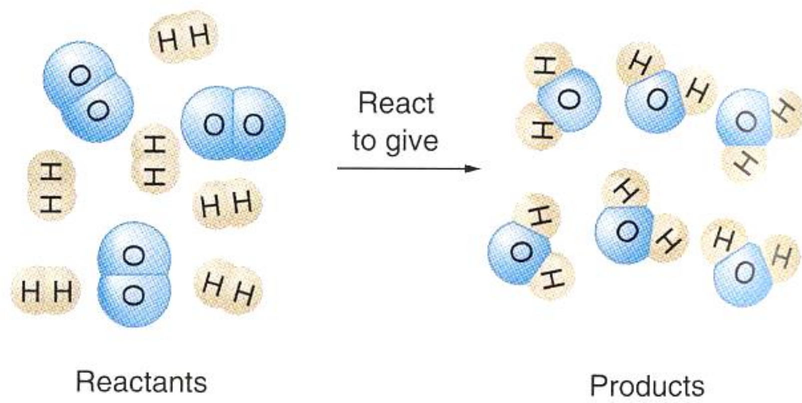
CHEMICAL REACTIONS

- A ***chemical reaction*** is a ***rearrangement*** of atoms in which some of the ***original bonds are broken*** and ***new bonds are formed*** to give ***different chemical structures***.

- In a ***chemical reaction***, atoms are ***neither created***, ***nor destroyed***.

- A ***chemical reaction***, as described above, is supported by ***Dalton's postulates***.



- A ***chemical reaction*** can be detected by one of the following ***evidences***:

    1. Change of color
    2. Formation of a solid
    3. Formation of gas
    4. Exchange of heat with surroundings

THE CHEMICAL EQUATION

- A ***chemical equation*** is a shorthand expression for a ***chemical reaction***.

Word equation: ***Aluminum*** combines with ***ferric oxide*** to form ***iron*** and ***aluminum oxide***.

Chemical Equation: $\mathbf{Al + Fe_2O_3 \rightarrow Fe + Al_2O_3}$

- A chemical equation consists of the following information:

1. ***Reactants*** separated from ***products*** by an arrow (→):

$$\underbrace{\mathbf{Al + Fe_2O_3}}_{\textbf{reactants}} \longrightarrow \underbrace{\mathbf{Fe + Al_2O_3}}_{\textbf{products}}$$

2. ***Coefficients*** are placed in front of substances to ***balance*** the equation:

$$\mathbf{2\,Al + Fe_2O_3 \longrightarrow 2\,Fe + Al_2O_3}$$

3. Reaction ***conditions*** are placed over the arrow:

$$\mathbf{2\,Al + Fe_2O_3 \xrightarrow{\Delta} 2\,Fe + Al_2O_3}$$

4. The **physical state** of the substances are indicated by symbols (s), (l), (g) and (aq):

$$\mathbf{2\,Al\,(s) + Fe_2O_3\,(s) \xrightarrow{\Delta} 2\,Fe\,(l) + Al_2O_3\,(s)}$$

WRITING & BALANCING EQUATIONS

- A ***balanced equation*** contains the ***same number of atoms*** on each side of the equation, and therefore obeys the ***law of conservation of mass***.
- Many equations are balanced by ***trial and error***; but it must be remembered that ***coefficients can be changed*** in order to balance an equation, but ***not subscripts*** of a correct formula.

The general procedure for balancing equations is:

1. ***Write the unbalanced equation***
   - Make sure the formula for each substance is correct

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

2. ***Balance by inspection***
   - ***Count*** and ***compare*** each element on both sides of the equation.

Reactant		**Product**
1 C	=	**1C**
4 H	≠	**2H**
2O	≠	**3O**

   - ***Balance*** elements that appear ***only in one substance*** first.

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

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

   - When finally done, check for the ***smallest coefficients*** possible.

Examples:

1) $AgNO_3 + H_2S \rightarrow Ag_2S + HNO_3$

	Ag	H	S	NO_3
Reactant				
Products				

___ $AgNO_3$ + ___H_2S $\rightarrow$ ___Ag_2S + ____ HNO_3

2) $Al(OH)_3 + H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2O$

	Al	H	O	SO_4
Reactant				
Products				

___ $Al(OH)_3$ + ___ H_2SO_4 $\rightarrow$ ___$Al_2(SO_4)_3$ + ____ H_2O

3) $Fe_3O_4 + H_2 \rightarrow Fe + H_2O$

	Fe	O	H
Reactant			
Products			

___Fe_3O_4 + ___ H_2 $\rightarrow$ ___ Fe + ___ H_2O

4) $C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$

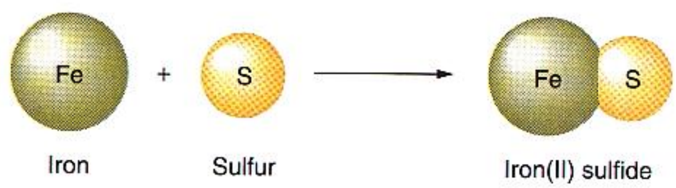
	C	H	O
Reactant			
Products			

___ C_4H_{10} + ___ O_2 $\rightarrow$ ___ CO_2 + ___ H_2O

TYPES OF CHEMICAL REACTIONS

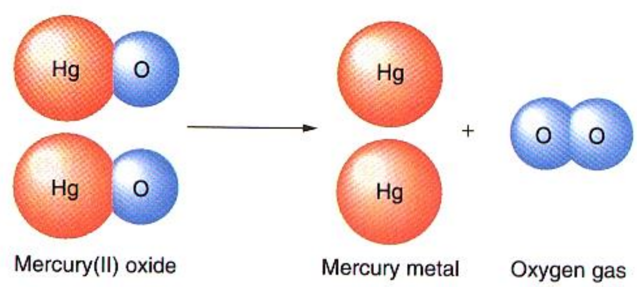
- Chemical reactions are classified into ***five types***:

1. Synthesis or Combination



- Two ***elements or compound*** combine to form another compound.

2. Decomposition



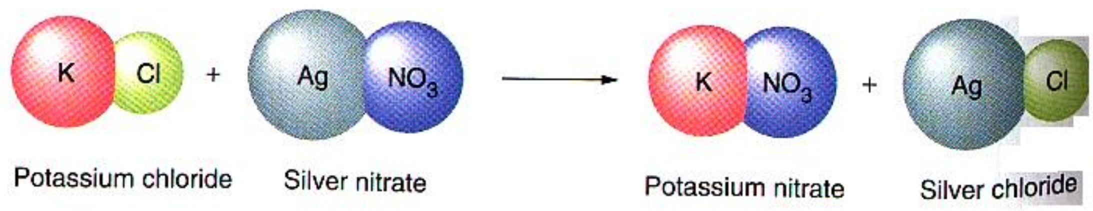
- A compound breaks up to form ***elements or simpler compound***.

