Chapter 9: An Overview of Chemical Reactions

9.1: Chemical Reactions are Represented by Chemical Equations

Reactants: - chemicals that goes into a reaction.

Products: - chemicals that are produced from a reaction.

$$\textbf{Reactants} \xrightarrow{\text{"yields"}} \textbf{Products}$$

States of Chemicals: - (*s*) solid, (*l*) liquid, (*g*) gas, (*aq*) aqueous – dissolved in water

Other Chemical Symbols:

1. **Heat is Added**: - $\xrightarrow{\textit{heat}}$ or $\xrightarrow{\Delta}$
2. **Catalyst is Added**: - $\xrightarrow{\textit{Name of Catalyst}}$

(**Catalyst**: - a chemical that is used to speed up a reaction but does not get consumed)

Example 1: Convert the following word equations into chemical equations without balancing.

a. Heating solid diphosphorous pentaoxide decomposes into phosphorous and oxygen.

$P_2O_{5\,(s)} \rightarrow P_{4\,(s)} + O_{2\,(g)}$	**Recall that phosphorus is a polyatomic element, and oxygen is a diatomic element.**

b. Hypochlorous acid ($HClO_{(aq)}$) is neutralized by barium hydroxide solution to form water and soluble barium hypochlorite.

$HClO_{(aq)} + Ba(OH)_{2\,(aq)} \rightarrow H_2O_{(l)} + Ba(ClO)_{2\,(aq)}$	**It is very important that the subscripts for these ionic compounds are correct.**

Balancing Chemical Equation: - the process by which we place **coefficient (numbers in front of reactants and products)** in an attempt to equate the number of atoms or polyatomic ions for the elements or compounds in a chemical equation.

Steps involve to Balance Chemical Equation

1. Write the chemical formulas of all reactants and products with their proper subscripts and state to form a skeletal equation.
2. Count the number of atoms / polyatomic ions (always view complex ion as a group) of each chemical formula. Balance the atoms / polyatomic ions by writing the coefficient in front of the reactant / product. Do NOT mess with any of the subscripts.
3. Always balance the atoms that appear in more than two chemicals last.
4. Verify that all atoms / polyatomic ions are balanced.

Example 2: Balance the following chemical equations

a. $N_{2\,(g)} + H_{2\,(g)} \rightarrow NH_{3\,(g)}$

$N_{2\,(g)} + \underline{3}\ H_{2\,(g)} \rightarrow \underline{2}\ NH_{3\,(g)}$

b. $Al_2O_{3\,(s)} \rightarrow Al_{(s)} + O_{2\,(g)}$

$\underline{2}\ Al_2O_{3\,(s)} \rightarrow \underline{4}\ Al_{(s)} + \underline{3}\ O_{2\,(g)}$

c. $Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$

$Na_{(s)} + HOH_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$
$\underline{2}\ Na_{(s)} + \underline{2}\ HOH_{(l)} \rightarrow \underline{2}\ NaOH_{(aq)} + H_{2\,(g)}$

Sometimes, it is easier to rewrite H_2O as HOH. This is especially true when a OH^- is part of the products. The first H atom in the HOH becomes $H_{2\,(g)}$ and the remaining part, OH, becomes the polyatomic ion, OH^-.

d. $Fe(NO_3)_{3\,(aq)} + Ba(OH)_{2\,(aq)} \rightarrow Fe(OH)_{3\,(s)} + Ba(NO_3)_{2\,(aq)}$

$\underline{2}\ Fe(NO_3)_{3\,(aq)} + \underline{3}\ Ba(OH)_{2\,(aq)} \rightarrow \underline{2}\ Fe(OH)_{3\,(s)} + \underline{3}\ Ba(NO_3)_{2\,(aq)}$

Look at each polyatomic ion as one item. There were $(NO_3)_3$ on the left hand side and $(NO_3)_2$ on the right hand side. Hence, we need to use 2 and 3 as coefficients to balance them. Similarly, there were $(OH)_2$ on the reactant side and $(OH)_3$ on the product side. Therefore, we are required to use 3 and 2 to balance them. Note that once the coefficients are in place, the Fe and Ba atoms are also balanced.

e. $C_2H_{6\,(g)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O_{(l)}$

For burning (adding O_2) with hydrocarbons (compounds containing carbon and hydrogen), we must *balance C, H, O in that order*. This is because oxygen atoms exist in more than two compounds on the product side. After we balance carbon and hydrogen, we have the following number of oxygen atoms.

$C_2H_{6\,(g)}$	+	$\underline{\ \ }\ O_{2\,(g)}$	$\rightarrow$	$\underline{2}\ CO_{2\,(g)}$	+	$\underline{3}\ H_2O_{(l)}$
		need 7 oxygen		4 oxygen		3 oxygen
$C_2H_{6\,(g)}$	+	$\underline{\tfrac{7}{2}}\ O_{2\,(g)}$	$\rightarrow$	$\underline{2}\ CO_{2\,(g)}$	+	$\underline{3}\ H_2O_{(l)}$

Multiply all coefficients by 2:

$2\ C_2H_{6\,(g)} + \underline{7}\ O_{2\,(g)} \rightarrow \underline{4}\ CO_{2\,(g)} + \underline{6}\ H_2O_{(l)}$

Since elemental oxygen is diatomic, the 7 oxygen needed on the reactant side needs to have a coefficient of 7/2.

This can easily be converted to a whole number by multiplying all the coefficients by 2.

Example 3: Rewrite the following sentences into balanced chemical equations.

a. Sulfur is burned in the presence of oxygen to form sulfur dioxide gas.

First, we write the skeletal equation.
(Recall that sulfur is a polyatomic element, and oxygen is a diatomic element.)

$$S_{8\,(s)} + O_{2\,(g)} \rightarrow SO_{2\,(g)}$$

Then, we balance the equation. $S_{8\,(s)} + \underline{8}\ O_{2\,(g)} \rightarrow \underline{8}\ SO_{2\,(g)}$

b. Hydrogen gas is added to propyne (C_3H_4) gas with the aid of a platinum catalyst to form propane gas.

First, we write the skeletal equation. (Propane is C_3H_8 – a formula that should have been memorized.)

$$C_3H_{4\,(g)} + H_{2\,(g)} \xrightarrow{Pt} C_3H_{8\,(g)}$$

Then, we balance the equation.

$C_3H_{4\,(g)}$	+	$\underline{2}\ H_{2\,(g)}$	$\xrightarrow{Pt}$	$C_3H_{8\,(g)}$
4 hydrogen		*need 4 more hydrogen*		8 hydrogen

There are 5 basic types of chemical reactions:

1. **Formation or Composition** **(Element + Element → Compound)**

   Example: $2\ Mg_{(s)} + O_{2\,(g)} \rightarrow 2\ MgO_{(s)}$

2. **Deformation or Decomposition** **(Compound → Element + Element)**

   Example: $2\ Al_2O_{3\,(s)} \rightarrow 4\ Al_{(s)} + 3\ O_{2\,(g)}$

3. **Single Replacement** **(Element + Compound → Element + Compound)**

   Example: $2\ AgNO_{3\,(aq)} + Cu_{(s)} \rightarrow 2\ Ag_{(s)} + Cu(NO_3)_{2\,(aq)}$

4. **Double Replacement** **(Compound + Compound → Compound + Compound)**

   Example: $AgNO_{3\,(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3\,(aq)}$

5. **Hydrocarbon Combustion** **(Hydrocarbon + Oxygen → Carbon Dioxide + Water)**

   Example: $CH_{4\,(g)} + 2O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\ H_2O_{(g)}$

Precipitation Reaction: - a reaction where a precipitate (new solid) is formed as a product.

Neutralization Reaction: - a reaction between an acid and a base where water is formed as a product.

To Predict Products and Balance Chemical Equations:

1. Write the correct chemical formulas for all products and reactants with proper subscripts. The presence of metals or ionic compounds indicates that we will need to use ions and charges to form any products.
2. For hydrocarbon combustion, balance in the order of C, H, and then O.
3. For other type of reactions, balance the equation for each type of cations and anions. Do NOT break up complex ions. Water may be written as HOH (H^+ and OH^-) in single and double replacement reactions.
4. Check with the Solubility Table (see Section 7.3) and the Table of Elements for the states of chemicals.

Example 4: Predict the product(s) along with the states, indicate the type of reaction, and balance the following chemical reactions.

a. Sulfur trioxide gas is produced from its elements.

Formation: $S_{8\,(s)} + \underline{12}\ O_{2\,(g)} \rightarrow \underline{8}\ SO_{3\,(g)}$

b. A solid piece of zinc is immersed in an iron (III) chloride solution.

Single Replacement: $\underline{3}\ Zn_{(s)} + \underline{2}\ FeCl_{3\,(aq)} \rightarrow \underline{3}\ ZnCl_{2\,(aq)} + \underline{2}\ Fe_{(s)}$

Zn^{2+} Fe^{3+} Cl^-

c. Propane ($C_3H_{8\,(g)}$) is burned in a gas barbecue.

Hydrocarbon Combustion: $C_3H_{8\,(g)} + \underline{5}\ O_{2\,(g)} \rightarrow \underline{3}\ CO_{2\,(g)} + \underline{4}\ H_2O_{(g)}$

d. Chlorine gas is bubbled through a copper (II) iodide solution.

Single Replacement: $Cl_{2\,(g)} + CuI_{2\,(aq)} \rightarrow CuCl_{2\,(aq)} + I_{2\,(s)}$

Cl^- Cu^{2+} I^-

e. Ammonia gas is decomposed into its elements.

Decomposition: $\underline{2}\ NH_{3\,(g)} \rightarrow N_{2\,(g)} + \underline{3}\ H_{2\,(g)}$

f. Sulfuric acid ($H_2SO_{4\,(aq)}$) is neutralized by sodium hydroxide solution.

Double Replacement: $H_2SO_{4\,(aq)} + \underline{2}\ NaOH_{(aq)} \rightarrow 2\ HOH_{(l)} + Na_2SO_{4\,(aq)}$

H^+ SO_4^{2-} Na^+ OH^-

g. Propanol ($C_3H_7OH_{(l)}$) is accidentally ignited.

Hydrocarbon Combustion: $C_3H_7OH_{(l)} + \frac{9}{2}\ O_{2\,(g)} \rightarrow \underline{3}\ CO_{2\,(g)} + \underline{4}\ H_2O_{(g)}$

(Multiply Coefficients by 2) $\underline{2}\ C_3H_7OH_{(l)} + \underline{9}\ O_{2\,(g)} \rightarrow \underline{6}\ CO_{2\,(g)} + \underline{8}\ H_2O_{(g)}$

h. Lead (II) nitrate solution is reacted with chromium (III) sulfate solution.

Double Replacement: $\underline{3}\ Pb(NO_3)_{2\,(aq)} + Cr_2(SO_4)_{3\,(aq)} \rightarrow \underline{2}\ Cr(NO_3)_{3\,(aq)} + \underline{3}\ PbSO_{4\,(s)}$

Pb^{2+} NO_3^- Cr^{3+} SO_4^{2-}

i. Octane ($C_8H_{18\,(l)}$) is combusted in an automobile.

