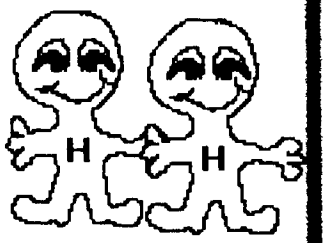
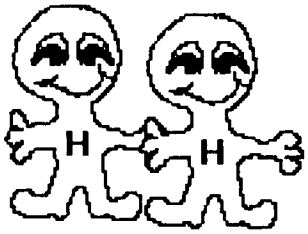
Chemical Equations

Confucius says:
A chemical equation is worth a thousand words!

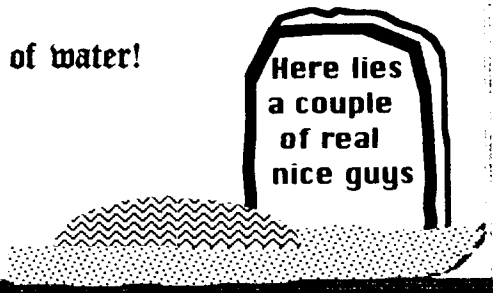
Once upon a time, there were two little hydrogen molecules floating around the room.





Little did they know, that a big, bad oxygen molecule was lurking amongst the air particles, waiting for the chance to shred their happy world.

Suddenly, the oxygen molecule attacked and sparks began to fly! In an instant, it was all over. The hydrogen molecules were ripped apart and forced into bondage by the oxygen atoms. All that remained was a small pool of water!



Unit 5

Unit 5: Chemical Equations

I. A balanced equation is worth a thousand words!!!

A **chemical equation** is a shorthand method of representing a chemical change that is either known or predicted to occur. The starting materials, called **reactants,** are listed first on the left side of the equation. An arrow is used to show the direction of change into the **products**, listed on the right side. Any special requirements needed to cause the reaction, such as heating or adding a catalyst, is listed above the arrow. All chemical equations <u>must</u> conform to the law of conservation of mass. This means that no atoms may be created or destroyed during a chemical change. Therefore, the total number of atoms of each element and their masses MUST always remain the same in the reactants and products.

Word equations provide only the names of the reactants and products involved in a chemical reaction.

Example:

$$\underset{\text{reactants}}{\text{table sugar + potassium chlorate}} \xrightarrow{H_2SO_4} \underset{\text{products}}{\text{carbon dioxide + water + potassium chloride}}$$

Balanced formula equations are more detailed, because the formulas show exactly which elements and how many atoms of each are involved in the reaction. **Coefficients** placed in front of the chemical formulas to balance the total number of atoms of each element also indicate the mole ratios between the reactants and products. These mole ratios can be converted into equivalent mass ratios, number of particle ratios, or volume ratios for gases.(Review the last unit to see how these values are calculated) The balanced formula equation for the above word equation is:

$$C_{12}H_{22}O_{11}(s) + 8KClO_3(s) \xrightarrow{H_2SO_4(aq)} 12CO_2(g) + 11H_2O(g) + 8KCl(s)$$

mole ratio:	1 mole	8 moles	12 moles	11 moles	8 moles
mass ratio:	342 g	981 g	528 g	198 g	597 g
volume ratio: (at STP for gases)	N/A	N/A	269 liters	246 liters	N/A

Note the phase symbols used to indicate the physical state of the reactants and products: (s) = solid; (aq) = aqueous, a water solution; (g) = gas; (l) = a pure liquid; (cr) = crystalline solid

II. Writing Balanced Formula Equations

Writing a balanced chemical equation involves basically two steps:

1. Formulas for all substances involved in the reaction must be written correctly. This is done by determining the formula and charge of the positive and negative ions in the compound, and adjusting the subscripts to make it electrically neutral. Atomic symbols are used to represent elements, and there are 7 elements (**HINClBrOF**) that exist as diatomic molecules in nature. Example: H_2, I_2, N_2, etc.
   *If an incorrect formula is written for a reactant or product, then the equation cannot be properly balanced.
2. Coefficients must be assigned to all formulas so that the number of atoms of each

kind of element in the reactants equals the number of atoms of each kind of element in the products. The values for the coefficients can usually be determined by inspection of the equation and trial and error. It is usually best to leave oxygens and hydrogens to the end, since they are commonly found in many of the reactants and products. Formulas of the substances may NOT be rewritten by changing subscripts in order to balance the equation.

Sample Problem 1:

To balance: $H_2SO_4 + Al_2(CO_3)_3 \text{ -----> } Al_2(SO_4)_3 + H_2O + CO_2$

Place a **3** in front of the H_2SO_4 to balance the S atoms:

$$\mathbf{3}H_2SO_4 + Al_2(CO_3)_3 \text{ -----> } Al_2(SO_4)_3 + H_2O + CO_2$$

The Al already balances with 2 atoms. Place a **3** in front of CO_2 to balance the C atoms:

$$3H_2SO_4 + Al_2(CO_3)_3 \text{ -----> } Al_2(SO_4)_3 + H_2O + \mathbf{3}CO_2$$

Place a **3** in front of the H_2O to balance the H atoms

$$3H_2SO_4 + Al_2(CO_3)_3 \text{ -----> } Al_2(SO_4)_3 + \mathbf{3}H_2O + 3CO_2$$

There are now 12 oxygen atoms in $\mathbf{3}H_2SO_4$, plus 9 more in $Al_2(CO_3)_3$. This equals the total oxygen atoms in the products from the 12 oxygens in $Al_2(SO_4)_3$, plus 3 O in $3H_2O$, plus 6 O in the $3CO_2$.

