AP Chemistry Summer Assignment

The summer assignment for AP Chemistry consists of several chapters. You must read each review chapter and complete all exercises within each review chapter. You may write the answers on notebook paper. Head each section with the exercise name, for example Exercise 1.1. Some answers consist of one word, whereas some require math, so show all work necessary. The assigned problems are due at the end of the first week of school. There are no excuses or exceptions. These assignments are preview of concepts you will be expected to master. If you have difficulty with these concepts, seek help from YouTube videos such as Khan Academy, Tyler DeWitt, Bozeman Science or Issac Teach. If you are still having trouble, you may want to reconsider your course selection. Again, these concepts are a preview.

Also included is a list of polyatomic ions that you must know on the first day. There will be a quiz! The list also contains some suggestions for making the process of memorization easier. Remember that most monoatomic ions have charges that are directly related to their placement on the periodic table. There are naming patterns that greatly simplify the learning of the polyatomic ions. Start memorizing them now, inevitability there will be some students who will procrastinate and try to cram and may be successful on the quiz, but they will forget the ions and struggle each time they need to be recalled.

I look forward to seeing you all at the beginning of the school year. Best of luck to you all!

Common Ions and Their Charges

A mastery of the common ions, their formulas and their charges, is essential to success in AP Chemistry. You are expected to know all of these ions on the first day of class, when I will give you a quiz on them. You will always be allowed a periodic table, which makes indentifying the ions on the left "automatic." For tips on learning these ions, see the opposite side of this page.

From the table:	
Cations	**Name**
H^+	Hydrogen
Li^+	Lithium
Na^+	Sodium
K^+	Potassium
Rb^+	Rubidium
Cs^+	Cesium
Be^{2+}	Beryllium
Mg^{2+}	Magnesium
Ca^{2+}	Calcium
Ba^{2+}	Barium
Sr^{2+}	Strontium
Al^{3+}	Aluminum
Anions	**Name**
H^-	Hydride
F^-	Fluoride
Cl^-	Chloride
Br^-	Bromide
I^-	Iodide
O^{2-}	Oxide
S^{2-}	Sulfide
Se^{2-}	Selenide
N^{3-}	Nitride
P^{3-}	Phosphide
As^{3-}	Arsenide
Type II Cations	**Name**
Fe^{3+}	Iron(III)
Fe^{2+}	Iron(II)
Cu^{2+}	Copper(II)
Cu^+	Copper(I)
Co^{3+}	Cobalt(III)
Co^{2+}	Cobalt(II)
Sn^{4+}	Tin(IV)
Sn^{2+}	Tin(II)
Pb^{4+}	Lead(IV)
Pb^{2+}	Lead(II)
Hg^{2+}	Mercury(II)

Ions to Memorize	
Cations	**Name**
Ag^+	Silver
Zn^{2+}	Zinc
Hg_2^{2+}	Mercury(I)
NH_4^+	Ammonium
Anions	**Name**
NO_2^-	Nitrite
NO_3^-	Nitrate
SO_3^{2-}	Sulfite
SO_4^{2-}	Sulfate
HSO_4^-	Hydrogen sulfate (bisulfate)
OH^-	Hydroxide
CN^-	Cyanide
PO_4^{3-}	Phosphate
HPO_4^{2-}	Hydrogen phosphate
$H_2PO_4^-$	Dihydrogen phosphate
NCS^-	Thiocyanate
CO_3^{2-}	Carbonate
HCO_3^-	Hydrogen carbonate (bicarbonate)
ClO^-	Hypochlorite
ClO_2^-	Chlorite
ClO_3^-	Chlorate
ClO_4^-	Perchlorate
BrO^-	Hypobromite
BrO_2^-	Bromite
BrO_3^-	Bromate
BrO_4^-	Perbromate
IO^-	Hypoiodite
IO_2^-	iodite
IO_3^-	iodate
IO_4^-	Periodate
$C_2H_3O_2^-$	Acetate
MnO_4^-	Permanganate
$Cr_2O_7^{2-}$	Dichromate
CrO_4^{2-}	Chromate
O_2^{2-}	Peroxide
$C_2O_4^{2-}$	Oxalate
NH_2^-	Amide
BO_3^{3-}	Borate
$S_2O_3^{2-}$	Thiosulfate

Tips for Learning the Ions

"From the Table"

These are ions can be organized into two groups.

1. Their place on the table suggests the charge on the ion, since the neutral atom gains or loses a predictable number of electrons in order to obtain a noble gas configuration. This was a focus in first year chemistry, so if you are unsure what this means, get help BEFORE the start of the year.
   a. All Group 1 Elements (alkali metals) lose one electron to form an ion with a 1+ charge
   b. All Group 2 Elements (alkaline earth metals) lose two electrons to form an ion with a 2+ charge
   c. Group 13 metals like aluminum lose three electrons to form an ion with a 3+ charge
   d. All Group 17 Elements (halogens) gain one electron to form an ion with a 1- charge
   e. All Group 16 nonmetals gain two electrons to form an ion with a 2- charge
   f. All Group 15 nonmetals gain three electrons to form an ion with a 3- charge

   Notice that cations keep their name (sodium ion, calcium ion) while anions get an "-ide" ending (chloride ion, oxide ion).

2. Metals that can form more than one ion will have their positive charge denoted by a roman numeral in parenthesis immediately next to the name of the

Polyatomic Anions

Most of the work on memorization occurs with these ions, but there are a number of patterns that can greatly reduce the amount of memorizing that one must do.

1. "ate" anions have one more oxygen then the "ite" ion, but the same charge. If you memorize the "ate" ions, then you should be able to derive the formula for the "ite" ion and vice-versa.
   a. sulfate is SO_4^{2-}, so sulfite has the same charge but one less oxygen (SO_3^{2-})
   b. nitrate is NO_3^-, so nitrite has the same charge but one less oxygen (NO_2^-)

2. If you know that a sufate ion is SO_4^{2-} then to get the formula for hydrogen sulfate ion, you add a hydrogen ion to the front of the formula. Since a hydrogen ion has a 1+ charge, the net charge on the new ion is less negative by one.
   a. Example:

PO_4^{3-}	→	HPO_4^{2-}	→	$H_2PO_4^-$
phosphate		hydrogen phosphate		dihydrogen phosphate

3. Learn the hypochlorite → chlorite → chlorate → perchlorate series, and you also know the series containing iodite/iodate as well as bromite/bromate.
   a. The relationship between the "ite" and "ate" ion is predictable, as always. Learn one and you know the other.
   b. The prefix "hypo" means "under" or "too little" (think "hypodermic", "hypothermic" or "hypoglycemia")
      i. Hypochlorite is "under" chlorite, meaning it has one less oxygen
   c. The prefix "hyper" means "above" or "too much" (think "hyperkinetic")
      i. the prefix "per" is derived from "hyper" so perchlorate (hyperchlorate) has one more oxygen than chlorate.
   d. Notice how this sequence increases in oxygen while retaining the same charge:

ClO^-	→	ClO_2^-	→	ClO_3^-	→	ClO_4^-
hypochlorite		chlorite		chlorate		perchlorate

Chemical reactions

·6·

When elements or compounds are placed together with other elements or compounds, they may or may not react with one another. Most reactions require the addition of energy to break the existing bonds in the reactants so new bonds can be made to form products. This added energy is called the *activation energy*. Besides having the right energy, the species reacting together must also collide with one another in the proper orientation to react. During a reaction there may be several collisions that occur that cause intermediate species to form. The accounting for these intermediate species is called the *reaction mechanism*. Even if two substances known to react are put together, a reaction may not occur.

When the substances do react, we write an equation to represent what occurs. An equation is an accounting of the atoms involved in the reaction. Since atoms are not created or destroyed in regular reactions (only nuclear ones can do that), the number of atoms on each side of the equation must remain the same—in other words, the equation must be balanced. This is called conservation of atoms (or conservation of mass). The *reactants* are what you start with and are found on the left side of the equation. After the reactants, an arrow with its head pointed toward the right is written. The *products* are what are formed in the reaction and are found on the right side of this arrow. The arrow can be read as "to form" or "yields," but in mathematical terms it is an equation, since all atoms listed on the left must equal those on the right. Note that it is always the *arrangement* of the atoms that changes in a chemical reaction, *not* the total number. There are a few basic steps to follow when writing chemical-reaction equations.

