Chemical Equations

1.0 Introduction

This exercise will provide you with practice at writing balanced chemical equations for the major types of chemical reactions:

- **combination**
- **decomposition**
- **complete oxidation or burning** of organic compounds
- **oxidation-reduction** (single replacement type only)
- **double replacement**, formation of a precipitate or molecular compound

CHEMICAL OVERVIEW

A chemist uses a balanced chemical equation to represent a chemical change. The general form of a chemical equation is :

Reactant 1 + Reactant 2 + ... $\longrightarrow$ Product 1 + Product 2 + ...

Substances on the left side of the reaction arrow are called the **reactants** while those on the right of the arrow are the **products**. Each reactant or product is separated by a plus sign. The reaction arrow separates the two sides of the chemical reaction, the left initial side represents the substances before they react while the right side of the reaction arrow represents the final changed substances produced by the chemical reaction. When reading the chemical equation, a chemist will often replace the arrow with the word "yields", "forms" , "produces", "react to form" Etc. to indicate the change that must occur during the conversion of the reactants into products.

Sometimes chemists will use special symbols to indicate the states of the substances in the chemical equation. The symbols **(s), (l), (g)** and **(aq)** represent a **solid, liquid, gas or water (aqueous) solution** respectively. Each of these symbols is placed immediately after the substance it identifies. Additional, information about the reaction conditions is often placed above and sometimes below the reaction arrow. Examples would be Fe, Pt or HCl used as a catalyst for the reaction. Also, conditions such as 300 atm. (pressure) or 200 °C (temperature) could be added to an equation. Examples of these symbols are :

$$2\ Na\ (s) + Cl_2\ (g) \xrightarrow[200\ atm.,\ 500\ ^{o}C]{} 2\ NaCl\ (s)$$

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

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

A chemical reaction will tell you two major things. First, it will tell you what substances are involved in the chemical reaction. To do this properly, all of the formulas for the chemicals in the equation must be correct. Second, a chemical equation has a quantitative significance. Since all chemical reactions obey the Law of Conservation of Mass, the number and types of atoms must balance on both sides of the chemical equation. Only when each element appears in equal numbers on **BOTH** sides of a chemical equation is the equation said to be balanced.

To write a properly balanced chemical equation, you must follow two INDEPENDENT steps:

1. **write the correct formula for each reactant and product in the reaction**
2. **use whole number coefficients to balance the number of each element appearing on both sides of the equation** (a coefficient = 1 is not written)

Remember to complete step 1) BEFORE attempting step 2) and NEVER change a correct substance formula in an attempt to "balance the equation". Also, NEVER add "extra" substances or atoms to either side of a reaction in the attempt to balance it.

REACTION TYPE EXAMPLES

A. Combination Reactions Combination reactions occur when two or more chemical substances combine to form new compounds. Chemists often call these reactions synthesis reactions and synthetic chemists attempt to "synthesize" new materials by combining older available ones. The substances which react can be elements or compounds or a combination of both and only one product is formed by the combination of the reactants. Sometimes you will only be given the name of the final product and must deduce the elements necessary to form it as in the description : ammonia is formed from it elements.

EXAMPLE 1: Magnesium oxide is formed from its elements.

Since magnesium oxide "is formed" it must be a product and goes on the right side of the reaction arrow. DO NOT make the mistake of placing it on the left side simply because it appears first in the sentence describing the reaction. Since the product is formed from "its elements", we are only allowed to place elements on the left (reactant) side of the reaction arrow.

First carry out step 1) and write each formula correctly

$Mg + O_2 \longrightarrow MgO$ (Note the ELEMENT oxygen MUST be diatomic)

Next use whole number coefficients to balance the final reaction equation and complete step 2

$2\,Mg + O_2 \longrightarrow 2\,MgO$

It would be a mistake to assume that the oxygen must have the formula "O" rather than its naturally occurring O_2 simply because the product is MgO. Obviously, it would be equally incorrect to try to form MgO_2 from the elements. Thus, it is important to complete step 1) and assign each reactant and product its proper formula BEFORE moving to step 2) and attempting to balance the reaction using whole number coefficients. Your instructor and the text will have some valuable tips to aid you in your assigning the proper formulas to the various elements and compounds.