Hydrocarbon Combustion: $C_8H_{18\,(l)} + \frac{25}{2}\ O_{2\,(g)} \rightarrow \underline{8}\ CO_{2\,(g)} + \underline{9}\ H_2O_{(g)}$

(Multiply Coefficients by 2) $\underline{2}\ C_8H_{18\,(l)} + \underline{25}\ O_{2\,(g)} \rightarrow \underline{16}\ CO_{2\,(g)} + \underline{18}\ H_2O_{(g)}$

Assignment

9.1 pg. 301 #1 to 6 and Worksheet: Writing and Balancing Chemical Equations

Worksheet: Writing and Balancing Chemical Reactions

1. Balance the following equations and indicate the type of reaction as formation, decomposition, single replacement, double replacement, hydrocarbon combustion, or other.

a. ____ $Cu_{(s)}$ + ____ $O_{2\,(g)}$ → ____ $CuO_{(s)}$
b. ____ $H_2O_{(l)}$ → ____ $H_{2\,(g)}$ + ____ $O_{2\,(g)}$
c. ____ $Fe_{(s)}$ + ____ $H_2O_{(g)}$ → ____ $H_{2\,(g)}$ + ____ $Fe_3O_{4\,(s)}$
d. ____ $AsCl_{3\,(aq)}$ + ____ $H_2S_{(aq)}$ → ____ $As_2S_{3\,(s)}$ + ____ $HCl_{(aq)}$
e. ____ $CuSO_4 \bullet 5\,H_2O_{(s)}$ → ____ $CuSO_{4\,(s)}$ + ____ $H_2O_{(g)}$
f. ____ $Fe_2O_{3\,(s)}$ + ____ $H_{2\,(g)}$ → ____ $Fe_{(s)}$ + ____ $H_2O_{(l)}$
g. ____ $CaCO_{3\,(s)}$ → ____ $CaO_{(s)}$ + ____ $CO_{2\,(g)}$
h. ____ $Fe_{(s)}$ + ____ $S_{8\,(s)}$ → ____ $FeS_{(s)}$
i. ____ $H_2S_{(aq)}$ + ____ $KOH_{(aq)}$ → ____ $H_2O_{(l)}$ + ____ $K_2S_{(aq)}$
j. ____ $NaCl_{(l)}$ → ____ $Na_{(l)}$ + ____ $Cl_{2\,(g)}$
k. ____ $Al_{(s)}$ + ____ $H_2SO_{4\,(aq)}$ → ____ $H_{2\,(g)}$ + ____ $Al_2(SO_4)_{3\,(aq)}$
l. ____ $H_3PO_{4\,(aq)}$ + ____ $NH_4OH_{(aq)}$ → ____ $H_2O_{(l)}$ + ____ $(NH_4)_3PO_{4\,(aq)}$
m. ____ $C_3H_{8\,(g)}$ + ____ $O_{2\,(g)}$ → ____ $CO_{2\,(g)}$ + ____ $H_2O_{(l)}$
n. ____ $Al_{(s)}$ + ____ $O_{2\,(g)}$ → ____ $Al_2O_{3\,(s)}$
o. ____ $CH_{4\,(g)}$ + ____ $O_{2\,(g)}$ → ____ $CO_{2\,(g)}$ + ____ $H_2O_{(l)}$
p. ____ $K_2SO_{4\,(aq)}$ + ____ $BaCl_{2\,(aq)}$ → ____ $KCl_{(aq)}$ + ____ $BaSO_{4\,(s)}$
q. ____ $C_5H_{12\,(l)}$ + ____ $O_{2\,(g)}$ → ____ $CO_{2\,(g)}$ + ____ $H_2O_{(g)}$
r. ____ $Ca(OH)_{2\,(aq)}$ + ____ $NH_4Cl_{(aq)}$ → ____ $NH_4OH_{(aq)}$ + ____ $CaCl_{2\,(aq)}$
s. ____ $V_2O_{5\,(s)}$ + ____ $Ca_{(s)}$ → ____ $CaO_{(s)}$ + ____ $V_{(s)}$
t. ____ $Na_{(s)}$ + ____ $ZnI_{2\,(aq)}$ → ____ $NaI_{(aq)}$ + ____ $Zn_{(s)}$
u. ____ $C_7H_6O_{3\,(l)}$ + ____ $O_{2\,(g)}$ → ____ $CO_{2\,(g)}$ + ____ $H_2O_{(l)}$
v. ____ $Ca_{(s)}$ + ____ $N_{2\,(g)}$ → ____ $Ca_3N_{2\,(s)}$
w. ____ $Fe_2O_{3\,(s)}$ + ____ $H_{2\,(g)}$ → ____ $Fe_{(s)}$ + ____ $H_2O_{(l)}$
x. ____ $C_{15}H_{30\,(l)}$ + ____ $O_{2\,(g)}$ → ____ $CO_{2\,(g)}$ + ____ $H_2O_{(g)}$
y. ____ $BN_{(s)}$ + ____ $F_{2\,(g)}$ → ____ $BF_{3\,(s)}$ + ____ $N_{2\,(g)}$
z. ____ $C_{12}H_{26\,(l)}$ + ____ $O_{2\,(g)}$ → ____ $CO_{2\,(g)}$ + ____ $H_2O_{(g)}$

2. Predict the product(s) along with the states, indicate the type of reaction, and balance the following chemical reactions.

a. A solution of lead (II) nitrate is mixed with a solution of sodium iodide.
b. Solid zinc sulfide reacts with oxygen in the air.
c. Liquid butane ($C_4H_{10\,(l)}$) is used as a fuel to ignite a lighter.
d. Barium hydroxide solution is neutralized by adding hydrochloric acid ($HCl_{(aq)}$).
e. Copper metal is placed in a solution of silver nitrate.
f. Sulfur burns in oxygen to make sulfur dioxide gas.
g. A solution of aluminum sulfate is mixed with a solution of calcium hydroxide.
h. Zinc metal is placed in sulfuric acid ($H_2SO_{4\,(aq)}$).
i. Aluminum powder is placed in a container filled with chlorine gas.
j. Sucrose undergoes cellular respiration.

Answers

Question 1

a. $\mathbf{2}\ Cu_{(s)} + O_{2\,(g)} \rightarrow \mathbf{2}\ CuO_{(s)}$ **(formation)**
b. $\mathbf{2}\ H_2O_{(l)} \rightarrow \mathbf{2}\ H_{2\,(g)} + O_{2\,(g)}$ **(decomposition)**
c. $\mathbf{3}\ Fe_{(s)} + \mathbf{4}\ H_2O_{(g)} \rightarrow \mathbf{4}\ H_{2\,(g)} + Fe_3O_{4\,(s)}$ **(single replacement)**
d. $\mathbf{2}\ AsCl_{3\,(aq)} + \mathbf{3}\ H_2S_{(aq)} \rightarrow As_2S_{3\,(s)} + \mathbf{6}\ HCl_{(aq)}$ **(double replacement)**
e. $CuSO_4 \bullet 5\ H_2O_{(s)} \rightarrow CuSO_{4\,(s)} + \mathbf{5}\ H_2O_{(g)}$ **(other – dehydration or decomposition)**
f. $Fe_2O_{3\,(s)} + \mathbf{3}\ H_{2\,(g)} \rightarrow \mathbf{2}\ Fe_{(s)} + \mathbf{3}\ H_2O_{(l)}$ **(single replacement)**
g. $CaCO_{3\,(s)} \rightarrow CaO_{(s)} + CO_{2\,(g)}$ **(other or decomposition)**
h. $\mathbf{8}\ Fe_{(s)} + S_{8\,(s)} \rightarrow \mathbf{8}\ FeS_{(s)}$ **(formation)**
i. $H_2S_{(aq)} + \mathbf{2}\ KOH_{(aq)} \rightarrow \mathbf{2}\ H_2O_{(l)} + K_2S_{(aq)}$ **(double replacement)**
j. $\mathbf{2}\ NaCl_{(l)} \rightarrow \mathbf{2}\ Na_{(l)} + Cl_{2\,(g)}$ **(decomposition)**
k. $\mathbf{2}\ Al_{(s)} + \mathbf{3}\ H_2SO_{4\,(aq)} \rightarrow \mathbf{3}\ H_{2\,(g)} + Al_2(SO_4)_{3\,(aq)}$ **(single replacement)**
l. $H_3PO_{4\,(aq)} + \mathbf{3}\ NH_4OH_{(aq)} \rightarrow \mathbf{3}\ H_2O_{(l)} + (NH_4)_3PO_{4\,(aq)}$ **(double replacement)**
m. $C_3H_{8\,(g)} + \mathbf{5}\ O_{2\,(g)} \rightarrow \mathbf{3}\ CO_{2\,(g)} + \mathbf{4}\ H_2O_{(l)}$ **(hydrocarbon combustion)**
n. $\mathbf{4}\ Al_{(s)} + \mathbf{3}\ O_{2\,(g)} \rightarrow \mathbf{2}\ Al_2O_{3\,(s)}$ **(formation)**
o. $CH_{4\,(g)} + \mathbf{2}\ O_{2\,(g)} \rightarrow CO_{2\,(g)} + \mathbf{2}\ H_2O_{(l)}$ **(hydrocarbon combustion)**
p. $K_2SO_{4\,(aq)} + BaCl_{2\,(aq)} \rightarrow \mathbf{2}\ KCl_{(aq)} + BaSO_{4\,(s)}$ **(double replacement)**
q. $C_5H_{12\,(l)} + \mathbf{8}\ O_{2\,(g)} \rightarrow \mathbf{5}\ CO_{2\,(g)} + \mathbf{6}\ H_2O_{(g)}$ **(hydrocarbon combustion)**
r. $Ca(OH)_{2\,(aq)} + \mathbf{2}\ NH_4Cl_{(aq)} \rightarrow \mathbf{2}\ NH_4OH_{(aq)} + CaCl_{2\,(aq)}$ **(double replacement)**
s. $V_2O_{5\,(s)} + \mathbf{5}\ Ca_{(s)} \rightarrow \mathbf{5}\ CaO_{(s)} + \mathbf{2}\ V_{(s)}$ **(single replacement)**
t. $\mathbf{2}\ Na_{(s)} + ZnI_{2\,(aq)} \rightarrow \mathbf{2}\ NaI_{(aq)} + Zn_{(s)}$ **(single replacement)**
u. $C_7H_6O_{3\,(l)} + \mathbf{7}\ O_{2\,(g)} \rightarrow \mathbf{7}\ CO_{2\,(g)} + \mathbf{3}\ H_2O_{(l)}$ **(hydrocarbon combustion)**
v. $\mathbf{3}\ Ca_{(s)} + N_{2\,(g)} \rightarrow Ca_3N_{2\,(s)}$ **(formation)**
w. $Fe_2O_{3\,(s)} + \mathbf{3}\ H_{2\,(g)} \rightarrow \mathbf{2}\ Fe_{(s)} + \mathbf{3}\ H_2O_{(l)}$ **(single replacement)**
x. $\mathbf{2}\ C_{15}H_{30\,(l)} + \mathbf{45}\ O_{2\,(g)} \rightarrow \mathbf{30}\ CO_{2\,(g)} + \mathbf{30}\ H_2O_{(g)}$ **(hydrocarbon combustion)**
y. $\mathbf{2}\ BN_{(s)} + \mathbf{3}\ F_{2\,(g)} \rightarrow \mathbf{2}\ BF_{3\,(s)} + N_{2\,(g)}$ **(single replacement)**
z. $\mathbf{2}\ C_{12}H_{26\,(l)} + \mathbf{37}\ O_{2\,(g)} \rightarrow \mathbf{24}\ CO_{2\,(g)} + \mathbf{26}\ H_2O_{(g)}$ **(hydrocarbon combustion)**

Question 2

a. $Pb(NO_3)_{2\,(aq)} + \mathbf{2}\ NaI_{(aq)} \rightarrow PbI_{2\,(s)} + \mathbf{2}\ NaNO_{3\,(aq)}$ **(double replacement)**
b. $\mathbf{8}\ ZnS_{(s)} + \mathbf{4}\ O_{2\,(g)} \rightarrow \mathbf{8}\ ZnO_{(s)} + S_{8\,(s)}$ **(single replacement)**
c. $\mathbf{2}\ C_4H_{10\,(l)} + \mathbf{13}\ O_{2\,(g)} \rightarrow \mathbf{8}\ CO_{2\,(g)} + \mathbf{10}\ H_2O_{(g)}$ **(hydrocarbon combustion)**
d. $Ba(OH)_{2\,(aq)} + \mathbf{2}\ HCl_{(aq)} \rightarrow BaCl_{2\,(aq)} + \mathbf{2}\ H_2O_{(l)}$ **(double replacement)**
e. $Cu_{(s)} + \mathbf{2}\ AgNO_{3\,(aq)} \rightarrow Cu(NO_3)_{2\,(aq)} + \mathbf{2}\ Ag_{(s)}$ **(single replacement)**
f. $S_{8\,(s)} + 8\ O_{2\,(g)} \rightarrow 8\ SO_{2\,(g)}$ **(formation)**
g. $Al_2(SO_4)_{3\,(aq)} + \mathbf{3}\ Ca(OH)_{2\,(aq)} \rightarrow \mathbf{2}\ Al(OH)_{3\,(s)} + \mathbf{3}\ CaSO_{4\,(s)}$ **(double replacement)**
h. $Zn_{(s)} + H_2SO_{4\,(aq)} \rightarrow ZnSO_{4\,(aq)} + H_{2\,(g)}$ **(single replacement)**
i. $\mathbf{2}\ Al_{(s)} + \mathbf{3}\ Cl_{2\,(g)} \rightarrow \mathbf{2}\ AlCl_{3\,(s)}$ **(formation)**
j. $C_{12}H_{22}O_{11\,(s)} + \mathbf{12}\ O_{2\,(g)} \rightarrow \mathbf{12}\ CO_{2\,(g)} + \mathbf{11}\ H_2O_{(l)}$ **(hydrocarbon combustion)**

9.2A: Chemists Use Relative Masses to Count Atoms and Molecules

Relative Mass: - the mass of one atom compares to another type of atom.
- the atomic mass in the Table of Elements are relative masses. It is based on carbon, which is assigned the mass of 12.00. This means that chlorine atom (with an atomic mass of 35.45) is roughly 1.5 times more massive than a carbon atom.