Writing and balancing chemical equations

First, write all the formulas for the elements and compounds involved, using the rules previously learned in Chapter 5. Once these are written, keep in mind *the formula cannot change.* Note it is imperative to know the *symbols* for the atoms of each element. The resulting equation with all the formulas in place is called a *skeleton equation.*

Second, count the number of each atom on each side of the equation. If they are not the same, add coefficients in front of the formulas to balance.

What are coefficients? They are numbers you put *in front* of a compound or element to say how many are needed, just like in a cooking recipe: 2 cups of flour. The coefficient is the amount and the formula is the compound. In the flour example, where the flour is the compound, the coefficient is the 2. In a chemical reaction, when you need two sodium chloride formula units you write 2NaCl. Drawing a picture with each type of atom represented can help make sure all atoms are accounted for on each side of the equation.

Sample Problem 1: Sodium phosphate solution is added to a solution of silver nitrate, forming a solution of sodium nitrate and a precipitate of silver phosphate.

First, write all the formulas:

Sodium is Na^+ and phosphate is PO_4^{3-}, making sodium phosphate $Na_3PO_4(aq)$.
Silver is Ag^+ and nitrate is NO_3^-, making silver nitrate $AgNO_3(aq)$.
Sodium is Na^+ and nitrate is NO_3^-, making sodium nitrate $NaNO_3(aq)$.
Silver is Ag^+ and phosphate is PO_4^{3-}, making silver phosphate $Ag_3PO_4(s)$.

Next, write all the formulas in an equation and balance the number of each atom on each side of the equation. The skeleton equation without coefficients is given in Figure 6.1.

$$Na_3PO_4(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + Ag_3PO_4(s)$$

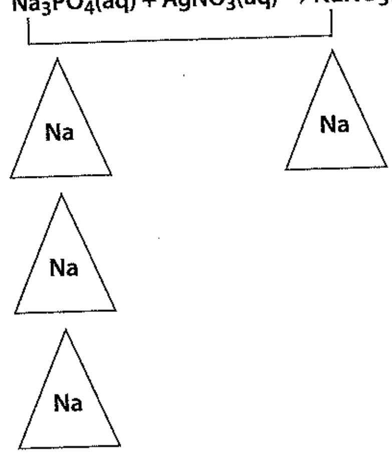

Figure 6.1

There are three sodium atoms on the left and only one sodium atom on the right. Adding a coefficient of 3 to the sodium nitrate on the right side of the equation balances the sodium atoms.

$$Na_3PO_4(aq) + AgNO_3(aq) \rightarrow 3NaNO_3(aq) + Ag_3PO_4(s)$$

The formula of a compound cannot change, only the number of units present.

Looking at the phosphorus atoms (see Figure 6.2), there is one on each side, so no additional coefficient is needed.

$$Na_3PO_4(aq) + AgNO_3(aq) \rightarrow 3NaNO_3(aq) + Ag_3PO_4(s)$$

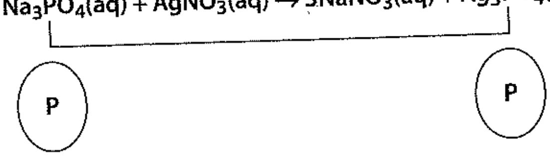

Figure 6.2

Next, look at the silver atoms (in Figure 6.3).

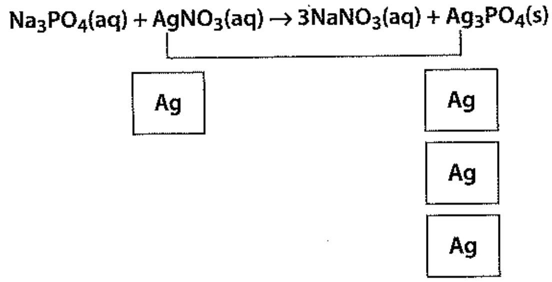

Figure 6.3

There is one on the left and three on the right. Adding a 3 in front of the silver nitrate gives

$$Na_3PO_4(aq) + 3AgNO_3(aq) \rightarrow 3NaNO_3(aq) + Ag_3PO_4(s).$$

Checking the nitrogen atoms, there are now three on the left and three on the right, since the coefficients distribute to *all* the atoms that follow in the formula (see Figure 6.4).

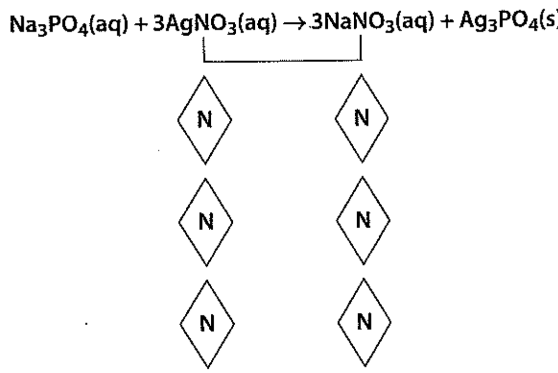

Figure 6.4

Now, checking the oxygen atoms in Figure 6.5, there are 13 on each side!

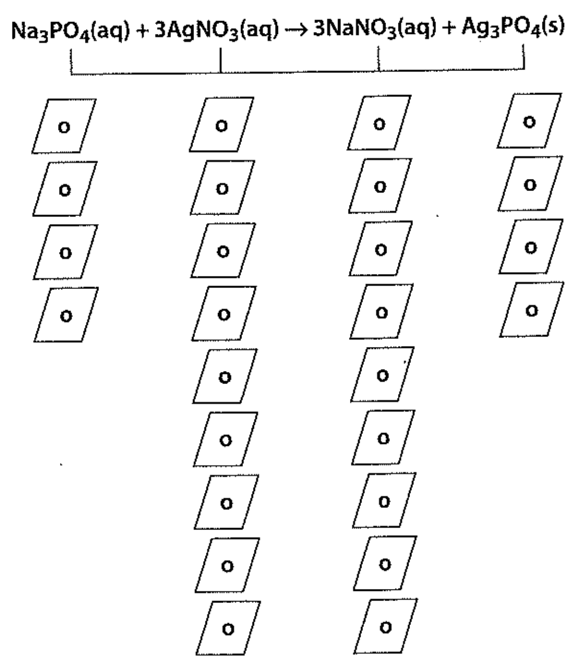

Figure 6.5

This may seem really long, but most equations balance very quickly once you have the formulas written correctly *and* you have worked many practice examples. One tip is to imagine you are building model buildings with a model set such as Legos. You take apart one building and build a different one, but you want to use up *all* the blocks, with none left over.

Sample Problem 2: Lead(II) nitrate solution is mixed with a solution of potassium iodide to form a precipitate of lead(II) iodide and a solution of potassium nitrate.

First, balance all compounds in the equation. See Chapter 5 if you need to review this.

Lead(II) nitrate: Pb^{2+} combined with NO_3^- is $Pb(NO_3)_2(aq)$.

Potassium iodide: K^+ combined with I^- is $KI(aq)$.

Lead(II) iodide: Pb^{2+} combined with I^- is $PbI_2(s)$.

Potassium nitrate: K^+ combined with NO_3^- is $KNO_3(aq)$.

Next, write the skeleton equation with all the compounds in it:

$$Pb(NO_3)_2(aq) + KI(aq) \rightarrow PbI_2(s) + KNO_3(aq)$$

Balance each atom to determine if a coefficient is needed.

Pb—1:1, no coefficient needed.

N—2:1, so add a coefficient of 2 to the product side: $2KNO_3$.

O—6:6, using the coefficient of 2 that was just added.

K—1:2, so a coefficient of 2 needs to be added to the reactant, making 2KI.

I—2:2.

Now, write the fully balanced equation:

$$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

Sample Problem 3: Zinc metal is added to a solution of copper(II) sulfate to form a solution of zinc sulfate and copper metal.

First, balance all compounds in the equation.

Zinc: Zn(s).

Copper(II) sulfate: Cu^{2+} combined with SO_4^{2-} is $CuSO_4(aq)$.