EXAMPLE 2: Write the equation for the formation of iron(III) oxide from its elements.

Complete step 1) by writing the formulas of the reactants (left) and the products (right).

$$Fe + O_2 \rightarrow Fe_2O_3$$

Next balance the iron and leave the oxygen for last. You will find it helpful to focus on only one element at a time as you attempt to balance a chemical reaction. If you get "stuck" it can often be helpful to erase your numbers and attempt to start your equation balancing by focusing on another element. Balancing the iron will give the result :

$$2\ Fe + O_2 \rightarrow Fe_2O_3$$

Now turn your attention to the oxygen part of the balancing. You may proceed by two possible steps.

1. Notice that there are 2 oxygen atoms that must balance with 3 oxygen atoms. If we are to add whole number coefficients we need to find the smallest number that can be "factored" by 2 and 3 which is 6. Thus, if we set up the final equation so that 6 oxygen atoms appear on both sides the equation should balance. This would give 3 times O_2 and 2 times Fe_2O_3 or written as a proper equation:

$$2\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$$

Now the 2 in front of the reactant iron would need to be changed to give 4 Fe atoms to react.

2. A second possible approach would be to notice that the product has 3 oxygen atoms in it which could be supplied by 3/2 of an O_2 (each 1/2 of an O_2 gives 1 O atom). If we write this equation down we get:

$$2\ Fe + 3/2\ O_2 \rightarrow Fe_2O_3$$

We should NOT leave the equation in this fractional form so we need to multiply the **ENTIRE** equation by 2 to "clear" the fraction. This will result in the balanced equation :

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

Notice that this is the same result that we would get by the first method. You may find it helpful to use the method of fractional balancing to balance some equations but always remember to clear the fractions in the final equation. (Also reduce your coefficients if they are not in their lowest form.)

B. Decomposition Reactions Decomposition reactions are often the reverse of the combination reactions and can be recognized by the conversion of one chemical into two or more products. One type of decomposition reaction that is readily reversible is the decomposition of a hydrate into its anhydrous salt (anhydrous = without water) and water. A hydrate such as $Na_2CO_3 \cdot 10H_2O$ can be decomposed into the anhydrous sodium carbonate salt and water by heating the compound according to the equation :

$$Na_2CO_3 \cdot 10H_2O \rightarrow Na_2CO_3 + 10\ H_2O$$

The addition of water to the salt will reform the hydrate which is a combination type reaction and is the reverse of the decomposition shown above. This type of reaction is a reversible reaction, while many decompositions are irreversible since the starting compound is too complex to be easily reformed from the decomposition products.

EXAMPLE 3: Calcium carbonate is decomposed into calcium oxide (lime) and carbon dioxide by heat.

First write the reactants and products.

$$CaCO_3 \rightarrow CaO + CO_2$$

Second balance the atoms on both sides of the equation. This time it is easy since all the coefficients are 1 !

