Department of Chemical Engineering
Review Sheet - Chemical Reactions

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Introduction

This document is intended to help you review the basics of writing and balancing equations, how to predict the products of several general types of inorganic reactions, and how to write and balance equations for the combustion of hydrocarbons. There are also practice exercises with answers.

It is most important for a chemical engineer to be able to write correctly balanced equations when the reactants and products are all specified. It is also essential for a chemical engineer to know how to predict the products of certain specific types of common reactions.

Essential Concepts about Chemical Reactions

<u>Energy of Chemical Reactions</u>
Chemical reactions always involve a change in energy. Energy is neither created or destroyed. Energy is absorbed or released in chemical reactions. Chemical reactions can be described as endothermic or exothermic reactions.

<u>Endothermic Reactions</u>
Chemical reactions in which energy is absorbed are endothermic. Energy is required for the reaction to occur. The energy absorbed is often heat energy or electrical energy. Adding electrical energy to metal oxides can separate them into the pure metal and oxygen. Adding electrical energy to sodium chloride can cause the table salt to break into its original sodium and chlorine parts.

<u>Exothermic Reactions</u>
Chemical reactions in which energy is released are exothermic. The energy that is released was originally stored in the chemical bonds of the reactants. Often the heat given off causes the product(s) to feel hot. Any reaction that involves combustion (burning) is an exothermic chemical reaction.

Key Principles Involving Chemical Reactions

The diatomic elements (when they stand alone) are always written as H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2

The sign "$\longrightarrow$" means "yields" and shows the direction of the action.

A small delta "$\xrightarrow{\Delta}$" above the arrow shows that heat has been added.

A double arrow "$\rightleftharpoons$" shows that the reaction is reversible and can go in both directions.

If a reactant or product is a solid, (s) is placed after the formula.

If a reactant or product is a gas, (g) is placed after it.

If a reactant or product is in water solution, (aq) is placed after it.

Before beginning to balance an equation, check each formula to see that it is correct. NEVER change a formula during the balancing of an equation. This means you have to be sure of the specific reactants and products before balancing the equation. You cannot deduce the products of a reaction by looking at "what's left over".

Balancing is done by placing coefficients in front of the formulas to insure the same number of atoms of each element on both sides of the arrow.

Overview of Types of Chemical Reactions

In this section, we quickly review the major types of chemical reactions. In following sections, we look at these categories in more detail.

A chemical reaction is a process that is usually characterized by a chemical change in which the starting materials (reactants) are different from the products. Chemical reactions tend to involve the motion of electrons, leading to the formation and breaking of chemical bonds. There are several different types of chemical reactions and more than one way of classifying them. Here are some common reaction types:

<u>Direct Combination or Synthesis Reaction</u>
In a synthesis reaction two or more chemical species combine to form a more complex product.

$A + B \longrightarrow AB$

The combination of iron and sulfur to form iron (II) sulfide is an example of a synthesis reaction:

$8\ Fe + S_8 \longrightarrow 8\ FeS$

<u>Chemical Decomposition or Analysis Reaction</u>
In a decomposition reaction a compound is broken into smaller chemical species.

$AB \longrightarrow A + B$

The electrolysis of water into oxygen and hydrogen gas is an example of a decomposition reaction:

$2\ H_2O \longrightarrow 2\ H_2 + O_2$

<u>Single Displacement or Substitution Reaction</u>
A substitution or single displacement reaction is characterized by one element being displaced from a compound by another element.

$A + BC \longrightarrow AC + B$

An example of a substitutions reaction occurs when zinc combines with hydrochloric acid. The zinc replaces the hydrogen:

$Zn + 2\ HCl \longrightarrow ZnCl_2 + H_2$

Double Displacement Reaction or Metathesis

In a double displacement or metathesis reaction two compounds exchange bonds or ions in order to form different compounds.

$AB + CD \longrightarrow AD + CB$

An example of a double displacement reaction occurs between sodium chloride and silver nitrate to form sodium nitrate and silver chloride.

