

Chapter 8 – Chemical Reactions

Part 1 – Notes: Describing Chemical Change

Objectives: Identify, define, and explain: chemical equation, chemical statement, catalyst, coefficient, balanced equation, yield, reactant, product, chemical reaction, diatomic element, Law of Conservation of Mass.
Differentiate between chemical and algebraic equations and between chemical reactions and chemical equations.
List various common indicators of a chemical reaction.
Write equations describing chemical reactions using appropriate symbols.
Write and balance chemical equations given names or formulas of reactants and products in the reaction.
List elements that exist as diatomic molecules in their free state.

Text Reference: Section 8.1 – pages 203-211

A chemical **reaction** is the process by which one or more substances change into new substance(s).

Example: solid calcium reacts with solid sulfur to form solid calcium sulfide.

A chemical **equation** is the written representation, in symbolic terms, of the overall change in a chemical reaction on an atomic level.

Example: $\text{Ca}_{(s)} + \text{S}_{(s)} \rightarrow \text{CaS}_{(s)}$

Observations that indicate a chemical reaction may have occurred include (but are not limited to):

A **reactant** is (one of) the initial substance(s) in a chemical reaction; it reacts to form something new. Reactants are always found on the **left** side of the equation.

A **product** is (one of) the resulting substance(s) in a chemical reaction; it is produced from the reactant(s). Products are always found on the **right** side of the equation. The product is always different from the reactant(s) and the properties of a product are different than the properties of the constituent reactant(s).

A “+” represents the combining of two substances in a chemical reaction. It is commonly read *plus, combines with, or reacts with*.

A “→” represents the split in the equation between reactants and products; it points in the direction of the final result. It is commonly read *yields, forms, breaks down into, combines to form, or decomposes into*. It is NEVER read as “equal.”

(s), (l), (g), and (aq) indicate . . .

Adding heat to an equation is indicated with . . .

A **catalyst** is represented by A catalyst is . . .

Chemical equations are said to be balanced; this means there are an equal number of atoms of each type of element present on opposite sides of the yield sign. Symbolic representations of chemical reactions need to be balanced in order to be a true *chemical equation*. If a symbolic representation is not balanced, it is called a *chemical statement* or a *skeleton equation*.

Law of Conservation of Mass –

When balancing equations, you must make sure that MASS is conserved and that there are the SAME number of each type of atom on both sides of the yield. Be careful not to confuse the subscripts and the coefficients when balancing equations. **NEVER change the subscript in a formula in an attempt to balance an equation.** Subscripts of a compound are there to balance the charges of the ions of indicate the correct number of atoms covalently bonded. The correct compound cannot be changed to get an equation to balance.

KEEP IN MIND THAT WHEN YOU USE COEFFICIENTS TO BALANCE EQUATIONS, THEY MUST BE IN THE SMALLEST WHOLE NUMBER RATIO.

Note: In this class, you will be given directions telling you to balance the chemical *equations*. In actuality, you are balancing *chemical statements* and making them *equations*.

Diatomic molecules: two atoms of the same element that form molecules when they are in a *free* or uncombined state

Elements that form diatomic molecule in their free state:

Simple Rules to Follow for Basic Balancing:

1. Balance metallic elements first.
2. Balance nonmetallic elements next, except oxygen and hydrogen.
3. Balance hydrogen and oxygen last.

Writing and Balancing Chemical Equations

1. Convert from compound names to compound formulas.
- Use the correct subscripts to make the formula correct. Once the compounds are correctly written, **DO NOT** change them!
2. Make sure the elements that form diatomic molecules are written with the subscript of two.
3. Balance the chemical statement with whole numbers coefficients in the smallest whole number ratio.
- Do NOT change the subscripts in the compound formulas.
- Simply add coefficients to illustrate the conservation of mass in the balanced chemical equation.

Chapter 8 – Chemical Reactions

Part 1 – Assignment: Describing Chemical Change

Balance the following chemical equations using coefficients with the smallest whole number ratio.