C. Complete Oxidation or Burning of Organic Compounds Other than a few carbon compounds such as the oxides of carbon, carbonates, and other simple carbon compounds, the compounds formed from carbon are called **organic compounds**. Most of these organic compounds contain hydrogen and those that consist of only carbon and hydrogen are known as "hydrocarbons". Fuels, oils, natural gas, butane and propane are examples of hydrocarbons. Other organic compounds consist of carbon and hydrogen combined with some oxygen, nitrogen or sulfur. Still others such as the "fluorocarbons" have fluorine while others have other halogen atoms. We will focus our attention to compounds with only carbon, hydrogen and oxygen in them so we can simplify the products that will be formed from the combustions. When a hydrocarbon or an organic compound containing oxygen such as alcohol or sugar is burned the products are carbon dioxide and water. This is true even if the burning occurs in the body of a living animal. Carbon dioxide is always produced if the compound is completely burned by an excess of O_2. If the compound is burned in a closed space (such as an automobile cylinder) where the O_2 is limited carbon monoxide may form instead of carbon dioxide, however, we will assume for these examples that the combustion will occur with an abundance of oxygen - producing CO_2. The description of these reactions will often contain the words **"burning", "completely oxidized", "oxidized" or "combusted"** which for our purposes will be considered synonyms. Also, the final reactions may have large coefficients if the organic compounds that are burning are large molecules. Don't let this distract you - the products will be CO_2 and H_2O when these compounds are combined with oxygen. Note, however, that an organic compound which decomposes by simple heating **without combining with O_2** may not produce carbon dioxide but might break down into carbon and water instead.

EXAMPLE 4: Write the equation for the complete combustion of ethyl alcohol, C_2H_5OH.

First we know the two reactants must be the organic compound as given and the O_2 that will react with it. Secondly, we know the products will be CO_2 and H_2O. Begin by writing all of these down in the equation.

$$C_2H_5OH + O_2 \rightarrow CO_2 + H_2O$$

Now begin balancing the equation. Combustion reactions usually balance best by starting with carbons, then balancing hydrogens and finally balancing oxygens. DO NOT even try balancing the oxygen atoms first ! Balancing the carbons gives :

$C_2H_5OH + O_2 \rightarrow \mathbf{2}\, CO_2 + H_2O$ 2 carbons on both sides

Next balancing the hydrogens gives the equation :

$C_2H_5OH + O_2 \rightarrow 2\, CO_2 + \mathbf{3}\, H_2O$ 6 hydrogens on both sides

Finally, the oxygens are balanced to give the final balanced equation :

$$C_2H_5OH \; + \; \mathbf{3}\, O_2 \; \rightarrow \; 2\, CO_2 \; + \; 3\, H_2O \qquad \text{7 oxygens both sides}$$

Be very careful not to "overlook" any oxygen atoms that are in the compound being burned !

EXAMPLE 5: Balance the equation for the burning of methyl ethyl ether, $CH_3OCH_2CH_3$.

Again we start by writing the basic equation :

$$CH_3OCH_2CH_3 \; + \; O_2 \; \rightarrow \; CO_2 \; + \; H_2O$$

Following through the balancing steps, we first balance the carbon atoms :

$$CH_3OCH_2CH_3 \; + \; O_2 \; \rightarrow \; \mathbf{3}\, CO_2 \; + \; H_2O \qquad \text{3 carbons each side}$$

Next we balance the hydrogen atoms :

$$CH_3OCH_2CH_3 \; + \; O_2 \; \rightarrow \; 3\, CO_2 \; + \; \mathbf{4}\, H_2O \qquad \text{8 hydrogens each side}$$

And finally, we must balance the oxygen atoms (don't forget those in the compound).

$$CH_3OCH_2CH_3 \; + \; O_2 \; \rightarrow \; 3\, CO_2 \; + \; 4\, H_2O$$

$$\text{1 O} \qquad \text{need 9 O} \qquad \text{6 O} \; + \; \text{4 O} \; = \text{10 O}$$

Looking at this equation, we find a small problem. We have 10 O atoms on the product side so we must match them by 10 O atoms in the reactants; however, 1 O atom is in the organic compound leaving us needing only nine more. Since O_2 always has **two** it appears that we can only get **even** numbers of oxygen atoms from it. The simple solution is to use fractions to balance the equation and to then clear the fraction by multiplying the ENTRE equation by 2.

$$CH_3OCH_2CH_3 \; + \; \mathbf{9/2}\, O_2 \; \rightarrow \; 3\, CO_2 \; + \; 4\, H_2O \qquad \text{10 oxygens each side}$$

Using 9/2 of an O_2 will give us 9 O atoms s needed ! Then we can clear the fraction by multiply the entire equation by a factor of 2 :

$$2\, CH_3OCH_2CH_3 \; + \; 9\, O_2 \; \rightarrow \; 6\, CO_2 \; + \; 8\, H_2O$$

This gives us the final balanced equation for the combustion.