3. Single Replacement

- A ***more reactive element*** replaces a ***less reactive element*** in a compound.
- These reactions will be discussed in more detail later in this chapter.

TYPES OF CHEMICAL REACTIONS

4. Double Replacement



- ***Two compounds*** interact to form two ***new compounds***.
- These reactions will be discussed in more detail later in this chapter.

5. Combustion Reactions:

- A reaction that involves ***oxygen*** as a reactant and ***produces large amounts of heat*** is classified as a combustion reaction.

$$CH_4\,(g) + 2\,O_2\,(g) \rightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

Examples:

Classify each of the reactions below:

1. $Mg + CuCl_2 \rightarrow MgCl_2 + Cu$ ______________________
2. $CaCO_3 \rightarrow CaO + CO_2$ ______________________
3. $2\,HCl + Ca(OH)_2 \rightarrow CaCl_2 + H_2O$ ______________________
4. $C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O$ ______________________
5. $4\,Fe + 3\,O_2 \rightarrow 2\,Fe_2O_3$ ______________________

DOUBLE REPLACEMENT REACTIONS

- Double replacement reactions can be subdivided into one of the following subgroups:

1. ***Precipitation:***
   In these reactions one of the products formed is an insoluble solid called a ***precipitate***. For example, when solutions of potassium chromate, K_2CrO_4 , and barium nitrate, $Ba(NO_3)_2$, are combined an insoluble salt barium chromate, $BaCrO_4$, is formed.

   $$K_2CrO_4\,(aq) + Ba(NO_3)_2\,(aq) \rightarrow \underbrace{BaCrO_4\,(s)}_{\text{Precipitate}} + 2KNO_3\,(aq)$$

   These reactions will be further discussed in Chapter 8

2. ***Neutralization:***
   The most important reaction of acids and bases is called **neutralization**. In these reactions an acid combines with a base to form a **salt and water**. For example:

   $$\underset{\text{acid}}{HCl\,(aq)} + \underset{\text{base}}{NaOH\,(aq)} \longrightarrow \underset{\text{salt}}{NaCl\,(aq)} + \underset{\text{water}}{H_2O\,(l)}$$

   Salts are ***ionic*** substances with the ***cation*** donated from the ***base*** and the ***anion*** donated from the ***acid***. In the laboratory, neutralization reactions are observed by an increase in temperature (exothermic reaction).

3. ***Unstable product:***
   Some chemical reactions ***produce gas*** because one of the products formed in the reaction is ***unstable***. Two such products are listed below:

Carbonic acid	$H_2CO_3\,(aq) \rightarrow CO_2\,(g) + H_2O\,(l)$
Sulfurous acid	$H_2SO_3\,(aq) \rightarrow SO_2\,(g) + H_2O\,(l)$

   When either of these products appears in a chemical reaction, they should be replaced with their decomposition products.

   $$2\,HCl\,(aq) + Na_2CO_3\,(aq) \rightarrow 2\,NaCl\,(aq) + \underset{\textbf{(unstable)}}{H_2CO_3\,(aq)}$$

   $$2\,HCl\,(aq) + Na_2CO_3\,(aq) \rightarrow 2\,NaCl\,(aq) + CO_2\,(g) + H_2O\,(l)$$

DOUBLE REPLACEMENT REACTIONS

Examples:
Complete and balance each neutralization reaction below:

1. HNO_3 (aq) + $Ba(OH)_2$ (aq) $\rightarrow$

2. H_2SO_4 (aq) + $NaOH$ (aq) $\rightarrow$

3. $HC_2H_3O_2$ (aq) + KOH (aq) $\rightarrow$

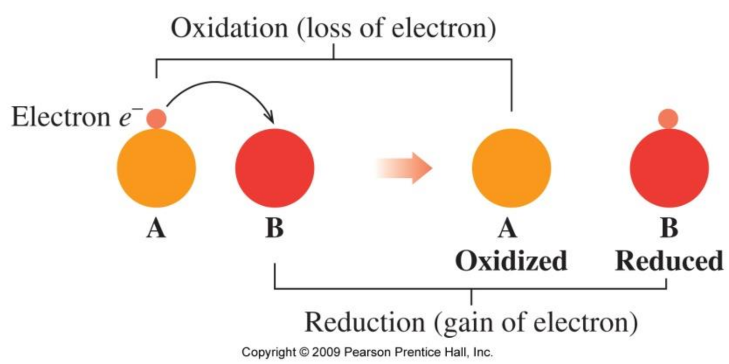
Complete and balance the unstable product reaction shown below:

4. HNO_3 (aq) + K_2SO_3 (aq) $\rightarrow$

OXIDATION-REDUCTION REACTIONS

- Reactions known as **oxidation** and **reduction** (redox) have many important applications in our everyday lives. Rusting of a nail or the reaction within your car batteries are two examples of redox reactions.

- In an oxidation-reduction reaction, electrons are transferred from one substance to another. If one substance loses electrons, another substance must gain electrons.



- **Oxidation** is defined as **loss** of electrons, and **reduction** is defined as **gain** of electrons. One way to remember these definitions is to use the following mnemonic:

OIL RIG

Oxidation **I**s **L**oss of electrons

Reduction **I**s **G**ain of electrons

- Combination, decomposition, single replacement and combustion reactions are all examples of redox reactions.

OXIDATION-REDUCTION REACTIONS

- In general, atoms of **metals lose electrons** to form cations, and are therefore **oxidized**, while atoms of **non-metals gain electrons** to form anions, and are therefore **reduced**.
- For example, in the formation of calcium sulfide from calcium and sulfur

$$Ca + S \rightarrow CaS$$

$Ca \rightarrow Ca^{2+} + 2\ e^-$ oxidized (loses electrons)

$S + 2\ e^- \rightarrow S^{2-}$ reduced (gains electrons)

- Therefore, the formation of calcium sulfide involves **two half-reactions** that occur simultaneously, one an **oxidation** and the other a **reduction**.
- Similarly, in the reaction of magnesium metal with hydrochloric acid

$$Mg + 2\ HCl \rightarrow MgCl_2 + H_2$$

$Mg \rightarrow Mg^{2+} + 2\ e^-$ oxidized (loses electrons)

$2\ H^+ + 2\ e^- \rightarrow H_2$ reduced (gains electrons)

- In every oxidation-reduction reaction, the number of electrons lost must be equal to the number of electrons gained.

