur.

b) How many sulfur atoms in 2.75 moles of sulfur?

c) What is the mass of 8.3×10^{19} atoms of sulfur?

d) How many moles in 12.0 g of lithium?

e) How many atoms in 100.0 g of barium?

A molecule is treated in much the same way as an atom.
Molecular mass is the mass, in amu's, of one molecule of a substance.
Relative Molecular mass, $\mathbf{M_r}$, is the mass of the substance relative to the mass of C-12. It has no units, but same numerical value as molecular mass.
Molar Mass is the mass of one mole of a substance. Units are g mol^{-1}.

Because ionic compounds do not form discrete molecules, the term **formula mass** can be substituted for molecular mass.

Ex 2:

a) Find the relative molecular mass of carbon tetrachloride.

b) Find the formula mass of ammonium phosphate.

Molecular conversions to moles are accomplished similarly to atomic conversions.

(Draw the equality diagram for water here)

Ex 3:

a) Find the mass of 2.5 moles of NaCl.

b) Find the number of molecules in 15 g of hydrochloric acid.

c) How many atoms in 6.6 moles of Aluminium oxide. (Careful, there are 5 atoms in one formula unit of aluminium oxide.)

d) How many hydrogen atoms in 7.0 g of water?

Follow Up Problems 3.1, 2
Problems 3.2, 8, 12, 14, 16

2 Percent Composition

The mole allows a convenient conversion from mass to the number of particles (atoms or molecules). This will be used often because we measure mass very easily.
A chemical formula provides the number of atoms in a molecule or formula unit.
These amounts can be converted to masses.
To determine the mass of each element in a compound we compare the mass of that element to the mass of the compound.
It is most convenient to assume a 1 mole sample:

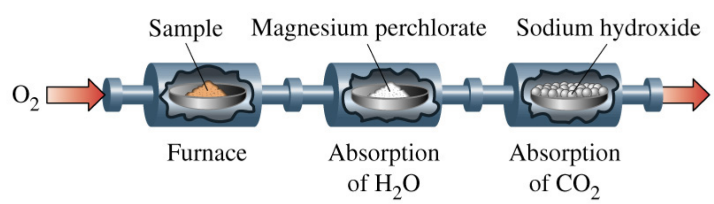
$$\%\ X = \left(\frac{\text{mass of X}}{\text{mass of compound}}\right) x\ 100\%$$

Ex 1:
Find the percent composition of each element in:

a) Ethanol C_2H_6O

b) Potassium sulfite

For an unknown compound, the first step in identifying the unknown is to perform a compositional analysis.
For carbohydrates (only have C, H, and O) this is performed by burning the sample completely.



The mass of each component can be determined.
They masses must be converted to moles because the formula is a ratio of the number of atoms.

Empirical Formula is the simplest ratio of the elements in a compound.
Molecular Formula is the actual number of elements present in a covalent compound.

Ex 2:
Give the empirical formula for:

a) glucose: $C_6H_{12}O_6$

b) sucrose: $C_{12}H_{24}O_{12}$

c) C_6H_6

d) $C_4H_5O_3$

What are the possible molecular formula for the empirical formula: C_2H_3O

To find the formula from the percent composition, we convert all mass percents to moles, and then find the ratio of moles.

Ex 3:
A sample is found to be 85.6% carbon and 14.4 % hydrogen. What is the empirical formula?

Ex 4: A compound is found to be 42.1 % sodium, 18.9 % phosphorous, and the rest oxygen. What is the empirical formula?

To find the molecular formula the molar mass of the compound must be compared to the molar mass of the empirical formula.

Why do ionic compound only have empirical formula?

Ex 5:
A compound has 49.2% C, 43.8% O and the rest hydrogen. Its molar mass is found to be 146.1 g mol^{-1}. What is the empirical formula and the molecular formula?

Follow Up Problems 3.3, 4, 5
Problems 3. 33, 35, 37, 39, 44

3 Chemical Equations

A chemical equation represents how atoms will rearrange from their original compound (reactants or reagents) to their final compounds (products)

$$\text{reactants}_{(\text{state})} \xrightarrow{\text{conditions}} \text{products}_{(\text{state})}$$

By convention, reactants ore on the left, products are on the right.
State symbols:
(*s*) solid
(*l*) liquid
(*g*) gas
(*aq*) aqueous, dissolved in water
Conditions are: temperature, catalysts, other reaction factors…

A chemical reaction must be balanced.

Each type of atom must be present in the same amount on both sides of the equation.
Consistent with conservation of mass and Dalton's atomic theory, atoms are not destroyed, only rearranged.