Sample Problem 2:

When balancing the reaction:

magnesium + hydrochloric acid ---> magnesium chloride + hydrogen gas

Write the correct formulas for the reactants and products based upon their names:

$$Mg + HCl \text{ ---> } MgCl_2 + H_2$$

There is only one Cl on the left and two on the right. A coefficient of **2** is placed in front of the HCl to balance the Cl. At the same time, this also balances the H atoms. Originally, there was only one H on the left, but two on the right, because hydrogen exists as a diatomic molecule. The balanced equation is:

$$Mg + \mathbf{2}HCl \text{ ---> } MgCl_2 + H_2$$

Problems: Write balanced formula equations for each of the following reactions:

1. $Na + H_2O \text{ --------> } NaOH + H_2$

Ans: 2:2:2:1

2. $Al(NO_3)_3 + KOH \text{ --------> } Al(OH)_3 + KNO_3$

Ans: 1:3:1:3

3. $UO_3 \text{ -------> } UO + O_2$

Ans: 1:1:1

4. $Hf + N_2 \text{ -------> } Hf_3N_4$

Ans: 3:2:1

5. Ga + sulfuric acid ---------> $Ga_2(SO_4)_3$ + H_2

Ans: 2:3:1:3

6. $PbCl_2$ + nitric acid ---------> $Pb(NO_3)_2$ + hydrochloric acid

Ans: 1:2:1:2

7. lithium bromide + silver chloride --------> silver bromide + lithium chloride

Ans: 1:1:1:1

8. phosphorus (V) iodide ---------> phosphorus + iodine

Ans: 2:2:5

9. oxygen + antimony(III) sulfide -----> diantimony tetroxide + sulfur dioxide

Ans: 5:1:1:3

10. copper + chlorine ---------> copper(II) chloride

Ans: 1:1:1

III. Balancing Complex Chemical Equations by Using Algebra....Imagine That!

In order to balance a complex chemical equation that is too difficult to solve by inspection, the coefficients can be determined algebraically as follows:

Step 1 : Write the correct chemical formulas for all reactants and products.

H_2SO_4 + $Al_2(CO_3)_3$ -----> $Al_2(SO_4)_3$ + H_2O + CO_2

Step 2 : Assign a variable(using a, b, c, etc.) for each reactant and product.

H_2SO_4 + $Al_2(CO_3)_3$ -----> $Al_2(SO_4)_3$ + H_2O + CO_2

a **b** **c** **d** **e**

Step 3 : Write an expression for each element, by counting the number of atoms present in each reactant or product and listing that number as a coefficient for the variable.

For hydrogen: $2a = 2d$
For sulfur: $a = 3c$
For oxygen: $4a + 9b = 12c + d + 2e$
For aluminum: $2b = 2c$
For carbon: $3b = e$

Step 4 : Let one variable have a value of 1, then solve algebraically for the values of the other variables.

Let a = 1: $2(1) = 2d$, therefore $d = 1$
For sulfur: $(1) = 3c$, therefore $c = 1/3$
For aluminum: $2b = 2(1/3)$, therefore $b = 1/3$
For carbon: $3(1/3) = e$, therefore $e = 1$
For oxygen: $4(1) + 9(1/3) = 12(1/3) + d + 2(1)$, therefore $d = 1$

Step 5 : Express all of the values as a whole number ratio, and use them as coefficients for balancing the equation.
Since b and c each have a value of 1/3, multiply all variable values by three. The final values for b and c would be 1, and the values for a, d and e would be 3.

$3H_2SO_4$ + $Al_2(CO_3)_3$ -----> $Al_2(SO_4)_3$ + $3H_2O$ + $3CO_2$

Step 6 : Recheck the balanced equation by counting the number of atoms of each element in the reactants and compare with the number of atoms of each in the products.
There are 6 hydrogen atoms on each side, 3 sulfur atoms, 21 oxygen atoms, 2 aluminum atoms and 3 carbon atoms. The equation is balanced.

Sample Problem 2 :

$$As_2O_3 + HNO_3 + H_2O \longrightarrow H_3AsO_4 + NO$$

a **b** **c** **d** **e**

For As: $2a = d$
For O: $3a + 3b + c = 4d + e$
For H: $b + 2c = 3d$
For N: $b = e$

Let $a = 1$ $\quad 2(1) = d$, therefore $d = 2$
$b + 2c = 3(2)$, solving for $c = (6 - b)/2$
$3(1) + 3b + (6 - b)/2 = 4(2) + b$, solving for $b = 4/3$, therefore $e = 4/3$ also.
Substituting 4/3 for the b variable, $c = (6 - 4/3)/2 = (14/3)/2 = 14/6$ or $7/3$
Multiply all values for the variables by 3 to change the fractions to whole numbers:
$a = 3$, $b = 4$, $c = 7$, $d = 6$, $e = 4$.

$$3As_2O_3 + 4HNO_3 + 7H_2O \longrightarrow 6H_3AsO_4 + 4NO$$

There are 6 arsenic atoms, 28 oxygen atoms, 18 hydrogen atoms and 4 nitrogen atoms on both sides of the equation. The reaction is balanced.

Problems:
Balance the following reactions using algebra:

1. $H_2O + P_4 + P_2I_4 \longrightarrow PH_4I + H_3PO_4$

Ans: 128 : 13 : 10 :40 : 32

2. $H_2SO_4 + K_2Cr_2O_7 + I_2 \longrightarrow K_2SO_4 + Cr_2(SO_4)_3 + H_2O + KIO_3$

Ans: 17 : 5 : 3 : 2 : 5 : 17 : 6

IV. Oxidation and Reduction Reactions

Many chemical reactions involve the atoms losing or gaining one or more negatively charged, subatomic particles, called **electrons**. The loss of electrons by an element is called **oxidation** and the gain of electrons is called **reduction**. (You can remember this difference by the expression, "**LEO** the lion goes **GER**".....**L**ose **E**lectrons **O**xidation, **G**ain **E**lectrons **R**eduction!) This type of reaction is sometimes called a **redox**

reaction. In any redox reaction, the total number of electrons lost <u>must equal</u> the total number of electrons gained. In order to determine which elements in the reaction are being oxidized or reduced, it is first necessary to assign oxidation numbers. You should recall that we used oxidation numbers when writing chemical formulas in order to balance the electrical charge of the compound. An **oxidation number** is a positive or negative number assigned to each atom in a compound to indicate if it more likely to gain electrons(becoming more negative in charge) or to lose electrons(becoming more positive in charge). The assigned oxidation number shows the <u>average</u> charge acquired by all of the atoms of a particular element, and as such, it may have a fractional value.

For example: A red-colored form of lead oxide has the empirical formula, Pb_3O_4. Since each oxygen atom will usually have a 2- oxidation number, each lead atom must have an average oxidation number of 8/3+. Actually, this compound is a mixture of lead (II) oxide and lead (IV) oxide, $2PbO \bullet PbO_2$.

<u>General Rules for Assigning Oxidation Numbers:</u>

1. The oxidation number of any uncombined atom in its elementary state is 0. This includes the 7 diatomic elements, HINClBrOF.
2. In ions containing only 1 atom, such as Na^{1+} or S^{2-}, the oxidation number is equal to the charge of the ion.
3. The oxidation number for oxygen is usually 2-. In peroxides it is 1-.
4. The oxidation number for hydrogen is 1+ in all its compounds, except in metallic hydrides, such as NaH, where it is 1-.
5. All other oxidation numbers are assigned so that the sum of oxidation numbers equals the net charge on the molecule or polyatomic ion.