Examples:

1. Identify each of the reactions below as oxidation or reduction:
   a) $Zn \rightarrow Zn^{2+} + 2\ e^-$ ________________________
   b) $Cl_2 + 2\ e^- \rightarrow 2\ Cl^-$ ________________________
   c) $Fe^{3+} + e^- \rightarrow Fe^{2+}$ ________________________
   d) $2\ Br^- \rightarrow B_2 + 2\ e^-$ ________________________

OXIDATION-REDUCTION IN BIOLOGICAL SYSTEMS

- Many important biological reactions involve oxidation and reduction. In these reactions, **oxidation** involves **addition of oxygen** or **loss of hydrogen**, and **reduction** involves **loss of oxygen** or **gain of hydrogen**.

- For example, poisonous methyl alcohol is metabolized by the body by the following reaction:

CH_3OH	$\rightarrow$	H_2CO	+	$2\ H\bullet$	**oxidation** (loss of hydrogen)
methyl alcohol		**formaldehyde**			

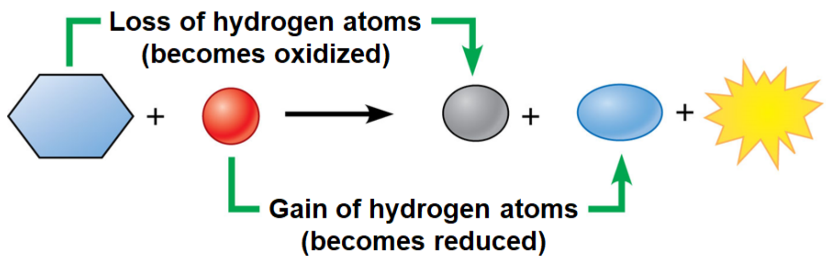
- The formaldehyde is further oxidized to formic acid and finally carbon dioxide and water by the following reactions:

$2\ H_2CO$	+	O_2	$\rightarrow$	$2\ H_2CO_2$	**oxidation** (gain of oxygen)
formaldehyde				**formic acid**	

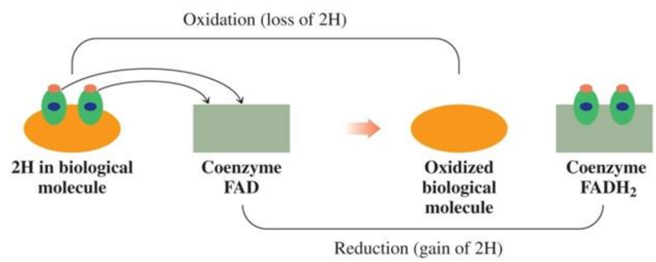
$2\ H_2CO_2 + O_2 \rightarrow CO_2 + H_2O$	**oxidation** (gain of oxygen)
formic acid	

- In many biochemical oxidation-reduction reactions, the transfer of hydrogen atoms produces energy in the cells.

- For example, cellular respiration is an oxidation-reduction process that transfers energy from the bonds in glucose to form ATP.



- The oxidation of a typical biochemical molecule can involve the transfer of hydrogen atoms to a proton acceptor such as coenzyme FAD to produce its reduced form $FADH_2$.



OXIDATION-REDUCTION IN BIOLOGICAL SYSTEMS

- In summary, the particular definition of oxidation-reduction depends on the process that occurs in the reaction. A summary of these definitions appears below:

Oxidation	
Always Involves	**May Involve**
Loss of electrons	Addition of oxygen
Electrons are a product	Loss of hydrogen

Reduction	
Always Involves	**May Involve**
Gain of electrons	Loss of oxygen
Electrons are a reactant	Gain of hydrogen

Examples:

1. Linoleic acid, an unsaturated fatty acid, can be converted to a saturated fatty acid by the reaction shown below. Is linoleic acid oxidized or reduced in this reaction?

$$C_{18}H_{32}O_2 \quad + \quad 2\ H_2 \quad \rightarrow \quad C_{18}H_{36}O_2$$
linoleic acid

2. The reaction of succinic acid provides energy for the ATP synthesis and is shown below:

$$C_4H_6O_4 \rightarrow C_4H_4O_4 + 2\ H\bullet$$
succinic acid

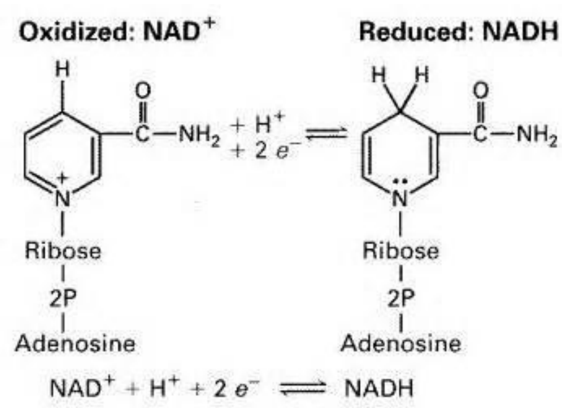
$$FAD + 2H\bullet \rightarrow FADH_2$$

a) Is succinic acid oxidized or reduced?

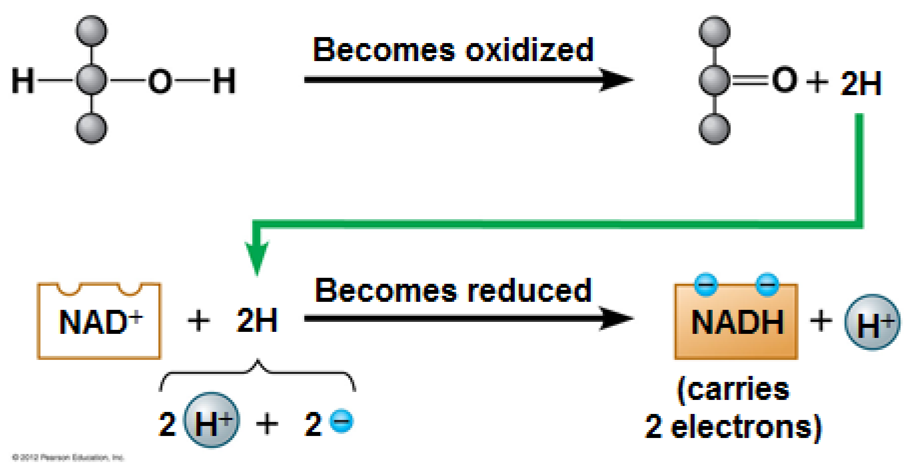
b) Is FAD oxidized or reduced?

ENZYMES IN BIOLOGICAL SYSTEMS

- In biochemical reactions, enzymes are necessary to oxidize glucose and other foods.
- For example, oxidation of glucose involves the transfer of hydrogen atoms and electrons to an enzyme, such as NAD^+ to produce its reduced form NADH.