$NaCl(aq) + AgNO_3(aq) \longrightarrow NaNO_3(aq) + AgCl(s)$

Acid-Base Reaction

An acid-base reaction is type of double displacement reaction that occurs between an acid and a base. The H^+ ion in the acid reacts with the OH^- ion in the base to form water and an ionic salt:

$HA + BOH \longrightarrow H_2O + BA$

The reaction between hydrobromic acid (HBr) and sodium hydroxide is an example of an acid-base reaction:

$HBr + NaOH \longrightarrow NaBr + H_2O$

Oxidation-Reduction or Redox Reaction

In a redox reaction the oxidation numbers of atoms are changed. Redox reactions may involve the transfer of electrons between chemical species.

The reaction that occurs when In which I_2 is reduced to I^- and $S_2O_3^{2-}$ (thiosulfate anion) is oxidized to $S_4O_6^{2-}$ provides an example of a redox reaction:

$2\ S_2O_3^{2-}(aq) + I_2(aq) \longrightarrow S_4O_6^{2-}(aq) + 2\ I^-(aq)$

Combustion

A combustion reaction is a type of redox reaction in which a combustible material combines with an oxidizer to form oxidized products and generate heat (exothermic reaction). Usually in a combustion reaction oxygen combines with another compound to form carbon dioxide and water. An example of a combustion reaction is the burning of naphthalene:

$C_{10}H_8 + 12\ O_2 \longrightarrow 10\ CO_2 + 4\ H_2O$

Isomerization

In an isomerization reaction, the structural arrangement of a compound is changed but its net atomic composition remains the same.

Hydrolysis Reaction

A hydrolysis reaction involves water (either as a reactant or a product). The general form for a hydrolysis

reaction is:

$X^-(aq) + H_2O(l) \text{<-->} HX(aq) + OH^-(aq)$

Special Notes on Decomposition Reactions

Some products are unstable and break down (decompose) as they are produced during the reaction. You need to be able to recognize these products when they occur and write the decomposition products in their places. For example, if a carbonate decomposes, one forms water and carbon dioxide not hydrogen gas and CO_3. These need to be memorized.

Examples of common decomposition reactions:

1. Metallic carbonates, when heated, form metallic oxides and $CO_{2(g)}$.

   $CaCO_{3(s)} \longrightarrow CaO_{(s)} + CO_{2(g)}$

2. Most metallic hydroxides, when heated, decompose into metallic oxides and water.

   $Ca(OH)_{2(s)} \longrightarrow CaO_{(s)} + H_2O_{(g)}$

3. Metallic chlorates, when heated, decompose into metallic chlorides and oxygen.

   $2KClO_{3(s)} \longrightarrow 2KCl_{(s)} + 3O_{2(g)}$

4. Some acids, when heated, decompose into nonmetallic oxides and water.

   $H_2SO_4 \longrightarrow H_2O_{(l)} + SO_{3(g)}$

5. Some oxides, when heated, decompose.

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

6. Some decomposition reactions are produced by electricity.

   $2H_2O_{(l)} \longrightarrow 2H_{2(g)} + O_{2(g)}$

   $2NaCl_{(l)} \longrightarrow 2Na_{(s)} + Cl_{2(g)}$

7. Carbonic acid, as in soft drinks, decomposes when it is formed.

   $H_2CO_{3(aq)} \longrightarrow H_2O_{(l)} + CO_{2(g)}$

8. Sulfurous acid also decomposes as it is formed.

   $H_2SO_{3(aq)} \longrightarrow H_2O_{(l)} + SO_{2(g)}$

9. Ammonium hydroxide decomposes as it is formed.

$NH_4OH_{(aq)} \longrightarrow NH_{3(g)} + H_2O_{(l)}$

Special Notes on Replacement Reactions

Note: Refer to the activity series for metals and nonmetals 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. In other words, a more active element takes the place of another element in a compound and sets the less active one free.