Sample Problem:

Assign the oxidation numbers for the elements in the following reaction, identify which element is oxidized and which is reduced, and determine how many electrons are gained or lost per atom. $2Al + 3Cu(NO_3)_2 \text{ -------> } 2Al(NO_3)_3 + 3Cu$

- The Al and Cu are atoms in their elementary state, and have an oxidation number of 0.

$$\overset{0}{2Al} + 3Cu(NO_3)_2 \text{ -------> } 2Al(NO_3)_3 + 3\overset{0}{Cu}$$

- The nitrate group(NO_3) is a polyatomic ion that always has a net charge of 1-. Therefore, the oxidation number of Cu is 2+ in $Cu(NO_3)_2$ and the Al is 3+ in $Al(NO_3)_3$.

$$\overset{0}{2Al} + 3\overset{2+}{Cu}(NO_3)_2 \text{ -------> } 2\overset{3+}{Al}(NO_3)_3 + 3\overset{0}{Cu}$$

- Since oxygen is usually 2-, the oxidation number of N must be 5+ in both compounds. This will balance the net charge on the molecules to be equal to 0. It will also combine with the three O^{2-} to equal the nitrate's net ionic charge of 1-.

$$\overset{0}{2Al} + 3\overset{2+}{Cu}(\overset{5+}{N}\overset{2-}{O_3})_2 \text{ -------> } 2\overset{3+}{Al}(\overset{5+}{N}\overset{2-}{O_3})_3 + 3\overset{0}{Cu}$$

- Notice that the oxidation numbers of the N and O atoms do not change. These are called **spectator ions(elements)** because they are not changed during the redox reaction. Sometimes only the elements being oxidized and reduced are shown in what is called the **net ionic reaction**.
  - <u>Oxidation reaction</u>: $2Al^0 \text{ -------> } 2Al^{3+} + 6e(\text{electrons})$
  - <u>Reduction reaction</u>: $6e + 3Cu^{2+} \text{ -------> } 3Cu^0$
  - <u>Net Ionic Equation</u>: $2Al^0 + 3Cu^{2+} \text{ -------> } 2Al^{3+} + 3Cu^0$

Each Al changes from a neutral atom with no charge, Al^0, to Al^{3+} by losing 3 electrons per atom. Since there are two Al atoms per formula, a total of 6 electrons are lost. Each Cu changes from an ion with a 2+ charge into a neutral atom with no charge(Cu^{2+} to Cu^0) by gaining 2 electrons per atom. Since there are three Cu atoms per formula, a total of 6 electrons are gained.

Problems:
Assign the oxidation numbers for the elements in the following reactions, identify which element is oxidized and which is reduced, and determine how many electrons are gained or lost per atom.

1. $2AgNO_3 + Cu \longrightarrow Cu(NO_3)_2 + 2Ag$

Ans: red: $2Ag^{1+} \longrightarrow 2Ag^0$
gain 1 electron/atom; 2 total
ox: $Cu^0 \longrightarrow Cu^{2+}$
lose 2 electron/atom; 2 total

2. $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$

Ans: red: $Fe_2^{3+} \longrightarrow 2Fe^0$
gain 3 electron/atom; 6 total
ox: $3C^{2+} \longrightarrow 3C^{4+}$
lose 2 electron/atom; 6 total

3. $2K_2Cr_2O_7 + 2H_2O + 3S \longrightarrow 4KOH + 2Cr_2O_3 + 3SO_2$

Ans: red: $2Cr_2\overset{6+}{O_7}{}^{2-} \longrightarrow 2Cr_2^{3+}$
gain 3 electron/atom; 12 total
ox: $3S^0 \longrightarrow 3S^{4+}$
lose 4 electron/atom; 12 total

V. Balancing Redox Equations by the Electron Transfer Method

Since the total number of electrons lost MUST equal the total number of electrons gained, a redox equation can be balanced by comparing the increases and decreases in oxidation numbers using the **oxidation number change method**. Coefficients are used to make the total increase in oxidation number equal to the total decrease in oxidation number. Other spectator ions in the equation can then be balanced by inspection.

Sample Problem:
Assign oxidation numbers to all the atoms in the equation:

$$\overset{1+\,1-}{KI}(aq) + \overset{1+\,7+\,2-}{KIO_4}(aq) + \overset{1+\,1-}{HCl}(aq) \longrightarrow \overset{1+\,1-}{KCl}(aq) + \overset{0}{I_2}(s) + \overset{1+\,2-}{H_2O}(l)$$

Identify which atoms are oxidized and reduced, then draw a line to connect the atoms that undergo oxidation and one for those that undergo reduction. On the line, write the number of electrons lost(or gained) per atom and total per formula. Note that the I^{1-} in the KI is losing one electron while the I^{7+} in KIO_4 is gaining seven electrons and both form elemental iodine, I_2.

gain 7 electrons per atom

$$\overset{1+\,1-}{KI}(aq) + \overset{1+\,7+\,2-}{KIO_4}(aq) + \overset{1+\,1-}{HCl}(aq) \longrightarrow \overset{1+\,1-}{KCl}(aq) + \overset{0}{I_2}(s) + \overset{1+\,2-}{H_2O}(l)$$

lose 1 electron per atom

Make the total increase in oxidation number equal to the total decrease in oxidation number by adjusting the coefficients. A coefficient of 7 is placed in front of KI and 4 in front of I_2 in order to balance the electron transfer.

gain 7 electrons per atom(and total)

$$\overset{1+\ 1-}{7KI(aq)} + \overset{1+\ 7+\ 2-}{KIO_4(aq)} + \overset{1+\ 1-}{HCl(aq)} \text{ -------> } \overset{1+\ 1-}{KCl(aq)} + \overset{0}{4I_2(s)} + \overset{1+\ 2-}{H_2O(l)}$$

lose 1 electron per atom x 7 = 7 electrons total

Balance the other spectator ions, <u>without changing the coefficients for the elements involved in the oxidation and reduction reactions!</u> An 8 is placed in front of KCl to balance the K atoms. Then an 8 is needed in front of HCl to balance the Cl atoms. Finally, a 4 is placed in front of the H_2O to balance the 8 hydrogen and 4 oxygen atoms found in the reactants.

$$7KI(aq) + KIO_4(aq) + 8HCl(aq) \text{ -------> } 8KCl(aq) + 4I_2(s) + 4H_2O(l)$$

Problems:

Balance the following reactions using the **oxidation number change method** described above:

1. $Cu + HNO_3 \text{ -----> } Cu(NO_3)_2 + NO + H_2O$

Ans: 3 : 8 : 3 : 2 : 4

2. $As_2O_3 + Cl_2 + H_2O \text{ ----> } H_3AsO_4 + HCl$

Ans: 1 : 2 : 5 : 2 : 4

Sometimes, only the net ionic equation for the reaction is given. In this case, the **half-reaction method**(also called the ion-electron method) is used to determine the complete balanced equation. The overall reaction is divided into two parts, the oxidation half-reaction and the reduction half-reaction. These parts are balanced separately, then added together to give the balanced equation. The following steps are used in this method:

Step 1: Write the equation in ionic form. Assign the oxidation numbers for the elements.