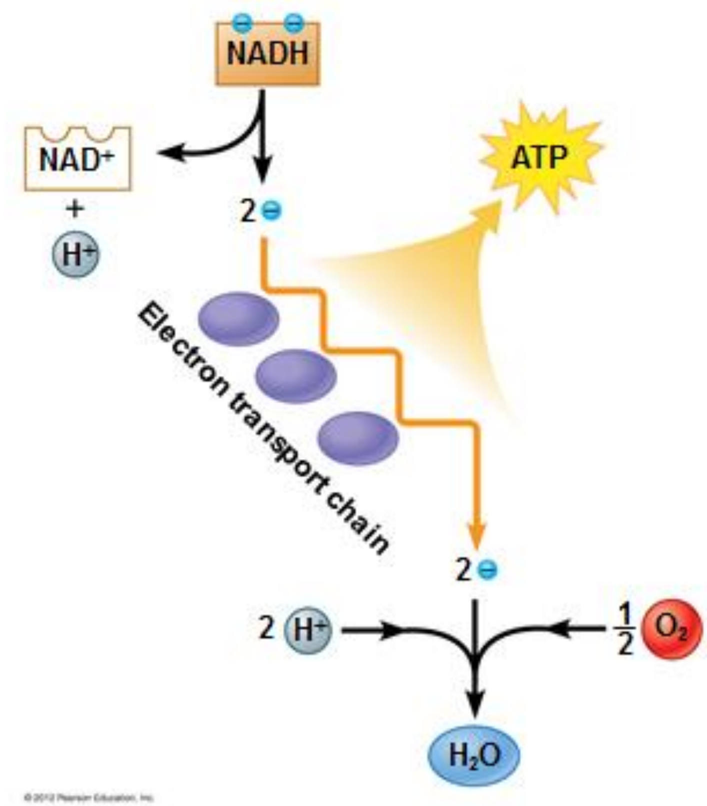
- Similarly, oxidation of methanol involves transfer of 2 hydrogen atoms and 2 electrons to NAD^+ to form the reduced form NADH.



- Molecules such as NAD^+ are called "***electron carriers***" since they carry electrons in their reduced form.

ENZYMES IN BIOLOGICAL SYSTEMS

- The electron carriers collectively are called ***electron transport chain***.
- As electrons are transported down the chain, ATP is generated.



THE MOLE CONCEPT

- Chemists find it more convenient to use ***mass relationships*** in the laboratory, while ***chemical reactions*** depend on the ***number of atoms*** present.

- In order to relate the ***mass and number of atoms***, chemists use the SI unit ***mole*** (abbreviated ***mol***).

- The number of ***particles*** in a ***mole*** is called ***Avogadro's number*** and is ***6.02x10 ^{23}***.

1 mol of H atoms.........contains: **6.02 x 10^{23}** H atoms

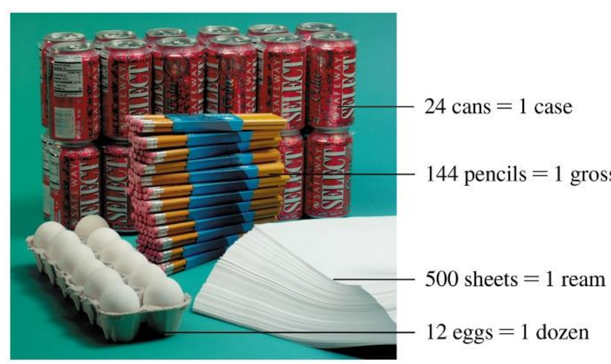
1 mol of H_2 molecules.........contains: **6.02 x 10^{23}** H_2 molecules
2 x(6.02x10^{23}) H atoms

1 mol of **H_2O** molecules....contains: 6.02x10^{23} H_2O molecules
2 x (6.02x10^{23}) H atoms
1 x (6.02x10^{23}) O atoms

1 mol of Na^+ ions...........contains: **6.02 x 10^{23}** Na^+ ions

- The **atomic mass** of one atom expressed in **amu** is numerically the same as the **mass of one mole** of atoms of the element expressed in **grams**.

Element	Mass of one atom	Mass of one mole of atoms
H	1.008 amu	1.008 grams
Mg	24.31 amu	24.31 grams
Cl	35.45 amu	35.45 grams



- Mole is a counting unit for atoms similar to other counting units used for other objects

MOLAR MASS

- The mass of one mole of a substance is called ***molar mass*** and is measured in ***grams***.

Mass of one mole of H_2O

$$\begin{array}{lll} 2 \text{ mol H} = 2\,(1.008\text{ g}) = & 2.016\text{ g} & \\ 1 \text{ mol O} = 1\,(16.00\text{ g}) = & \underline{16.00\text{ g}} & \\ & 18.02\text{ g} & \leftarrow \textbf{Molar Mass} \end{array}$$

Mass of one mole of $Ca(OH)_2$

$$\begin{array}{lll} 1 \text{ mol Ca} = 1\,(40.08\text{ g}) = & 40.08\text{ g} & \\ 2 \text{ mol O} = 1\,(16.00\text{ g}) = & 32.00\text{ g} & \\ 2 \text{ mol H} = 2\,(1.01\text{ g}) = & \underline{2.02\text{ g}} & \\ & 74.10\text{ g} & \leftarrow \textbf{Molar Mass} \end{array}$$

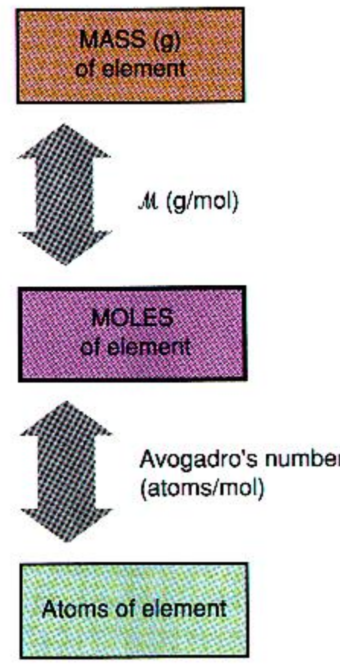
<u>Examples:</u>
Calculate the molar mass of each compound shown below:

1. Lithium carbonate (Li_2CO_3)

2. Salicylic acid ($C_7H_6O_3$)

CALCULATIONS USING THE MOLE CONCEPT

When solving problems involving ***mass-mole-number*** relationships of elements or compounds, we can use:



- The ***molar mass*** to convert between mass and moles.
- ***Avogadro's number*** (6.02×10^{23}) to convert between moles and number of entities.

