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 a review of concepts from Chemistry I Honors. 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 review and not AP Chemistry material.

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

Matter

Classification, properties, and changes

·1·

Classification of matter

All matter can be classified as a substance or mixture depending on its composition. The word *substance* refers to elements or compounds with an exact chemical composition. *Elements* are composed of only one type of atom, such as hydrogen, H_2; *compounds* are composed of two or more types of atoms maintaining an exact ratio, such as water, H_2O.

In contrast, *mixtures*, such as lemonade, do not have such an exact composition, though overall they may have approximately fixed ratios of components. Mixtures can be identified as homogeneous or heterogeneous. *Homogeneous mixtures*, like clear lemonade, are also called *solutions*, since all the component parts are uniformly distributed and appear as one thing, which may have a certain concentration, but can be made to have almost any composition. Another example is air. The amount of different gases in air varies by place, time of day, and weather. The individual components in the mixtures retain their identity, such as the elements nitrogen (N_2) and oxygen (O_2) and the compound carbon dioxide (CO_2). The air also can be classified as heterogeneous.

Heterogeneous mixtures usually have visibly different components, which may be seen with the aid of instruments rather than the naked eye. This might be when you can see particulate matter such as black soot coming out the tailpipe of a truck into the air, or when using a microscope you can see different types of blood cells unevenly distributed. Other examples of heterogeneous mixtures would be Jell-O with fruit in it that can be seen, and uncut fruitcake that contains both external (seen) and internal (unseen) variations. Figure 1.1 shows how matter can be classified by its components.

Elements are substances that contain only one type of atom. *Atoms* cannot be separated into smaller units by normal means, but they can be broken down by nuclear reactions. (This will be discussed in Chapter 3.) Compounds, which have formulas to represent them, can, through chemical reactions, be broken into the atoms of elements composing them.

Mixtures are different, since they are not held together through the bonding of atoms; physical methods like filtering, sifting, and heating can recover the individual parts of the mixture. Marshmallows can be removed from hot cocoa. A saltwater solution can be separated into pure water and salt using distillation.