Step 2: Write the separate oxidation and reduction half-reactions.

Step 3: Adjust the coefficients to get an equal number of atoms of all elements EXCEPT oxygen and hydrogen.

Step 4: Balance the hydrogen and oxygen atoms by adding H^{1+} and H_2O if the solution is acidic. Use OH^{1-} and H_2O for reactions in a basic solution.

Step 5: Balance the electric charge on both sides of each half-reaction by adding electrons.

Step 6: Multiply each half-reaction by an appropriate number to make the electron changes equal.

Step 7: Combine the two half-reactions and simplify(subtract substances that appear on both the reactant and product sides of the equation). Check to see if the numbers of atoms of each element and total charges of reactants equals the atoms and charges of the products.

Sample Problem:

Balance: $H_2SO_3 + V^{2+} \text{-------}> S_2O_3^{2-} + V^{3+}$ (acid solution)

Step 1: $\overset{1+\ 4+\ 2-}{H_2SO_3} + \overset{2+}{V^{2+}} \text{-------}> \overset{2+\ 2-}{S_2O_3^{2-}} + \overset{3+}{V^{3+}}$

Step 2: oxidation $\overset{2+}{V^{2+}} \text{-------}> \overset{3+}{V^{3+}}$

reduction $\overset{4+}{H_2SO_3} \text{-------}> \overset{2+}{S_2O_3^{2-}}$

Step 3: reduction $\mathbf{2}H_2SO_3 \text{-------}> S_2O_3^{2-}$

(no change for the oxidation half-reaction)

Step 4: reduction $2H_2SO_3 + \mathbf{2H^{1+}} \text{-------}> S_2O_3^{2-} + \mathbf{3H_2O}$

Step 5: oxidation $V^{2+} \text{-------}> V^{3+} + \mathbf{1e^-}$

reduction $2H_2SO_3 + 2H^{1+} + \mathbf{4e^-} \text{-------}> S_2O_3^{2-} + 3H_2O$

Step 6: oxidation $\mathbf{4}V^{2+} \text{-------}> \mathbf{4}V^{3+} + \mathbf{4}e^-$

reduction $2H_2SO_3 + 2H^{1+} + 4e^- \text{-------}> S_2O_3^{2-} + 3H_2O$

Step 7: $4V^{2+} + 2H_2SO_3 + 2H^{1+} \text{-------}> 4V^{3+} + S_2O_3^{2-} + 3H_2O$

Problems:

Balance the following reactions using the **half-reaction method** described above:

1. $Cl^{1-} + MnO_4^{1-} \text{----}> Cl_2 + Mn^{2+}$ (acid solution)

Ans: $10Cl^{1-} + 2MnO_4^{1-} + 16H^{1+} \text{----}> 5Cl_2 + 2Mn^{2+} + 8H_2O$

2. $S + Cr_2O_7^{2-} \text{----}> SO_2 + Cr_2O_3$ (basic solution)

Ans: $3S + 2Cr_2O_7^{2-} + 2H_2O \text{----}> 3SO_2 + 2Cr_2O_3 + 4OH^{1-}$

VI. Predicting Products in a Reaction

Most chemical reactions can be classified as being a particular type of reaction based upon the nature of the reactants involved. Each type follows a specific pattern, which can be used to predict what the products will be or even if the reaction will occur at all. Once the products have been determined, a balanced chemical equation can be written using the previously described methods.

Types of Reactions

Group I

Combination or Synthesis with elements

A combination(or synthesis) reaction involves 2 or more substances reacting to form a single substance. It follows the pattern: A + X -----> AX (a single compound is formed) Often the reaction involves 2 elements.

•Rules:

1) Combine the most metallic element with the least metallic element, and then balance the equation. The most metallic elements are located in the lower left corner of the periodic table. The least metallic elements are in the upper right corner.

   Example: sodium plus iodine ------->

   $Na + I_2$ ---------> NaI (unbalanced) becomes $2Na + I_2$ ------> 2NaI (balanced)

2) If there is a variable oxidation number metal or nonmetal , use the most common oxidation number as the probable one to use in determining the formula.

   Example: Iron + bromine --------->

   $Fe + Br_2$ ------> $FeBr_3$ (unbalanced, most common oxidation number of Fe is 3+)

   $2Fe + 3Br_2$ --------> $2 FeBr_3$ (balanced)

   Example: sulfur plus oxygen --------->

   $S + O_2$ --> SO_3 (unbalanced, most common positive oxidation number of S is 6+)

   $2S + 3O_2$ -------> $2SO_3$ (balanced)

Combination or Synthesis with compounds

2 or more simple compounds react to form a more complex compound. One of the reactants is often water.

•There are certain common patterns to follow:

1) nonmetallic oxides + water ------> ternary acids (H containing, ending in oxygen)

   Example: sulfur trioxide plus water ----->

   $SO_3 + H_2O$ ---------> H_2SO_4 (balanced)

2) metallic oxides + water ----------> bases (hydroxides)

   Example: sodium oxide plus water ----->

   $Na_2O + H_2O$ ---> NaOH (unbalanced) $Na_2O + H_2O$ ---> 2NaOH (balanced)

Combination or Synthesis with elemental oxygen and a compound

Elemental oxygen can react with an oxide compound containing a variable oxidation number metal or nonmetal, and further oxidize the positive element by adding extra oxygen atoms.

Example: carbon (II) oxide plus oxygen ----->

$CO + O_2$ -----> CO_2 (unbalanced) $2CO + O_2$ -----> $2CO_2$ (balanced)

Group II

Decomposition reactions

A single compound is broken down into 2 or more simple substances . This can involve forming any combination of elements and compounds and it follows the pattern:

AX -----> A + X Energy in the form of heat, light, or electricity is usually required. (Note the symbol, Δ = heat)

• There are certain patterns to follow for decomposition reactions:

1) a single <u>binary compound</u> decomposes into 2 elements

Example: water $\xrightarrow{\text{electricity}}$

(decomposition by an electric current in a process called <u>electrolysis</u>.)