D. Oxidation-Reduction Reactions To the chemist, the words **oxidize** and **oxidation** have a broader meaning than to react with oxygen. When we analyze the typical reactions of oxygen, we find that combining oxygen with another element almost always results in the formation of O^{2-} or the oxide ion. That is, the oxygen **takes** electrons from the other atom. Thus, the term **oxidation** has become generalized to mean **"loss of electrons"**. When an element is oxidized it looses one or more electrons. If one element looses an electron, it is only because another element has taken it or gained an electron. The gaining of an electron is called **reduction**. Since these two processes occur together, we always have **oxidation-reduction** reactions. These are often called "**redox**" reactions for a shorter label.

Although there are many types of redox reactions, including those which produce energy in batteries and our bodies, we are only going to focus on the **single replacement** type. In this type of redox reaction, one element replaces another element in a compound. The element going into the compound looses electrons or is oxidized; while the element being replaced is reduced or gains electrons. One common example of this type of reaction is a metal being "dissolved" by an acid.

EXAMPLE 6: Write the equation for the reaction of zinc metal with hydrochloric acid which releases hydrogen gas. (This is the sizzling when acids react with metals.)

First, write the general reactants and products :

$$Zn + HCl \rightarrow H_2 + ZnCl_2$$

Since hydrogen gas comes out, it must have been replaced by the zinc giving $ZnCl_2$ as the other product. It is very important to determine **all** of the reactants and product in the first step.

Now the balancing is fairly simple. Since the zinc atoms already balance one to one, try either the chlorine or the hydrogen atoms next. A flip of the coin says "heads" for hydrogen. (Sometimes it really does not matter when you start - just do it.) Balance of the hydrogen atoms gives:

$$Zn + \mathbf{2}\,HCl \rightarrow H_2 + ZnCl_2 \qquad \text{2 hydrogen atom each side}$$

Sometimes one action has a dual result. In this case placing the 2 coefficient before the HCl balanced **both** the hydrogen atoms and the chlorine atoms!

EXAMPLE 7: Write the equation for the reaction between aluminum and nickel nitrate.

This time only the reactants are given so write these down. It is often **much easier** to "see" a possible chemical reaction if you are looking at the formulas rather than the words.

$$Al + Ni(NO_3)_2 \rightarrow ????$$

Looking at this, we could possibly decompose the salt but into what ? Why is the aluminum there if this is a simple decomposition? Perhaps we could form something from nickel ion and aluminum ? Not very likely is it ? Let's transfer electrons (**redox**) and replace the nickel ion with aluminum. In the process aluminum looses electrons to form an ion (is **oxidized**) and the nickel ion gains electrons to form the metal (is **reduced**). One pattern in this is that "more reactive" metals will replace "less reactive" metals. How do you know the "less reactive" metals ? Think about jewelry and coins. Are these made of more reactive or less reactive metals ? Nickel coins make sense aluminum coins are rare. Thus, the more reactive aluminum will take the place of the less reactive nickel to give :

$$Al + Ni(NO_3)_2 \rightarrow Ni + Al(NO_3)_3$$

Now we could begin to balance each individual atom as we have done in pervious examples or we can balance groups of atoms. The second choice is easier in this case. In practice, whenever you see a group of atoms, such as the nitrate ion in this example, that remains unchanged from reactants to products, it will be easier to balance the reaction by balancing the group rather than individual atoms. But how can we balance 2 nitrates with 3 nitrates ? We will need the smallest number that two and three are factors of or 6 on each side. This gives the equation:

$$Al + \mathbf{3}\,Ni(NO_3)_2 \rightarrow Ni + \mathbf{2}\,Al(NO_3)_3 \qquad \text{6 nitrate ions each side}$$