Formula Mass: - the sum of all the atomic masses in a chemical formula.
- it is the same as the **molar mass (*M*)**, which is the mass per mole (6.02×10^{23} molecules) of a substance.

Avogadro's Number: - a group of (6.02×10^{23}) molecules = 1 mole

Stoichiometry: - the calculation of quantities in a chemical reaction.
- the coefficients of various reactants and /or products form **mole ratios**.
- these mole ratio hold for moles, molecules or atoms.

Mole Ratio: - a ratio form between the coefficient of the required chemical amount to the given chemical amount. $\left(\dfrac{\textbf{require coefficient}}{\textbf{given coefficient}}\right)$

Example 1: Interpret the chemical equation $4\ NH_{3\,(g)} + 7\ O_{2\,(g)} \rightarrow 4\ NO_{2\,(g)} + 6\ H_2O_{(g)}$ in terms of
a. moles.
b. molecules.
c. masses.

	$4\ NH_{3\,(g)}$	$7\ O_{2\,(g)}$	$4\ NO_{2\,(g)}$	$6\ H_2O_{(g)}$
a.	**4 moles of NH_3**	**7 moles of O_2**	**4 moles of NO_2**	**6 moles of H_2O**
b.	**4 molecules of NH_3**	**7 molecules of O_2**	**4 molecules of NO_2**	**6 molecules of H_2O**
c.	$m = nM$ $m = (4\ \cancel{\text{mol}})(17.04\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 68.16\ \text{g}}$	$m = nM$ $m = (7\ \cancel{\text{mol}})(32.00\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 224.0\ \text{g}}$	$m = nM$ $m = (4\ \cancel{\text{mol}})(46.01\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 184.0\ \text{g}}$	$m = nM$ $m = (6\ \cancel{\text{mol}})(18.02\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 108.1\ \text{g}}$

Example 2: 4.35 mol of $PCl_{5\,(g)}$ is decomposed into its elements. Write a balance equation and determined the amount of chlorine produced.

$$\underline{4}\ PCl_{5\,(g)} \rightarrow P_{4\,(s)} + \underline{10}\ Cl_{2\,(g)}$$

4.35 mol ? mol

$$n_{Cl_2} = 4.35\ \cancel{\text{mol PCl}_5} \times \frac{\textbf{10 mol Cl}_2}{\textbf{4}\ \cancel{\textbf{mol PCl}_5}} = 10.875\ \text{mol Cl}_2 \qquad \boldsymbol{n_{Cl_2} = 10.9\ \text{mol}}$$

Gravimetric Stoichiometry: - stoichiometry that involves quantities of masses.

Gravimetric Stoichiometry Procedure:

1. Predict the products and balance the chemical equation.
2. Put all the information given under the appropriate chemicals and determine the molar masses of the chemical involved.
3. Find the moles of the given chemical. $\left(n = \frac{m}{M}\right)$
4. Find the mole of the required chemical using mole ratio.
$$\left(\text{mol of require} = \text{mol of given} \times \frac{\text{require coefficient}}{\text{given coefficient}}\right)$$
5. Convert mole of the required chemical to its mass equivalence. **($m = nM$)**

Example 3: Determine the mass of carbon dioxide formed when 50.0 kg of butane ($C_4H_{10\,(l)}$) is burned.

$2\ C_4H_{10\,(g)}$	+	$13\ O_{2\,(g)}$	→	$8\ CO_{2\,(g)}$	+	$10\ H_2O_{(g)}$
50.0 kg				? g		
58.14 g/mol				M = 44.01 g/mol		

① $n_{C_4H_{10}} = \dfrac{50.0\ \cancel{\text{k}}\text{g}}{58.14\ \cancel{\text{g}}\text{/mol}} = 0.8599931201\ \text{kmol}\ C_4H_{10}$

② $n_{CO_2} = 0.8599931201\ \text{k}\cancel{\text{mol}\ C_4H_{10}} \times \dfrac{8\ \text{mol}\ CO_2}{2\ \cancel{\text{mol}\ C_4H_{10}}} = 3.43997248\ \text{kmol}\ CO_2$

③ $m_{CO_2} = nM = (3.43997248\ \text{k}\cancel{\text{mol}}\ CO_2)(44.01\ \text{g/}\cancel{\text{mol}})$

$\boldsymbol{m_{CO_2} = 151\ \text{kg}}$

Example 4: Barium chloride solution was mixed with an excess sodium phosphate solution. What was the mass of barium chloride solid needed in the original solution to form 3.21 g of precipitate?

$3\ BaCl_{2\,(aq)}$	+	$2\ Na_3PO_{4\,(aq)}$	→	$Ba_3(PO_4)_{2\,(s)}$	+	$6\ NaCl_{(aq)}$
? g				3.21 g		
208.23 g/mol				M = 601.93 g/mol		

① $n_{Ba_3(PO_4)_2} = \dfrac{3.21\ \cancel{\text{g}}}{601.93\ \cancel{\text{g}}\text{/mol}} = 0.005332846\ \text{mol}\ Ba_3(PO_4)_2$

② $n_{BaCl_2} = 0.005332846\ \cancel{\text{mol}\ Ba_3(PO_4)_2} \times \dfrac{3\ \text{mol}\ BaCl_2}{1\ \cancel{\text{mol}\ Ba_3(PO_4)_2}} = 0.015998538\ \text{mol}\ BaCl_2$

③ $m_{BaCl_2} = nM = (0.015998538\ \cancel{\text{mol}}\ BaCl_2)(208.23\ \text{g/}\cancel{\text{mol}})$

$\boldsymbol{m_{BaCl_2} = 3.33\ \text{g}}$

Assignment

9.2A pg. 301–302 #7 to 16 and Worksheet: Gravimetric Stoichiometry

9.2A Worksheet: Gravimetric Stoichiometry

1. Liquid hexane (C_6H_{14}) undergoes combustion. Determine the mole ratio of:
a. hexane to oxygen **b.** oxygen to carbon dioxide **c.** oxygen to water vapour
d. hexane to carbon dioxide **e.** hexane to water vapour

2. Magnesium reacts with hydrochloric acid according to the following balanced chemical equation:

$$Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$$

If two moles of hydrochloric acid react with excess magnesium, how many moles of hydrogen gas will be produced?

3. Calcium carbonate combines with hydrochloric acid, $HCl_{(aq)}$, to produce calcium chloride, water, and carbon dioxide gas.

$$CaCO_{3\,(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_2O_{(l)} + CO_{2\,(g)}$$

a. How many moles of $HCl_{(aq)}$ are required to react with 2.5 mol of calcium carbonate?
b. How many moles of carbon dioxide would be produced if 2.5 mol of calcium carbonate is used?

4. Aluminum reacts with hydrochloric acid, $HCl_{(aq)}$, to produce aluminum chloride and hydrogen gas.
a. Write a balanced equation for the reaction.
b. Calculate the number of moles of $HCl_{(aq)}$ required to react with 0.87 mol of Al.

5. Glucose ($C_6H_{12}O_6$) combines with O_2 in the body to produce carbon dioxide and water.
a. Write a balanced equation for this reaction.
b. How many moles of O_2 are required to combine with 0.25 mol of glucose?
c. How many moles of CO_2 and H_2O would be produced if 0.25 mol of glucose is reacted?

6. Zinc metal reacts readily with nitric acid, $HNO_{3\,(aq)}$ to produce a flammable gas. Identify this gas, and calculate the moles of gas produced if 0.36 mol of zinc react with an excess amount of $HNO_{3\,(aq)}$.

7. One way to produce methane gas is to react aluminum carbide with water in the following reaction.

$$Al_4C_{3\,(s)} + 12\ H_2O_{(g)} \rightarrow 4\ Al(OH)_{3\,(s)} + 3\ CH_{4\,(g)}$$

Calculate the mass of methane gas produced when 1.73 mol of aluminum carbide completely react with water?

8. Glucose can quickly oxidized using sodium chlorite to produce carbon dioxide, water vapour, and salt.

$$C_6H_{12}O_{6\,(s)} + 6\ NaClO_{2\,(l)} \rightarrow 6\ CO_{2\,(g)} + 6\ H_2O_{(g)} + 6\ NaCl_{(s)}$$

If 323 grams of carbon dioxide came from the above reaction, how many moles of glucose reacted?

9. What mass of hydrogen peroxide (H_2O_2) must decompose to produce 0.77 g of water?

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

10. Determine the mass of calcium hydroxide produced when excess calcium carbide (CaC_2) reacts with 0.640 kg of water according to the following balanced chemical equation:

$$CaC_{2\,(s)} + 2\ H_2O_{(l)} \rightarrow Ca(OH)_{2\,(aq)} + C_2H_{2\,(g)}$$

11. Determine the mass of lithium hydroxide produced when 0.38 g of lithium nitride reacts with water according to the following **unbalanced** chemical equation:

$$Li_3N_{(s)} + H_2O_{(l)} \rightarrow NH_{3\,(g)} + LiOH_{(aq)}$$

12. Determine the mass of antimony produced when 0.46 g of antimony (III) oxide reacts with excess carbon according to the following **unbalanced** equation:

$$Sb_2O_{3\,(s)} + C_{(s)} \rightarrow Sb_{(s)} + CO_{(g)}$$

13. 400 g of ammonia gas is combined with excess oxygen to produce nitrogen monoxide gas and steam.
- **a.** Write a balanced equation for this reaction.
- **b.** Calculate the mass of oxygen used.
- **c.** Determine the mass of water vapour produced.

14. What mass of sodium chloride is produced when 0.29 g of chlorine gas reacts with excess sodium iodide?

15. Determine the mass of carbon dioxide produced when 0.85 g of butane (C_4H_{10}) is burned.

16. In the *combustion* of carbon monoxide to produce a single product, what mass of $CO_{(g)}$ is required to produce 0.69 g of $CO_{2\,(g)}$?

17. Determine the mass of the precipitate produced when 0.73 g of sodium hydroxide reacts with a concentrated nickel (II) nitrate solution.

18. Iron (II) sulfide is produced from its elements. Determine the mass of the product formed when 16.0 g of sulfur is used in the reaction.

19. Copper metal is placed in a concentrated silver nitrate solution. How much metal copper is needed in order to produce 90.0 g of precipitate?

20. Glucose is produced via the process of photosynthesis. An average human requires 120 g of glucose everyday. How much carbon dioxide is require to produced this much glucose?

Answers

1a. $\dfrac{2\text{ mol }C_6H_{14}}{19\text{ mol }O_2}$ **b.** $\dfrac{19\text{ mol }O_2}{12\text{ mol }CO_2}$ **c.** $\dfrac{19\text{ mol }O_2}{14\text{ mol }H_2O}$ **d.** $\dfrac{2\text{ mol }C_6H_{14}}{12\text{ mol }CO_2} = \dfrac{1\text{ mol }C_6H_{14}}{6\text{ mol }CO_2}$

e. $\dfrac{2\text{ mol }C_6H_{14}}{14\text{ mol }H_2O} = \dfrac{1\text{ mol }C_6H_{14}}{7\text{ mol }H_2O}$ **2.** 1 mol $H_{2\,(g)}$ **3a.** 5.0 mol HCl **b.** 2.5 mol CO_2

4a. $2\,Al_{(s)} + 6\,HCl_{(aq)} \rightarrow 2\,AlCl_{3\,(aq)} + 3\,H_{2\,(g)}$ **b.** 2.61 mol of $HCl_{(aq)}$

5a. $C_6H_{12}O_{6\,(s)} + 6\,O_{2\,(g)} \rightarrow 6\,CO_{2\,(g)} + 6\,H_2O_{(l)}$ **b.** 1.50 mol of O_2 **c.** 1.50 mol CO_2; 1.5 mol H_2O

6. 0.36 mol H_2 **7.** 83.3 g CH_4 **8.** 1.22 mol $C_6H_{12}O_6$ **9.** 1.45 g H_2O_2

10. 1.32 kg $Ca(OH)_2$ **11.** 0.78 g LiOH **12.** 0.38 g Sb

13a. $4\,NH_{3\,(g)} + 5\,O_{2\,(g)} \rightarrow 4\,NO_{(g)} + 6\,H_2O_{(g)}$ **b.** 939 g O_2 **c.** 635 g H_2O

14. 0.48 g NaCl **15.** 2.57 g CO_2 **16.** 0.44 g CO **17.** 0.85 g $Ni(OH)_{2\,(s)}$

18. 43.9 g FeS **19.** 26.5 g Cu **20.** 176 g CO_2

9.2B: Solution Stoichiometry and Calculations Involving Limiting Reagents

Steps to Solve a Neutralization Reaction:

1. Write a balanced molecular equation.
2. Put the given information underneath the proper chemicals.
3. Using $n = CV$, convert the given information to moles.
4. Determine the moles of the required chemical by using the mole ratio $\left(\dfrac{\text{Require Coefficient}}{\text{Given Coefficient}}\right)$.
5. Covert moles of the required chemical to concentration or volume $\left(C = \dfrac{n}{V} \text{ or } V = \dfrac{n}{C}\right)$.