Basic form: $A + BX \longrightarrow AX + B$ or $AX + Y \longrightarrow AY + X$

Examples of replacement reactions:

1. Replacement of a metal in a compound by a more active metal.

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

2. Replacement of hydrogen in water by an active metal.

   $2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$

   $Mg_{(s)} + H_2O_{(g)} \longrightarrow MgO_{(s)} + H_{2(g)}$

3. Replacement of hydrogen in acids by active metals.

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

4. Replacement of nonmetals by more active nonmetals.

   $Cl_{2(g)} + 2NaBr_{(aq)} \longrightarrow 2NaCl_{(aq)} + Br_{2(l)}$

Special Notes on Ionic Reactions

Ionic reactions occur between ions in aqueous solution. A reaction will occur when a pair of ions come together to produce at least one of the following:

- a precipitate
- a gas
- water or some other non-ionized substance

Basic form: $AX + BY \longrightarrow AY + BX$

Examples of ionic reactions:

1. Formation of precipitate.

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

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

2. Formation of a gas.

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

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

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

4. Formation of a product which decomposes.

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

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”.

Special Notes about Combustion (of Hydrocarbons) Reactions

When a hydrocarbon is burned with sufficient oxygen supply, the products are always carbon dioxide and water vapor. If the supply of oxygen is low or restricted, then carbon monoxide will be produced.

Special Combustion Terminology:

- Complete combustion means the higher oxidation number is attained.
- Incomplete combustion means the lower oxidation number is attained.
- The phrase "to burn" means to add oxygen unless told otherwise.

Examples of ionic reactions:

1. Hydrocarbon $(C_xH_y) + O_{2(g)} \longrightarrow CO_{2(g)} + H_2O_{(g)}$

2. $CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$

3. $2C_4H_{10(g)} + 13O_{2(g)} \longrightarrow 8CO_{2(g)} + 10H_2O_{(g)}$

Special Notes about Dissociation Reactions

Dissociation is commonly mistaken as decomposition, but there is a difference. When the compound is broken down, it is broken down into ions rather than atoms, so there will be a charge on the product side of the equation. An example of dissociation is as follows:

$$AB \longrightarrow A^{+} + B^{-}$$

Another Summary of Common Types of Chemical Reactions

1. When <u>two elements</u> react, a combination reaction occurs (think: could any <u>other</u> type of reaction occur?), producing a binary compound (that is, one consisting of only two <u>types</u> of atoms). If a <u>metal and a nonmetal react</u>, the product is <u>ionic</u> with a formula determined by the charges on the ions the elements form. If <u>two nonmetals</u> react, the product is a molecule with <u>polar covalent</u> bonds, with a formula consistent with the normal valences of the atoms involved. Some pairs of elements may react only slowly and require heating for significant reaction to occur.

Examples:

$K + S_8 \longrightarrow K_2S$ (ionic)
$Ca + O_2 \longrightarrow CaO$ (ionic)
$Al + I_2 \longrightarrow AlI_3$ (ionic)
$H_2 + O_2 \longrightarrow H_2O$ (covalent)
$I_2 + Cl_2 \longrightarrow ICl, ICl_3$, or ICl_5 (covalent)
(exact product depends on relative amounts of I_2 and Cl_2)

Note: The above reactions are not balanced, nor were they intended to be. They are meant only to show the correct formula for the reactants and products.

2. Reaction of a <u>metal oxide with water</u> produces a <u>metal hydroxide</u>; that is, a strong base. Reaction of a <u>nonmetal oxide with water</u> produces an <u>oxyacid</u> in which the nonmetal is in the same oxidation state as in the oxide you started with. Both of these are combination reactions, and both can be reversed by heating the products. Metal hydroxides decompose on heating to give the metal oxide and water, and oxyacids decompose on heating to give water and the nonmetal oxide in the appropriate oxidation state.

Examples:

$Na_2O + H_2O \longrightarrow NaOH$
$MgO + H_2O \longrightarrow Mg(OH)_2$
$SO_2 + H_2O \longrightarrow H_2SO_3$
$Cl_2O_5 + H_2O \longrightarrow HClO_3$
$HNO_3 \xrightarrow{\Delta} N_2O_5 + H_2O$
$Fe(OH)_3 \xrightarrow{\Delta} Fe_2O_3 + H_2O$

3. Reaction of a <u>metal oxide with a nonmetal oxide gives an oxysalt</u>; reaction of a <u>metal hydroxide with a nonmetal oxide produces a "hydrogen" oxysalt</u>. This is essentially a reaction of the O^{2-} or OH^{-} in the metal compound with the molecular nonmetal oxide. This combination reaction occurs <u>only</u> if no water

is present; in the presence of water, the nonmetal and metal oxides react with the water to produce acid and hydroxide, respectively (as shown in (2) above), then these react as in (4) below.