Now it is a simple matter to complete the balance by balancing the metal atoms on both sides giving a final balanced equation of :

$$2\,Al + 3\,Ni(NO_3)_2 \rightarrow 3\,Ni + 2\,Al(NO_3)_3$$

E. Double Replacement Reactions When two ionic substances such as salts are placed together, there is a possibility that they will "exchange partners". Of course, not all salts will do this, so it is possible to mix ionic materials with no reaction as a result. If you continue your chemical studies, you will begin to learn more about "**driving forces**" or processes which compel chemicals to react, but for now you will be told that a reaction occurs. When one of these exchange type reactions takes place, we have what is called a **double replacement reaction**. The exchange process will occur because of the driving force. Driving forces include the formation of a precipitate (one combination of ions is insoluble in water), the formation of a gas (one combination escapes as a gas) or the formation of a stable molecular compound.

One special instance of the last driving force, is the reaction of an acid with a base to make water. A simple definition of an acid is a substance which can donate or release a hydrogen ion. A simple definition of a base is a substance which can accept or react with a hydrogen ion. Thus, the reaction between the acid, HCl, and the base, NaOH, occurs because the hydrogen ion released by the acid reacts with the hydroxide ion of the base to form a very stable molecular compound, H_2O. This formation of water leaves the sodium ion and the chloride ion to form sodium chloride salt. **Many acids and bases follow this pattern of forming water and a salt**. This reaction type is so important that we give it a special name: **neutralization**. Neutralizations may be special cases, but they still represent double replacement reaction types when two chemicals exchange ions.

EXAMPLE 8: Write the equation for the precipitation reaction that occurs between sodium hydroxide and iron(III) nitrate.

In this case you are told there is a driving force of a precipitate forming. This tells you that an exchange of ions between these two chemicals produces at least one insoluble compound. Be careful not to assume that because you have a base (NaOH) that you automatically have an acid-base neutralization. You must have **both** an acid and a base for neutralization. For this reaction, proceed by writing an exchange reaction:

$$NaOH + Fe(NO_3)_3 \rightarrow NaNO_3 + Fe(OH)_3$$

Now begin your balance by noting that the iron atoms and sodium atoms already balance one to one. Using our previous example, we will balance the nitrate ions and hydroxide ions as entire groups rather than trying to count individual atoms. Starting with the nitrate ion we see 3 on the reactant side and 1 nitrate on the product side which we balance by the equation :

$$NaOH + Fe(NO_3)_3 \rightarrow \mathbf{3}\,NaNO_3 + Fe(OH)_3 \qquad \text{3 nitrate ions each side}$$

Now balancing the hydroxide ions will give the equation:

$$\mathbf{3}\,NaOH + Fe(NO_3)_3 \rightarrow 3\,NaNO_3 + Fe(OH)_3 \qquad \text{3 hydroxide ions each side}$$

This is the final balanced equation we wanted. If you study this reaction, you should see that we could have started by balancing the hydroxide ions rather than the nitrate ions. Also, after balancing the nitrate ions a balance of the sodium ions would have automatically balanced the hydroxide ions. Thus, balancing one chemical species often balances another and sometimes it does not matter where you start to balance. Just remember if you get stuck, back up and start your balance with another chemical because sometimes it does matter where you start. For your information, the driving force for this reaction is the formation of iron(III) hydroxide precipitate.

EXAMPLE 9: Write the equation for the neutralization of phosphoric acid with calcium hydroxide.

Since we know this is an acid-base neutralization, we know to expect water and a salt as the products of the reaction. Writing the basic equation gives:

$$H_3PO_4 + Ca(OH)_2 \rightarrow H_2O + Ca_3(PO_4)_2$$

Now balancing everything as groups and remembering that $H^+ + OH^-$ gives HOH, we can start the balance. If you start with the hydrogen ion you will probably "get stuck". It can be done but it is more difficult. Instead start with either the calcium ion or the phosphate ion group.