Example 1: 21.50 mL of 0.0350 mol/L of potassium hydroxide solution is used to neutralize 10.0 mL of sulfuric acid, $H_2SO_{4\,(aq)}$. Determine the molar concentration of sulfuric acid.

$2\,KOH_{(aq)}$	+	$H_2SO_{4\,(aq)}$	→	$2\,HOH_{(l)}$	+	$K_2SO_{4\,(aq)}$
21.50 mL		10.0 mL				
0.0350 mol/L		? mol/L				

① $n_{KOH} = CV = (0.0350 \text{ mol/}\cancel{\text{L}})(21.50 \text{ m}\cancel{\text{L}}) = 0.7525 \text{ mmol}$

② $n_{H_2SO_4} = 0.7525 \text{ m}\cancel{\text{mol KOH}} \times \dfrac{1 \text{ mol } H_2SO_4}{2\,\cancel{\text{mol KOH}}} = 0.37625 \text{ mmol } H_2SO_4$

③ $C_{H_2SO_4} = \dfrac{n}{V} = \dfrac{0.37625\,\cancel{\text{m}}\text{mol}}{10.0\,\cancel{\text{m}}\text{L}} = 0.037625 \text{ mol/L}$

$[H_2SO_4] = 0.0376$ mol/L

Excess: - the reactant with more than enough amount for the reaction.

Limiting Reagent: - the reactant with the smaller amount (after taken account of the mole ratio) for the reaction.

Note: A limiting reagent question will always have enough information to find the moles of both reactants.

Steps to deal with Limiting Reagent Problems:

1. Assume one of the reactants is the limiting reagent and determine its mole amount.
2. Determine the mole amount of the other reactant.
3. Use the mole amount of the assumed limiting reagent and the mole ratio, calculate the mole amount of the other reactant actually needed.
4. If the mole amount of the other reactant is smaller than what is needed, then our assumption was wrong. The other reactant is the limiting reagent.
5. If the mole amount of the other reactant is bigger than what is needed, then our assumption was correct. It means that the other reactant is the excess.

Example 2: 5.00 g of phosphorus is reacted with 15.00 g of chlorine gas to produce phosphorus trichloride. Determine the mass of the product produced.

$P_{4\,(s)}$	+	$6\,Cl_{2\,(g)}$	$\rightarrow$	$4\,PCl_{3\,(s)}$
5.00 g		15.00 g		? g
M = 123.88 g/mol		M = 70.90 g/mol		M = 137.32 g/mol

Since there is enough information to determine the moles of two reactants, we need to determine which one is the limiting reagent.

① $n_{P_4} = \dfrac{m}{M} = \dfrac{5.00\ \cancel{g}}{123.88\ \cancel{g}/mol} = 0.04036\ldots$ mol P_4 ② $n_{Cl_2} = \dfrac{m}{M} = \dfrac{15.00\ \cancel{g}}{70.90\ \cancel{g}/mol} = 0.2115\ldots$ mol Cl_2

Let's assume P_4 is the limiting reagent. Calculate the mol Cl_2 actually needed.

③ $n_{Cl_2} = 0.04036\ \cancel{mol\ P_4} \times \dfrac{6\ mol\ Cl_2}{1\ \cancel{mol\ P_4}} = 0.24216\ldots$mol Cl_2 needed

But we don't have 0.2416 mol of Cl_2, we only have 0.2115… mol of Cl_2. Therefore, Cl_2 is the limiting reagent. (*Note: the limiting reagent is NOT always the chemical with the smaller number of moles. You have to always compare like we did above.)*

Now, we calculate the moles of PCl_3 formed by using moles of limiting reagent Cl_2.

④ $n_{PCl_3} = 0.2115655853\ \cancel{mol\ Cl_2} \times \dfrac{4\ mol\ PCl_3}{6\ \cancel{mol\ Cl_2}} = 0.1410437236$ mol PCl_3

Finally, we determine the mass of PCl_3 produced.

⑤ $m_{PCl_3} = nM = (0.1410437236\ \cancel{mol}\ PCl_3)(137.32\ g/\cancel{mol}) = 19.36812412$ g $\boldsymbol{m_{PCl_3} = 19.4\ g}$

Steps to Solve a Precipitation Reaction in an Aqueous Environment:

1. Write a balanced molecular equation. Identify the precipitate.
2. Put the given information underneath the proper chemicals. Identify the limiting reagent if any.
3. Using $n = CV$, convert all given information to moles.
4. Identify and use the information of the limiting reagent if necessary.
5. Determine the moles of precipitate form by using the mole ratio $\left(\dfrac{\text{Require Coefficient}}{\text{Given Coefficient}}\right)$.
6. Covert moles of precipitate to mass ($m = nM$).

Example 3: A 30.0 mL of 0.200 M of lead (II) nitrate solution is reacted with 20.0 mL of 0.150 mol/L of sodium sulfate solution. What is the mass of the precipitate formed at the end of this reaction?

$Pb(NO_3)_{2\,(aq)}$	+	$Na_2SO_{4\,(aq)}$	$\rightarrow$	$PbSO_{4\,(s)}$	+	$2\,NaNO_{3\,(aq)}$
30.0 mL		20.0 mL		? g		
0.200 mol/L		0.150 mol/L		M = 303.28 g/mol		

Since there is enough information to determine the moles of two reactants, we need to determine which one is the limiting reagent.

① $n_{Pb(NO_3)_2} = CV = (0.200\ \text{mol}/\cancel{\text{L}})(30.0\ \text{m}\cancel{\text{L}}) = 6\ \text{mmol}$

② $n_{Na_2SO_4} = CV = (0.150\ \text{mol}/\cancel{\text{L}})(20.0\ \text{m}\cancel{\text{L}}) = 3\ \text{mmol}$

Since the two reactants are of 1:1 ratio, the chemical wth the smaller moles becomes the limiting reagent. ***(Note: Again, the limiting reagent is NOT always the chemical with the smaller number of moles. We don't have to do the "required over given" to compare only because they are of 1:1 ratio.)***

③ Limiting Reagent is Na_2SO_4

Now, we calculate the moles of $PbSO_4$ formed by using moles of limiting reagent Na_2SO_4.

④ $n_{PbSO_4} = 3\ \text{m}\cancel{\text{mol}\ Na_2SO_4} \times \dfrac{1\ \text{mol}\ PbSO_4}{1\ \cancel{\text{mol}\ Na_2SO_4}} = 3\ \text{mmol}\ PbSO_4$

Finally, we determine the mass of $PbSO_4$ produced.

⑤ $m_{PbSO_4} = nM = (3\ \text{m}\cancel{\text{mol}}\ PbSO_4)(303.28\ \text{g}/\cancel{\text{mol}}) = 909.84\ \text{mg} = 0.90984\ \text{g}$

$$m_{PbSO_4} = \mathbf{0.910\ g}$$

Assignment

9.2B Worksheet: Solution Stoichiometry and Limiting Reagent Worksheet

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1. Determine the mass of aluminum required to completely react with 40.0 mL of 2.00 mol/L hydrochloric acid ($HCl_{(aq)}$).
2. What is the volume of a 3.00 mol/L phosphoric acid, $H_3PO_{4\,(aq)}$, is required to react 7.50 g of zinc?
3. How many grams of sodium can be reacted with 250 mL of a 5.00 mol/L solution of sulfuric acid, $H_2SO_{4\,(aq)}$?
4. Find the minimum volume of a 0.150 mol/L lead (II) nitrate solution needed to completely react with 120 mL of 0.200 mol/L of sodium iodide solution.
5. How many Litres of a 0.850 mol/L of calcium nitrate solution will be required to react with 130 g of sodium carbonate powder?
6. If 50.0 mL of a 2.50 mol/L silver nitrate solution is added to excess potassium chloride solution, what is the mass of precipitate formed?
7. What is the mass of precipitate form when excess sodium sulfate reacts with 30.0 mL of 0.400 mol/L of barium nitrate solution?
8. If 425 mL of 0.800 mol/L of hydrochloric acid, $HCl_{(aq)}$, is neutralized with 300 mL of strontium hydroxide solution, determine the molar concentration of the strontium hydroxide solution?
9. What mass of water will be produced if 4.25 g of hydrogen gas reacts with 43.23 g of oxygen gas?
10. A mass of 25.0 g of propane gas is burned in the presence of 13.50 g of oxygen gas.
    - **a.** Determine the limiting reagent.
    - **b.** Determine the mass of carbon dioxide produced.
    - **c.** Find the mass of H_2O produced.
11. A neutralization reaction is performed between 125 g of solid sodium hydroxide with 300 mL of 10.5 mol/L nitric acid, $HNO_{3\,(aq)}$.
    - **a.** Determine the limiting reagent.
    - **b.** Determine the mass of water produced.
12. Determine the mass of precipitate formed when 60.0 mL of a 2.50 mol/L sodium sulfate solution and 80.0 mL of a 3.00 mol/L silver nitrate solution are mixed together.
13. Solid calcium hydroxide is used to neutralize acid spills. Suppose 23.0 g of calcium hydroxide powder is used to neutralize 15.0 mL of 1.50 mol/L of phosphoric acid.
    - **a.** What is the mass of precipitate formed.
    - **b.** Was there enough calcium hydroxide to completely neutralize the acid? Explain.
14. Acids react with metal to form hydrogen gas. Find the mass of hydrogen gas formed when 4.50 g of zinc reacts with 30.0 mL of 3.50 mol/L of hydrobromic acid, $HBr_{(aq)}$.
15. A 65.0 mL of 0.300 mol/L barium sulfide solution is reacted with 45.0 mL of 0.600 mol/L of lithium sulfite solution. Calculate the mass of the precipitate formed.

Answers

1. 0.719 g	**2.** 25.5 mL	**3.** 57.5 g	**4.** 80.0 mL	**5.** 1.84 L	**6.** 17.9 g	**7.** 2.80 g
8. 0.567 M	**9.** 37.9 g	**10a.** $O_{2\,(g)}$	**b.** 11.14 g	**c.** 6.08 g	**11a.** NaOH	**b.** 56.3 g
12. 37.4 g	**13a** 32.09 g	**b.** No, since $H_3PO_{4\,(aq)}$ is the excess reagent, some acid would be left over.				
14. 0.106 g	**14.** 4.24 g					

9.2C: The Properties of Gases and Gas Laws

Properties of Gases:

1. **Compressibility**: - a main property of gas where the amount of volume can decrease under increase pressure.

2. **Lack of Particle Interaction**: - Unlike liquids and solids, gas particles have no attractive and repulsive forces between them as assumed in the kinetic theory.

3. **Rapid and Constant Motion**: - gas particles move in a straight line and independent of each other.
   - at collisions, gas particles are completely **elastic** (total kinetic energy remains constant before and after the collision).

Variables to Describe a Gas:

1. **Pressure (*P*)**: - the amount of force per unit of area, measures in **kiloPascal (kPa)**.
   - in a pressurized container, a pressure can be felt as the particles pushed on the inside wall of the container.

2. **Volume (*V*)**: - the amount of space the gas is occupied; measures in **Litre (L)**.

3. **Temperature (*T*)**: - the average of kinetic energy of the gas; measures in **Kelvin (K)**.

4. **Moles (*n*)**: - the amount of gas particle in a closed system; measures in **moles (mol)**.

Factors Affecting Gas Pressure

1. **Amount of Gas**: as amount of gas particles increases, there are more particles pushing on the inside wall of the container, causing an increase in pressure.

   Amount of Gas ↑ **Pressure ↑**

2. **Volume**: - as the volume decreases, there is less room for the gas particles to move about. This causes the particles to push "harder" onto the inside wall of the container, causing an increase in pressure.

