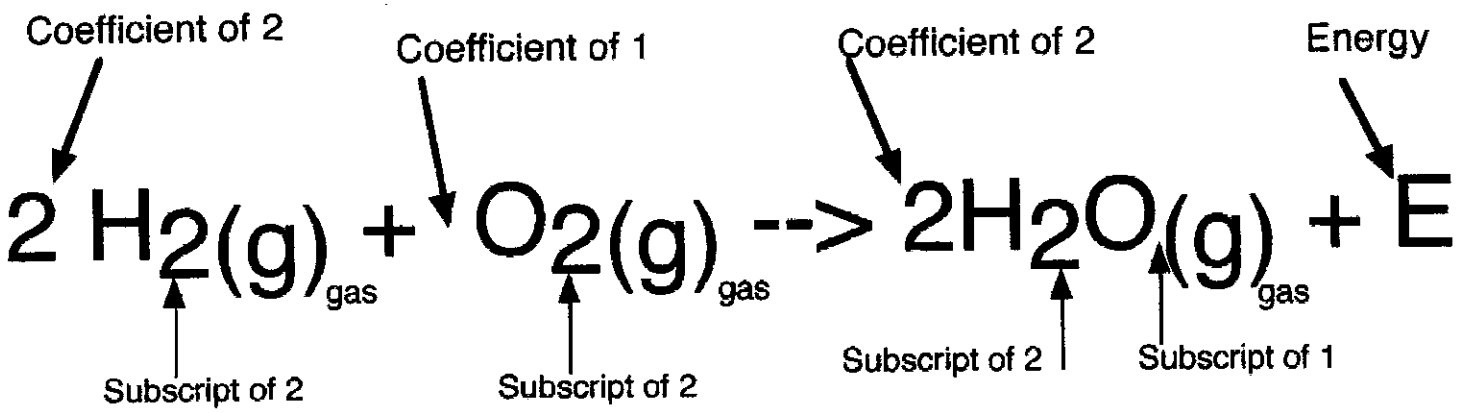
Understanding Chemical Equations

Chemical equations represent chemical reactions. A chemical equation uses numbers and atomic symbols to describe the rearrangement of atoms as substances (compounds or elements) recombine to form new substances (elements or compounds).
A chemical reaction is the actual process. A chemical equation is a representation of that process.

These symbols represent the states of the substances involved in the chemical equation.

(s) solid (l) liquid (g) gas
(aq) aqueous - mixed with water



The large number that preceeds (comes before) an element, compound, or ion is called a **coefficient**. A coefficient describes the number of **moles** of the element or compound that follows it. Its effect stops when it comes to either a **+** or an **—>**. If no coefficient is shown it is assumed to be **1**. For the above reaction, 2 moles of diatomic hydrogen gas react with 1 mole of diatomic oxygen gas to produce 2 moles of water (H_2O) gas and energy (E).

Remember! Any process (chemical or physical) that releases energy is **exothermic**.
The small number that follows an element or group of atoms in parenthesis is called a **subscript**. A subscript describes the number of moles of each item that preceeds (comes before) it. If no subscript is shown it is assumed to be **1**. A subscript that follows a parentheses means that there is that # of everything in the parentheses.
Ex: $(NO_3)_2$ This means that there are 2 (NO_3) s for a total of 2 Ns and 2 X 3 = 6 O s.

A **reactant** is a substance consumed in a chemical reaction.
A **product** is a substance produced by a chemical reaction.

Reactants Product

$$2A + B \rightarrow A_2B$$

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Four basic types of chemical equations: 2

A. Synthesis:

Two or more elements or compounds may combine to form a more complex (bigger) compound.

Basic form: A + X --> AX

Examples of synthesis reactions:

1. $2Mg_{(s)} + O_{2(g)} \text{----->} 2MgO_{(s)}$

2. $C_{(s)} + O_{2(g)} \text{----->} CO_{2(g)}$

3. $MgO_{(s)} + H_2O_{(l)} \text{----->} Mg(OH)_{2(s)}$

4. $CO_{2(g)} + H_2O_{(l)} \text{----->} H_2CO_{3(aq)}$

5. $2\,Na_{(s)} + Cl_{2(g)} \text{----->} 2NaCl_{(s)}$

6. $2P_{(s)} + 3Cl_{2(g)} \text{----->} 2PCl_{3(g)}$

7. $N_{2(g)} + 3H_{2(g)} \text{----->} 2NH_{3(g)}$

8. $NH_{3(g)} + H_2O_{(l)} \text{----->} NH_4OH_{(aq)}$

B. Decomposition:

A single compound breaks down into its smaller parts (elements or compounds).