$H_2O \xrightarrow{\text{electricity}} H_2 + O_2$ (unbalanced) $2H_2O \xrightarrow{\text{electricity}} 2H_2 + O_2$ (balanced)

Example: mercury (II) oxide $\xrightarrow{\text{heat}}$

$HgO \xrightarrow{\Delta} Hg + O_2$ (unbalanced) $2HgO \xrightarrow{\Delta} 2Hg + O_2$ (balanced)

2) <u>Metallic Carbonates</u> form metal oxides + carbon dioxide

Example: calcium carbonate $\xrightarrow{\text{heat}}$

$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$ (balanced)

3) <u>Metallic Hydroxides</u> form metal oxides + water

Example: copper (II) hydroxide $\xrightarrow{\text{heat}}$

$Cu(OH)_2 \xrightarrow{\Delta} CuO + H_2O$ (balanced)

4) <u>Metallic Chlorates</u> form metal chloride + oxygen

Example: potassium chlorate $\xrightarrow{\text{heat}}$

$KClO_3 \xrightarrow{\Delta} KCl + O_2$ (unbalanced) $2KClO_3 \xrightarrow{\Delta} 2KCl + 3O_2$(balanced)

5) <u>Oxyacids</u> form nonmetal oxides + water

Example: carbonic acid $\xrightarrow{\text{heat}}$

$H_2CO_3 \xrightarrow{\Delta} CO_2 + H_2O$ (balanced)

6) Some <u>Metallic Oxides</u> form (metal or metal oxide) + oxygen

Example: mercury (II) oxide $\xrightarrow{\text{heat}}$

$HgO \xrightarrow{\Delta} Hg + O_2$ (unbalanced) $2HgO \xrightarrow{\Delta} 2Hg + O_2$ (balanced)

Example: lead (IV) oxide $\xrightarrow{\text{heat}}$

$PbO_2 \xrightarrow{\Delta} PbO + O_2$ (unbalanced) $2PbO_2 \xrightarrow{\Delta} 2PbO + O_2$ (balanced)

(**Note**: some peroxides follow this same pattern)

Example: hydrogen peroxide $\xrightarrow{\text{heat}}$

$H_2O_2 \xrightarrow{\Delta} H_2O + O_2$ (unbalanced) $2H_2O_2 \xrightarrow{\Delta} 2H_2O + O_2$ (balanced)

Group III

Single Replacement or Displacement Reactions

Single replacement reactions occur when a more active free element replaces a less active element in a compound. (check an activity series chart to determine if the free element **is** more reactive) It follows the pattern: B + AX --->BX + A **or** Y + AX--> AY + X Note that either the positive or negative ion can be replaced to form a different element and a new compound. If the positive element formed has more than one oxidation number, assume complete oxidation occurs and <u>use the highest oxidation state</u>. There are 4 basic types of single replacement reactions:

1) Replacement of a metal by a more active metal
   Example: copper plus silver nitrate ------->
   $Cu(s) + AgNO_3(aq)$ -----> $Cu(NO_3)_2(aq) + Ag(s)$ (unbalanced)
   $Cu(s) + 2AgNO_3(aq)$ -----> $Cu(NO_3)_2(aq) + 2Ag(s)$ (balanced)

(Remember to use the highest oxidation state for Cu, which is the copper (II) ion)

2) Replacement of hydrogen in acids by a metal
   Example: zinc plus hydrochloric acid ------->
   $Zn(s) + HCl(aq)$ -----> $ZnCl_2(aq) + H_2(g)$ (unbalanced)
   $Zn(s) + 2HCl(aq)$ -----> $ZnCl_2(aq) + H_2(g)$ (balanced)

3) Replacement of hydrogen in water by very active metals (primarily Group IA metals)
   Example: sodium plus water ------->
   $Na(s) + H_2O(aq)$ -----> $NaOH(aq) + H_2(g)$ (unbalanced)
   $2Na(s) + 2H_2O(aq)$ -----> $2NaOH(aq) + H_2(g)$ (balanced)

4) Replacement of a halogen(Group VIIA element) by a more active halogen
   Example: chlorine gas plus sodium iodide ------->
   $Cl_2(g) + NaI(aq)$ -----> $NaCl(aq) + I_2(aq)$ (unbalanced)
   $Cl_2(g) + 2NaI(aq)$ -----> $2NaCl(aq) + I_2(aq)$ (balanced)

ACTIVITY SERIES

<u>metals</u>		<u>nonmetals</u>	
Li	**most active**	F_2	**most active**
K		Cl_2	
Ba		Br_2	
Ca		I_2	**least active**
Na			
Mg			
Al			
Zn			
Fe			
Ni			
Sn			
Pb			
(H)*			
Cu			
Hg			
Ag			
Au	**least active**		

*Metals from Li to Na will replace H in water and in acids from Mg to Pb, they will only replace the H in acids

The uncombined element must be more active than the combined element in order to replace it. If the uncombined element is **higher** on the series, that indicates that it is more active, and is capable of replacing it. Thus Na would replace Zn or Ag from a compound. By contrast, Na will not replace Li or Ca from a compound. F_2 would replace Cl_2, but Cl_2 will not replace F_2.

Group IV

Double Replacement or Displacement Reactions

Involves only an exchange of positive ions between compounds, without any oxidation or reduction taking place. A double replacement reaction can occur only if the reactants are two compounds that contain four different ions and the ions are free to move, such as in a solution. Generally, the reaction takes place between 2 ionic compounds in aqueous solution and it follows the pattern: $A^+X^- + B^+Y^- \text{-------}> A^+Y^- + B^+X^-$ Two new compounds are formed, at least one of which must be separable from the remaining ions in the solution. This can happen in one of 3 ways:

1) the formation of an insoluble solid(precipitate), which can be separated by filtration
   Example: potassium iodide plus lead (II) nitrate ------->
   $KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + KNO_3(aq)$ (unbalanced)
   $2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$ (balanced)

2) if water is one of the products, then it can be separated by distillation or evaporation
   Example: sodium hydroxide plus hydrochloric acid ------->
   $NaOH(aq) + HCl(aq) \text{----}> NaCl(aq) + H_2O(l)$ (balanced)

3) if one of the products is an insoluble gas, it will bubble out of the solution
   Example: hydrochloric acid plus iron (II) sulfide ------->
   $HCl(aq) + FeS(s) \rightarrow H_2S(g) + FeCl_2(aq)$ (unbalanced)
   $2HCl(aq) + FeS(s) \rightarrow H_2S(g) + FeCl_2(aq)$ (balanced)

When working with water solutions, it is helpful to have a few rules concerning which substances are soluble, and which will form precipitates. The more common **solubility rules** are listed below:

1. All common salts of the Group IA(Li, Na, K, etc) elements and the ammonium ion are soluble.
2. All common acetates and nitrates are soluble.
3. All binary compounds of Group VIIA elements(other than F) with metals are soluble, except those of silver, mercury (I), and lead.
4. All sulfates are soluble except those of barium, strontium, lead, calcium, silver, and mercury (I).
5. Except for those in Rule #1, carbonates, hydroxides, oxides, and phosphates are insoluble.