   Volume ↑ **Pressure ↓**

3. **Temperature**: - as the temperature increases, gas particles move faster and push harder on the inside wall of the container, causing an increase in pressure.

   Temperature ↑ **Pressure ↑**

The Gas Laws

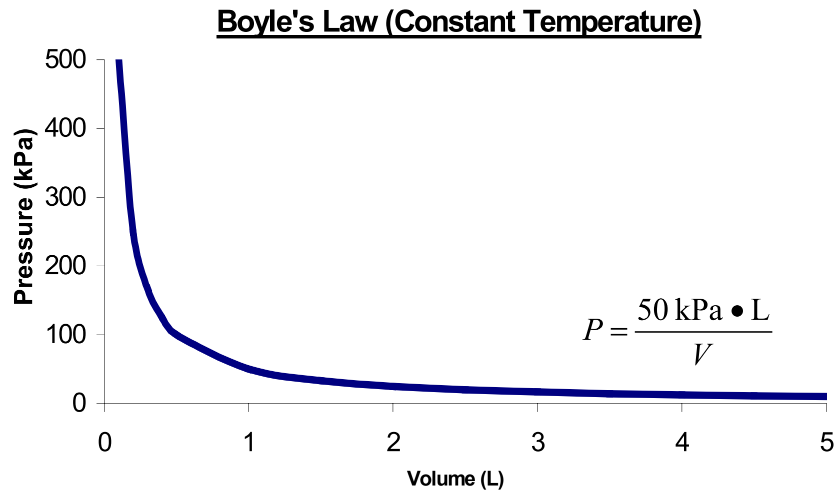
1. **Boyle's Law**: - at a constant temperature, **pressure is inversely proportional to the volume**.

$$P \propto \frac{1}{V}$$

$$P = \frac{k}{V} \quad \text{or} \quad PV = k$$

k = constant of proportion

Volume ↑ Pressure ↓



Boyle's Law

$$P_1V_1 = P_2V_2$$

P_1 = Pressure of Initial Condition **V_1 = Volume of Initial Condition**
P_2 = Pressure of Final Condition **V_2 = Volume of Final Condition**

Example 1: A gas cylinder changed its volume from 2.50 L to 6.25 L. If it were at 101.325 kPa initially, what would be its final pressure?

P_1 = 101.325 kPa
P_2 = ?
V_1 = 2.50 L
V_2 = 6.25 L

$$P_1V_1 = P_2V_2 \qquad \frac{P_1V_1}{V_2} = P_2$$

$$P_2 = \frac{(101.325\text{ kPa})(2.50\cancel{\text{L}})}{(6.25\cancel{\text{L}})}$$

$$P_2 = 40.5\text{ kPa}$$

As Volume ↑, Pressure ↓

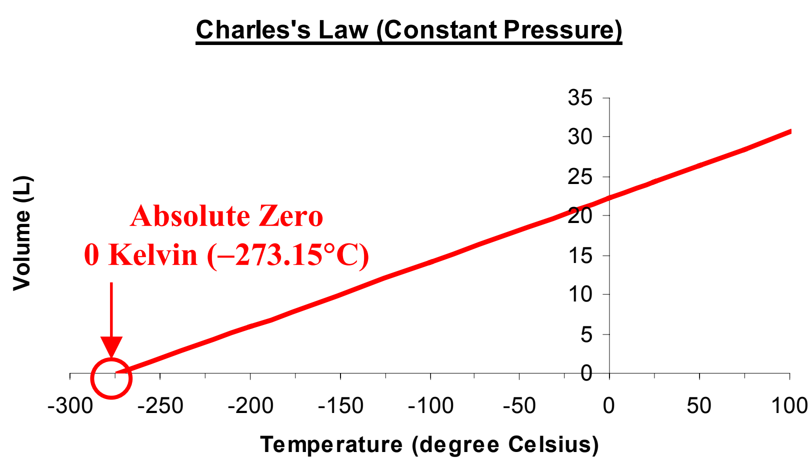
2. **Charles's Law**: - at constant pressure, **volume is directly proportional to the temperature**.

$$V \propto T$$

$$V = kT \quad \text{or} \quad \frac{V}{T} = k$$

k = constant of proportion

Temperature ↑ Volume ↑



Charles's Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

T_1 = Temperature of Initial Condition **V_1 = Volume of Initial Condition**
T_2 = Temperature of Final Condition **V_2 = Volume of Final Condition**

Example 2: A balloon is has a volume of 3.25 L at 25.0°C. Determine the volume of the same balloon when the temperature is dropped to 5.00°C.

V_1 = 3.25 L
T_1 = 25.0°C = 298.15 K
V_2 = ?
T_2 = 5.00°C = 278.15 K

(Change °C to K)

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad \frac{V_1 T_2}{T_1} = V_2$$

As Temp ↓, Volume ↓

$$V_2 = \frac{(3.25\ \text{L})(278.15\ \cancel{\text{K}})}{(298.15\ \cancel{\text{K}})}$$

$$V_2 = 3.03\ \text{L}$$

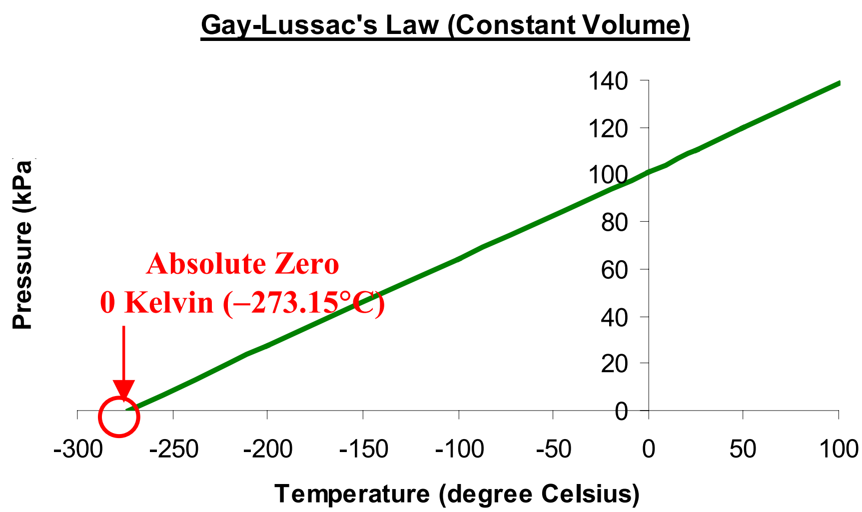
3. **Gay-Lussac's Law**: - at constant volume, **pressure is directly proportional to the temperature**.

$$P \propto T$$

$$P = kT \quad \text{or} \quad \frac{P}{T} = k$$

k = constant of proportion

Temperature ↑ Pressure ↑



Gay-Lussac's Law

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

T_1 = Temperature of Initial Condition **P_1 = Pressure of Initial Condition**
T_2 = Temperature of Final Condition **P_2 = Pressure of Final Condition**

Example 3: A canister is has a pressure of 8.00 atm at 15.0°C. Calculate its pressure if the temperature was to increase to 100.0°C. **(Don't do this at home!)**

P_1 = 8.00 atm
T_1 = 15.0°C = 288.15 K
P_2 = ?
T_2 = 100.0°C = 373.15 K

(Change °C to K)
(P can be in atm)

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \qquad \frac{P_1 T_2}{T_1} = P_2$$

As Temp ↑, Pressure ↑

$$P_2 = \frac{(8.00\ \text{atm})(373.15\ \cancel{\text{K}})}{(288.15\ \cancel{\text{K}})}$$

$$P_2 = 10.4\ \text{atm}$$

4. **Combined Gas Law**: - a formula that summarizes Boyle's Charles's and Guy-Lussac's Gas Laws.
   - allow the user of the formula to determine the change in conditions of the same amount of gas.

Combined Gas Law

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

P_1 = Pressure of Initial Condition	**P_2 = Pressure of Final Condition**
V_1 = Volume of Initial Condition	**V_2 = Volume of Final Condition**
T_1 = Temperature of Initial Condition	**T_2 = Temperature of Final Condition**

Example 4: A high altitude weather balloon has a volume of 57.2 L at SATP. Determine its volume at its maximum height where the pressure is 37.0 kPa and the temperature is −45.0°C.

V_1 = 57.2 L

At SATP:
P_1 = 100 kPa
T_1 = 25.0°C = 298.15 K

V_2 = ?
P_2 = 37.0 kPa
T_2 = −45.0°C = 228.15 K

(Change °C to K)

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{P_1V_1T_2}{T_1P_2} = V_2$$

$$V_2 = \frac{(100\ \cancel{\text{kPa}})(57.2\ \text{L})(228.15\ \cancel{\text{K}})}{(298.15\ \cancel{\text{K}})(37.0\ \cancel{\text{kPa}})} = 118.2986978\ \text{L}$$

$$V_2 = 118\ \text{L}$$

Because P↓ much more than T↓, the final Volume ↑.

Mole-Volume Relationship

1. **Standard Temperature and Pressure (STP)**: - the amount of any gas at 0°C and 101.325 kPa (Earth's atmospheric pressure at sea level).
2. **Standard Ambient Temperature and Pressure (SATP)**: - the amount of any gas at 25°C and 100 kPa.

STP = 22.4 L/mol @ 0°C and 101.325 kPa (1 atm)
SATP = 24.8 L/mol @ 25°C and 100 kPa

Note: The amount of gas is determined by temperature, pressure and volume. The type of gas particles have no effect on these variables. (*Avogadro's Hypothesis)*

Example 5: Determine the amount of oxygen gas in a 5.00 L container under STP and SATP.

a. STP

STP = 22.4 L/mol

$$n = 5.00\ \cancel{\text{L}} \times \frac{1\ \text{mol}}{22.4\ \cancel{\text{L}}}$$

$$n = 0.223\ \text{mol}$$

b. SATP

SATP = 24.8 L/mol

$$n = 5.00\ \cancel{\text{L}} \times \frac{1\ \text{mol}}{24.8\ \cancel{\text{L}}}$$

$$n = 0.202\ \text{mol}$$

Example 6: Determine the volume of 3.50 g of nitrogen gas under STP and SATP.

a. STP

$$n = \frac{m}{M} = \frac{3.50\ \cancel{g}}{28.014\ \cancel{g}/\text{mol}} = 0.1249375312\ \text{mol}$$

STP = 22.4 L/mol

$V = (0.1249375312\ \cancel{\text{mol}})(22.4\ \text{L}/\cancel{\text{mol}})$

$V = 2.80\ \text{L}$

b. SATP

$$n = \frac{m}{M} = \frac{3.50\ \cancel{g}}{28.014\ \cancel{g}/\text{mol}} = 0.1249375312\ \text{mol}$$

SATP = 24.8 L/mol

$V = (0.1249375312\ \cancel{\text{mol}})(24.8\ \text{L}/\cancel{\text{mol}})$

$V = 3.10\ \text{L}$

Example 7: Calculate the molar mass of a gaseous compound containing carbon and hydrogen if its density is 0.645 g/L at SATP.

SATP = 24.8 L/mol

$D = 0.645$ g/L

M = ? (g/mol)

$M = D \times (\text{SATP})$

$M = (0.645\ \text{g}/\cancel{\text{L}})\ (24.8\ \cancel{\text{L}}/\text{mol})$

$M = 15.996$ g/mol

$M = 16.0\ \text{g/mol}$

Example 8: Calculate the density in g/L of a sulfur dioxide gas at STP.

STP = 22.4 L/mol

M = 64.06 g/mol (SO_2)

D = ? (g/L)

$M = D \times (\text{STP})$

$$D = \frac{M}{(\text{STP})} = \frac{(64.06\ \text{g}/\cancel{\text{mol}})}{(22.4\ \text{L}/\cancel{\text{mol}})} = 2.859821429\ \text{g/L}$$

$D = 2.86\ \text{g/L}$

Ideal Gas Law: - a formula that relates pressure, volume, amount, and temperature of an **ideal gas** (gaseous volume does not account for total particles volumes) at one specific condition.

Ideal Gas Law

$$PV = nRT$$

P = Pressure (kPa) **V = Volume (L)**

n = Amount of Gas (mol) **T = Temperature (K)**

R = Gas Constant = 8.314 (L • kPa)/(K • mol)

Example 9: Determine the mass of propane if it is in a 200 L container at 15.0°C and at 32.0 atm.