<u>Examples:</u>

1. How many **moles** of iron are present in 25.0 g of iron?

$$25.0 \text{ g Fe} \left(\frac{1 \text{ mole}}{55.85 \text{ g}}\right) = 0.448 \text{ mol Fe}$$

2. What is the mass of 5.00 mole of water?

3. How many magnesium **atoms** are present in 5.00 g of Mg?

$$5.00 \cancel{\text{ g}} \text{ Mg}\left(\frac{1 \cancel{\text{ mol}}}{24.3 \cancel{\text{ g}}}\right)\left(\frac{6.02\text{x}10^{23} \text{ atoms}}{1 \cancel{\text{ mol}}}\right) = 1.24\text{x}10^{23} \text{ atoms Mg}$$

4. How many molecules of HCl are present in 25.0 g of HCl?

MOLES OF ELEMENTS IN A FORMULA

- The subscripts in a chemical formula of a compound indicate the number of atoms of each type of element. For example, in a molecule of aspirin, $C_9H_8O_4$, there are 9 carbon atoms, 8 hydrogen atoms and 4 oxygen atoms.

- The subscript also indicates the number of moles of each element in one mole of the compound. For example, one mole of aspirin contains 9 moles of carbon atoms, 8 moles of hydrogen atoms and 4 moles of oxygen atoms.

$C_9H_8O_4$

	Carbon	**Hydrogen**	**Oxygen**
Atoms in 1 molecule	9 atoms C	8 atoms H	4 atoms O
Moles of atoms in 1 mole	9 moles C	8 moles H	4 moles O

- Using the subscripts from the aspirin formula, one can write the following conversion factors for each of the elements in 1 mole of aspirin:

$$\frac{\textbf{9 moles C}}{\textbf{1 mole } \mathbf{C_9H_8O_4}} \qquad \frac{\textbf{8 moles H}}{\textbf{1 mole } \mathbf{C_9H_8O_4}} \qquad \frac{\textbf{4 moles O}}{\textbf{1 mole } \mathbf{C_9H_8O_4}}$$

<u>Examples:</u>

1. Determine the moles of C atoms in 1 mole of each of the following substances:

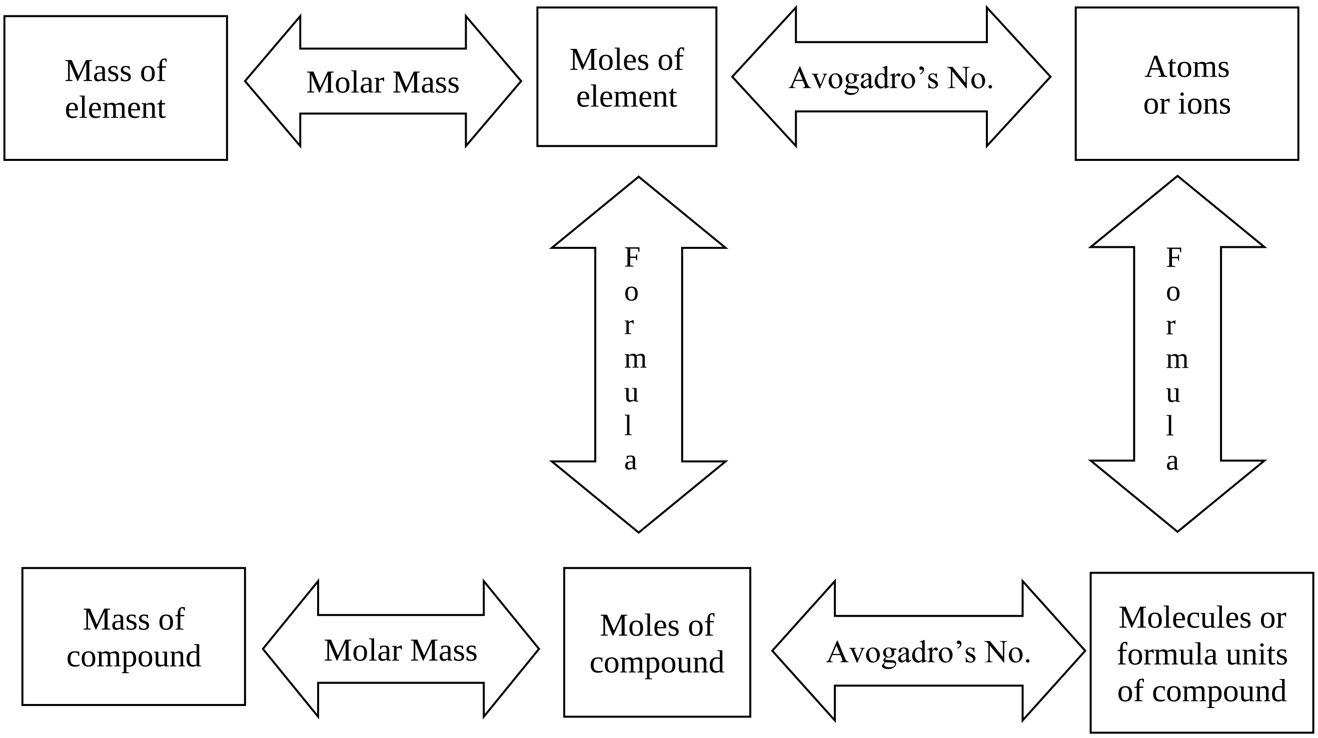
   a) Acetaminophen used in Tylenol, $C_8H_9NO_2$

   b) Zinc dietary supplement, $Zn(C_2H_3O_2)_2$

2. How many carbon atoms are present in 1.50 moles of aspirin, $C_9H_8O_4$?

SUMMARY OF MASS-MOLE CALCULATIONS

Mass **Moles** **Particles**



STOICHIOMETRY

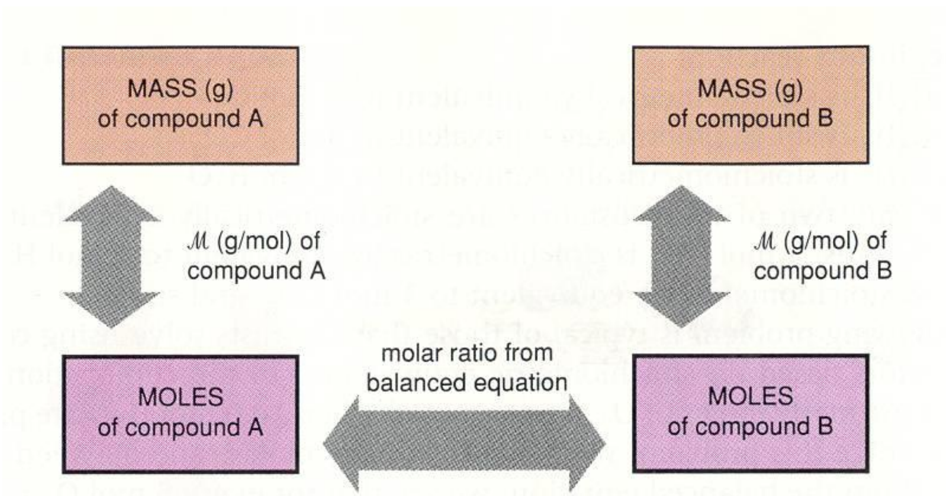
- ***Stoichiometry*** is the quantitative relationship between the ***reactants and products*** in a balanced ***chemical equation***.