Basic form: AX -----> A + X

Examples of decomposition reactions:

1. $CaCO_{3(s)} \text{----->} CaO_{(s)} + CO_{2(g)}$
2. $Ca(OH)_{2(s)} \text{----->} CaO_{(s)} + H_2O_{(g)}$
3. $2KClO_{3(s)} \text{----->} 2KCl_{(s)} + 3O_{2(g)}$
4. $H_2SO_4 \text{----->} H_2O_{(l)} + SO_{3(g)}$
5. $2HgO_{(s)} \text{----->} 2Hg_{(l)} + O_{2(g)}$
6. $2H_2O_{(l)} \text{----->} 2H_{2(g)} + O_{2(g)}$
7. $2NaCl_{(l)} \text{----->} 2Na_{(s)} + Cl_{2(g)}$

C. Single Replacement:

A more active element takes the place of another element in a compound and sets the less active one free. Use Table J to determine which element is more active.

Basic form: A + BX ——> AX + B

Examples of single replacement reactions:

1. Replacement of a metal in a compound by a more active metal.
EX. $Fe_{(s)} + CuSO_{4(aq)} \longrightarrow FeSO_{4(aq)} + Cu_{(s)}$

2. Replacement of hydrogen in water by an active metal from Group 1 or group 2.
EX. $2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$
or
EX. $Mg_{(s)} + H_2O_{(g)} \longrightarrow MgO_{(s)} + H2_{(g)}$

3. Replacement of hydrogen in acids by active metals.
EX. $Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2(g)}$

4. Replacement of nonmetals by more active nonmetals.
EX. $Cl_{2(g)} + 2NaBr_{(aq)} \longrightarrow 2NaCl_{(aq)} + Br_{2(l)}$

NOTE: Refer to the **activity series for metals "Table J"** to predict products of replacement reactions. If the free element is above the element to be replaced in the compound, then the reaction will occur. If it is below, then no reaction occurs.

D. Double Replacement:

Basic form: AX + BY -----> AY + BX

Double replacement reactions take place between two ionic compounds dissolved in water (aq). In a double replacement reaction the cation (+ ion) of one compound replaces the cation (+ ion) in the other compound to produce two new compounds.

The cation always comes first in a compound.

It occurs between ions in aqueous solution (aq). A reaction will occur when a pair of ions come together to produce at least one of the following:

1. a precipitate (an insoluble solid)

2. a gas (g)

3. water or some other non-ionized substance.

Basic form: AX + BY -----> AY + BX

Examples of Double Replacement reactions:

1. Formation of precipitate (a solid).

EX. $NaCl_{(aq)} + AgNO_{3(aq)} \longrightarrow NaNO_{3(aq)} + AgCl_{(s)}$

or

EX. $BaCl_{2(aq)} + Na_2SO_{4(aq)} \longrightarrow 2NaCl_{(aq)} + BaSO_{4(s)}$

2. Formation of a gas.

EX. $HCl_{(aq)} + FeS_{(s)} \longrightarrow FeCl_{2(aq)} + H_2S_{(g)}$

3. Formation of water. (If the reaction is between an acid and a base it is called a neutralization reaction.)

EX. $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

NOTE: Use the **solubility rules** to decide whether a product of an ionic reaction is insoluble in water and will thus form a precipitate. If a compound is soluble in water then it should be shown as being in aqueous solution, or left as separate ions. It is, in fact, often more desirable to show only those ions that are actually taking part in the actual reaction. Equations of this type are called **net ionic equations**.

Describe the following chemical equations as **synthesis**, **decomposition**, **single replacement**, or **double replacement**.