Zinc sulfate: Zn^{2+} combined with SO_4^{2-} is $ZnSO_4(aq)$.

Copper: Cu(s).

Next, write the skeleton equation with all the compounds in it:

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

Next, balance each atom to determine if a coefficient is needed.

Zn—1:1, no coefficient needed.

Cu—1:1, no coefficient needed.

S—1:1, no coefficient needed.

O—4:4, no coefficient needed.

Now write the fully balanced equation:

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

EXERCISE

6·1

Write skeleton and balanced equations for each of the following chemical reactions.

1. Solutions of barium chloride and sodium carbonate react to form aqueous sodium chloride and a precipitate of barium carbonate.

2. Solutions of potassium hydroxide and iron(III) chloride are mixed and iron(III) hydroxide precipitates in a solution of potassium chloride.

3. Copper wire is placed in a solution of silver nitrate, and metallic silver forms in a solution of copper(II) nitrate.

4. Chlorine gas is bubbled into a solution of potassium bromide, forming bromine liquid and potassium chloride solution.

5. Solutions of barium chloride and zinc sulfate are mixed to form a solution of zinc chloride and a precipitate of barium sulfate.

6. Solutions of aluminum nitrate and sodium hydroxide are mixed to form a solution of sodium nitrate and a precipitate of aluminum hydroxide.

7. Aluminum metal is placed in a solution of copper(II) chloride, and a solution of aluminum chloride forms along with copper metal.

8. Solutions of barium chloride and sodium sulfate are mixed to form a solution of sodium chloride and a precipitate of barium sulfate.

9. Sulfur trioxide gas is heated and decomposes to sulfur dioxide gas and oxygen gas.

10. Chlorine gas is bubbled into a solution of lithium iodide to form lithium chloride solution and iodine solid.

Combination, or synthesis

There are several basic types of reactions. *Combination*, or *synthesis*, occurs when two or more species combine and make one product. These are easy to recognize, since only one product is formed. If two elements are combined, the only possible reaction is combination. The general format appears as $A + B \rightarrow AB$. There are a few rules that will help with predictions.

- Metal oxides combining with carbon dioxide yield metal carbonates, e.g., solid calcium oxide reacts with carbon dioxide gas to form solid calcium carbonate.

$$CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$$

- An element plus an element yields a compound. The element listed first in the compound will be the one forming a positive charge. If oxygen is one of the elements, it is always last; if hydrogen is one of the elements, it is usually first. Hydrogen gas reacts with oxygen gas to form liquid water.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Solid magnesium ribbon, when ignited, will burn with oxygen in the air to form solid magnesium oxide.

$$Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

- Nonmetal oxides plus water yield an acid, e.g., carbon dioxide combined with water yields carbonic acid.

$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

EXERCISE
6·2

Predict the product of each of the following reactions and write a balanced equation for the reaction.

1. Iron filings react with oxygen in the air. Hint: the iron(III) compound is formed.

2. When magnesium metal is ignited in nitrogen gas.

3. Sulfur trioxide gas reacts with water droplets in the air to form an acid.

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4. Calcium metal reacts with oxygen in the air.

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5. Hydrogen gas reacts with chlorine gas.

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6. Sodium metal reacts with chlorine gas.

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7. Sodium oxide solid reacts with water.

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8. Silver metal reacts with oxygen in the air.

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9. Dinitrogen trioxide reacts with water vapor in the air.

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10. Aluminum metal reacts with oxygen in the air.

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Decomposition

The reverse of combination, *decomposition* starts with one reactant that breaks apart, often because the substance is heated. Look for one reactant before the arrow. The general format appears as AB → A + B. A decomposition reaction is often the reverse of a synthesis reaction. For example, if carbon dioxide and water combine to make carbonic acid in a synthesis reaction, when carbonic acid decomposes it forms water and carbon dioxide.

- Metal carbonates decompose to metal oxides and carbon dioxide; e.g., solid magnesium carbonate decomposes when heated to solid magnesium oxide and carbon dioxide.

$$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$$

Solid sodium carbonate decomposes when heated into solid sodium oxide and carbon dioxide gas.

$$Na_2CO_3(s) \rightarrow Na_2O(s) + CO_2(g)$$

- A compound of two elements breaks into the two elements; e.g., solid magnesium nitride decomposes upon heating into magnesium metal and nitrogen gas.

$$Mg_3N_2(s) \rightarrow 3Mg(s) + N_2\,(g)$$

- Acids decompose into water and a nonmetal oxide; e.g., aqueous sulfuric acid decomposes into liquid water and sulfur trioxide gas.

$$H_2SO_4(aq) \rightarrow H_2O(l) + SO_3(g)$$

- Metal chlorates decompose into metal chlorides and oxygen gas; e.g., solid potassium chlorate decomposes into solid potassium chloride and oxygen gas.

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

- Hydrogen peroxide decomposes into water and oxygen gas; e.g., liquid hydrogen peroxide decomposes to oxygen gas and liquid water.

$$2H_2O_2(l) \rightarrow O_2(g) + 2H_2O(l)$$

- Bases decompose to form a metallic oxide and liquid water; e.g., upon heating gently, solid magnesium hydroxide decomposes into solid magnesium oxide and water vapor.

$$Mg(OH)_2(s) \rightarrow MgO(s) + H_2O(l)$$

EXERCISE

6·3

Predict the products of each of the following reactions and write a balanced equation for what occurs.

1. Liquid water decomposes when an electrical current is added.

2. When heated, solid potassium chlorate releases a gas and leaves a new solid salt behind.

3. Solid mercury(II) oxide decomposes when heated.

4. Carbonic acid solution decomposes upon heating.

5. Calcium hydroxide solution decomposes upon heating.

6. Lithium carbonate solid decomposes upon heating.

7. When heated, solid sodium chlorate decomposes.

8. Liquid sodium chloride is decomposed by electrolysis.

9. When heated, solid silver oxide decomposes.

10. When heated gently, aqueous sulfurous acid decomposes.

Single replacement

When an element combines with a compound to form a different element and a new compound, a single element is replaced on the product side. Hence, it is called *single replacement* or sometimes single displacement. The general format appears as $A + BC \rightarrow AC + B$ or $A + BC \rightarrow BA + C$. In the first example here, the reactant metallic element replaces a metal in the compound, and in the second example, the reactant nonmetallic element replaces the nonmetal in the compound.

Example 1: A solid sodium pellet is placed in a solution of copper(II) nitrate and reacts to form a solution of sodium nitrate and solid copper.

$$2Na(s) + Cu(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + Cu(s)$$

Example 2: Chlorine gas is bubbled into a solution of sodium bromide to form a solution of sodium chloride and liquid bromine.

$$Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(l)$$

Not every metal can replace another metal. Metals have a reactivity order (see Figure 6.6), where a metal can only replace a metal ion below it in the reactivity series. This means lithium can replace any metal ion in solution.

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

Figure 6.6

On the other end of the series, gold cannot replace any other metal ion in solution.

Halogens also have a reactivity series (see Figure 6.7) and can similarly only replace ones that are below them in the series. Fluorine can replace all other halide ions in solution, while iodine, at the bottom, cannot replace any others.

Halogen series
F_2
Cl_2
Br_2
I_2

Figure 6.7

Many metals are capable of replacing hydrogen in acids, some can replace hydrogen in steam (very hot water), and some metals can even replace hydrogen in cold water. When you look carefully at the metals, you will observe fewer choices each time but they are the same metals! Those that can do all three are bolded.

- Metals replacing hydrogen in acids are **Li**, **K**, **Ba**, **Ca**, **Na**, Mg, Al, Zn, Cr, Fe, Cd, Co, Ni, Sn, and Pb. Look carefully again at the metal reactivity series in Figure 6.6 on the previous page. Do you notice a trend?

  Example 1: A solid zinc pellet is placed in a solution of hydrochloric acid to form a solution of zinc chloride and hydrogen gas.

  $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

  Example 2: Aluminum foil is placed in a solution of hydrochloric acid to form a solution of aluminum chloride and hydrogen gas.

  $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$

- Metals replacing hydrogen in steam are **Li**, **K**, **Ba**, **Ca**, **Na**, Mg, Al, Zn, Cr, Fe, and Cd. Again, look at Figure 6.6 to see the trend.