Group V

Combustion Reactions

During combustion, oxygen gas reacts with organic hydrocarbons(compounds made of primarily hydrogen and carbon atoms), often producing energy in the form of heat and light. It follows the pattern: $C_xH_y + O_2 \text{---}> xCO_2 + y/2\ H_2O$ Carbon dioxide and water are formed unless there is incomplete combustion, in which case C and/or CO may

be additional products. Example: methane (CH_4) plus oxygen -------->

$CH_4 + O_2$ ----> $CO_2 + H_2O$ (unbalanced) $CH_4 + 2O_2$ ---> $\mathbf{1}CO_2 + \mathbf{2}H_2O$ (balanced)

Example: benzene (C_6H_6) plus oxygen ------->

$C_6H_6 + O_2$ -------> $CO_2 + H_2O$ (unbalanced)

$C_6H_6 + O_2$ -------> $\mathbf{6}CO_2 + \mathbf{3}H_2O$ (unbalanced)

$C_6H_6 + \mathbf{7.5}O_2$ ------> $6CO_2 + 3H_2O$ (balanced, 2 x equation for whole numbers)

$\mathbf{2}C_6H_6 + \mathbf{15}O_2$ -------> $\mathbf{12}CO_2 + \mathbf{6}H_2O$ (balanced)

Example: methyl alcohol(CH_3OH) plus oxygen ------->

$CH_3OH + O_2$ --> $CO_2 + H_2O$ unbalanced, becomes $CH_3OH + \mathbf{1.5}O_2$ --> $\mathbf{1}CO_2 + \mathbf{2}H_2O$

$\mathbf{2}CH_3OH + \mathbf{3}O_2$ ----> $\mathbf{2}CO_2 + \mathbf{4}H_2O$ (double coefficients for whole numbers, balanced)

Problems:

Given the following reactants, determine the type of reaction that would occur, if any. If no reaction is predicted to occur, state your reasons for this prediction. For all probable reactions, list the names of the products and write a balanced formula equation for each:

1. copper(II) oxide + hydrogen gas ---->

Ans: SR: $CuO + H_2$ -----> $Cu + H_2O$
copper + water

2. zinc + nickel (III) chloride ----->

Ans: SR: $3Zn + 2NiCl_3$ -----> $3ZnCl_2 + 2Ni$
zinc chloride + nickel

3. sodium carbonate ------->

Ans: Decomp: Na_2CO_3 -----> $Na_2O + CO_2$
sodium oxide + carbon (IV) oxide

4. sulfur + copper ------->

Ans: Comp: $S + Cu$ -----> CuS
copper (II) sulfide

5. silver nitrate + sulfuric acid ------->

Ans: DR: $2AgNO_3 + H_2SO_4$ ----> $Ag_2SO_4(s) + 2HNO_3$
silver sulfate + nitric acid

6. water + sulfur (VI) oxide ------->

=**Ans: Comp:** $H_2O + SO_3$ -----> H_2SO_4
sulfuric acid

7. calcium carbonate + hydrochloric acid ------->

=**Ans: DR & D:** $CaCO_3 + 2HCl$ ---> $CaCl_2 + H_2O + CO_2(g)$
=calcium chloride + water + carbon (IV) oxide

8. ammonium nitrate + lithium chloride ------->

Ans: no rxn: both new compounds are soluble
ammonium chloride + lithium nitrate

9. octane(C_8H_{18}) + oxygen gas ------->

Ans: Combust: $2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$
carbon (IV) oxide + water

10. aluminum + nitric acid ------->

Ans: SR: $2Al + 6HNO_3 \longrightarrow 2Al(NO_3)_3 + 3H_2$
aluminum nitrate + hydrogen

11. potassium + water ------->

Ans: SR: $2K + 2H_2O \longrightarrow 2KOH + H_2$
potassium hydroxide + hydrogen

12. iodine + sodium bromide ------->

Ans: no rxn: iodine is less active than Br

13. lithium hydroxide + phosphoric acid ------->

Ans: DR: $3LiOH + H_3PO_4 \longrightarrow Li_3PO_4(aq) + 3H_2O$
lithium phosphate + water

14. oxygen gas + sucrose($C_{12}H_{22}O_{11}$) ------->

Ans: Combust: $12O_2 + C_{12}H_{22}O_{11} \longrightarrow 12CO_2 + 11H_2O$
carbon (IV) oxide + water

VII. Stoichiometry: Mole/Mass/Volume Relationships

Stoichiometry is the branch of chemistry that deals with the numerical relationships and mathematical proportions of reactants and products in chemical reactions. For any chemical equation the mole, mass, atom, and volume ratios must remain constant. The coefficients from a balanced chemical equation provide definite ratios between the number of moles of reactants and products involved. One mole of any chemical substance contains 6.02×10^{23} particles (molecules, atoms, etc.) and has a mass equal to the sum of atomic weights of all elements in the formula. For gases, 1 mole also has a volume of 22.4 liters at STP. Therefore, the mole ratio can be converted into a mass or number of particles ratio or a volume ratio. These ratios can be used as conversion factors to determine the exact amounts of reactants that combine and products that form during a chemical change. Dimensional analysis is used to convert the units of the given quantity into proportional units of the unknown quantity.

Sample Problem 1:
How many grams of water will be formed by the complete burning of 50.0 grams of hydrogen in air? How many moles of oxygen gas are needed for complete combustion? At STP, what volume of hydrogen gas is used?

The equation provides the conversion factors used to solve the problems.