V = 200 L

T = 15.0°C = 288.15 K

$$P = 32.0\ \cancel{\text{atm}} \times \frac{101.325\ \text{kPa}}{1\ \cancel{\text{atm}}}$$

P = 3242.4 kPa

[need to change atm to kPa because R is in (L • kPa)/(K • mol)]

R = 8.314 (L • kPa)/(K • mol)

m = ?

n = ? (need to find n first)

$PV = nRT$

$$\frac{PV}{RT} = n$$

$$n = \frac{(3242.4\ \cancel{\text{kPa}})(200\ \cancel{\text{L}})}{\left(8.314\ \frac{\cancel{\text{L}} \bullet \cancel{\text{kPa}}}{\cancel{\text{K}} \bullet \text{mol}}\right)(288.15\ \cancel{\text{K}})}$$

n = 270.670642 mol

For propane, C_3H_8, M = 44.11 g/mol

$m = nM = (270.670642\ \cancel{\text{mol}})(44.11\ \text{g}/\cancel{\text{mol}})$

m = 11939.28202 g

$m = 1.19 \times 10^4\ \text{g} = 11.9\ \text{kg}$

Example 10: It is known that air is less dense at higher altitude. Calculate the density of pure oxygen gas near the top of Mount Everest where the temperature is −30.0°C and 31.0 kPa. (By comparison, pure oxygen has a density of 1.29 g/L at SATP.)

T = −30.0°C = 243.15 K
P = 31.0 kPa
R = 8.314 (L • kPa)/(K • mol)
M = 32.00 g/mol

$D = \dfrac{m}{V} = ?$

Substitute $\dfrac{m}{M}$ for n.

Then, solve for $\dfrac{m}{V}$.

$$PV = nRT$$

$$PV = \frac{m}{M}RT$$

$$\frac{PM}{RT} = \frac{m}{V}$$

$$D = \frac{PM}{RT} = \frac{(31.0\ \cancel{\text{kPa}})(32.0\ \text{g}/\cancel{\text{mol}})}{\left(8.314\dfrac{\text{L}\bullet\cancel{\text{kPa}}}{\cancel{\text{K}}\bullet\cancel{\text{mol}}}\right)(243.15\ \cancel{\text{K}})} = 0.4907127905\ \text{g/L}$$

$$D = 0.491\ \text{g/L}$$

Assignment
9.2C Worksheet: Gas Laws Worksheet

9.2C Worksheet: Gas Laws Worksheet

1. A balloon filled with helium gas has a volume of 600 mL at a pressure of 1.5 atm. The balloon is released and reaches and altitude of 7.0 km where the pressure is 0.4 atm. Assuming that the temperature remained the same the entire time, what volume does the gas occupy at this height?
2. At a given pressure, a sample of sulfur dioxide gas has a volume of 36.0 L when the temperature is 30.0°C. What will be its volume if the temperature is raised to 75.0°C when pressure is kept constant?
3. At 18.0°C, a sample of hydrogen gas has a volume of 48.0 L. To what temperature must this as be heated to change the volume to 74.0 L is pressure was not change?
4. Before a long road trip, the pressure in the automobile tire is 210 kPa at 22.0°C. At the end of the trip. the pressure gauge reads 230 kPa. Assuming constant volume, what is the new temperature of the air inside the tire?
5. A sample of gas has a pressure of 190 kPa and occupies a volume of 8.50 L. If the gas is compressed to a volume of 3.25 L, what will its pressure be, assuming constant temperature?
6. A flask of helium gas has a pressure of 1.85 atm at 28.0°C. At what temperature should the balloon be lowered to so the pressure will decrease to 0.82 atm if the volume is kept constant?
7. A sample of oxygen gas has a volume of 40.0 L at 75.0 kPa. Suppose temperature remain constant throughout, what will the volume of the oxygen gas be at a pressure of 30 kPa?
8. A sample of nitrogen gas has a volume of 50.0 L at STP. What will the new pressure be if it is allowed to expand to 67.0 L if the temperature stayed the same?
9. A sample of carbon dioxide gas at 110°C has a volume of 35.0 L. What is the volume of this gas at −30.0°C if pressure remained the same?
10. At 130°C, the pressure of a sample of air is 1.10 atm. At 500°C, what will be the new pressure if the volume remains the same?
11. A helium balloon has a volume of 1.75 L at 35°C and 120 kPa. What is the new volume if it's pressure is increased to 150 kPa and the temperature is lowered to 10.0°C?

12. A sample of argon gas has a volume of 300 mL at 30.0°C and 1.50 atm. What will be the resulting temperature is the gas is allowed to expand to 750 mL and pressure lowered to 0.725 atm?

13. A sample of methane gas has a volume of 80.0 L at 45.0°C and 200 kPa. What will be its final volume if the gas is heated to 550°C with the volume expanded to 850 kPa?

14. Determine the volume for the following gases at STP and SATP
a. 0.80 mol of carbon dioxide **b.** 1.35 mol of sulfur trioxide
c. 6.20 mol of carbon monoxide **d.** 3.25 g of propane
e. 18.4 g of methane **f.** 9.42 g of ammonia

15. Calculate the amount and mass of the following gases.
a. 7.35 L of nitrogen monoxide at STP **b.** 34.2 L of sulfur dioxide at SATP
c. 355 mL of fluorine at SATP **d.** 16.5 L of neon at STP

16. Find the molar mass of the following gases under the stated density and condition.
a. 1.82 g/L at STP **b.** 0.378 g/L at SATP **c.** 0.827 g/L at STP **d.** 2.19 g/L at SATP

17. Find the density of the following gas in g/L.
a. nitrogen dioxide at SATP **b.** krypton at STP **c.** oxygen at SATP **d.** nitrogen at STP

18. A sample of hydrogen gas with a volume of 1.63 L is at SATP. What is its final volume when its pressure is 180 kPa and its temperature is raised to 42°C?

19. A sample of air has a volume of 50.0 L at 86.1 kPa at –40.0°C. What will be its volume at STP?

20. Determine the volume of 56.0 g of butane gas, $C_4H_{10\,(g)}$, at 65.0°C and 75.0 kPa.

21. To what temperature must 40.0 g of nitrogen gas be heated at 120 kPa to occupy a volume of 36 L?

22. What is the mass of 37.0 L of carbon dioxide gas at 42.0°C at 138 kPa?

23. A 0.538 g sample of an ideal gas has a volume of 360 mL at 108 kPa and 23.0°C. What is its molar mass? If the gas consists of sulfur and hydrogen, what is the identity of the gas?

24. A 7.50 L container is occupied by 90.0 g of oxygen gas. What is the pressure inside the container if its temperature is 12.0°C?

25. Find the density of nitrogen dioxide gas at a temperature of –10.0°C and a pressure of 75.0 kPa.

26. A 28.4 g of an unknown noble gas has a volume of 5.30 L at 98.0 kPa and 16.0°C. Find its molar mass and identify this noble gas.

27. Determine the density of carbon monoxide gas at a temperature of 37.0°C and a pressure of 120 kPa.

Answers
1. 2.25×10^3 mL or 2.25 L **2.** 41.3 L **3.** 449 K or 176°C **4.** 323 K or 50.1°C
5. 497 kPa **6.** 133 K or –140°C **7.** 100 L **8.** 0.746 atm or 75.6 kPa
9. 22.2 L **10.** 2.11 atm **11.** 1.29 L **12.** 366 K or 93.2°C **13.** 48.7 L
14a. At STP: 17.9 L; At SATP:19.8 L **b.** At STP: 30.2 L; At SATP:33.4 L **c.** At STP: 139 L; At SATP:154 L
d. At STP: 1.65 L; At SATP:1.83 L **e.** At STP: 25.7 L; At SATP:28.4 L **f.** At STP: 12.4 L; At SATP:13.7 L
15a. 0.328 mol; 9.85 g **b.** 1.38 mol; 88.4 g **c.** 0.0143 mol; 0.544 g **d.** 0.734 mol; 14.9 g
16a. 40.8 g/mol **b.** 9.37 g/L **c.** 18.5 g/L **d.** 54.3 g/L
17a. 1.85 g/L **b.** 3.74 g/L **c.** 1.29 g/L **d.** 1.25 g/L
18. 0.957 L or 957 mL **19.** 49.7 L **20.** 22.6 L **21.** 363 K or 90.8°C **22.** 85.8 g
23. 34.1 g/mol; $H_2S_{(g)}$ **24.** 889 kPa **25.** 1.58 g/L **26.** 161 g/mol; $Xe_{(g)}$ **27.** 1.30 g/L

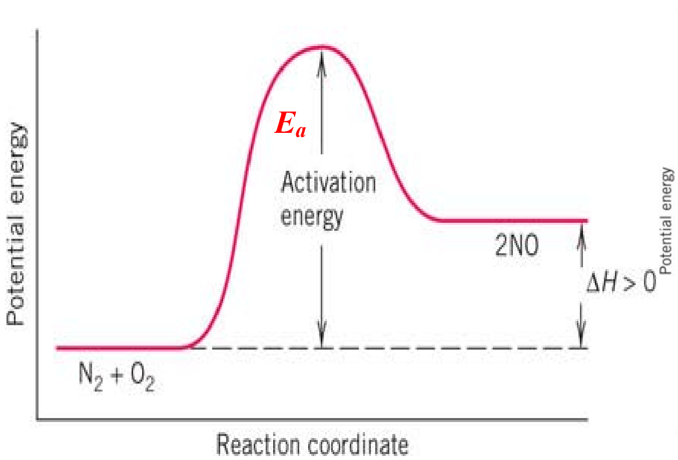
9.3: Reaction Rate is influenced by Concentration and Temperature

Collision Model: - a model that state for a reaction to occur, molecules must collide with each other.

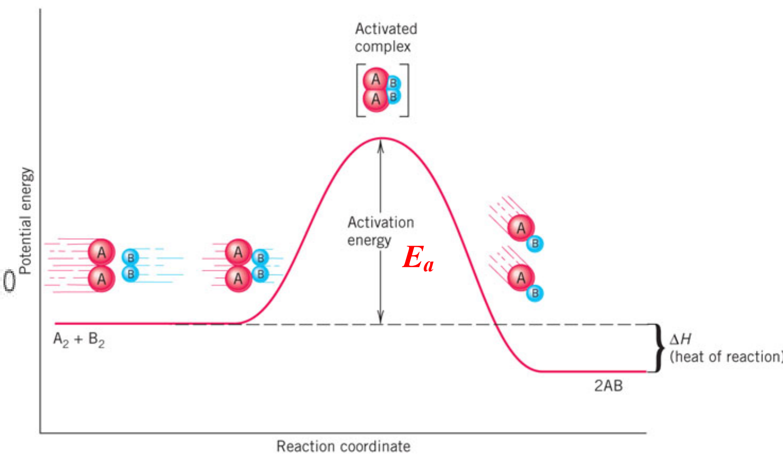
Factors Affecting the Collision Model:

1. **Activation Energy** (**E_a**): - the threshold energy molecules needed to overcome to cause a chemical reaction that was first proposed by Svante Arrhenius.
   - **E_a is the highest energy (top of the hill - E_{max}) minus the sum of energy of the reactants ($\Sigma H_{reactants}$) on the potential energy diagram.**

   Activated Complex: - sometimes refer to as **transition state**. It is the transitional molecule found at the top "hill" of the activation energy.