  Example 1: Solid iron is placed in steam to form an iron(II) hydroxide solution and hydrogen gas.

  $Fe(s) + 2H_2O(l) \rightarrow Fe(OH)_2(aq) + H_2(g)$

  Example 2: Magnesium ribbon is placed in steam, reacting to form a magnesium hydroxide solution and hydrogen gas.

  $Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$

- Metals replacing hydrogen in cold water are **Li**, **K**, **Ba**, **Ca**, and **Na**.

  Example 1: Solid sodium is placed in cold water to form a solution of sodium hydroxide and hydrogen gas.

  $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

  Example 2: Solid calcium is placed in water to form a solution of calcium hydroxide and hydrogen gas.

  $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

EXERCISE

6·4

Predict the products of each of the following reactions and write a balanced equation for what occurs.

1. A piece of lithium metal is added to cold water.

2. Fluorine gas is added to a solution of potassium iodide.

3. A zinc bar is placed in steam.

4. Sulfuric acid is added to a sheet of lead metal.

5. Zinc pellets are placed in a solution of hydrochloric acid.

6. Chlorine gas is bubbled into a solution of potassium bromide.

7. Potassium metal is added to water.

8. Copper wire is added to a solution of silver nitrate. Hint: one of the products contains copper(II).

9. Fluorine gas is bubbled into a solution of sodium chloride.

10. Iron filings are added to a solution of copper(II) nitrate. Hint: one of the products contains iron(II).

Double replacement

When two compounds combine and form two new compounds by switching anions, it is called *double replacement* or double displacement. There are several types of double displacement reactions. One type are reactions that form precipitates. This type of reaction often occurs in solution (two dissolved substances), and one or both products formed are a solid called a *precipitate*. There are rules to determine which ones form precipitates. Right now, let's just look at the format: the general format is $AB + CD \rightarrow AD + CB$.

Example 1: Solutions of silver nitrate and sodium chloride are mixed to form a precipitate of silver chloride and a solution of sodium nitrate.

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

Example 2: Solutions of zinc sulfate and barium nitrate are mixed.

$$ZnSO_4(aq) + Ba(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + BaSO_4(s)$$

How does one tell if a precipitate forms? Solubility rules are used, but rather than have you derive these rules from experiments or memorize them, Table 6.1 will be used as a reference. Use the first column to locate the cation and then move your finger across until the column with the desired anion is reached. At the intersection of the cation and anion, *S* indicates solubility, which means the compound stays in solution (think *S* = *s*tays *s*olution); and *I* indicates insolubility, so the compound is forming a precipitate (think *I* = *i*n prec*i*pitate)—but is noted as a solid, *(s)*, in the balanced equation. This can be very confusing. The *S* on the solubility table does not mean a solid is formed. Solubility tables differ in format, but all are based on the same principles.

Table 6.1 Solubility Table of Selected Ions in Aqueous Solutions

	Br^-	Cl^-	ClO_3^-	I^-	NO_3^-	OH^-	PO_4^{3-}	SO_4^{2-}
Ag^+	I	I	S	I	S	I	I	I
Al^{3+}	S	S	S	S	S	I	I	S
Ba^{2+}	S	S	S	S	S	S	I	I
Ca^{2+}	S	S	S	S	S	I	I	I
Cu^{2+}	S	S	S	S	S	I	I	S
Hg^{2+}	I	I	S	I	S	I	I	I
K^+	S	S	S	S	S	S	S	S
Li^+	S	S	S	S	S	S	S	S
Mg^{2+}	S	S	S	S	S	I	I	S
Na^+	S	S	S	S	S	S	S	S
Ni^{2+}	S	S	S	S	S	I	I	S
Pb^{2+}	I	I	S	I	S	I	I	I
Sr^{2+}	S	S	S	S	S	I	I	I
Zn^{2+}	S	S	S	S	S	I	I	S

EXERCISE
6·5

Predict the products of each of the following reactions and write a balanced equation for what occurs.

1. Solutions of ammonium iodide and silver nitrate react.

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2. A sodium sulfate solution is mixed with a solution of lead(II) nitrate.

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3. A solution of copper(II) chloride is added to a potassium hydroxide solution.

____

4. Solutions of aluminum nitrate and sodium phosphate react.

5. Solutions of sodium chloride and lead(II) nitrate react.

6. Solutions of potassium hydroxide and nickel(II) nitrate react.

7. Solutions of barium chlorate and potassium sulfate are mixed.

8. Lithium hydroxide solution is added to a solution of magnesium chloride.

9. Strontium bromide solution is added to a solution of nickel(II) sulfate.

10. Zinc iodide solution is added to a solution of potassium hydroxide.

Combustion

Combustion reactions occur when hydrocarbons burn in the presence of oxygen gas. It will be assumed all combustion reactions are complete, meaning the only products formed are carbon dioxide and water. During incomplete combustion, such as the burning of a candle, carbon monoxide and carbon soot also form.

Example 1: Propane burns in air to form carbon dioxide and water vapor.

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

Example 2: Methane burns in air to form carbon dioxide and water vapor.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

EXERCISE

6·6

Write a balanced equation for each of the following complete combustion reactions.

1. Butane gas burns in air.

2. Heptane gas burns in air.

3. Decane gas burns in air.

4. Octane gas burns in air.

5. Ethane gas burns in air.

Acid-base

Acid-base reactions are a unique form of double displacement reactions that generally occur without a precipitate forming. The acids and bases react as a hydrogen ion is transferred from the acid to the base, making a new product. There are different definitions of acids, but the most commonly used definition of an acid is a proton donor. Remember, an H^+ (hydrogen ion) is only a proton. More about acid-base reactions is included in Chapter 10.

Example 1: Solutions of hydrochloric acid and sodium hydroxide are mixed to form water and a solution of sodium chloride.

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$

Example 2: Solutions of hydrobromic acid and barium hydroxide are mixed to form a solution of barium bromide and water.

$$2HBr(aq) + Ba(OH)_2(aq) \rightarrow 2H_2O(l) + BaBr_2(aq).$$

Example 3: Some products when formed are unstable and will subsequently decompose into other products. For example, the weak acid carbonic acid decomposes to form carbon dioxide gas and liquid water.

$$H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l)$$

Example 4: Another example would be ammonium hydroxide, which decomposes to form ammonia gas and liquid water.

$$NH_4OH(aq) \rightarrow NH_3(g) + H_2O(l)$$

EXERCISE

6·7

Predict the products of each of the following reactions and write a balanced equation for what occurs. Assume the reactions have sufficient reactant to fully react.

1. Solid nickel(III) hydroxide is added to a solution of hydroiodic acid.

2. Solutions of sulfuric acid and potassium hydroxide are mixed.

3. A solution of calcium hydroxide is added to a solution of nitric acid.

____

4. Solutions of sulfuric acid and sodium hydroxide are mixed.

____

5. Solutions of hydrofluoric acid and potassium hydroxide are mixed.

____

6. Solutions of phosphoric acid and lithium hydroxide are mixed.

____

7. Solutions of calcium hydroxide and hydrochloric acid are mixed.

____

8. Magnesium hydroxide solution and phosphoric acid solution are mixed.

____

9. Barium carbonate solid is stirred into hydrochloric acid solution.

____

10. Sulfuric acid solution is added to a solution of lithium hydroxide.

____

Ionic and net ionic equations

An equation indicates the reactants and products in the overall reaction but is not an indication of the steps of the reaction or the factors driving the reaction to occur. To help understand this, there are other forms of equation writing. An ionic equation goes one additional step and, for water-based solutions, indicates in the equation if ions are formed. Depending on the solution, the species formed may be ionized. Ionizing species include strong acids, strong bases, and ionic compounds that dissolve in water. Refer back to Table 6.1 on page 84 to determine if a specific ionic compound is soluble. Non-ionizing species include weak acids, weak bases, and molecular compounds. These species are not separated in ionic equations. All water-based solutions are designated with *(aq)* behind them in the equation. So this reaction notation differs from a balanced equation, since it shows the actual species that are present in the solution. Note that some species—those that are liquids, solids, or gases—remain the same.