1. $Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2(g)}$

2. $2Ca_{(s)} + O_{2(g)} \longrightarrow 2CaO_{(s)}$

3. $2HgO_{(s)} \longrightarrow 2Hg_{(l)} + O_{2(g)}$

4. $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$

5. $Fe_{(s)} + CuSO_{4(aq)} \longrightarrow FeSO_{4(aq)} + Cu_{(s)}$

6. $SrCO_{3(s)} \longrightarrow SrO_{(s)} + CO_{2(g)}$

7. $LiCl_{(aq)} + AgNO_{3(aq)} \longrightarrow LiNO_{3(aq)} + AgCl_{(s)}$

8. $HCl_{(aq)} + KOH_{(aq)} \longrightarrow KCl_{(aq)} + H_2O_{(l)}$

9. $2NaClO_{3(s)} \longrightarrow 2NaCl_{(s)} + 3O_{2(g)}$

10. $HCl_{(aq)} + CrS_{(s)} \longrightarrow CrCl_{2(aq)} + H_2S_{(g)}$

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Balancing Chemical Equations

A balanced chemical equation obeys the **Law of Conservaqtion of Mass, Law of Conservation of Energy, and Law of Conservation of Charge.**
A balanced chemical equation must have equal #s of moles of each type of atom on each side of the arrow.
Below is a balanced chemical equation

Ex: **$2H_2 + O_2 \rightarrow 2H_2O$**

Notice!
There are 4 moles of hydrogen atoms on the left side of the arrow and 4 moles of hydrogen atoms on the right side of the arrow.

There are 2 moles of oxygen atoms on the left side of the arrow and 2 moles of oxygen atoms on the right side of the arrow.

Balancing a Chemical Equation
Balancing a chemical equation requires a trial and error method.
To balance a chemical equation: You **can** change coefficients.
You **can't** change subscripts.
You will need to take inventory of the atoms that are present in the equation to start with.
Step 1: Put the numbers of moles on top of each atom on each side.
Step 2: Change 1 coefficient at a time and see how it affects the rest of the equation.
Step 3: Count polyatomic ions as single units only if they remain intact on both sides of the arrow.
If an ionic compound contains more than 2 elements it contains a polyatomic ion. *Look in Table E to identify polyatomic ions.*

Ex: Balance this chemical equation.

$$\_\_\_N_2 + \_\_\_H_2 \rightarrow \_\_\_NH_3$$

Take inventory of what you have to start with. Put the numbers of each atom on top.

$$\overset{2}{\_\_\_N_2} + \overset{2}{\_\_\_H_2} \rightarrow \overset{1\ 3}{\_\_\_NH_3}$$

Placing a coefficient of 3 in front of H_2 will affect the number of hydrogen atoms on the left side of the arrow.

$$\overset{2}{\_\_\_N_2} + \underline{\mathbf{3}}\ \overset{\mathbf{6}}{H_2} \rightarrow \overset{1\ 3}{\_\_\_NH_3}$$

Now you need 6 hydrogens on the right side of the arrow to match the left side.
Place a 2 in front of the NH_3.

$$\overset{2}{\_\_\_N_2} + \underline{\mathbf{3}}\ \overset{\mathbf{6}}{H_2} \rightarrow \underline{\mathbf{2}}\ \overset{\mathbf{2\ 6}}{NH_3}$$ **Balanced**

Note packet # 18

Ex: Balance this chemical equation.

$$\_\_H_2 + \_\_O_2 \rightarrow \_\_H_2O$$

Take inventory of what you have to start with. Put the numbers of each atom on top.

$$\overset{2}{\_\_H_2} + \overset{2}{\_\_O_2} \rightarrow \overset{2\ 1}{\_\_H_2O}$$

Placing a coefficient of 2 in front of H_2O will affect the number of hydrogen and oxygen atoms on the right side of the arrow.

$$\overset{2}{\_\_H_2} + \overset{2}{\_\_O_2} \rightarrow \underline{2}\ \overset{4\ 2}{H_2O}$$

Now you need 4 hydrogens on the left side of the arrow to match the right side.
Place a 2 in front of the H_2.

$$\underline{2}\ \overset{4}{H_2} + \overset{2}{\_\_O_2} \rightarrow \underline{2}\ \overset{4\ 2}{H_2O}$$ **Balanced**

Ex: Balance this chemical equation.

$$\_\_Al(s) + \_\_Fe(NO_3)_2(aq) \rightarrow \_\_Al(NO_3)_3(aq) + \_\_Fe(s)$$

Take inventory of what you have to start with. Put the numbers of each atom on top.
Count each (NO_3) as an individual unit as long as it is on both sides of the equation.