Write a balanced equation:	$2H_2(g)$	+	$O_2(g)$	----->	$2H_2O(g)$
Determine the mole ratios:	2 moles		1 mole		2 moles
Multiply #moles x molar mass:	x 2 g/mole		x 32g/mole		x 18g/mole
Equals the equation mass:	4 g H_2	+	32 g O_2	----->	36 g H_2O
Multiply #moles x 22.4 L	44.8 L H_2	+	22.4 L O_2	---->	44.8 L H_2O = gas volume

$$x \text{ grams } H_2O = 50.0 \text{ grams } H_2 \frac{(36 \text{ grams } H_2O)}{(4 \text{ grams } H_2)} = 450 \text{ grams } H_2O$$

$$x \text{ moles } O_2 = 50.0 \text{ grams } H_2 \frac{(1 \text{ mole } O_2)}{(4 \text{ grams } H_2)} = 12.5 \text{ moles } O_2$$

$$x \text{ liters } H_2 = 50.0 \text{ grams } H_2 \frac{(44.8 \text{ L } H_2)}{(4 \text{ grams } H_2)} = 560 \text{ L } H_2$$

The number of grams of O_2 that react can also be easily determined by:

$$x \text{ grams } O_2 = 50.0 \text{ grams } H_2 \frac{(32 \text{ grams } O_2)}{(4 \text{ grams } H_2)} = 400. \text{ grams } O_2$$

Note the similarities between all four calculations.....only the quantities used in the numerator of the conversion factors differ.

Sample Problem 2:
In the decomposition of potassium chlorate, 82.6 grams of oxygen are formed. How many grams of potassium chloride are produced?

Even though the equation provides many mole, mass, particle, and volume ratios, focus on the ratio(s) needed to convert the given quantity into the unknown quantity.

		Given: 82.6 g O_2		Unknown: x g KCl
$2KClO_3$	------>	$3O_2$	+	$2KCl$
		3 moles		2 moles
		x 32 g/mole		x 74 g/mole
		96 g		148 g

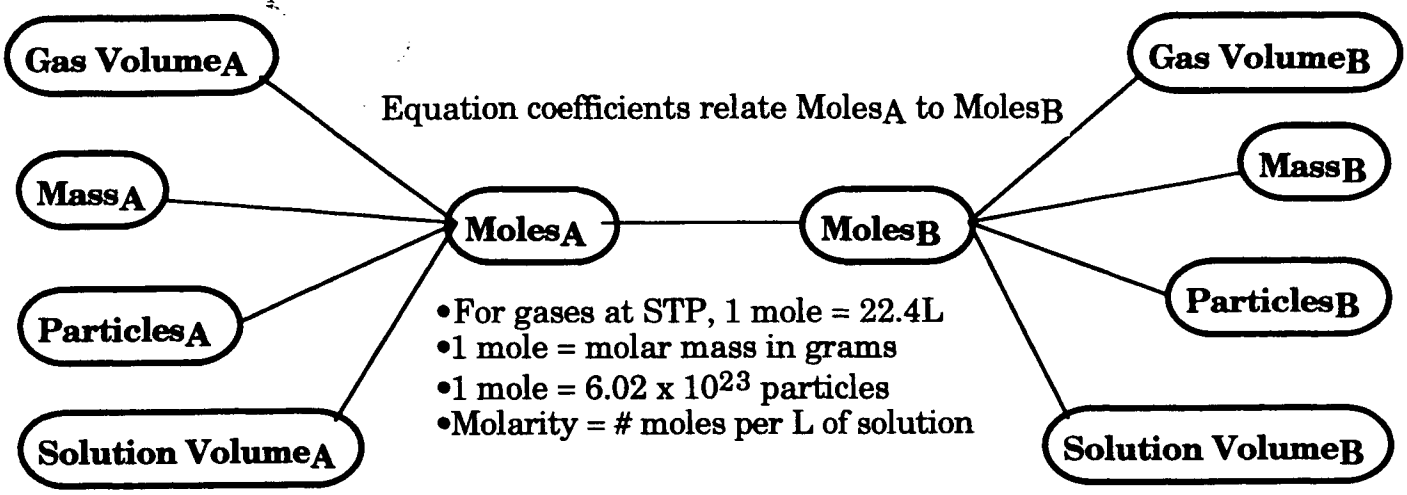
$$x \text{ grams KCl} = 82.6 \text{ grams } O_2 \frac{(148 \text{ g KCl})}{(96 \text{ g } O_2)} = 127 \text{ grams KCl}$$

Multiple conversion factors can be used to solve the same problem:

$$x \text{ grams KCl} = 82.6 \text{ grams } O_2 \frac{(1 \text{ mole } O_2)}{(32 \text{ g } O_2)} \frac{(2 \text{ mole KCl})}{(3 \text{ mole } O_2)} \frac{(74 \text{ g KCl})}{(1 \text{ mole KCl})} = 127\text{g KCl}$$

The guide on the next page can be used to help you determine the steps needed to make conversions between the units of reactants and products. Remember that you need to cancel out the given units and convert them into an equivalent amount of matter expressed by the new units.

Mole Map for Solving Stoichiometry Problems



Problems:

Solve each of the following problems, using the appropriate conversion factors:

1. How many moles of hydrogen gas molecules can be produced from the reaction of 72.0g of Na with an excess of water?

Ans: 1.57 moles H_2

2. An excess of nitrogen reacts with 6.57 g of hydrogen. How many liters of ammonia(NH_3) gas are produced at STP?

Ans: 49 L NH_3

3. How many grams of oxygen are required to completely burn 84.9 g of carbon?

Ans: 226 g O_2

4. How many grams of hydrochloric acid are required to react completely with 44.7 grams of calcium hydroxide?

Ans: 44.0 g HCl

5. How many grams of hydrogen are produced when 4.77 grams of aluminum react with excess sulfuric acid?

Ans: 0.53 g H_2

VIII. Exothermic vs Endothermic

Energy changes occur whenever a chemical reaction takes place. An equation that includes the amount of heat produced or absorbed by a reaction is called a **thermochemical equation**. At the start of any reaction, each of the reactants has a

certain heat content or **enthalpy**, symbolized by **H**. The change in the enthalpy, **ΔH**, of the reactants equals the amount of heat produced or absorbed. A negative ΔH value indicates that excess heat is released as a product during an exothermic reaction. A positive ΔH value indicates that heat is absorbed as a reactant during an endothermic reaction. Typical experimental conditions for measuring the heat energy change is 25°C and 1 atm pressure. The heat term can be treated like any other reactant or product, and provides still another proportion for problem solving.