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

would be changed to

$$Zn(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Zn^{2+}(aq) + 2Cl^-(aq) + H_2(g).$$

In the last equation, $2Cl^-$ exist on both sides of the equation. These are called *spectator ions* and were not actually involved in the reaction. In a net ionic equation, the spectators are cancelled out. This is important, as the net equation then shows the actual reaction driving the overall equation.

$$Zn(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Zn^{2+}(aq) + 2Cl^-(aq) + H_2(g)$$

would thus be changed to

$$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g).$$

For the reaction $ZnSO_4(aq) + Ba(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + BaSO_4(s)$, the ionic equation would be

$$Zn^{2+}(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq) \rightarrow Zn^{2+}(aq) + 2NO_3^-(aq) + BaSO_4(s)$$

and the net ionic equation would be

$$SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$$

after the spectator ions of $Zn^{2+}(aq)$ and $2NO_3^-(aq)$ were cancelled.

For the reaction $HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$, the ionic equation would be

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Na^+(aq) + Cl^-(aq)$$

and the net ionic equation would be

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l).$$

EXERCISE

6·8

For each of the following, write a balanced equation, a balanced ionic equation, and a balanced net ionic equation.

1. Solutions of sodium phosphate and silver nitrate react.

____

____

____

2. A solution of potassium hydroxide is added to a solution of strontium chloride.

____

____

____

3. Bromine liquid is added to a solution of potassium iodide.

____

____

____

4. Solutions of hydrochloric acid and potassium hydroxide are mixed.

5. Solutions of sodium sulfide and silver chlorate are mixed.

6. Chlorine gas is bubbled into a solution of lithium bromide.

7. Solutions of ammonium phosphate and barium hydroxide are mixed.

8. Copper wire is placed in a solution of silver nitrate.

9. Solutions of calcium hydroxide and sulfuric acid are mixed.

EXERCISE
6·9

Identify the type of each of the following reactions, using S for synthesis or combination, D for decomposition, SD for single displacement, DD for double displacement, and C for combustion.

1. $H_3PO_4(aq) + 3NaOH(aq) \rightarrow Na_3PO_4(aq) + 3H_2O(l)$ ______
2. $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$ ______

3. $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$ ____________

4. $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$ ____________

5. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$ ____________

EXERCISE

6·10

Predict the products of each of the following reactions and write a balanced equation for what occurs.

1. A piece of cadmium metal reacts with hydrochloric acid.

____

2. Pentane gas undergoes complete combustion in air.

____

3. Nitrogen gas reacts with hydrogen gas.

____

4. Solutions of barium chloride and aluminum sulfate are mixed.

____

5. Solid copper(II) hydroxide decomposes.

____

Mass and mole relationships

·7·

Now that we can write a formula, we need to be able to calculate formula and molecular mass. This is important so the skill of dimensional analysis with balanced equations can be practiced.

Calculating formula and molecular mass

To calculate a formula mass, we need the periodic table. The formula NaCl is composed of one sodium atom and one chlorine atom. On the periodic table, Na has a mass of 22.99 amu and Cl has a mass of 35.45 amu; adding these together gives the mass of one unit of NaCl as 58.44 amu. Some chemists now use the dalton (Da) as the term for the atomic mass unit. We will use the amu. Sulfur trioxide, SO_3, is made of one sulfur atom and three oxygen atoms and has a molecular mass of 32.07 amu + 3(16.00) amu = 80.07 amu.

Why was a different term for the mass used? There is a difference between formula mass and molecular mass. Formula mass is the term used when the atoms are held together by an ionic bond, and molecular mass is the term used for a molecule held together by covalent bonds. In general, if the compound has a metal and a nonmetal in it, it has a formula mass, and if it is composed of only nonmetals, it has a molecular mass.

EXERCISE

7·1

Indicate if you are determining the formula or molecular mass of each of the following substances and then record the calculated value.

	TYPE OF MASS	CALCULATED VALUE
1. MgO	______	______
2. N_2O_5	______	______
3. CaF_2	______	______
4. CCl_4	______	______
5. $AlPO_4$	______	______
6. Ag_2CO_3	______	______
7. $Au_2(C_2O_4)_3$	______	______

8. HgO ____________ ____________
9. HF ____________ ____________
10. NH_3 ____________ ____________
11. SO_3 ____________ ____________
12. Na_2SO_4 ____________ ____________
13. $Cu(NO_3)_2$ ____________ ____________
14. $BaCO_3$ ____________ ____________
15. $(NH_4)_3PO_4$ ____________ ____________
16. KCl ____________ ____________
17. H_2SO_4 ____________ ____________
18. CO ____________ ____________
19. KOH ____________ ____________
20. $Mg_3(PO_4)_2$ ____________ ____________

Calculating molar mass

In chemistry experiments the mass of objects is measured in grams—so how many grams are in 58.44 amu? Using the conversion factor of 1 amu to 1.661×10^{-24} g,

$$58.44 \text{ amu} \times \frac{1.661 \times 10^{-24} \text{ g}}{1 \text{ amu}} = 9.707 \times 10^{-23} \text{ g}.$$

This number is so small, our balances can't mass it! A typical balance will measure to one hundredth of a gram. So how do we mass substances? If a larger number of each species are used and always the same number of them, they can compare on a larger scale. This is similar to using the word *dozen* to represent 12 of anything. The word *dozen* represents 12 of something, such as a dozen eggs or a dozen oranges. But to get a mass big enough to measure in chemistry, a dozen won't do.

We use a number called Avogadro's number, which is equal to 6.022×10^{23} particles of a substance; this is called the mole (or 1 mol). (Why such a strange number? The French chemist Jean Baptiste Perrin experimentally determined the number of atoms in 32 g of oxygen molecules.) So the word *mole* is like the word *dozen*—it represents a certain number of things. There is a unique relationship between the amu of one formula unit and the mass of a mole:

$$58.44 \text{ amu} \times \frac{1.661 \times 10^{-24} \text{ g}}{1 \text{ amu}} \times \frac{6.022 \times 10^{23}}{1 \text{ mol}} = 58.44 \text{ g/mol}.$$

What does this mean? If we add a formula unit of a species we get an answer in daltons; if we add a mole of them, the numeric value *is the same* but the unit is g/mol. This means we do not have to do conversions from amu to grams and then to moles in every problem. We simply need to know if we are adding the mass of one unit (amu) or 1 mole of them to choose the proper label. The mass of 1 mole is called molar mass and is given the units of g/mol or gmol^{-1}.

EXERCISE

7·2

Calculate the molar mass of the following substances. Save your work—it will help with Exercise 7-3.

1. CO ______________________
2. SiO_2 ______________________
3. N_2O_3 ______________________
4. $CuSO_4$ ______________________
5. $(NH_4)_3PO_3$ ______________________
6. NO ______________________
7. $NaOH$ ______________________
8. FeS ______________________
9. $CuCl_2$ ______________________
10. $Cu(OH)_2$ ______________________
11. I_2 ______________________
12. N_2O_4 ______________________
13. $(NH_4)_2SO_4$ ______________________
14. HNO_2 ______________________
15. H_2O ______________________
16. O_2 ______________________
17. $AgC_2H_3O_2$ ______________________
18. MgS ______________________
19. N_2O_5 ______________________
20. $Ca(OH)_2$ ______________________

Calculating percent composition

The percent composition is the relative amount of each atom in the compound. No matter the amount of a compound present, the percent of each atom making up the compound remains the same. To calculate the percent composition, there are four simple steps:

1. Find the total mass of each atom in one unit of the compound.
2. Find the molar mass of the compound.
3. Divide each atom's mass by the molar mass.
4. Multiply by 100.

$$\% \text{ of each atom} = \frac{\#\text{ of each atom} \times \text{mass of the atom}}{\text{molar mass of the compound}} \times 100$$

Double-check by adding the percentages together. The total should be 100.

Sample Problem 1: In our sodium chloride example from before, sodium contributed 22.99 g/mol and chlorine contributed 35.45 g/mol to the total molar mass of 58.44 g/mol. Find the percent composition of NaCl.

$$\% \text{ of Na} = \frac{1 \times 22.99 \text{ g}}{58.44 \text{ g/mol}} \times 100$$

$$\% \text{ of Na} = 39.34\%$$

$$\% \text{ of Cl} = \frac{1 \times 35.45 \text{ g}}{58.44 \text{ g/mol}} \times 100$$

$$\% \text{ of Cl} = 60.66\%$$

Double-check: 39.34% + 60.66% = 100%.