$$\overset{1}{\_\_Al(s)} + \overset{1\ 2}{\_\_Fe(NO_3)_2(aq)} \rightarrow \overset{1\ 3}{\_\_Al(NO_3)_3(aq)} + \overset{1}{\_\_Fe(s)}$$

Trial and error.

$$\overset{1}{\_\_Al(s)} + \underline{3}\ \overset{3\ 6}{Fe(NO_3)_2(aq)} \rightarrow \overset{1\ 3}{\_\_Al(NO_3)_3(aq)} + \overset{1}{\_\_Fe(s)}$$

$$\overset{1}{\_\_Al(s)} + \underline{3}\ \overset{3\ 6}{Fe(NO_3)_2(aq)} \rightarrow \underline{2}\ \overset{2\ 6}{Al(NO_3)_3(aq)} + \overset{1}{\_\_Fe(s)}$$

$$\underline{2}\ \overset{2}{Al(s)} + \underline{3}\ \overset{3\ 6}{Fe(NO_3)_2(aq)} \rightarrow \underline{2}\ \overset{2\ 6}{Al(NO_3)_3(aq)} + \underline{3}\ \overset{3}{Fe(s)}$$

Balanced

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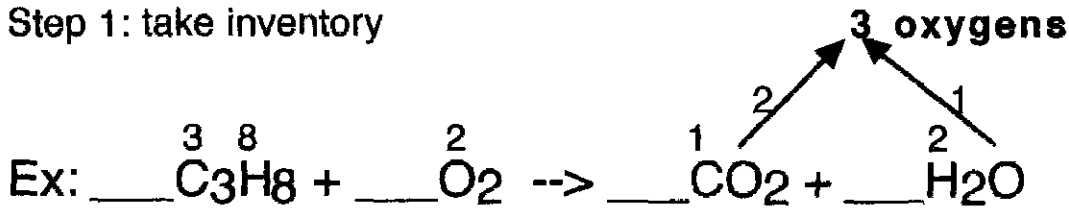
Balancing Combustion Reactions More Difficult Problem! 9

If you have to balance a combustion reaction, you will notice that oxygen appears in 2 compounds on the product side (carbon dioxide and water). It makes your counting of atoms more tricky. Always use pencil.

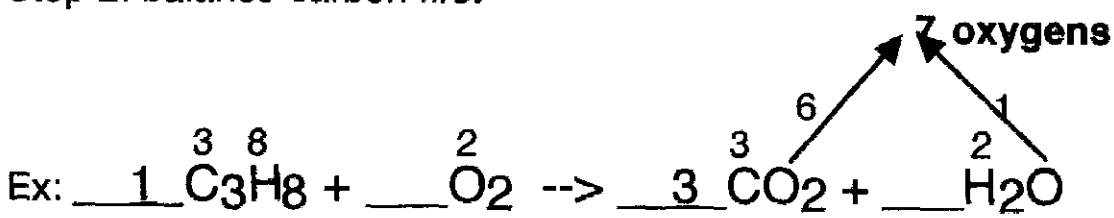
Strategy for combustion reactions:

Step 1: take inventory
Step 2: balance carbon first
Step 3: balance hydrogen second
Step 4: balance oxygen last

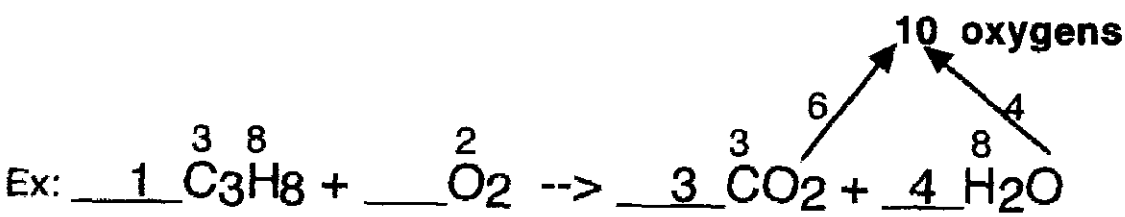
Step 1: take inventory



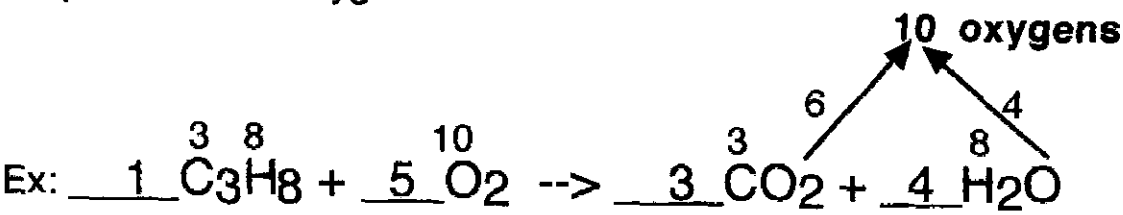
Step 2: balance carbon first



Step 3: balance hydrogen second



Step 4: balance oxygen last



Balanced!!!!!!!!!!!!!!!!!!