Examples:

$CaO(s) + SO_3(g) \longrightarrow CaSO_4(s)$
$NaOH(s) + CO_2(g) \longrightarrow NaHCO_3(s)$

4. Reaction of an <u>acid with a base gives a salt plus water</u>. The cation in the salt comes from the base; the anion comes from the acid. The base may be a metal hydroxide, a metal oxide, or a weak base such as NH_3. The acid and/or base may be pure solids, liquids, or gases, or in aqueous solution. The oxidation states of the anion of the acid and cation of the base normally remain unchanged.

Examples:

$HCl(aq) + Ca(OH)_2(aq) \longrightarrow CaCl_2(aq) + H_2O(l)$
$H_2SO_4(aq) + Fe(OH)_3(s) \longrightarrow Fe_2(SO_4)_3(aq) + H_2O(l)$
$NH_3(g) + HC_2H_3O_2(l) \longrightarrow NH_4C_2H_3O_2(s)$
$Al_2O_3(s) + HClO_4(aq) \longrightarrow Al(ClO_4)_3(aq) + H_2O(l)$

5. <u>Ammonium salts</u> react with <u>metal hydroxides and oxides in</u> an acid-base reaction to produce <u>ammonia</u>. This is essentially the reverse of one of the reaction types mentioned in (4) above. Either or both of the reactants may be a pure material or in aqueous solution.

Examples:

$NH_4Cl(aq) + KOH(aq) \longrightarrow NH_3(g) + H_2O(l) + KCl(aq)$
$NH_4NO_3(s) + CaO(s) \longrightarrow NH_3(g) + H_2O(l) + Ca(NO_3)_2(s)$

6. Reaction of the <u>salt of a weak acid</u> (that is, a compound containing the <u>anion of a weak acid</u>) with a <u>strong acid</u> produces the <u>weak acid</u> and a <u>salt</u>. This is another example of an acid-base reaction, in addition to the ones given in (4) and (5) above. The original salt of the weak acid may be either a pure solid or in aqueous solution. The cation in the salt formed as the product comes from the weak acid salt; the anion in the product salt comes from the strong acid. In many cases, the weak acid produced is unstable and decomposes to give the oxide of a nonmetal and water (see (2) above). This is especially true if the nonmetal oxide is a compound of limited solubility in water such as SO_2, CO_2, or the nitrogen oxides. The best-known examples of this type of reaction involve carbonates, bicarbonates, sulfides, and sulfites, but <u>many</u> other examples are known as well. Normally, these reactions do not involve oxidation or reduction.

Examples:

$BaCO_3(s) + HBr(aq) \longrightarrow BaBr_2(aq) + H_2O(l) + CO_2(g)$
$NaHCO_3(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + CO_2(g) + H_2O(l)$
$MgS(s) + HCl(aq) \longrightarrow H_2S(g) + MgCl_2(aq)$
$K_2SO_3(aq) + HNO_3(aq) \longrightarrow KNO_3(aq) + SO_2(g) + H_2O(l)$
$Ca_3(PO_4)_2(s) + HCl(aq) \longrightarrow CaCl_2(aq) + H_3PO_4(aq)$

$Zn(C_2H_3O_2)_2(aq) + HBr(aq) \longrightarrow ZnBr_2(aq) + HC_2H_3O_2(aq)$

7. Reaction of solutions of <u>two soluble salts</u> with one another can give a <u>precipitate</u> of an insoluble salt formed by a double replacement reaction (also called a <u>metathesis</u>). Whether or not a precipitate forms depends on the exact combination of salts used. To make a prediction as to whether a reaction will take place or not, <u>you must know the solubility rules</u> for common salts (Ebbing 4/e, page 104; lab manual, Appendix 7). Some combinations of salts may give oxidation-reduction reactions (see (11) below), but most do not.