Sample Problem 2: If we have a 155.0 g sample of sodium chloride, how many grams are Na and how many are Cl? NaCl?

Using our percentages we can determine how many grams of the sample are sodium atoms and how many grams are chlorine atoms. Sodium is 39.34% of the sample, or $155.0 \times 0.3934 = 60.98$ g, and chlorine is 60.66% of the sample, or $155.0 \times 0.6066 = 94.02$ g.

Sample Problem 3: What is the percent composition of each atom in calcium phosphate, $Ca_3(PO_4)_2$?

In this formula there are three calcium atoms, each with a mass of 40.08 g/mol, two phosphorus atoms, each with a mass of 30.97 g/mol, and eight oxygen atoms, each with a mass of 16.00 g/mol.

$$\% \text{ of Ca} = \frac{3 \times 40.08 \text{ g}}{310.18 \text{ g/mol}} \times 100$$

$$\% \text{ of Ca} = 38.76\%$$

$$\% \text{ of P} = \frac{2 \times 30.97 \text{ g}}{310.18 \text{ g/mol}} \times 100$$

$$\% \text{ of P} = 19.97\%$$

$$\% \text{ of O} = \frac{8 \times 16.00}{310.18 \text{ g/mol}} \times 100$$

$$\% \text{ of O} = 41.27\%$$

Double-check: 38.76% + 19.97% + 41.27% = 100%.

EXERCISE

7·3

Calculate the percentage of each atom in the following substances.

1. CO ______________________
2. SiO_2 ______________________
3. N_2O_3 ______________________
4. $CuSO_4$ ______________________
5. $(NH_4)_3PO_3$ ______________________
6. NO ______________________
7. NaOH ______________________
8. FeS ______________________
9. $CuCl_2$ ______________________
10. $Cu(OH)_2$ ______________________
11. I_2 ______________________
12. N_2O_4 ______________________
13. $(NH_4)_2SO_4$ ______________________
14. HNO_2 ______________________
15. H_2O ______________________
16. O_2 ______________________
17. $AgC_2H_3O_2$ ______________________
18. MgS ______________________
19. N_2O_5 ______________________
20. $Ca(OH)_2$ ______________________

Molar conversions between grams and particles

Do we always mass one mole? No! We need to practice changing different amounts of a substance into moles and converting moles into grams. We can also calculate how many particles we have. The particles could be molecules, ions, or formula units, depending on the species given.

To convert grams to moles, the given grams are divided by the molar mass. Set up in dimensional analysis, it looks like this:

$$\text{given grams of substance} \times \frac{1 \text{ mol of substance}}{\text{molar mass of substance}} = \text{moles of substance.}$$

To convert from moles to grams, the setup looks like this:

$$\text{given moles of substance} \times \frac{\text{molar mass of substance}}{1 \text{ mol of substance}} = \text{grams of substance.}$$

Figure 7.1 shows the conversion relationships between grams and moles.

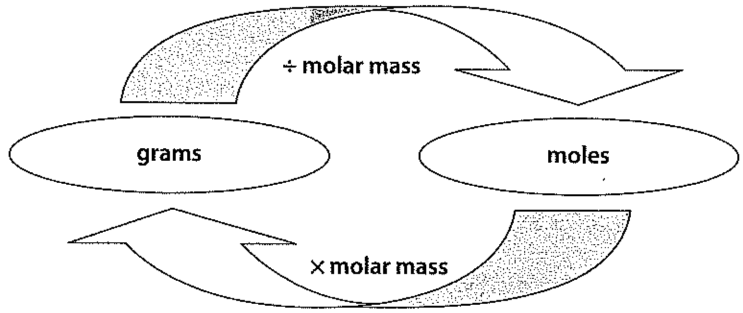

Figure 7.1

If we want to convert between particles and moles, we need to use Avogadro's number in the ratio. Converting from moles to particles looks like this:

$$\text{given moles of substance} \times \frac{6.022 \times 10^{23} \text{particles}}{1 \text{ mol of substance}} = \text{particles of substance.}$$

The conversion from particles to moles looks like this:

$$\text{given particles of substance} \times \frac{1 \text{ mol of substance}}{6.022 \times 10^{23} \text{particles}} = \text{moles of substance.}$$

Figure 7.2 shows the conversion relationships between moles and particles.

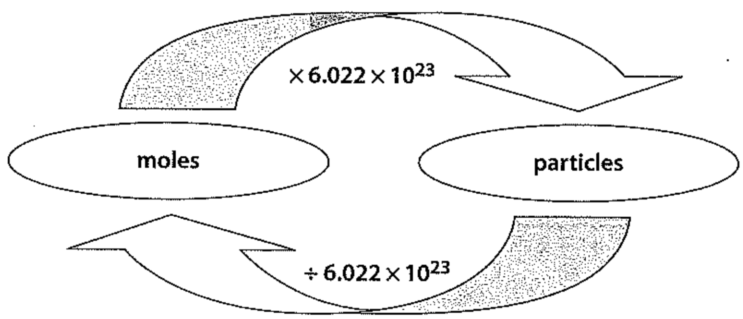

Figure 7.2

It is a good idea to draw a plan of the steps that need to be taken to solve a problem. Combining the steps above, we can draw one map to make a plan for any of these conversions (see Figure 7.3).

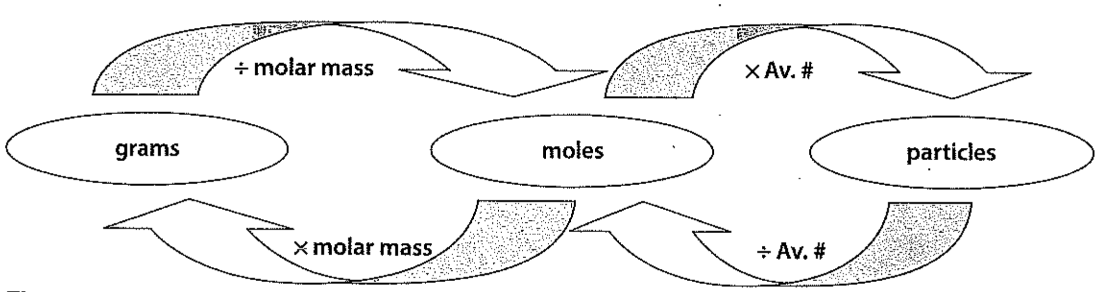

Figure 7.3

One additional step is to change to atoms from particles (see Figure 7.4). This requires knowing the number of atoms in the particle. For instance, the number of atoms in carbon dioxide, CO_2, is three. If we only want to know how many oxygen atoms are present in the particles of CO_2, we would use 2 in the conversion.

$$\text{number of particles} \times \frac{\text{number of atoms}}{1 \text{ particle}} = \text{number of atoms}$$

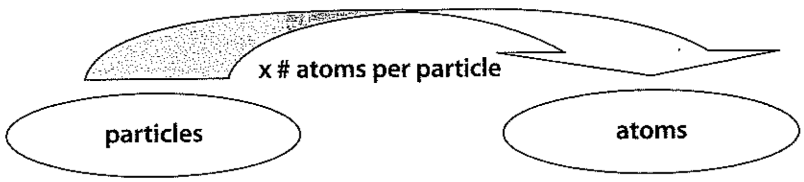

Figure 7.4

Sample Problem 4: Given 71.4 g of $Mg_3(PO_4)_2$, how many moles are present?