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Don't forget! If you don't see a coefficient it is 1.

Balance the following chemical equations. Trial and error, **use pencil!**

1. ___Fe + ___O_2 --> ___Fe_2O_3

2. ___N_2O_5 --> ___N_2 + ___O_2

3. ___CH_4 + ___O_2 --> ___CO_2 + ___H_2O

4. ___Zn + ___HCl --> ___$ZnCl_2$ + ___H_2

5. ___$Ba(OH)_2$ + ___Na_2SO_4 --> ___$BaSO_4$ + ___$NaOH$

6. ___Na +___O_2 --> ___Na_2O

7. ___Fe + ___H_2O --> ___Fe_3O_4 + ___H_2

8. ___$FeCl_2$ + ___Na_2CO_3 --> ___$FeCO_3$ + ___$NaCl$

9. ___$AgNO_3$ + ___$NaCl$ --> ___$NaNO_3$ + ___$AgCl$

Challenge problems 10 - 13 Pencil!!!!!!!!

10. ___C_8H_{18} + ___O_2 --> ___CO_2 + ___H_2O

11. ___$Fe_2(SO_4)_3$ + ___$K(SCN)$ ----> ___$K_3Fe(SCN)_6$ + ___K_2SO_4

12. ___$C_{12}H_{22}O_{11(s)}$ + ___$O_{2(g)}$ -> ___$CO_{2(g)}$ + ___$H_2O_{(g)}$

13. ___Fe_2O_3 + ___CO -> ___Fe_3O_4 + ___CO_2

Mole Ratios Using Balanced Chemical Equations

The **coefficients** in a balanced chemical equation indicate the number of **moles** of each reactant and product.

A balanced chemical equation can be used to solve mole ratio problems.

To solve these problems we must use a balanced chemical equation and set up a ratio.

Ex: $2H_2 + O_2 \rightarrow 2H_2O$ This **balanced** chemical equation tells us that 2 moles of diatomic hydrogen react with 1 mole of diatomic oxygen to produce 2 moles of water.

Ex: How many moles of oxygen are required to produce 10 moles of water?

To solve: Set up a ratio.

Remember! If you don't see a coefficient, it is 1.

$$\underset{}{2H_2} + \overset{X}{1O_2} \rightarrow \overset{10}{2H_2O}$$

$\frac{X}{1} = \frac{10}{2}$ Cross-multiply and divide. $2X = 10$ $X =$ **5 moles of oxygen**

Ex: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$

How many moles of oxygen are consumed (used) in the production of 30 moles of Fe_2O_3?

To solve: Set up a ratio.

$$\underset{}{4Fe} + \overset{X}{3O_2} \rightarrow \overset{30}{2Fe_2O_3}$$

$\frac{X}{3} = \frac{30}{2}$ Cross-multiply and divide. $2X = 90$ $X =$ **45 moles of oxygen**

Use the following balanced chemical equations to solve the following mole ratio problems.

1. $N_2 + 3H_2 \rightarrow 2NH_3$

How many moles of hydrogen are required to produce 15 moles of ammonia?

2. $FeCl_2 + Na_2CO_3 \rightarrow FeCO_3 + 2\ NaCl$

How many moles of sodium carbonate are required to produce 7 moles of sodium chloride?

3. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$ *C_3H_8 is propane*

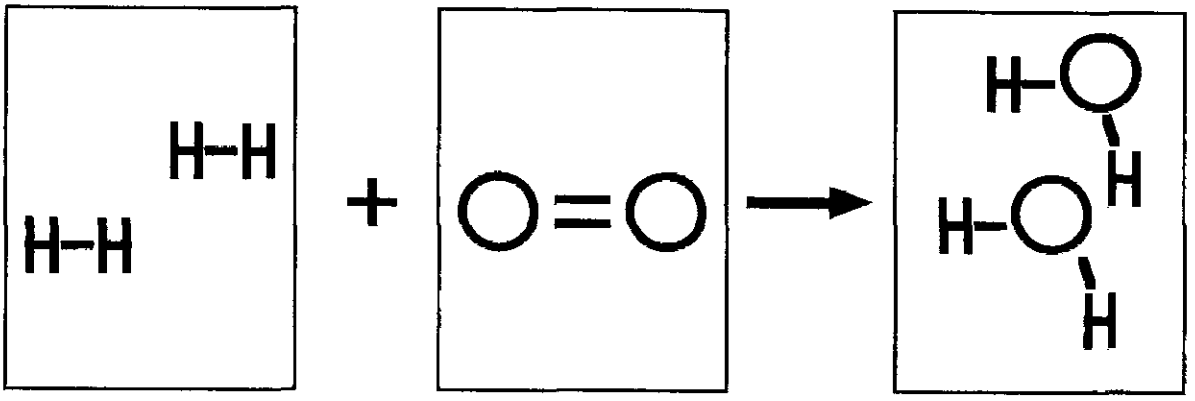
How many moles of propane are consumed in the production of 12 moles of water?

4. $2N_2O_5 \rightarrow 2N_2 + 5O_2$

How many moles of N_2 are produced when 25 moles of O_2 are produced?

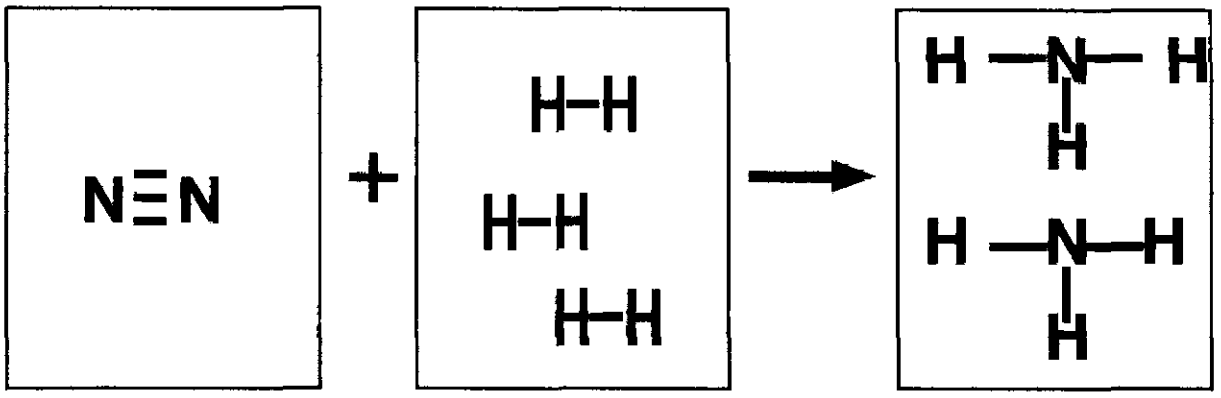
Chemical Reaction Diagrams are representations of the reacting particles during a chemical reaction. A reaction diagram must obey the law of conservation of mass.

$$2H_2 + O_2 \rightarrow 2H_2O$$



The particle diagram above represents 2 diatomic molecules of hydrogen reacting with 1 diatomic molecule of oxygen to form 2 water molecules.