Example: Burning coal is an exothermic reaction that can be written as:

$$C(s) + O_2(g) \text{ -------> } CO_2(g) + 394 \text{ kJ}$$

or $$C(s) + O_2(g) \text{ -------> } CO_2(g) \quad \Delta H = -394 \text{ kJ}$$

The endothermic decomposition of sodium carbonate by heating can be written as:

$$Na_2CO_3(s) + 1131 \text{ kJ} \text{ -------> } Na_2O(s) + CO_2(g)$$

or $$Na_2CO_3(s) \text{ -------> } Na_2O(s) + CO_2(g) \quad \Delta H = +1131 \text{ kJ}$$

Sample Problem:
If 5.00 kg of coal is needed to heat a home for 1 day, how many kJ of heat is produced?

$$x \text{ kJ} = 5.00 \text{ kg C} \frac{(1000 \text{ g C})(1 \text{ mole C})(394 \text{ kJ})}{(1 \text{ kg C})(12.0 \text{ g C})(1 \text{ mole C})} = 164{,}000 \text{ kJ heat released}$$

Problem:
How many grams of sodium carbonate can be decomposed using the heat produced by burning 5.00 kg of coal?

Ans: 15400 g Na_2CO_3

IX. Solving Complex Stoichiometry Problems

Even the most complicated problems involving chemical reactions can be solved by using dimensional analysis and the equation ratios between reactants and products. **Molarity** of reactants or products that are in solution can be used as an additional conversion factor to determine the number of moles present or the volume of solution needed. If the reactants are not mixed in the correct proportions, then one of the reactants will have an **excess** amount. This means that when the reaction has completely stopped, some of the reactant will still remain unchanged. The reactant that is all used up is called the **limiting factor**, since it controls the amount of products that can be formed. When this material is gone, the reaction stops. Once again, the equation ratios can be used to determine if an excess exists.

Sample Problem:
If 20.0 grams of zinc metal is placed in a beaker containing 100. mL of 1.50 M hydrochloric acid and allowed to react overnight, what will remain in the beaker the next day?

A metal with an acid will undergo a single replacement reaction to form hydrogen gas and the metallic salt. zinc + hydrochloric acid ----> hydrogen gas + zinc chloride

$$Zn + 2HCl \text{ ----> } H_2 + ZnCl_2$$

Given: 20.0 g Zn and 100. mL of 1.50M HCl.....

Determine if either reactant is in excess by calculating how much of one reactant is actually needed to combine with the given quantity of the other reactant.

$$\text{x g Zn needed} = \frac{100.\text{ mL HCl}}{(1000\text{ mL 1.50M HCl})}\left(\frac{1.50\text{ moles HCl}}{2\text{ moles HCl}}\right)\left(\frac{1\text{ mole Zn}}{1\text{ mole Zn}}\right)\left(\frac{65.4\text{ g Zn}}{}\right) = 4.90\text{ g Zn needed}$$

OR

$$\text{x mL HCl needed} = \frac{20.0\text{ g Zn}}{(65.4\text{ g Zn})}\left(\frac{1\text{ mole Zn}}{1\text{ mole Zn}}\right)\left(\frac{2\text{ moles HCl}}{1.50\text{ moles HCl}}\right)\left(\frac{1000\text{mL 1.50M HCl}}{}\right) = 408\text{mL HCl needed}$$

Compare the needed amount to that which is actually available:

20.0 g Zn available - 4.90 g Zn needed to react = 15.1 g Zn in excess(unreacted)
15.1 g Zn will remain unchanged in the beaker.

The HCl is the limiting factor and determines the amount of products that can be formed. A volume of 408 mL of 1.50M HCl would be required to completely react with 20.0 g Zn.

$$\text{x g } H_2\text{ lost} = \frac{100.\text{ mL HCl}}{(1000\text{ mL 1.50M HCl})}\left(\frac{1.50\text{ moles HCl}}{2\text{ moles HCl}}\right)\left(\frac{1\text{ mole } H_2}{1\text{ mole } H_2}\right)\left(\frac{2\text{ g } H_2}{}\right) = 0.15\text{ g } H_2\text{ lost}$$

$$\text{x mole } ZnCl_2\text{ dissolved} = \frac{100.\text{ mL HCl}}{(1000\text{ mL 1.50M HCl})}\left(\frac{1.50\text{ moles HCl}}{2\text{ moles HCl}}\right)\left(\frac{1\text{ mole } ZnCl_2}{}\right) = 0.075\text{ mole } ZnCl_2\text{ dissolved}$$

Assuming 100mL of solution remains, the molarity of the dissolved $ZnCl_2$ is 0.75M.

Problems:

Solve each of the following problems, using the appropriate conversion factors:

1. A 21.10 g sample of limestone contains 87.3% of calcium carbonate by mass. When this sample is reacted with an excess amount of hydrochloric acid, calcium chloride solution, water, and carbon (IV) oxide gas are produced. What volume of $CO_2(g)$ will be formed at STP?

Ans: 4.12 L CO_2

2. When an iron nail is placed in a beaker with 75.0 mL of sulfuric acid overnight, some of the nail remains the next day. If the nail lost a mass of 3.81 g, what was the original molarity of the sulfuric acid used?

Ans: 1.37 M H_2SO_4

3. A pure copper penny with a mass of 3.21 g is placed in a beaker containing 215 mL of 0.150 M silver nitrate solution and allowed to react overnight. How many grams of silver metal can be filtered out of the mixture on the next day?

Ans: 3.48 g Ag

4. If 5.00 L of hydrogen iodide gas is bubbled through 2.00 L of 0.0275 M lead (II) nitrate solution, how many grams of lead (II) iodide will precipitate out?

Ans: 25.4 g $PbI_2(s)$

Unit 5 Objectives

Having studied the unit notes and done the problems, you should be able to:

1. List the various parts of a chemical equation including symbols representing the physical state of substances in the reaction.

2. Balance equations by inspection, given the names or formulas for reactants and products.

3. Assign oxidation numbers to elements in a reaction to determine which element is oxidized and which is reduced.

4. Balance a complex redox reaction using algebra, the oxidation number change method, and/or the half-reaction method.

5. Classify a reaction as one of five basic types; synthesis, decomposition, single displacement, double displacement or combustion.

6. Use generalizations to predict the products of simple reactions, given only the reactants.

7. Determine the mole, mass, atom or gas volume ratios between reactants and products from a balanced formula equation

8. Given the mass, moles, gas or solution volume of a reactant or product, calculate the mass, moles, gas or solution volume of another reactant or product using the ratios from a chemical equation.

9. Calculate the heat lost or gained during a chemical reaction, given the thermochemical equation and amount of reactant used or product formed.

10. Determine which reactant is in excess and then use the limiting factor to determine the amount of product(s) that will be formed.