- A balanced chemical equation provides several important information about the reactants and products in a chemical reaction. For example:

1N_2(g)	**+**	**3 H_2(g)**	**⟶ 2 NH_3(g)**
1 molecule		**3 molecules**	**2 molecules**
100 molecules		**300 molecules**	**200 molecules**
1 million molecules		**3 million molecules**	**2 million molecules**

1 mole N_2 + 3 moles H_2 ⟶ 2 moles NH_3

This is the **MOLE RATIO** between **REACTANTS** and **PRODUCTS**



Summary of Stoichiometric Calculations in Chemistry

<u>Examples:</u>

1. Determine each mole ration based on the reaction shown below:

$$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$$

a) $\dfrac{\text{mol } O_2}{\text{mol } CO_2} =$

b) $\dfrac{\text{mol } C_4H_{10}}{\text{mol } H_2O} =$

STOICHIOMETRIC CALCULATIONS

Mole-Mole Calculations:

- Relates moles of reactants and products in a balanced chemical equation

Examples:

1. How many moles of nitrogen will react with 2.4 moles of hydrogen to produce ammonia as shown in the reaction below?

$$1\ N_2(g) \quad + \quad 3\ H_2(g) \longrightarrow 2\ NH_3(g)$$

$$2.4\ \cancel{\text{mol } H_2} \times \frac{1 \text{ mol } N_2}{3\ \cancel{\text{mol } H_2}} = 0.80 \text{ mol } N_2$$

Mole Ratio

2. How many moles of ammonia can be produced from 32 moles of hydrogen? (Assume excess nitrogen present)

$$32 \text{ mol } H_2 \times \underline{\qquad\qquad} = \qquad\quad \text{mol } NH_3$$

3. In one experiment, 6.80 mol of ammonia are prepared. How many moles of hydrogen were used up in this experiment?

STOICHIOMETRIC CALCULATIONS

Mass-Mole Calculations:

- Relates moles and mass of reactants or products in a balanced chemical equation

Examples:

1. How many grams of ammonia can be produced from the reaction of 1.8 moles of nitrogen with excess hydrogen as shown below?

$$1\ N_2(g) \quad + \quad 3\ H_2(g) \longrightarrow 2\ NH_3(g)$$

$$1.8\ \cancel{\text{mol } N_2} \times \frac{2\ \cancel{\text{mol } NH_3}}{1\ \cancel{\text{mol } N_2}} \times \frac{17.0\text{ g } NH_3}{1\ \cancel{\text{mol } NH_3}} = 61\text{ g } NH_3$$

Mole Ratio **Molar Mass**

2. How many moles of hydrogen gas are required to produce 75.0 g of ammonia?

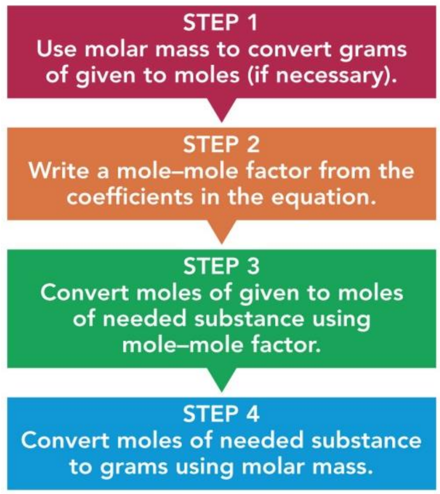
$$75.0\text{ g } NH_3 \times \underline{\qquad\qquad} \times \underline{\qquad\qquad} = \qquad\qquad \text{mol } H_2$$

3. How many moles of ammonia can be produced from the reaction of 125 g of nitrogen as shown above?

STOICHIOMETRIC CALCULATIONS

Mass-Mass Calculations:

- Relates mass of reactants and products in a balanced chemical equation



Examples:

1. What mass of oxygen will be required to react completely with 96.1 g of propane, C_3H_8, according to the equation below?

$$1\ C_3H_8\ (g)\ +\ 5\ O_2\ (g) \longrightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$$

$$96.1\ \cancel{g\ C_3H_8} \times \underbrace{\frac{1\ \cancel{mol\ C_3H_8}}{44.1\ \cancel{g\ C_3H_8}}}_{\text{Molar Mass}} \times \underbrace{\frac{5\ \cancel{mol\ O_2}}{1\ \cancel{mol\ C_3H_8}}}_{\text{Mole Ratio}} \times \underbrace{\frac{32.0\ g\ O_2}{1\ \cancel{mol\ O_2}}}_{\text{Molar Mass}} = 349\ g\ O_2$$

2. What mass of carbon dioxide will be produced from the reaction of 175 g of propane, as shown above?

$$175\ g\ C_3H_8 \times \frac{\qquad}{\qquad} \times \frac{\qquad}{\qquad} \times \frac{\qquad}{\qquad} = \qquad\qquad g\ CO_2$$