1. $\text{NH}_4\text{NO}_2 \rightarrow \text{H}_2\text{O} + \text{N}_2$
2. $\text{H}_2 + \text{N}_2 \rightarrow \text{NH}_3$
3. $\text{MgCO}_3 \rightarrow \text{MgO} + \text{CO}_2$
4. $\text{P}_4 + \text{Cl}_2 \rightarrow \text{PCl}_5$
5. $\text{CrO}_3 \rightarrow \text{Cr}_2\text{O}_3 + \text{O}_2$
6. $\text{IF}_5 + \text{H}_2\text{O} \rightarrow \text{HF} + \text{HIO}_3$
7. $\text{NH}_3 + \text{O}_2 \rightarrow \text{NO} + \text{H}_2\text{O}$
8. $\text{HgSO}_4 + \text{Hg} + \text{NaCl} \rightarrow \text{Hg}_2\text{Cl}_2 + \text{Na}_2\text{SO}_4$
9. $\text{HBrO}_3 \rightarrow \text{Br}_2\text{O}_5 + \text{H}_2\text{O}$
10. $\text{NO}_2 + \text{H}_2\text{O} \rightarrow \text{HNO}_3 + \text{NO}$
11. $\text{NH}_4\text{NO}_3 \rightarrow \text{N}_2\text{O} + \text{H}_2\text{O}$
12. $\text{B}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{BO}_3$

Write and balance the following equations.

13. calcium phosphate + aluminum sulfate \rightarrow calcium sulfate + aluminum phosphate
14. potassium hydroxide \rightarrow potassium oxide + water
15. potassium iodide + lead (II) nitrate \rightarrow potassium nitrate + lead (II) iodide
16. iron (III) oxide + carbon \rightarrow carbon monoxide + iron

Chapter 8 – Chemical Reactions

Part 2 – Notes: Classifying Reactions

Objectives: Identify, define, and explain: combination reaction, synthesis reaction, decomposition reaction, single replacement reaction, double replacement reaction, combustion reaction, and rapid oxidation.
Explain the importance of classifying reactions.
Identify and classify a reaction as synthesis, decomposition, single replacement, double replacement, or combustion.

Text Reference: Section 8.2 (part) – pages 212-221

It is useful to classify chemical reactions because . . .

There are FIVE basic types of reactions that we will need to be able to classify.

A **synthesis** reaction is one in which a product is being created (or synthesized) from two or more elements. It is also a reaction where a more complex compound is created from two or more simpler complexes. There are always two or more reactants but only *one* product. Synthesis reactions are also known as **composition** reactions or **combination** reactions.

Example:

In other words:

A **decomposition** reaction is one in which the single reactant is broken down into two or more elements or simpler compounds. It is the reverse of a synthesis reaction. It has only *one* reactant but two or more products. In such a reaction, the single reactant is decomposed into its constituent parts.

Example:

In other words:

A **single replacement** reaction is one in which an element that is a reactant replaces an element in the other reactant. In other words, *one* partner is switched. Key point: a resulting ionic compound must have a + and a – ion. The switch cannot result in two positive ions or two negative ions forming a compound.

Example:

In other words:

A **double replacement** reaction (also known as a double displacement reaction or metathesis reaction) is one in which the positive ions of the two reactants switch places. As products, each ion has a new and different partner that it originally had.

Example:

In other words:

There are 3 types of products that drive double replacements reactions. They are:

A **combustion** reaction is a **rapid oxidation** that usually produced a flame. For our purposes (but not always), oxygen (from the air) is a reactant. **Hydrocarbons** are *compounds that contain hydrogen and carbon*. (Combustion reactions may also include hydrocarbons that also contain oxygen.) When hydrocarbons combust (react with oxygen), the products are *always* carbon dioxide and water vapor.

Example:

In other words:

Chapter 8 – Chemical Reactions

Part 2 – Assignment: Classifying Reactions

For each of the following, write a balanced chemical equation. Then, classify the reaction as SYN, DEC, SR, DR, or COMB.