Starting with the calcium ion gives the equation :

$$H_3PO_4 + \mathbf{3}\ Ca(OH)_2 \rightarrow H_2O + Ca_3(PO_4)_2 \quad \text{3 calcium atoms each side}$$

Now try the phosphate group balance to get :

$$\mathbf{2}\ H_3PO_4 + 3\ Ca(OH)_2 \rightarrow H_2O + Ca_3(PO_4)_2 \quad \text{2 phosphate groups each side}$$

Finally, note that 6 hydrogen ions and 6 hydroxide ions will form 6 water molecules to give the balanced equation :

$$2\ H_3PO_4 + 3\ Ca(OH)_2 \rightarrow \mathbf{6}\ H_2O + Ca_3(PO_4)_2$$

EXAMPLE 10: Write the equation for the reaction between sulfuric acid and sodium acetate, $NaC_2H_3O_2$.

First write the reactants which are given:

$$H_2SO_4 + NaC_2H_3O_2 \rightarrow ????$$

Now for the products. Is this a simple neutralization that forms water ? No, since there is no source of hydroxide to react with the acid. First, try writing the reaction that would result if this was just an exchange reaction. If you have forgotten the charge of the acetate ion, it can be derived since you know the charge of the sodium ion. Writing the reaction as an exchange gives:

$$H_2SO_4 + NaC_2H_3O_2 \rightarrow Na_2SO_4 + HC_2H_3O_2$$

Now the equation can be balanced by starting with either the sodium or the hydrogen ions. Again keep the acetate and sulfate as groups and make no attempt to count individual atoms. By random choice let's start with the sodium ions. This gives the partial balance of :

$$H_2SO_4 + \mathbf{2}\ NaC_2H_3O_2 \rightarrow Na_2SO_4 + HC_2H_3O_2 \quad \text{2 sodium atoms each side}$$

Next we can either balance the acetate groups or the hydrogen ions. Looking at either one will compel us to write the following equation :

$$H_2SO_4 + 2\ NaC_2H_3O_2 \rightarrow Na_2SO_4 + \mathbf{2}\ HC_2H_3O_2$$

In either case this is the final balanced reaction equation that we seek. The driving force behind this reaction is the formation of the acetic acid molecule. It turns out that most acids DO NOT ionize or separate into ions in water. There are only 6 or 7 so called **strong acids** which do. Other acids are **weak** because they form stable molecules and do not ionize in water solutions.

Report Sheet Name:____________

EQUATION-WRITING EXERCISE: Write balanced chemical equations for each reaction given.

A) Combination Reactions

1) Sodium chloride is formed by direct combination of its elements.

2) Sulfur dioxide is formed when sulfur is burned in air.

B) Decomposition Reactions

3) Ammonium nitrite decomposes into nitrogen and water.

4) When heated, potassium chlorate decomposes into oxygen and potassium chloride.

C) Complete Oxidation or Burning of Organic Compounds

5) Propane, C_3H_8, burns in air.

6) Acetone (nail polish remover), CH_3COCH_3, is completely oxidized.

D) Oxidation-Reduction Reactions (Single Replacement)

7) Hydrogen is released when aluminum reacts with hydrochloric acid.

8) Magnesium reacts with silver nitrate solution.

E) Ion Combination Reactions (Double Replacement)

9) Barium sulfate precipitates when barium chloride and sodium sulfate react.

10) Carbon dioxide and water are two of the three products from the reaction of sulfuric acid with sodium hydrogen carbonate.

11) Magnesium hydroxide precipitates when magnesium sulfate and sodium hydroxide mix.

12) Calcium carbonate forms from calcium chloride and sodium carbonate solutions.

F) Mixed Reactions

13) Hydrobromic acid reacts with potassium hydroxide.

14) Aluminum reacts with phosphoric acid.

15) Sugar, $C_{12}H_{22}O_{11}$, is burned in air.

16) Ammonia is formed from its elements.