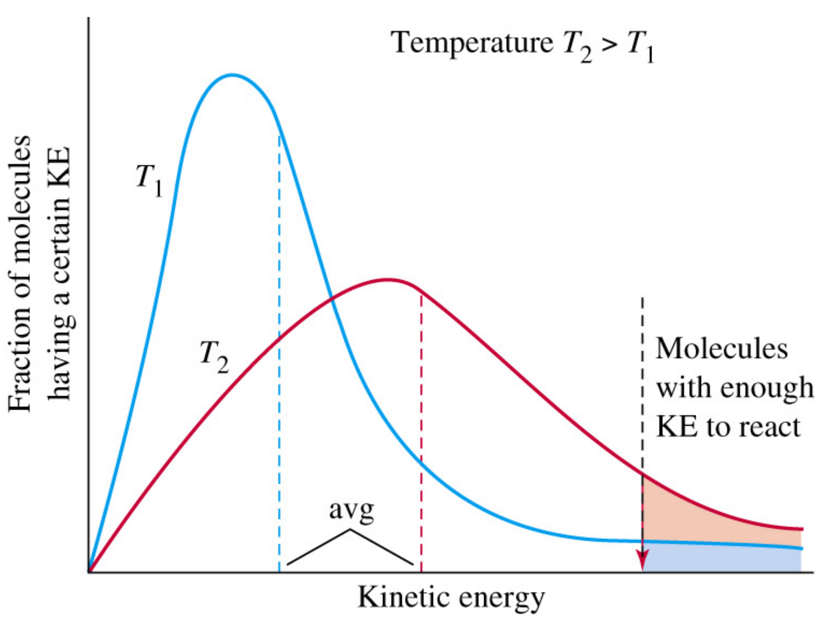


Activation Energy of an Endothermic Reaction



Activation Energy of an Exothermic Reaction

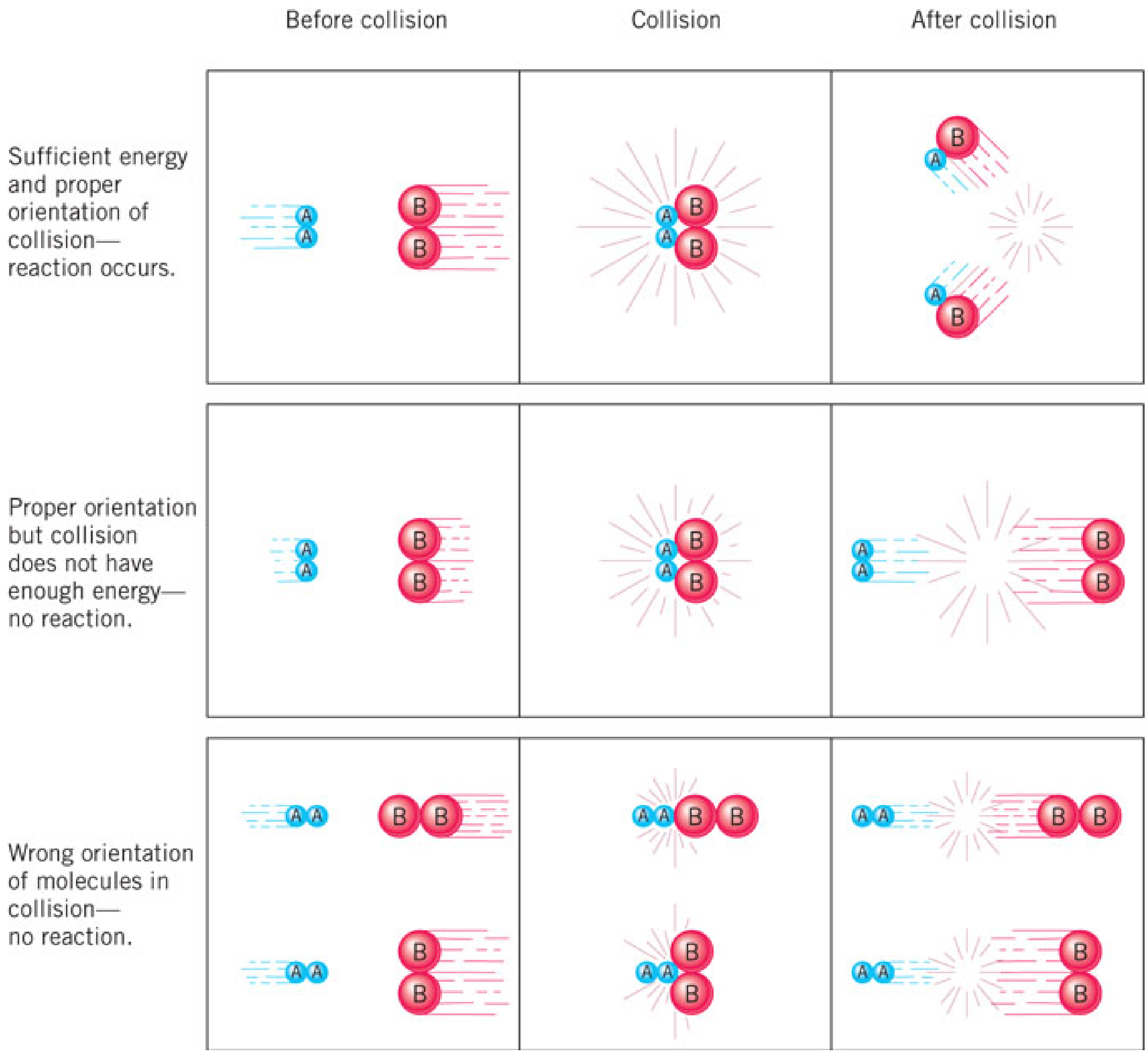
2. **Temperature** (**T**): - the effective number of collisions increases exponentially with temperature.



Because $T_2 > T_1$, there are more molecules colliding with enough Kinetic Energy causing a reaction to occur.

In general, the Higher the Temperature, the Faster the Rate of Reaction.

3. **Molecular Orientation**: - the number of ways molecules collide that will lead to a reaction.



4. **Particle Size**: - the **Smaller the Particle Size (the Larger the Surface Area exposed), the Faster the Reaction Rate.**

Example 1: Grain sugar dissolves faster than equal mass of sugar cubes because of smaller particle size and therefore increased surface area of the grain sugar.

5. **Concentration**: - the **higher the Concentration (the more molecules in an available space, the higher the chance of collision), the Faster the Reaction Rate.**

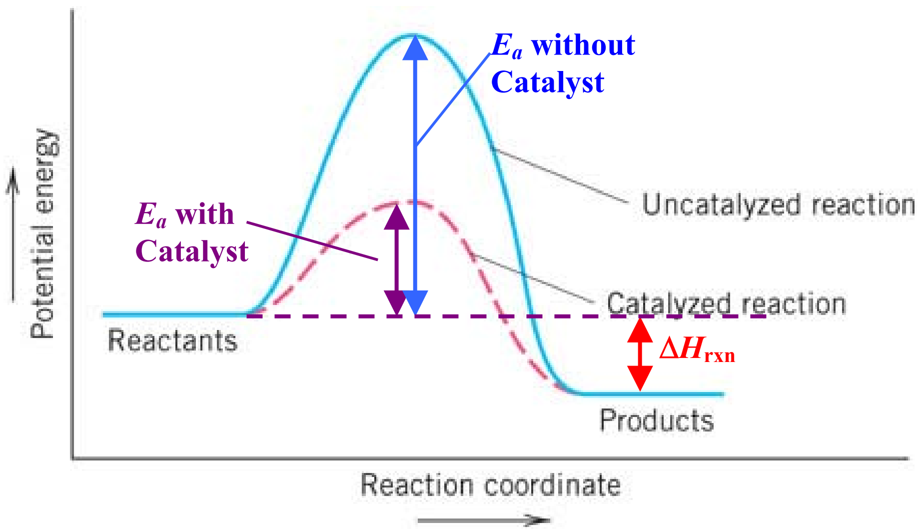
Assignment
9.3 pg. 302 #17 to 22

9.4: Catalysts Increases the Rate of Chemical Reactions

Other Factors Affecting the Rate of Chemical Reactions:

1. **Catalyst**: - a substance that speeds up the reaction **without** being consumed in the reaction.
   - unlike intermediates, catalyst is used and recycled in the reaction.
   - lowers activation energy by providing an alternate reaction pathway.
   (**ΔE_a is lowered but ΔH_{rxn} remains the same** – will be explained in the next section.)
   - in general, **the Addition of a Catalyst INCREASES the Rate of Reaction**.

Example 1: Enzyme is a catalyst in the body that speeds up certain bodily reaction.



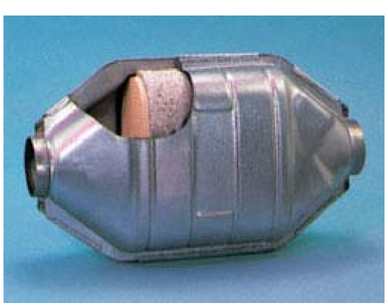
Example 2: Ammonia is formed from its elements using heterogeneous catalyst such as $Pt_{(s)}$:

$$N_{2\,(g)} + 3\,H_{2\,(g)} \xrightarrow{Pt} 2\,NH_{3\,(g)}$$

(Check out Video at http://www.dac.neu.edu/physics/b.maheswaran/phy1121/data/ch11/anim/anim11-5b.mov)

2. **Inhibitor**: - a substance that "inhibit" the function of a catalyst to speed up the reaction.
   - in general, the **Addition of an Inhibitor DECREASES the Reaction Rate**.

Example 3: The catalytic converter converts $NO_{(g)}$ (result of burning nitrogen at high temperature) to $N_{2\,(g)}$ and $O_{2\,(g)}$. The $O_{2\,(g)}$ along with the catalytic converter is used to produce $CO_{2\,(g)}$ from $CO_{(g)}$.

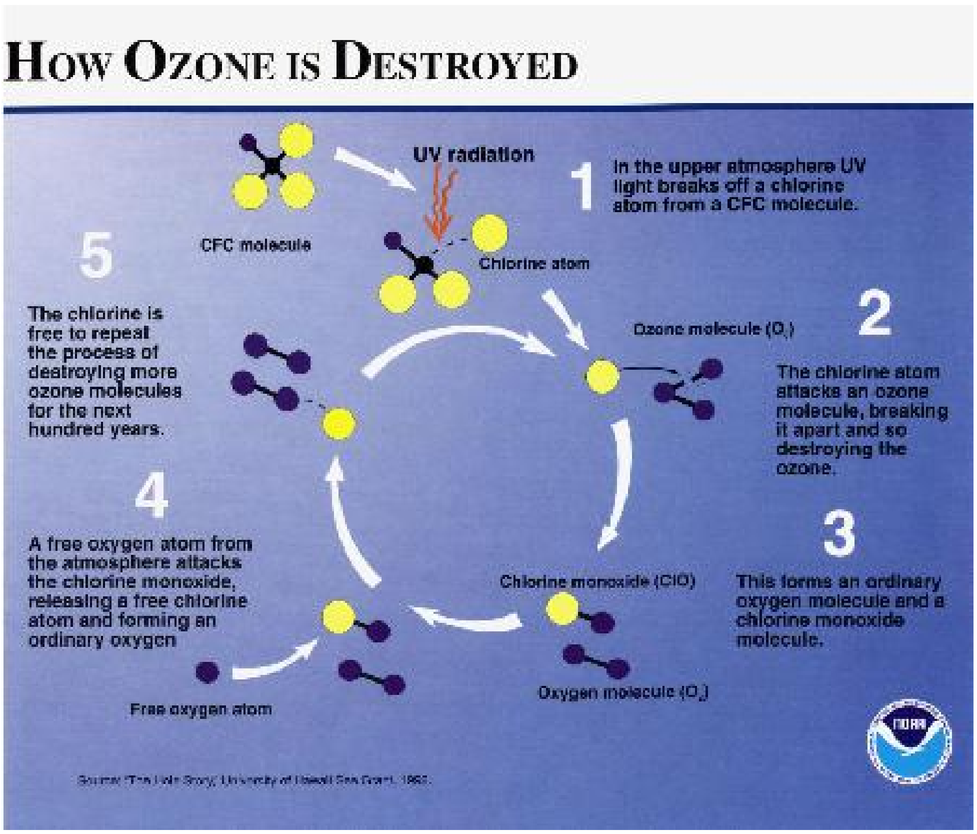


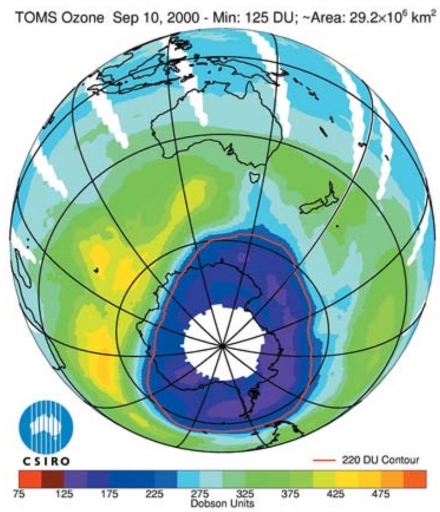
$$2\,NO_{(g)} \xrightarrow{\text{catalytic converter}} N_{2\,(g)} + O_{2\,(g)}$$
$$2\,CO_{(g)} + O_{2\,(g)} \xrightarrow{\text{catalytic converter}} 2\,CO_{2\,(g)}$$

(Left): A catalytic converter for most modern vehicle. Leaded gasoline, an inhibitor, deactivates catalytic converter. Hence, it is not legally used in vehicles

Example 4: The destruction of ozone in upper atmosphere can be attributed to $NO_{(g)}$ and CFCs acting as catalysts. $NO_{(g)}$ is produced from the combustion of $N_{2\,(g)}$ at high temperature commonly found in internal combustion engine (high-altitude aircraft produces lots of NO gas). CFCs (Chloro-Fluoro-Carbon compounds) are found in aerosol can propellants, refrigerators, and air conditioners. They break down to form $Cl_{(g)}$ with the presence of light.

$$\begin{array}{l} NO_{(g)} + O_{3\,(g)} \rightarrow NO_{2\,(g)} + O_{2\,(g)} \\ \underline{O_{(g)} + NO_{2\,(g)} \rightarrow NO_{(g)} + O_{2\,(g)}} \\ O_{(g)} + O_{3\,(g)} \rightarrow 2\,O_{2\,(g)} \end{array} \qquad \begin{array}{l} Cl_{(g)} + O_{3\,(g)} \rightarrow ClO_{(g)} + O_{2\,(g)} \\ \underline{O_{(g)} + ClO_{(g)} \rightarrow Cl_{(g)} + O_{2\,(g)}} \\ O_{(g)} + O_{3\,(g)} \rightarrow 2\,O_{2\,(g)} \end{array}$$





(Above) The Ozone Hole over the South Pole (Sept 2000). A similar hole is present over the Arctic. (Left) Process of Ozone Depletion. Ozone blocks harmful UV rays that can otherwise cause skin cancer.