To balance an equation, coefficients are used. These are numbers in front of each compound that indicate multiples of the compound.

$$2H_{2(g)} + O_{2(g)} \xrightarrow{\text{Pt}} 2H_2O_{(l)}$$

This means 2 molecules of hydrogen gas react with 1 molecule of oxygen gas with a platinum catalyst to form 2 molecules of liquid water.
It also means that 2 <u>moles</u> of hydrogen gas react with 1 <u>mole</u> of oxygen gas to produce 2 <u>moles</u> of liquid water.
Hints:

- 7 elements are diatomic: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2.
- If these elements are found on their own, they will be diatomic. This only applies to these elements not in compounds.
- Balance elements last.
- If the atoms in a polyatomic ion are only present in these ions, they can be treated as one thing.

Ex 1: Balanced the following:

a) $Na_{(s)}$ + $Cl_{2(g)}$ → $NaCl_{(s)}$

b) $C_3H_{8(l)}$ + $O_{2(g)}$ → $CO_{2(g)}$ + $H_2O_{(g)}$

c) $Al_2(SO_4)_3$ + $BaCl_2$ → $AlCl_3$ + $Ba(SO_4)$

d) $Ca_3(PO_4)_2$+ SiO_2+ C→ $CaSiO_3$+ CO+ P

e) Silver metal reacts with sulfur trioxide gas to produce silver sulfide and oxygen gas

Follow Up Problems 3.7
Problems 3.51, 53, 55

4 Stoichiometry

Stoichiometry is a branch of chemistry that deals with calculating the amounts of reactants and products involved in a chemical reaction.
To perform these calculations, we use a balanced equation and molar amounts of reactants and products.

Steps for a stoichiometric calculation:

1) Determine the balanced reaction.
2) Convert the amount you are given into moles.
3) Convert from moles of the known to moles of the unknown based on the coefficients in the balanced chemical reaction.
4) Convert from moles of the unknown to the units required.

Ex 1:
For the combustion of hydrogen:

$$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$$

a) How much water vapour is produced by 5.0 g of hydrogen?

b) What mass of hydrogen is needed to react with 235 g of oxygen gas?

c) What mass of oxygen is required to produce 2.5 moles of water vapour?

Ex 2: Aluminium chloride reacts with tin(IV)oxide to produce aluminium oxide and tin(IV)chloride.

a) What mass of aluminium chloride is required to react with 1.00 g of the tin oxide?

b) What mass of each product would you expect from the reaction in (a)?

Follow Up Problems
3.8
Problems
3.61, 63, 65, 67

5 Limiting Reactant and Percent Yield

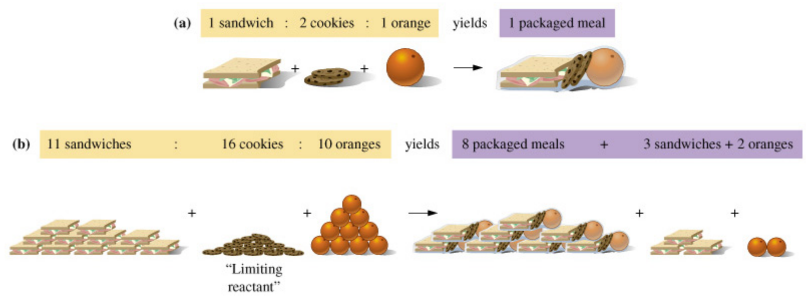
For the reaction: $2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$

2 moles of hydrogen (4.0 g) should react with 1 mole of oxygen (32.0 g) to produce 2 moles of water (36.0 g)

We must consider 2 possibilities.

1) What if there is a shortage of one of the reactants?
2) What if the reaction does not react as we expect and produces too little of too much product.

If one of the reactants is exhausted, the reaction stops. We call this reactant the limiting reactant (or limiting reagent).
The reactant that is left over is said to be in excess.



The questions that are commonly asked in reference to limiting reactants are:

1) Which reactant is the limiting reactant?
2) How much of the other reactant(s) is(are) in excess?
3) How much product is expected?

The easiest method to determine the limiting reactant is to calculate the amount of product formed be each reactant.

The reactant that produces the least product in the limiting reactant.

The limiting reactant can then be used to calculate the amount of the excess reactant that is needed.

Ex 1:
Hydrochloric acid reacts with magnesium hydroxide to produce water and aqueous magnesium chloride. How much water is produced if 20.0 g of hydrochloric acid reacts with 10.0 g of magnesium hydroxide? How much of the excess reactant is left over?

In a chemical reaction we write the reaction that we expect.
However, nature may not be accommodating.

- The reaction may stop before it is complete
- Other reactions may take place
- Samples may not be pure
- Other problems

To express how close reality is to our expectations, we use percent yield:

$$\%\text{ yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \text{x } 100\%$$

Actual yield is the result from the experiment.
Theoretical yield is the result from your stoichiometric calculations.

Ex 2:
A reaction of 2.00 g of hydrogen with excess oxygen produced 15.8 g of water. What is the yield of this experiment?

Ex 3:
24.0 g of Iodine react with 0.20 g of hydrogen to produce 11.8 g of hydrogen iodide. What is the percent yield of this reaction?

Follow Up Problems 3.10, 11, 12
Problems 3.71, 73, 75, 77, 81, 83