- _____ 1. iron + sulfur \rightarrow iron (II) sulfide
- _____ 2. zinc + copper (II) sulfate \rightarrow zinc sulfate + copper
- _____ 3. silver nitrate + sodium bromate \rightarrow silver bromate + sodium nitrate
- _____ 4. potassium chlorate \rightarrow potassium chloride + oxygen
- _____ 5. mercury (II) oxide \rightarrow mercury + oxygen
- _____ 6. aluminum sulfate + sodium bicarbonate \rightarrow aluminum hydroxide + sodium sulfate + carbon dioxide
- _____ 7. magnesium nitride + water \rightarrow magnesium hydroxide + ammonia (NH₃)
- _____ 8. copper (II) oxide + ammonia (NH₃) \rightarrow copper + water + nitrogen
- _____ 9. iron + water \rightarrow iron (III) oxide + hydrogen
- _____ 10. iron (III) chloride + potassium hydroxide \rightarrow potassium chloride + iron (III) hydroxide
- _____ 11. aluminum + hydrogen sulfate \rightarrow aluminum sulfate + hydrogen
- _____ 12. aluminum sulfate + calcium hydroxide \rightarrow aluminum hydroxide + calcium sulfate

13. What is the purpose of a catalyst?
14. Explain why it is useful to classify reaction by their type.
15. How do you predict the correct formula for the combination reaction between a nonmetal and a Group 2 metal?
16. What is characteristic for every synthesis reaction?

Chapter 8 – Chemical Reactions

Part 3 – Notes: Predicting Products of Reactions

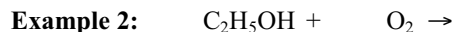
Objectives: Predict the products of synthesis, decomposition, single replacement, double replacement, and combustion reactions.
List three types of products that drive double replacement reactions.

Text Reference: Section 8.2 pages – (212-221) and 222-224

Since you are able to classify a reaction, you will use the patterns from the classifications to predict the products of the reaction.

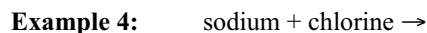
Combustion Reactions:

The products of the combustion of a hydrocarbon are *always* carbon dioxide and water vapor.



Synthesis Reactions:

Since a synthesis reaction is one where two or more elements (or simple compounds) combine to form a single compound, the resulting product must contain all elements/ions present in the reactants without any leftover or additional elements/ions. The product is a single product with balanced charges and the equation must be able to be balanced with correct coefficients.



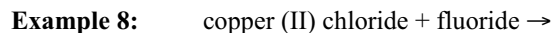
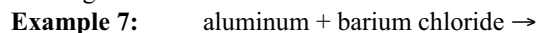
Decomposition Reactions:

Since a decomposition reaction is one where a single reactant breaks into two or more elements (or simple compounds), the resulting products must contain all elements/ions present in the reactant with no leftover. The rules are the same for a synthesis – since they are opposite reactants.



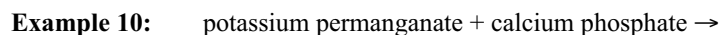
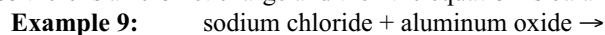
Single Replacement Reactions:

In a single replacement reaction, either the positive or negative ion of one reactant is replaced by the other reactant (an element that forms either a positive or negative ion). Only positive ions may replace other positive ions and only negative ions may replace other negative ions.



Double Replacement Reactions (Metathesis Reactions):

In a double replacement reaction, each reactant ion gets a new partner; the positive ions switch thereby giving everyone a new partner. There are two reactants, each compound is made of two ions (one positive and one negative). When positive ions switch, each positive ion must end up with a negative ion (not one to which it had been previously attached). The compounds are written so there is a zero net charge and then the equation is balanced.



What types of products drive double replacement reactions?

Chapter 3 – Chemical Reactions

Part 3 – Assignment: Predicting Products of Reactions

For each of the following, classify the reaction, predict the products, and balance the equation.

1. $\text{KNO}_3 \rightarrow$
2. $\text{Fe}_2\text{O}_3 + \text{H}_2 \rightarrow$
3. $\text{Zn} + \text{O}_2 \rightarrow$
4. $\text{CuS} + \text{KF} \rightarrow$
5. $\text{Al} + \text{Cl}_2 \rightarrow$
6. $\text{Ba}(\text{ClO}_3)_2 + \text{NaOH} \rightarrow$
7. $\text{Fe} + \text{O}_2 \rightarrow$
8. $\text{AuCl}_3 \rightarrow$
9. $\text{ZnI}_2 + \text{Br}_2 \rightarrow$
10. $\text{NH}_3 \rightarrow$

For each of the following, write the reactants, predict the products, classify the reaction, and balance the equation.