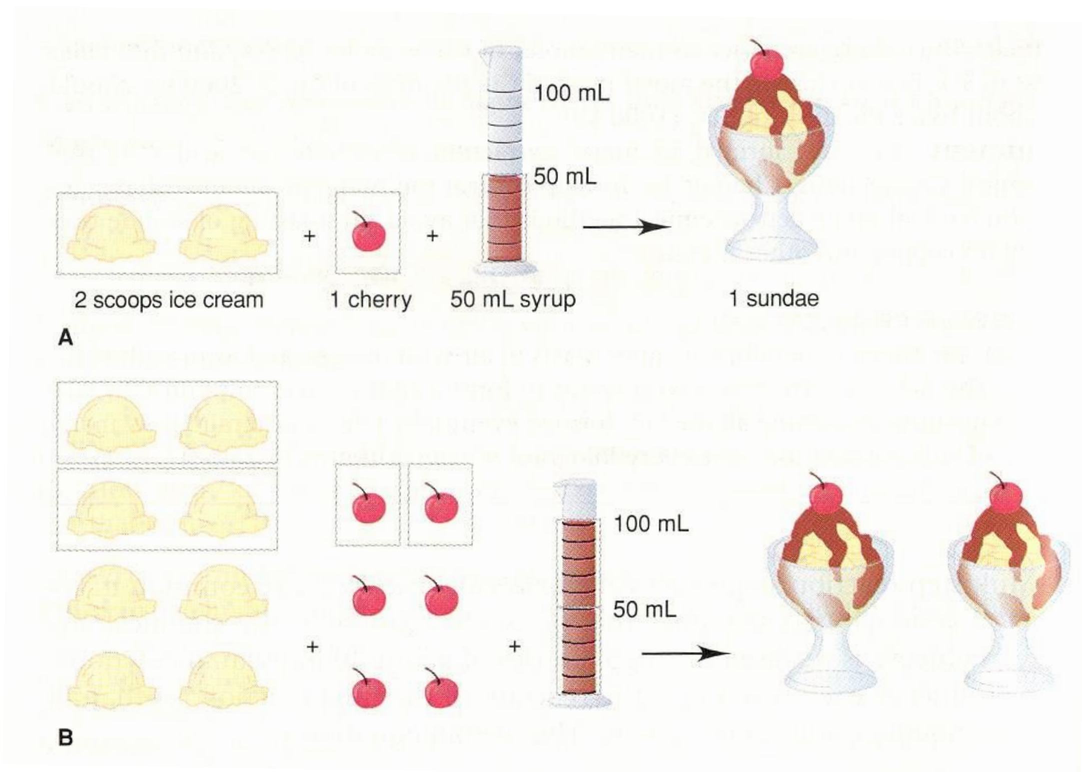
LIMITING REACTANT

- When ***2 or more reactants*** are combined in ***non-stoichiometric*** ratios, the amount of ***product*** produced is ***limited*** by the reactant that is ***not in excess***.

- This reactant is referred to as ***limiting reactant***.

- When doing stoichiometric problems of this type, the limiting reactant must be determined first before proceeding with the calculations.

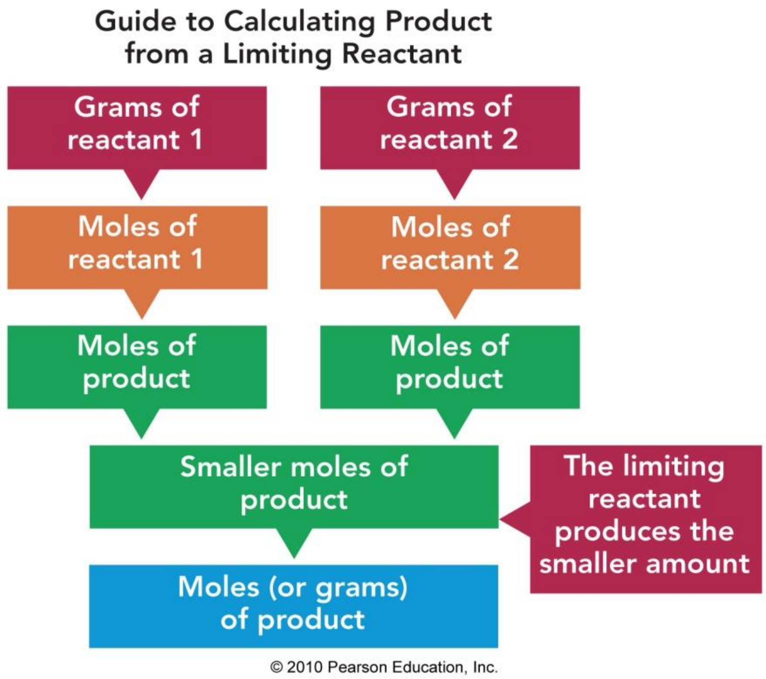
Analogy:



The number of sundaes possible is ***limited*** by the amount of syrup, the ***limiting reactant***.

- When solving limiting reactant problems, assume each reactant is limiting reactant and calculate the desired quantity based on that assumption.
- Compare your answers for each assumption; the lower value is the correct assumption.

LIMITING REACTANT



Examples:

1. How many moles of H_2O can be produced by reacting 4.0 mol of hydrogen and 3.0 mol of oxygen gases as shown below:

$$2\ H_2\ (g)\ +\ 1\ O_2\ (g) \longrightarrow 2\ H_2O\ (g)$$

Assume hydrogen is the limiting reactant:

$$4.0\ \text{mol}\ H_2 \times \frac{\qquad\qquad}{\qquad\qquad} = \qquad\qquad \text{mol}\ H_2O$$

Assume oxygen is the limiting reactant:

$$3.0\ \text{mol}\ O_2 \times \frac{\qquad\qquad}{\qquad\qquad} = \qquad\qquad \text{mol}\ H_2O$$

Correct limiting reactant is __________ and ___________ mol of H_2O are produced.

2. A fuel mixture used in the early days of rocketry was a mixture of N_2H_4 and N_2O_4, as shown below. How many grams of N_2 gas is produced when 100 g of N_2H_4 and 200 g of N_2O_4 are mixed?

$$2\ N_2H_4\ (l) + 1\ N_2O_4\ (l) \longrightarrow 3\ N_2\ (g) + 4\ H_2O\ (g)$$

$$100\ \cancel{g\ N_2H_4} \times \frac{1\ \cancel{mol\ N_2H_4}}{32.04\ \cancel{g\ N_2H_4}} \times \frac{3\ mol\ N_2}{2\ \cancel{mol\ N_2H_4}} = 4.68\ mol\ N_2$$

Limiting Reactant

$$200\ \cancel{g\ N_2O_4} \times \frac{1\ \cancel{mol\ N_2O_4}}{92.00\ \cancel{g\ N_2O_4}} \times \frac{3\ mol\ N_2}{1\ \cancel{mol\ N_2O_4}} = 6.52\ mol\ N_2$$

$$4.68\ \cancel{mol\ N_2} \times \frac{28.0\ g\ N_2}{1\ \cancel{mol\ N_2}} = 131\ g\ N_2$$

Notes:

1. Even though mass of N_2O_4 is greater than N_2H_4, there are fewer moles of it due to its larger molar mass.
2. Limiting reactant calculations must always be done when amount of both reactants are given.

LIMITING REACTANT

Examples:

3. How many moles of Fe_3O_4 can be produced by reacting 16.8 g of Fe with 10.0 g of H_2O as shown below:

$$3\,Fe\,(s) + 4\,H_2O\,(g) \xrightarrow{\Delta} Fe_3O_4\,(s) + 4\,H_2\,(g)$$

Assume Fe is the limiting reactant:

16.8 g Fe x ______________ x ______________ = mol Fe_3O_4

Assume H_2O is the limiting reactant:

10.0 g H_2O x ______________ x ______________ = mol Fe_3O_4

Correct limiting reactant is __________ and ___________ mol of Fe_3O_4 are produced.

4. How many grams of AgBr can be produced when 50.0 g of $MgBr_2$ is mixed with 100.0 g of $AgNO_3$, as shown below:

$$MgBr_2\,(aq) + 2\,AgNO_3\,(aq) \longrightarrow 2\,AgBr\,(s) + Mg(NO_3)_2\,(aq)$$

PERCENT YIELD

- The amount of product calculated through stoichiometric ratios is the ***maximum*** amount of product that can be produced during the reaction, and is thus called ***theoretical yield***.

- The ***actual yield*** of a product in a chemical reaction is the actual amount obtained from the reaction.