Examples:

$CaCl_2(aq) + K_2CO_3(aq) \longrightarrow CaCO_3(s) + KCl(aq)$
$AgNO_3(aq) + FeCl_3(aq) \longrightarrow AgCl(s) + Fe(NO_3)_3(aq)$

but: $NiSO_4(aq) + MgI_2(aq) \longrightarrow$ no reaction

(NiI_2 and $MgSO_4$ are both soluble)
$Al(NO_3)_3(aq) + Pb(C_2H_3O_2)_2(aq) \longrightarrow$ no reaction
($Al(C_2H_3O_2)_3$ and $Pb(NO_3)_2$ are both soluble)

8. <u>Heating an oxysalt</u> produces a <u>metal oxide plus a nonmetal oxide</u> or a <u>metal salt plus oxygen</u>, or some combination of these two decomposition reactions.

Examples:

$KClO_3(s) \xrightarrow{\Delta} KCl(s) + O_2(g)$
$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$
$Pb(NO_3)_2(s) \xrightarrow{\Delta} PbO(s) + NO(g) + NO_2(g) + O_2(g)$

9. <u>Heating a hydrated material</u> initially causes a decomposition reaction to produce the anhydrous compound and water. Further heating may yield further decomposition, depending on the material. (See (2) and (8) above.) Most binary compounds are stable to heat.

Examples:

$H_2C_2O_4 \cdot 2H_2O(s) \xrightarrow{\Delta} H_2O(g) + H_2C_2O_4(s)$; followed by
$H_2C_2O_4(s) \xrightarrow{\Delta} H_2O(g) + CO(g) + CO_2(g)$
$CaCl_2 \cdot 6H_2O(s) \xrightarrow{\Delta} H_2O(g) + CaCl_2(s)$; followed by
$CaCl_2(s) \xrightarrow{\Delta}$ no reaction
$CuSO_4 \cdot 5H_2O(s) \xrightarrow{\Delta} H_2O(g) + CuSO_4(s)$; followed by
$CuSO_4(s) \xrightarrow{\Delta} CuO(s) + SO_3(g)$ (requires strong heating)

10. Reaction of an <u>element with a compound</u> often gives a <u>single replacement</u> reaction in which a nonmetallic element <u>can</u> replace a combined nonmetal, and a metallic element <u>can</u> replace a combined

metal, or hydrogen from an acid. As a general rule, a more active (reactive) element will replace a less active (reactive) element from its compounds. In general (but with many exceptions), the most reactive nonmetals are found to the upper right in the periodic table, and the most reactive metals are found to the lower left. The order of reactivity of the halogens is $F_2>Cl_2>Br_2>I_2$. For hydrogen and the more common metals, the order of reactivity (the activity series) is

Li>K>Ca>Na>Mg>Al>Zn>Cr>Fe>Ni>Sn>Pb>H_2>Cu>Hg>Ag>Pt>Au

In these two series, one element can replace another one to its right in the series. Metals to the left of H_2 can replace H^+ from acids. The very reactive metals (Li, K, Na, Ca) can replace H^+ from cold water; metals of intermediate reactivity (Mg, Al) can replace H^+ from hot water or steam. Any single replacement reaction can also be categorized as an oxidation-reduction (redox) reaction.

Examples:

$Al(s) + NiSO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + Ni(s)$
$Fe(s) + HBr(aq) \longrightarrow FeBr_3(aq) + H_2(g)$
$Cl_2(g) + KI(aq) \longrightarrow KCl(aq) + I_2(s)$
$Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$
$Zn(s) + Cu(NO_3)_2(aq) \longrightarrow Cu(s) + Zn(NO_3)_2(aq)$

but: $Ag(s) + HClO_4(aq) \longrightarrow$ no reaction

$Br_2(l) + ZnCl_2(aq) \longrightarrow$ no reaction
$Sn(s) + H_2O(l) \longrightarrow$ no reaction
$Pb(s) + CrF_3(aq) \longrightarrow$ no reaction

11. Compounds containing one or more atoms in high oxidation states often act as oxidizing agents; compounds containing atoms in low oxidation states often act as reducing agents. For most elements, the (old) group number of the atom in the periodic table gives the highest oxidation state possible for that element. For nonmetals, the lowest oxidation state possible is given by the (old) group number minus eight. Elemental metals most often act as reducing agents (they are oxidized); nonmetals frequently act as oxidizing agents (they are reduced).