11. sodium oxide + iodine \rightarrow
12. calcium carbonate \rightarrow
13. calcium sulfate + sodium nitrate \rightarrow
14. hydrogen sulfite + calcium \rightarrow
15. aluminum oxide + magnesium fluoride \rightarrow

Chapter 8 – Chemical Reactions

Part 4A – Notes: Predicting Products and the Activity Series

Objectives: Define and explain activity series.
Use the activity series to predict if a reaction will occur.

Text Reference: Section 8.2 – page 217

In order to correctly predict the products of a single replacement reaction, the chemical activity of the elements that will potentially switch places needs to be considered. In order for an element to replace another, it must be more active.

A more active element will replace a less active element. If this is not the case, then no reaction occurs.

Examine the activity series and note any patterns:

ACTIVITY SERIES

METALS		NONMETALS
Lithium	More	
Potassium	Active	
Calcium	↓	
Sodium	↓	
Magnesium	↓	Fluorine
Aluminum	↓	
Zinc	↓	
Chromium	↓	Chlorine
Iron	↓	
Nickel	Decreasing	
Tin	Activity	Bromine
Lead	↓	
HYDROGEN	↓	
Copper	↓	Iodine
Mercury	↓	
Silver	↓	
Platinum	Less	
Gold	Active	

For example, for a set of reactants: **magnesium** and **copper sulfate**, you need to determine if the magnesium will replace the copper ion. You check the activity series. You find that *the magnesium is more active than the copper* and therefore the magnesium will replace the copper and the reaction will proceed.

Examples: Will the following reactions occur? If yes, what are the products? If no, write *NR*.

1. chromium + lead (II) chloride
2. zinc + potassium hydroxide
3. magnesium + sulfuric acid
4. iodine + sodium chloride
5. fluorine + sodium chloride

Chapter 8 – Chemical Reactions
Part 4B – Notes: Aqueous Reactions

Objectives: Identify, define, and explain: complete ionic equation, spectator ion, net ionic equation, solubility rules, aqueous reaction, and precipitate.
 Write and balance complete ionic equations and net ionic equations.
 Use solubility rules to predict the precipitate formed in a double replacement aqueous reaction.

Text Reference: Section 8.3 – pages 225-228

Deciding whether a double replacement reaction will occur naturally is actually a matter of predicting whether an insoluble product can form. Solubility rules are used to determine this. The solubility rules are a listing of the solubility of various substances in water. Some compounds are listed as soluble and other compounds are listed as insoluble. If one of the products formed is insoluble, it is this insoluble product that is the precipitate that forms in the reaction. If no insoluble precipitate forms, then there is no aqueous reaction.

Negative Ion (Anion)	Plus	Positive Ion (Cation)	Cmpd Is. . .	
Any anion	+	Alkali metal ions	"	Soluble
Any anion	+	Ammonium ion (NH ₄ ⁺)	"	Soluble
Nitrate (NO ₃ ⁻¹)	+	Any cation	"	Soluble
Acetate (C ₂ H ₃ O ₂ ⁻¹)	+	Any cation <i>except</i> Ag ⁺¹	"	Soluble
Chloride (Cl ⁻¹) or Bromide (Br ⁻¹) or Iodide (I ⁻¹)	+	Ag ⁺¹ , Pb ⁺² , Hg ₂ ⁺² , or Cu ⁺¹	"	Not soluble
	+	Any other cation		Soluble
Sulfate (SO ₄ ⁻¹)	+	Ca ⁺² , Sr ⁺² , Ba ⁺² , Ra ⁺² , Ag ⁺¹ , or Pb ⁺²	"	Not soluble
	+	Any other cation		Soluble
Sulfide (S ⁻²)	+	Alkali ions or ammonium	"	Soluble
	+	Be ⁺² , Mg ⁺² , Ca ⁺² , Sr ⁺² , Ba ⁺² , or Ra ⁺²		Soluble
	+	Any other cation		Not soluble
Hydroxide (OH ⁻¹)	+	Alkali ion or ammonium	"	Soluble
	+	Sr ⁺² , Ba ⁺² , or Ra ⁺²		Slightly soluble
	+	Any other cation		Not soluble
Phosphate (PO ₄ ⁻³) or Carbonate (CO ₃ ⁻²) or Sulfite (SO ₃ ⁻²)	+	Alkali ions or ammonium	"	Soluble
	+	Any other cation		Not soluble

Examples:

- Mix silver nitrate and sodium bromide. Does a reaction occur?
- Mix sulfuric acid and barium chloride. Does a reaction occur?
- Mix lithium chloride and copper (II) acetate. Does a reaction occur?