Factors Affecting Reaction Rate

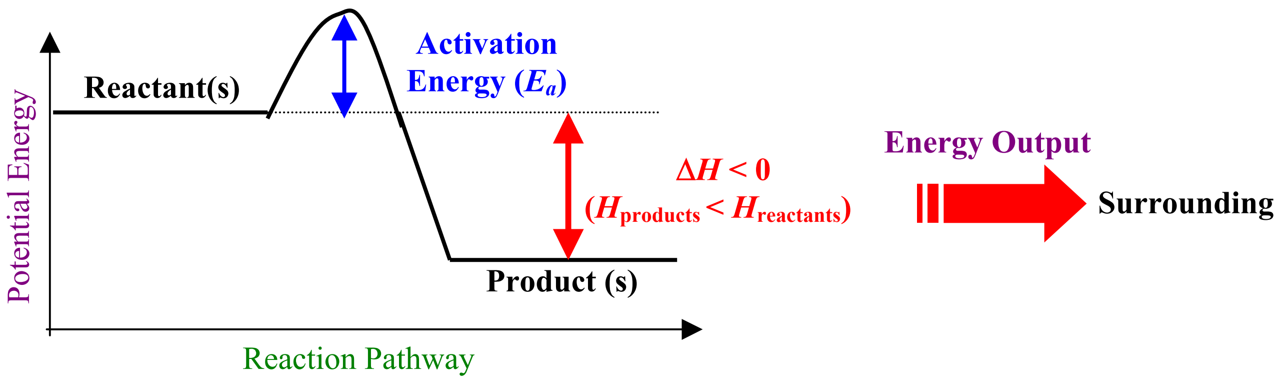
1.	**Activation Energy (E_a) ↓**	**Reaction Rate ↑**
2.	**Temperature (T) ↑**	**Reaction Rate ↑**
3.	**Molecular Orientations ↑**	**Reaction Rate ↑**
4.	**Particle Size ↓ (Surface Area ↑)**	**Reaction Rate ↑**
5.	**Concentration ↑**	**Reaction Rate ↑**
6.	**Catalyst ↑**	**Reaction Rate ↑**
7.	**Inhibitor ↑**	**Reaction Rate ↓**

Assignment
9.4 pg. 302 #23 to 28

9.5: Chemical Reactions can be either Exothermic or Endothermic

Exothermic Process ($\Delta H < 0$): - when **energy (heat) flows "out" of the system** into the surrounding.
(Surrounding gets Warmer.)

Potential Energy Diagram for Exothermic Process



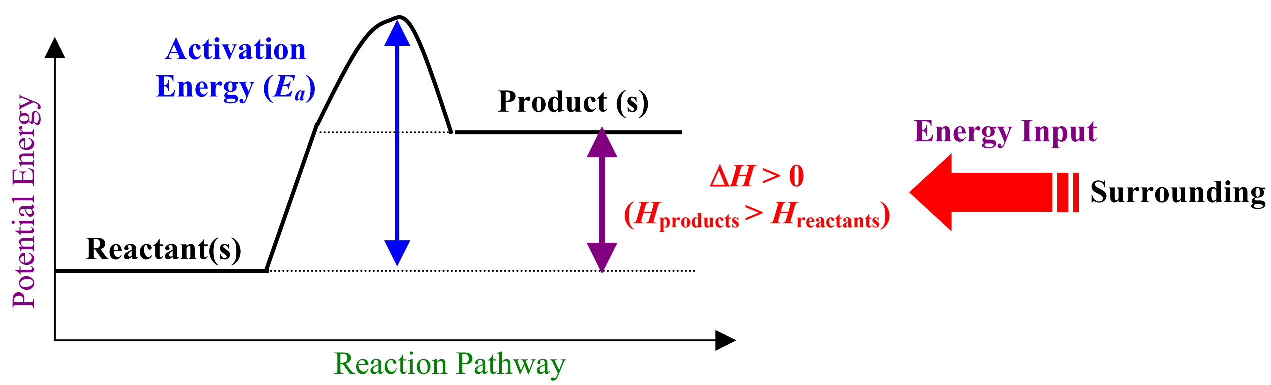
Reactant(s) → Product(s) + Heat **OR** **Reactant(s) → Product(s)** $\Delta H = -$____ kJ

Example: Methane undergoes combustion at constant pressure.

$CH_{4\,(g)} + 2\,O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\,H_2O_{(g)} + 803$ kJ or $CH_{4\,(g)} + 2\,O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\,H_2O_{(g)}$ $\Delta H = -803$ kJ

Endothermic Process ($\Delta H > 0$): - when **energy (heat) flows into the system** from the surrounding.
(Surrounding gets Colder.)

Potential Energy Diagram for Endothermic Process



Reactant(s) + Heat → Product(s) **OR** **Reactant(s) → Product(s)** $\Delta H = +$____ kJ

Example: Water is vaporized from its liquid state.

$H_2O_{(l)} + 40.7$ kJ $\rightarrow H_2O_{(g)}$ or $H_2O_{(l)} \rightarrow H_2O_{(g)}$ $\Delta H = +40.7$ kJ

Bond Energy: - the energy to required breaking or released from forming a particular chemical bond.

$\Delta E_{bond} > 0$ (*Endothermic*) when *Breaking a Bond*; $\Delta E_{bond} < 0$ (*Exothermic*) when *Forming a Bond*.

Selected Bond Energies (kJ/mol)							
Single Bonds						Multiple Bonds	
H – H	436	N – H	391	O – H	464	C = C	614
H – F	569	N – N	159	O – O	138	C ≡ C	837
H – Cl	431	N – F	272	O – F	190		
H – Br	363	N – Cl	200	O – Cl	203	O = O	498
H – I	295	N – Br	243	O – I	234		
		N – O	201			C = O	803
C – H	414			S – H	339	C ≡ O	1072
C – C	347	F – F	154	S – F	327		
C – N	389	F – Cl	253	S – Cl	253	N = N	418
C – O	358	F – Br	237	S – Br	218	N ≡ N	946
C – F	485	Cl – Cl	243	S – S	266		
C – Cl	339	Cl – Br	218			N = O	631
C – Br	276	Br – Br	193	Si – Si	340		
C – I	240	I – I	149	Si – H	393	C = N	615
C – S	259	I – Cl	208	Si – C	360	C ≡ N	891
		I – Br	175	Si – O	452		

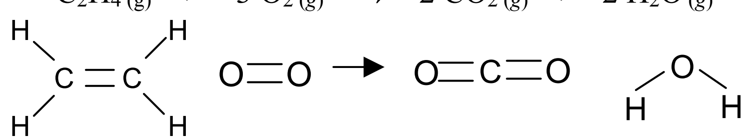
Determining the Molar Heat of Formation / Reaction of Covalent Compounds

$\Delta H = \Sigma D$ (breaking old bonds) + ΣD (forming new bonds)

(ΣD means sum of bond energy)

Example 1: Ethene is burned as described in the following equation.

$$C_2H_{4\,(g)} + 3\ O_{2\,(g)} \rightarrow 2\ CO_{2\,(g)} + 2\ H_2O_{(g)}$$



Using the bond energy data above, determine the heat of reaction for the combustion of ethene.

ΣD (breaking bonds):

1 C = C bond (1 ~~mol~~)(614 kJ/~~mol~~)
4 C – H bonds (4 ~~mol~~)(414 kJ/~~mol~~)
3 O = O bonds + (3 ~~mol~~)(498 kJ/~~mol~~)

ΣD (breaking bonds) = +3764 kJ

ΣD (forming bonds):

4 C = O bond (4 ~~mol~~)(–803 kJ/~~mol~~)
4 O – H bonds + (4 ~~mol~~)(–464 kJ/~~mol~~)

ΣD (forming bonds) = –5068 kJ

$\Delta H = \Sigma D$ (breaking bonds) + ΣD (forming bonds)

ΔH = (3764 kJ) + (–5068 kJ) **$\Delta H = -1304$ kJ**

$$C_2H_{4\,(g)} + 3\ O_{2\,(g)} \rightarrow 2\ CO_{2\,(g)} + 2\ H_2O_{(g)} + 1304\ \text{kJ}$$

Assignment

9.5 pg. 302 #29 to 33 and Worksheet: Changes in Heat of Chemical Reactions
Hands-On Activity: Warming and Cooling Water Mixtures (pg. 293)

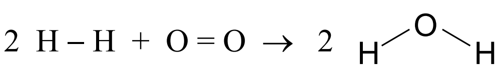
Assignment

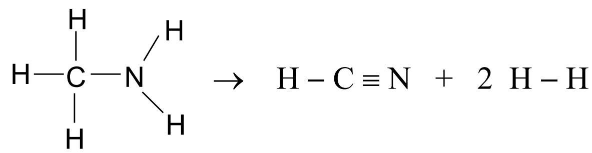
Chapter 9 Review: pg. 301 #1 to 13, 15 (Key Terms and Matching Definitions)
pg. 303–304 (Exercises) #1 to 32
pg. 304–305 (Problems) #1 to 5

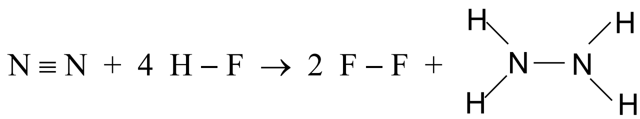
Worksheet: Changes in Heat of Chemical Reactions

1. Using the Bond Energies Table on pg. 112, determine the Δ*H* of the following reactions.

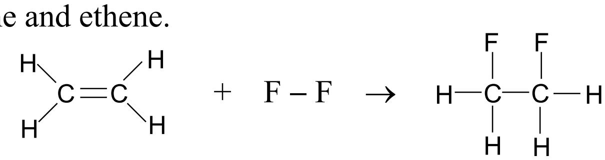
a. H – H + F – F → 2 H – F

b. 

c. 

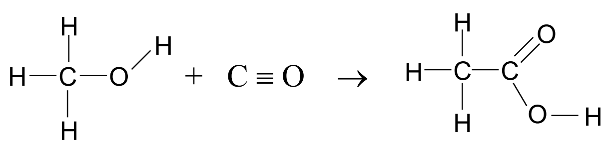
d. 

2. Halogens can be added readily across a double bond of a hydrocarbon molecule. Consider the following reaction between chlorine and ethene.



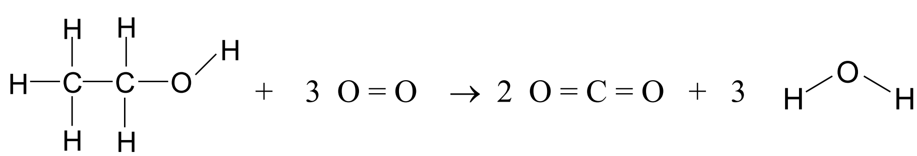
Determine the Δ*H* of the above reaction by using the Bond Energies Table on pg. 112.

3. Acetic acid, commonly known as vinegar, can be manufactured from the reaction of methanol and carbon monoxide.



Using the Bond Energies Table on pg. 112, calculate the Δ*H* of the above reaction.

4. One substitute for gasoline is ethanol. Find the Δ*H* of the combustion of ethanol using the following by using the Bond Energies Table on pg. 112.



Answers

1a. −548 kJ **b.** −486 kJ **c.** 236 kJ **d.** 1191 kJ **2.** −549 kJ **3.** −78 kJ **4.** −1263 kJ