For the representative elements (i.e., those in the first two and last six columns of the periodic table), oxidation states most often are two units apart. For example, Sn forms Sn(II) and Sn(IV); Br forms Br^{1-}, Br(I), Br(III), Br(V), and Br(VII). For the transition elements, (i.e., those in the "center" ten columns of the periodic table), oxidation states are often one unit apart, but can be in almost any relationship to one another. For the transition elements, the common oxidation states (charges on their ions) must be memorized. For example, Fe forms Fe^{2+} and Fe^{3+}; Cu forms Cu^+ and Cu^{2+}, etc. Some of the transition elements form oxyanions as well as cations. For example, Mn forms Mn^{2+}, Mn^{3+}, MnO_4^{2-}, and MnO_4^-; Cr forms Cr^{2+}, Cr^{3+}, CrO_4^{2-}, and $Cr_2O_7^{2-}$.

Any atom in its highest possible oxidation state can only act as an oxidizing agent; any atom in its lowest possible oxidation state can only act as a reducing agent. Atoms in intermediate oxidation states can be either oxidized or reduced; that is, they can act as either reducing or oxidizing agents. Some of the

oxidizing agents most commonly encountered are MnO_4^-, CrO_4^{2-}, $Cr_2O_7^{2-}$, HNO_3, H_2O_2, and the halogens. Some of the more common reducing agents are elemental H_2, metals, carbon, and I^-.

In predicting products of oxidation-reduction reactions, <u>oxidation and reduction must occur simultaneously</u>! It is impossible for oxidation to occur without reduction or vice versa.

Examples:

$Sn^{2+}(aq) + F_2(g) \longrightarrow Sn^{4+}(aq) + F^-(aq)$
$Mn^{2+}(aq) + BiO_3^-(aq) \longrightarrow Bi^{3+}(aq) + MnO_4^-(aq)$
(note that the Bi is in its highest possible oxidation state in BiO_3^-)
$K(s) + P_4O_{10}(s) \longrightarrow K_3PO_3(s)$
(note that P is reduced from P(V) to P(III))
$MnO_4^-(aq) + I^-(aq) \longrightarrow Mn^{2+}(aq) + I_2(aq)$
$CuS(s) + HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + S_8(s) + NO_2(g)$
(note $S^{2-} \longrightarrow S^0$ and N(V) $\longrightarrow$ N(IV))
$Fe_2O_3(s) + C(s) \longrightarrow CO_2(g) + Fe(s)$

Practice Problems

In this section, a variety of practice problems are provided. Write the formula for each material correctly and then balance the equation. There are some reactions that require completion. For each reaction tell what type of reaction it is.