Chapter 8 – Chemical Reactions

Part 4 – Assignment: Writing and Balancing Reactions

For each of the following, write the reactant(s), formula, predict and write the product formula, then classify and balance.

1. potassium phosphate + magnesium chloride →
2. copper (II) nitrate + ammonium hydroxide →
3. lead (II) hydroxide →
4. ammonium sulfide + lead (II) nitrate →
5. sodium hydroxide + calcium nitrate →
6. water + diphosphorus pentoxide → phosphoric acid
7. chromium (II) chloride + sulfuric acid →
8. calcium chloride + chromium (III) nitrate →
9. iron (III) hydroxide →
10. potassium chlorate + lead (II) nitrate →
11. aluminum fluoride + sulfuric acid →
12. zinc arsenate + hydrochloric acid →
13. sodium carbonate →
14. copper (II) chlorate →
15. lithium + water →

The following reactants are mixed in water. Write and balance the following as complete and net ionic equations.

16. copper (II) sulfate reacts with iron (III) chloride in an aqueous environment
17. iron (III) sulfate reacts with ammonium sulfide in an aqueous environment
18. potassium carbonate reacts with strontium nitrate in an aqueous environment
19. sodium iodide reacts with silver nitrate in an aqueous environment
20. aluminum sulfate reacts with calcium chloride in an aqueous environment

Chapter 8 – Chemical Reactions

Part 5 – Notes and Questions: Endothermic and Exothermic Reactions

- Objectives:** Identify, define, and explain: endothermic reaction, exothermic reaction, and relative energy changes.
Explain the relative energies of reactants versus products of endothermic and exothermic reactions.
Classify a reaction as endothermic or exothermic.
Qualitatively describe the change to the system and the surroundings in an endothermic or exothermic reaction.

Text Reference: Section 11.1 – pages 293-295

Reactions may be endothermic or exothermic.

1. What is an endothermic reaction?

How do you know when a chemical equation represents an endothermic reaction?

2. What happens to the reaction system and the surroundings during an endothermic reaction? (Respond in terms of heat.)

3. Discuss the relative energies for reactants and products in an endothermic reaction.

4. What is an exothermic reaction?

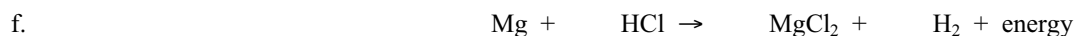
How do you know when a chemical equation represents an exothermic reaction?

5. What happens to the reaction system and the surroundings during an exothermic reaction? (Respond in terms of heat.)

6. Discuss the relative energies for reactants and products in an exothermic reaction.

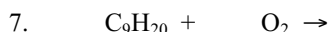
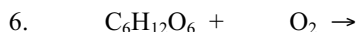
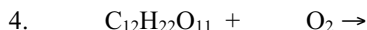
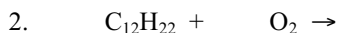
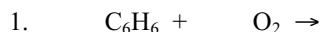
7. Based on energetic considerations only, which is a more probable reaction type: endothermic or exothermic? Why?

8. Classify the following as endothermic or exothermic. Then balance the equation.



Chapter 8 – Chemical Reactions
Part 5 – Assignment: Combustion Reactions

For each of the following sets of reactants, write the products and balance the equation.



9. Text problem – page 234 #65. Alkanes are hydrocarbon molecules that have the general formula C_nH_{2n+2} . The graph shows the number of oxygen, carbon dioxide, and water molecules needed to balance the equations for the complete combustion of every alkane having from one to ten carbon atoms.

Combustion of an alkane: Alkane + $\underline{\hspace{1cm}}$ $O_2 \rightarrow \underline{\hspace{1cm}}$ $CO_2 + \underline{\hspace{1cm}}$ H_2O

a. Use the graph to write balanced equations for the combustion of C_5H_{12} and C_9H_{20} .

b. Extrapolate the graph and write balanced equations for the combustion of $C_{12}H_{26}$ and $C_{17}H_{36}$.

c. The coefficient for O_2 in the general equation is $n + \sqrt{[(n + 1)/2]}$
 What are the coefficients for CO_2 and H_2O , with respect to n ?