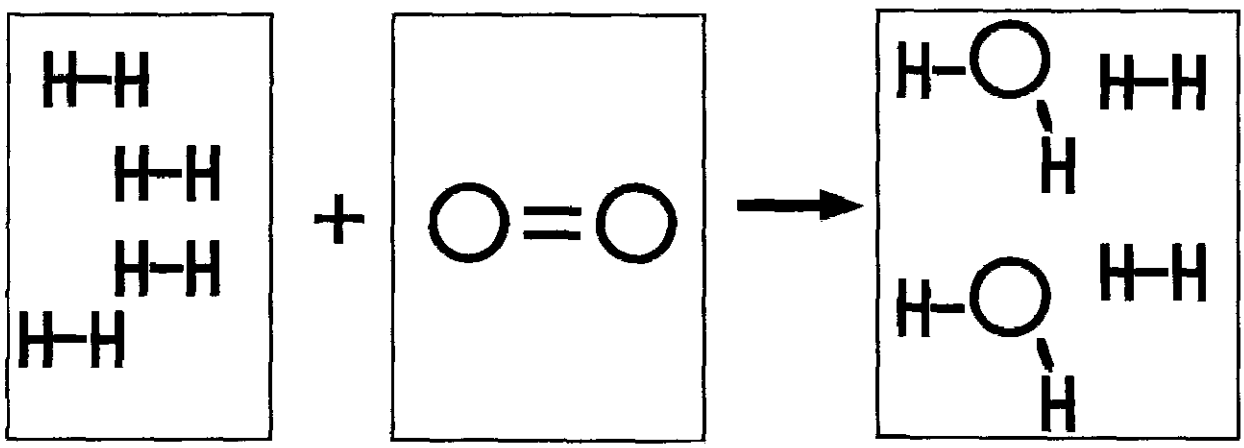
$$N_2 + 3H_2 \rightarrow 2NH_3$$



The particle diagram above represents 1 diatomic molecule of nitrogen reacting with 3 diatomic molecules of hydrogen to form 2 ammonia molecules.

Chemical reactions must obey the Law of Conservation of Mass. Therefore, all the particles must be accounted for whether they react or not.

$$2\ H_2 + O_2 \rightarrow 2H_2O$$



4 diatomic molecules of hydrogen

1 diatomic molecule of oxygen

2 H_2O molecules and 2 diatomic molecules of hydrogen.

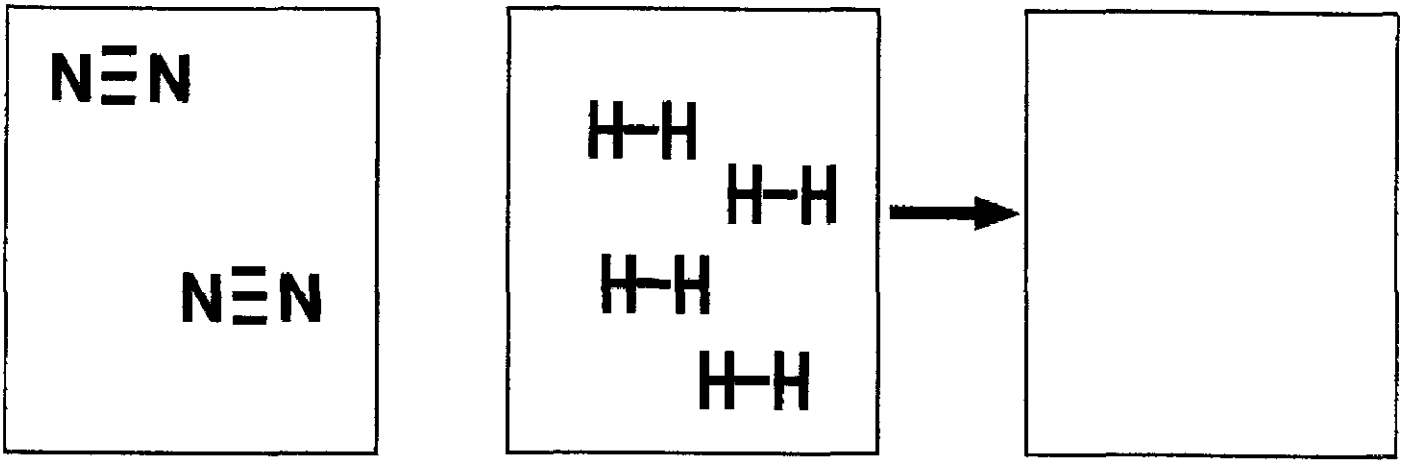
These particle diagrams show 4 diatomic molecules of hydrogen interacting with 1 diatomic molecule of oxygen to form 2 H_2O molecules.

Notice! There are 2 diatomic molecules of hydrogen that did not react.
Conservation of mass states that all atoms must be accounted for.

Complete the following particle diagram. Make sure to obey the Law of Conservation of Mass

Nitrogen and hydrogen react according to this balanced chemical reaction.

$$N_2 + 3H_2 \rightarrow 2NH_3$$



Complete the following particle diagram. Make sure to obey the Law of Conservation of Mass

Phosphorous and chlorine react according to this balanced chemical reaction.

$$2P + 3Cl_2 \rightarrow 2PCl_3$$