First we need to calculate the molar mass of $Mg_3(PO_4)_2$. From the periodic table we find that magnesium is 24.31 $gmol^{-1}$, phosphorus is 30.97 $gmol^{-1}$, and oxygen is 16.00 $gmol^{-1}$.

$$(24.31 \text{ gmol}^{-1} \times 3) + (30.97 \text{gmol}^{-1} \times 2) + (16.00 \text{ gmol}^{-1} \times 8) = 262.87 \text{ gmol}^{-1}$$

$$71.4 \text{ g } Mg_3(PO_4)_2 \times \frac{1 \text{ mol } Mg_3(PO_4)_2}{262.87 \text{ g } Mg_3(PO_4)_2} = 0.272 \text{ mol } Mg_3(PO_4)_2$$

Sample Problem 5: Given 1.58 mol of $Mg_3(PO4)_2$, how many grams are present?

$$1.58 \text{ mol } Mg_3(PO_4)_2 \times \frac{262.87 \text{ g } Mg_3(PO_4)_2}{1 \text{ mol } Mg_3(PO_4)_2} = 415 \text{ g } Mg_3(PO_4)_2$$

Sample Problem 6: Given 97.3 g of $Mg_3(PO_4)_2$, how many particles of $Mg_3(PO_4)_2$ are present?

This is a two-step problem:

$$97.3 \text{ g } Mg_3(PO_4)_2 \times \frac{1 \text{ mol } Mg_3(PO_4)_2}{262.87 \text{ g } Mg_3(PO_4)_2} \times \frac{6.022 \times 10^{23} \text{ particles}}{1 \text{ mol } Mg_3(PO_4)_2} = 2.23 \times 10^{23} \text{ particles } Mg_3(PO_4)_2$$

Sample Problem 7: Given 23.8 g of $Ca_3(PO_4)_2$, how many atoms of oxygen are present?

$$23.8 \text{ g } Ca_3(PO_4)_2 \times \frac{1 \text{ mol } Ca_3(PO_4)_2}{310.18 \text{ g } Ca_3(PO_4)_2} \times \frac{6.022 \times 10^{23} \text{ particles}}{1 \text{ mol } Ca_3(PO_4)_2}$$

$$= 4.62 \times 10^{22} \text{ particles } Ca_3(PO_4)_2 \times \frac{8 \text{ atoms of O}}{1 \text{ particle } Ca_3(PO_4)_2}$$

$$= 3.70 \times 10^{23} \text{ atoms of O}$$

EXERCISE

7·4

Solve the following problems.

1. Given 4.62 mol of sodium hydroxide, NaOH, how many grams are present?

2. If there are 5.13 g of NaOH, how many moles are present?

3. How many particles are present in 79.8 g of NaOH?

4. How many grams of potassium sulfide are present if there are 1.79×10^{24} particles present? Yes, find the formula first!

5. How many atoms of chlorine are present in 1.50 mol of chlorine gas, Cl_2?

6. How many atoms of oxygen are present in 37.8 g of calcium carbonate? Yes, find the formula first!

7. How many moles are in 32.0 g of SO_2?

8. Calculate the grams present in 4.00 mol of KI.

9. Calculate the number of moles in 68.0 g of Na_2S.

10. Calculate the mass in grams of 2.49×10^{21} molecules of water.

11. Calculate the number of formula units in 10.0 g of K_2SO_4.

12. Calculate the number of atoms in 32.0 grams of oxygen gas, O_2.

13. Calculate the number of moles in 6.43×10^{22} molecules of SO_2.

14. Calculate the number of grams in 12.0 mol of S.

15. Calculate the total number of atoms in 2.74 mol of NaCl.

____

16. Calculate how many carbon atoms are present in 2.55 mol of CO_2.

____

17. Calculate how many nitrogen atoms are present in 13.8 g of $Ca(NO_3)_2$.

____

18. Calculate the number of moles in 55.5 g of Al.

____

19. Calculate the number of moles in 89.3 g of $MgSO_4$.

____

20. Calculate the number of bromine atoms in 149.0 g of Br_2.

____

Stoichiometry

Stoichiometry is a big word for expressing the number of particles to a mass by using dimensional analysis with moles. You need a *balanced* equation and sharpened math skills! Problems can be asked in a variety of ways, but one of the steps will require changing from moles of one species to moles of another species by using a balanced equation. The plot thickens! Yes, you need several skills you have already learned: calculating molar masses, doing dimensional analysis, and balancing equations. Thanks to the understanding that a chemical equation such as

$$2Na + Cl_2 \rightarrow 2NaCl$$

can be expressed as "two moles of sodium atoms combined with one mole of chlorine molecules can make two moles of sodium chloride," the balanced equation gives us the *ratio* between species in moles! The ratio used depends on what is asked; the ratio needs to be written so the substance you are changing from is on the bottom, to cancel, and the substance you are changing to is on the top. For instance, the ratio $\frac{1 \text{ mol } Cl_2}{2 \text{ mol NaCl}}$ could also be written

$$\frac{2 \text{ mol NaCl}}{1 \text{ mol } Cl_2},$$

depending on the problem being solved. Looking at Figure 7.5, once moles of the starting substance are calculated, this can be multiplied by the ratio from the balanced equation, where the coefficient of the starting material is on the bottom and the coefficient of the substance being converted to is on the top, giving you the moles of the new substance.

$$\boxed{\text{Moles of A}} \times \frac{\text{coefficient of B}}{\text{coefficient of A}} = \boxed{\text{Moles of B}}$$

Figure 7.5

EXERCISE

7·5

Write expressions describing the relationships between moles of reactants and moles of products. For example, $\frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O}$ is one relationship for the equation $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$.

1. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

 ____

2. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

 ____

3. $Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$

 ____

4. $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

 ____

5. $Cl_2(g) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(l)$

 ____

Using this skill, we can convert from one species to another. Now you are ready for stoichiometry!

Mole-to-mole conversions

To do a conversion, we always need a balanced equation. Then we look at the problem to find out what two species need to be in the ratio. Then set up the ratio so units cancel.

- Write a balanced equation.
- Read the problem for the species being used.
- Set up the ratio so units cancel.

Sample Problem 8: Using $2Na + Cl_2 \rightarrow 2NaCl$, if we have 2 mol of sodium and excess chlorine, how many moles of sodium chloride can we make?

The ratio between sodium and sodium chloride in the balanced equation is 2:2.

$$2 \text{ mol of Na} \times \frac{2 \text{ mol NaCl}}{2 \text{ mol Na}} = 2 \text{ mol NaCl}$$

EXERCISE

7·6

Use the equation $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$ to answer the following questions.

1. How many moles of O_2 do we need to react with 56 mol of Fe?

 ____

2. If all 56 mol of Fe from question 1 react, how many moles of Fe_2O_3 are produced?

 ____

3. How many moles of Fe are necessary to make 25 mol of Fe_2O_3?

4. How many moles of O_2 are required to make 32 mol of Fe_2O_3?

5. If 120 mol of O_2 is used, how many moles of Fe_2O_3 could be made?

6. How many moles of Fe are necessary to react completely with 66 mol of O_2?

7. Calculate the number of moles of O_2 needed to react with 2.7 mol of Fe.

8. Calculate the number of moles of Fe_2O_3 formed if 7.7 mol of Fe fully reacts.

9. Calculate the number of moles of each reactant needed to form 5.8 mol of product.

10. Calculate the number of moles of product formed if 2.7 mol of O_2 reacts with excess Fe.

Other stoichiometric calculations

A variety of problems arise if you are given a unit other than moles to start with, and you want to convert to a unit other than moles. However, in all cases the important step is the stoichiometry step of converting from moles of a given substance to moles of a needed substance. Here are some samples of other types of problems, with a chance to practice them. Use the following equation for sample problems 9 and 10:

$$Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$$

Sample Problem 9: Given 32.1 g of Cu, how many grams of Ag can be made?

$$32.1\text{ g Cu} \times \frac{1\text{ mol Cu}}{63.55\text{ g}} \times \frac{2\text{ mol Ag}}{1\text{ mol Cu}} \times \frac{107.9\text{ g Ag}}{1\text{ mol Ag}} = 109\text{ g Ag}$$

Sample Problem 10: If 45.0 g of Ag needs to be produced, how many moles of Cu are needed?

$$45.0\text{ g Ag} \times \frac{1\text{ mol Ag}}{107.9\text{ g Ag}} \times \frac{1\text{ mol Cu}}{2\text{ mol Ag}} = 0.209\text{ mol Cu}$$

As you can see from these two problems, the ratios are arranged so units cancel and only the desired units remain. In the first problem, the silver was on the top in the ratio, and in the second problem it was on the bottom. Specifics of gas stoichiometry and solution stoichiometry will be covered in Chapters 8 and 9.