1. Sulfur trioxide and water combine to make sulfuric acid.

2. Lead II nitrate and sodium iodide react to make lead iodide and sodium nitrate.

3. Calcium fluoride and sulfuric acid make calcium sulfate and hydrogen fluoride (hydrofluoric acid)

4. Calcium carbonate will come apart when you heat it to leave calcium oxide and carbon dioxide.

5. Ammonia gas when it is pressed into water will make ammonium hydroxide.

6. Sodium hydroxide neutralizes carbonic acid

7. Zinc sulfide and oxygen become zinc oxide and sulfur.

8. Lithium oxide and water make lithium hydroxide

9. Aluminum hydroxide and sulfuric acid neutralize to make water and aluminum sulfate.

10. Sulfur burns in oxygen to make sulfur dioxide.

11. Barium hydroxide and sulfuric acid make water and barium sulfate.

12. Aluminum sulfate and calcium hydroxide become aluminum hydroxide and calcium sulfate.

13. Copper metal and silver nitrate react to form silver metal and copper II nitrate.

14. Sodium metal and chlorine react to make sodium chloride.

15. Calcium phosphate and sulfuric acid make calcium sulfate and phosphoric acid.

16. Phosphoric acid plus sodium hydroxide.

17. Propane burns (with oxygen)

18. Zinc and copper II sulfate yield zinc sulfate and copper metal

19. Sulfuric acid reacts with zinc

20. Acetic acid ionizes.

21. Methane is reacted with steam to get hydrogen and carbon dioxide

22. Calcium oxide and aluminum make aluminum oxide and calcium

23. Chlorine gas and sodium bromide yield sodium chloride and bromine

Answers to Practice Problems

1. $SO_3 + H_2O \longrightarrow H_2SO_4$	SYNTHESIS
2. $Pb(NO_3)_2 + 2NaI \longrightarrow PbI_2\downarrow + 2NaNO_3$	DOUBLE REPLACEMENT
3. $CaF_2 + H_2SO_4 \longrightarrow CaSO_4\downarrow + 2\ HF$	DOUBLE REPLACEMENT
4. $CaCO_3 \longrightarrow CaO + CO_2$	DECOMPOSITION
5. $NH_3 + H_2O \longrightarrow NH_4OH$	SYNTHESIS
6. $2\ NaOH + H_2CO_3 \longrightarrow Na_2CO_3 + 2\ H_2O$	DOUBLE REPLACEMENT OR ACID-BASE NEUTRALIZATION
7. $2\ ZnS + O_2 \longrightarrow 2\ ZnO + 2\ S$	ANIONIC SINGLE REPLACEMENT
8. $Li_2O + H_2O \longrightarrow 2\ LiOH$	SYNTHESIS
9. $2\ Al(OH)_3 + 3\ H_2SO_4 \longrightarrow 6\ H_2O + Al_2(SO_4)_3$	DOUBLE REPLACEMENT OR ACID-BASE NEUTRALIZATION
10. $S + O_2 \longrightarrow SO_2$	SYNTHESIS

11. $Ba(OH)_2 + H_2SO_4 \longrightarrow 2\ H_2O + BaSO_4$	DOUBLE REPLACEMENT OR ACID-BASE NEUTRALIZATION
12. $Al_2(SO_4)_3 + 3\ Ca(OH)_2 \longrightarrow 2\ Al(OH)_3\downarrow + 3\ CaSO_4\downarrow$	DOUBLE REPLACEMENT
13. $Cu + 2AgNO_3 \longrightarrow 2Ag + Cu(NO_3)_2$	CATIONIC SINGLE REPLACEMENT
14. $2Na + Cl_2 \longrightarrow 2\ NaCl$	SYNTHESIS
15. $Ca_3(PO_4)_2 + 3\ H_2SO_4 \longrightarrow 3\ CaSO_4 + 2\ H_3PO_4$	DOUBLE REPLACEMENT
16. $H_3(PO_4) + 3\ NaOH \longrightarrow Na_3PO_4 + 3\ H_2O$	DOUBLE REPLACEMENT (NEUTRALIZATION)
17. $C_3H_8 + 5\ O_2 \longrightarrow 4\ H_2O + 3\ CO_2$	BURNING OF A HYDROCARBON
18. $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$	CATIONIC SINGLE REPLACEMENT
19. $H_2SO_4 + Zn \longrightarrow ZnSO_4 + H_2$	CATIONIC SINGLE REPLACEMENT
20. $HC_2H_3O_2 \rightleftharpoons H^+ + (C_2H_3O_2)^-$	IONIZATION (REVERSIBLE)
21. $2\ H_2O + CH_4 \longrightarrow 4\ H_2 + CO_2$	WATER-GAS SHIFT
22. $3\ CaO + 2\ Al \longrightarrow Al_2O_3 + 3\ Ca$	CATIONIC SINGLE REPLACEMENT
23. $Cl_2 + 2\ NaBr \longrightarrow 2\ NaCl + Br_2$	ANIONIC SINGLE REPLACEMENT