- The ***percent yield*** of a reaction is obtained as follows:

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \text{x}100 = \text{Percent yield}$$

Example:

1. In an experiment forming ethanol, the theoretical yield is 50.5 g and the actual yield is 46.8 g. What is the percent yield for this reaction?

$$\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \text{ x } 100 = \frac{46.8 \text{ g}}{50.5 \text{ g}} \text{ x } 100 = 92.7\ \%$$

2. Silicon carbide can be formed from the reaction of sand (SiO_2) with carbon as shown below:

$$SiO_2\ (s) + 3\ C\ (s) \longrightarrow SiC\ (s) + 2\ CO\ (g)$$

When 100 g of sand are processed, 51.4 of SiC is produced. What is the percent yield of SiC in this reaction?

$$100 \text{ g } SiO_2 \text{ x } \underline{\qquad\qquad} \text{ x } \underline{\qquad\qquad} \text{ x } \underline{\qquad\qquad} =$$

$$\% \text{ yield SiC} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \text{ x } 100 =$$

PERCENT YIELD

Examples:

3. In an experiment to prepare aspirin, the theoretical yield is 153.7 g and the actual yield is 124.0 g. What is the percent yield of this reaction?

$$\text{Percent yield} = \frac{\qquad\qquad}{\qquad\qquad} \times 100 =$$

4. Carbon tetrachloride (CCl_4) was prepared by reacting 100.0 g of Cl_2 with excess carbon disulfide (CS_2), as shown below. If 65.0 g was prepared in this reaction, calculate the percent yield.

$$CS_2 + 3\ Cl_2 \rightarrow CCl_4 + S_2Cl_2$$

Calculate the theoretical yield:

$$100.0\text{ g } Cl_2 \times \frac{\qquad\qquad}{\qquad\qquad} \times \frac{\qquad\qquad}{\qquad\qquad} \times \frac{\qquad\qquad}{\qquad\qquad} = \qquad\qquad \text{g } CCl_4$$

Calculate percent yield:

$$\text{Percent yield} = \frac{\qquad\qquad}{\qquad\qquad} \times 100 =$$

ENERGY CHANGES IN CHEMICAL REACTIONS

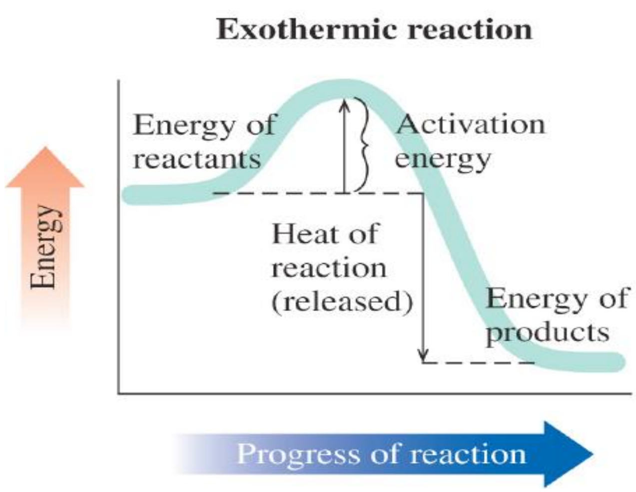
- ***Heat*** is energy change that is lost or gained when a chemical reaction takes place.

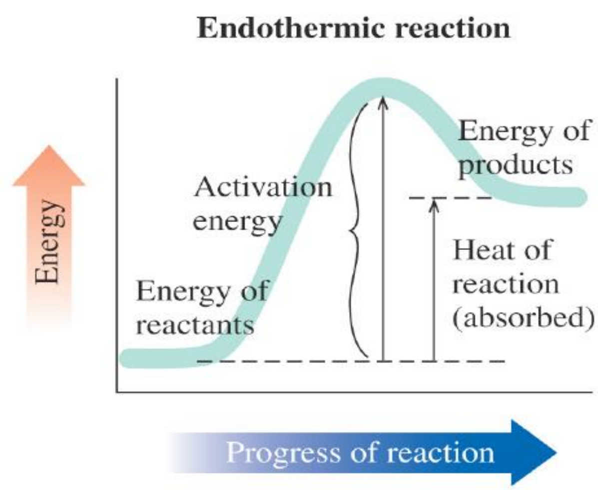
- The ***system*** is the reactants and products that we are observing. The ***surroundings*** are all the things that contain and interact with the system, such as the reaction flask, the laboratory room and the air in the room.

- The ***direction of heat flow*** depends whether the products in a reaction have more or less energy than the reactants.

- For a chemical reaction to occur, the molecules of the reactants must collide with each other with the proper ***energy and orientation***. The minimum amount of energy required for a chemical reaction to occur is called the ***activation energy***.

- The ***heat of reaction*** is the amount of heat absorbed or released during a reaction and is designated by the symbol ΔH.





- When ***heat is released*** during a chemical reaction, it is said to be ***exothermic***. For exothermic reactions, ΔH is negative and is included on the right side of the equation.

  $3\ H_2\ (g) + N_2\ (g) \rightarrow 2\ NH_3\ (g) + 22.0\ \text{kcal}$ $\quad\quad \Delta H = -22.0\ \text{kcal}$

- When ***heat is gained*** during a chemical reaction, it is said to be ***endothermic***. For endothermic reactions, ΔH is positive and is included on the left side of the equation.

  $2\ H_2O\ (l) + 137\ \text{kcal} \rightarrow 2\ H_2\ (g) + O_2\ (g)$ $\quad\quad \Delta H = +137\ \text{kcal}$

CALCULATING HEAT IN A REACTION

- The value of ΔH in a chemical reaction refers to the heat lost or gained for the number of moles of reactants and products in a balanced chemical equation. For example, based on the chemical equation shown below:

$$2\ H_2O\ (l) \rightarrow 2\ H_2\ (g) + O_2\ (g) \qquad \Delta H = +137\ \text{kcal}$$

The following conversion factors can be written:

$$\frac{137\ \text{kcal}}{2\ \text{mol}\ H_2O} \quad \text{or} \quad \frac{137\ \text{kcal}}{2\ \text{mol}\ H_2} \quad \text{or} \quad \frac{137\ \text{kcal}}{1\ \text{mol}\ O_2}$$

- These conversion factors can be used to calculate the amount of heat associated with a particular chemical reaction, based on given amounts of reactants and products.

Examples:

1. Given the reaction shown below, how much heat, in kJ, is released when 50.0 g of NH_3 form?

$$3\ H_2\ (g) + N_2\ (g) \rightarrow 2\ NH_3\ (g) \qquad \Delta H = -92.2\ \text{kJ}$$

2. How many kJ of heat are needed to react 25.0 g of HgO according to the reaction shown below:

$$2\ HgO\ (s) \rightarrow 2\ Hg\ (l) + O_2\ (g) \qquad \Delta H = 182\ \text{kJ}$$