EXERCISE

7·7

Using the equation $2Al(s) + 6HNO_3(aq) \rightarrow 2Al(NO_3)_3(aq) + 3H_2(g)$, answer the following questions.

1. How many grams of Al react completely with 12.0 mol of nitric acid?

2. If 3.45 mol of Al reacts with excess nitric acid, how many grams of $Al(NO_3)_3$ are produced?

3. How many grams of H_2 are produced from 15.7 g of Al reacting with an excess of nitric acid?

4. If 0.750 mol of $Al(NO_3)_3$ is needed, how many grams of HNO_3 are required if plenty of Al is available?

5. If 9.82 g of Al reacts with excess HNO_3, how many grams of $Al(NO_3)_3$ are made?

6. How many moles of Al are necessary to react completely with 65.3 g of HNO_3?

7. If 5.34×10^{23} atoms of Al are present in excess nitric acid, how many grams of H_2 can be formed?

8. How many grams of aluminum are required to completely react with excess nitric acid to form 13.0 g of aluminum nitrate?

9. If 5.34×10^{22} molecules of hydrogen gas are formed, how many atoms of aluminum are needed?

10. How many grams of aluminum nitrate are formed if 10.0 g of nitric acid reacts with excess aluminum?

11. If excess nitric acid reacts with 2.4 mol of aluminum, how many molecules of hydrogen gas are formed?

12. What mass of aluminum is needed to fully react with 57.0 g of nitric acid?

13. Calculate the number of grams of aluminum nitrate that can be formed if excess aluminum reacts with 3.74 mol of nitric acid.

____

14. Calculate the moles formed of each product if excess nitric acid reacts with 12.7 g of Al.

____

15. To form 374 mol of H_2, how many atoms of Al are needed with excess nitric acid?

____

Limiting reagents

In the problems so far, one of the reactants has been in excess, so the limiting reagent has always been known. The limiting reagent is the reactant that, because of its "limited" amount, limits the amount of product that can be made. From this point on these two terms, *limiting reagent* or *limiting reactant*, will be used to describe the chemical limiting the amount of product produced. The problems now are going to differ in that both reactants will have a given amount and we will have to determine which one is limiting. *This means two stoichiometry problems per question!*

There are multiple ways to up these problems, depending on how the question is asked. Always doing them as two full stoichiometry problems often answers two questions: (1) which reactant is the limiting reagent, and (2) how much product can be made. Whichever reactant makes the smallest amount of product is the limiting reagent, since it will all be used and limits how much product can be made!

Sample Problem 11: Using the equation $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, given 37.4 g of H_2 and 50.7 g of O_2, identify the limiting reagent and calculate the number of grams of water made.

Remember to first calculate the molar masses of H_2, O_2, and H_2O, and set up the ratios being used to cancel units.

$$37.4 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 334 \text{ g H}_2\text{O}$$

$$50.7 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2\text{mol H}_2\text{O}}{1 \text{ mol O}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 57.1 \text{ g H}_2\text{O}$$

The O_2 made less water, so it is the limiting reagent, and the amount of product that can be made is 57.1 g of H_2O. Don't be tricked by thinking the one with fewer grams at the start is limiting!

EXERCISE 7·8

Using the equation $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ and the information given in each problem, determine which reactant is the limiting reagent and how much zinc chloride (in grams) is produced in each problem.

1. 2.50 g of Zn is added to 5.00 g of HCl.

____

2. 10.0 g of Zn is added to 10.0 g of HCl.

____

3. 16.3 g of Zn is added to 8.15 g of HCl.

____

4. 1.2×10^{22} atoms of Zn are added to 15.00 g of HCl.

____

5. 2.3 mol of Zn reacts with 2.3 mol of HCl.

____

6. 2.7 mol of Zn is added to 34.0 g of HCl.

____

7. 34.7 g of Zn is added to 3.42×10^{22} formula units of HCl.

____

8. 4.5 mol of Zn reacts with 12.3 g of HCl.

____

9. 12.45 g of Zn reacts with 37.4 g of HCl.

____

10. 3.7×10^{23} atoms of Zn are added to 4.7×10^{23} formula units of HCl.

____

Percent yield

For a variety of reasons, when reactions occur they often do not make the amount expected. The amount calculated through stoichiometry is the maximum yield, 100%. Percent yield is a ratio of how many grams are actually made (actual yield) in comparison to the maximum amount (in grams) that could have been made (expected or theoretical yield), expressed as a percentage.

$$\frac{\text{grams of actual yield}}{\text{grams of theoretical yield}} \times 100 = \text{percent yield}$$

Sample Problem 12: If the expected yield is 42.1 g of S and 28.3 g S is actually made, what is the percent yield?

$$\frac{28.3 \text{ g S}}{42.1 \text{ g S}} \times 100 = 67.2\%$$

Don't forget to use significant figures!

EXERCISE

7·9

Perform the following calculations.

1. If 42.6 g of ammonia is made in a reaction where 55.0 g is expected, what is the percent yield?

2. If a reaction consistently has an 85% yield, and 23.7 g of the product is made, what is the theoretical yield of the product?

3. In the reaction $2NaOH(aq) + NiSO_4(aq) \rightarrow Ni(OH)_2(s) + Na_2SO_4(aq)$, with an excess of $NiSO_4$, what is the percent yield if 20.0 g of NaOH makes 12.5 g of $Ni(OH)_2$?

4. If 40.0 g of N_2 reacts with 20.0 g of H_2 to make 45.9 g of NH_3 in the reaction $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, what is the percent yield of NH_3?

5. Calculate the percent yield if 34.0 g of a single product is expected but only 26.9 g is formed.

6. If 82.0 g of N_2 reacts with 40.0 g of H_2 to make 45.9 g of NH_3 in the reaction $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, what is the percent yield of NH_3?

7. Calculate the percent yield if a reaction forms 82.0 g of product when 96.0 g is expected.

8. If 2.7×10^{23} molecules of ammonia are made in a reaction where 3.4×10^{23} molecules were expected, what was the percent yield?

9. If 2.4 mol of ammonia are made in a reaction where 3.7 mol were expected, what was the percent yield?

10. If 3.72×10^{22} molecules of ammonia are made in a reaction where 1.5 mol were expected, what was the percent yield?

EXERCISE

7·10

Answer the following questions.

1. What is the difference between formula mass and molecular mass?

2. What is the difference between formula/molecular mass and molar mass?

3. Calculate the formula mass of:

a. $CaCl_2$ ______________________

b. AgI ______________________

c. Fe_2O_3 ______________________

4. Calculate the molecular mass of:

a. HF ______________________

b. NI_3 ______________________

c. P_4O_{10} ______________________

5. Calculate the molar mass of:

a. Barium nitride (Ba_3N_2) ______________________

b. Dinitrogen pentoxide (N_2O_5) ______________________

c. Potassium permanganate ($KMnO_4$) ______________________

6. Using $Al_2(SO_4)_3$, find the following:

a. How many grams are in 3.00 mol of $Al_2(SO_4)_3$?

____

b. How many moles are in 157 g of $Al_2(SO_4)_3$?

____

c. How many formula units are in 157 g of $Al_2(SO_4)_3$?

____

d. How many grams are in 9.72×10^{25} formula units of $Al_2(SO_4)_3$?

____

7. According to the reaction $CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$, how many grams of copper can be produced from 3.16 g of zinc with an excess of $CuSO_4$?

____

8. If 5.25 g of nitrogen gas reacts with 7.52 g of hydrogen gas to make ammonia in the reaction $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, answer the following:

a. What is the limiting reactant? ______________________

b. How much product is made? ______________________

c. How much leftover reactant is present? ______________________

9. In a reaction where 82.1 g of product is expected but only 46.3 g is produced, what is the percent yield?

____

10. What is the percent yield of Ag_2CO_3 if 25.0 g of $AgNO_3$ reacts with 25.0 g of Na_2CO_3 according to the reaction $2AgNO_3(aq) + Na_2CO_3(aq) \rightarrow Ag_2CO_3(s) + 2NaNO_3(aq)$ and 18.8 g of Ag_2CO_3 is actually produced?

____