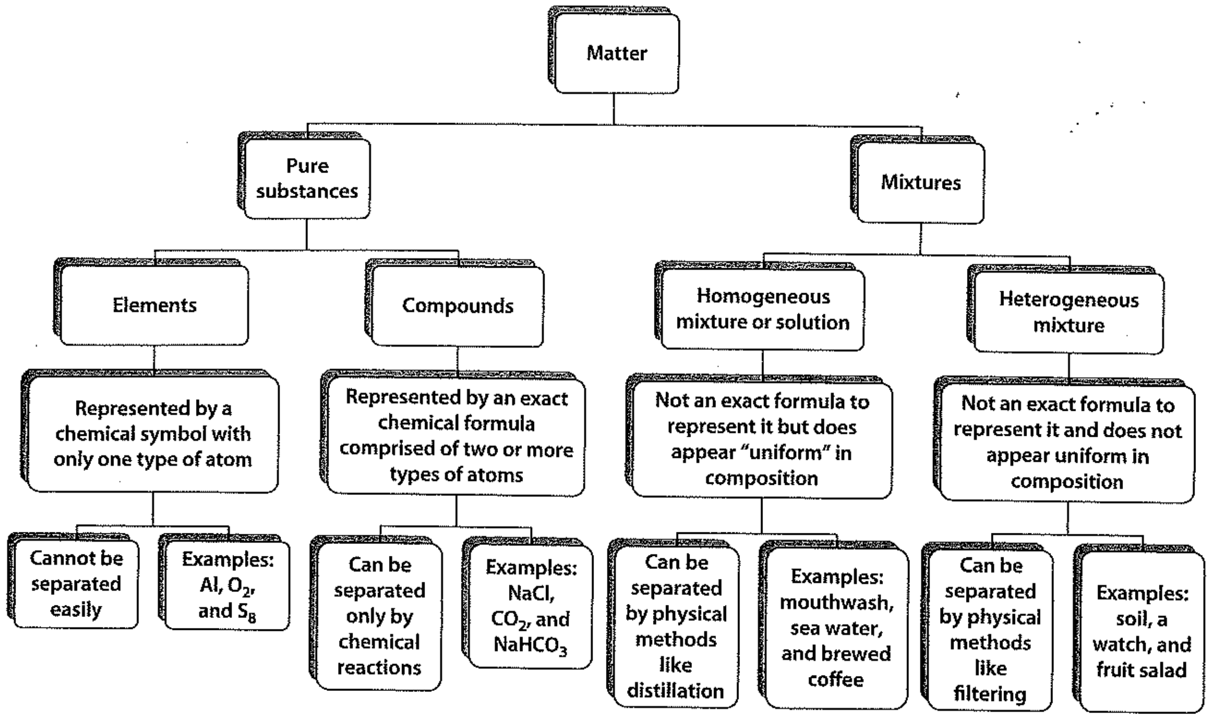


Figure 1.1

A question you can ask to help determine what type of matter you have might be, "Is there a way to get back the original components?" Can iron filings be separated from sand using a magnet? Yes—you can see the black iron filings in the light-colored sand—so that is a heterogeneous mixture.

EXERCISE

1·1

Classify each of the following examples of matter as an element, a compound, a homogeneous mixture, or a heterogeneous mixture.

1. Hot black coffee ______________________
2. Copper ______________________
3. Baking soda ______________________
4. Mixed vegetables ______________________
5. Soil ______________________
6. Chocolate ice cream ______________________
7. Sulfur ______________________
8. Non-iodized table salt ______________________
9. Italian salad dressing ______________________
10. Helium ______________________
11. Ocean water ______________________

12. Dry ice ______________________
13. Aluminum foil ______________________
14. A cheeseburger ______________________
15. Vitamin C ______________________

Properties of matter

Substances can be distinguished by their properties. These properties can be either physical or chemical. *Physical properties* can be observed or measured without changing the formula of the substance. *Chemical properties* are observed when a substance changes or does not change in a chemical reaction. Physical properties include melting point, boiling point, color, electrical conductivity, malleability, ductility, luster, density, and phase (solid, liquid, gas, or plasma—although we will not be studying plasmas in this book) at a set temperature and pressure. *Malleability* is how easily a metal can be molded into shapes, like a flat sheet, and ductility is its ability to be drawn into a wire. For example, lead is very soft and malleable and copper is very ductile.

Some physical properties depend on how much matter is present. These are called *extensive* properties; examples include mass and volume. If more matter is present, the mass will be greater and the space occupied will be greater. Properties that do not depend on the amount of matter present are called *intensive*. Temperature and density are examples of intensive properties. (Density will be discussed in a later section.)

Chemical properties are indications of either reactivity or non-reactivity with another substance. Iron can oxidize (rust) in air, indicating a chemical change. If it rusts, it changes from Fe to Fe_2O_3. Methane (CH_4) can burn with the oxygen gas in the air with a flame to change to into carbon dioxide and water. A substance with acidic properties can react with a base, chemically changing to form a salt and water. In each case of chemical change, the linking of the atoms changes (e.g., when CH_4 becomes CO_2 the carbon is now linked or bonded differently than before the change).

EXERCISE
1·2

Identify each of the following properties as being either physical or chemical.

1. Green color ______________________
2. Mass ______________________
3. Flammability ______________________
4. Freezing point ______________________
5. Reactivity with acid to form hydrogen ______________________
6. Boiling point at 90°C ______________________
7. Malleability ______________________
8. Resistance to corrosion ______________________
9. Conductivity of electricity ______________________
10. Dissolution in water ______________________

Identify each of the following properties as being either intensive or extensive.

11. Color
12. Length
13. Melting point
14. Height
15. Solubility in water at 20°C
16. Mass
17. Ductility
18. Boiling point
19. Density
20. Conductivity

Changes in matter

Physical change results when the amount or phase of a substance is changed but it is still the same substance. For instance, cutting a piece of copper in half does not change the copper to silver—there are just two pieces of copper. The physical form has changed but it is still the same substance. Another example is freezing water: ice is still H_2O, but instead of it being a liquid it is a solid.

In contrast, a chemical change changes the formula of the substance, since it has undergone a chemical reaction with another substance (see Figure 1.2). When iron (Fe) reacts with the oxygen (O_2) in the air to make iron(III) oxide (Fe_2O_3)—or, as it is more commonly known, rust—this is an example of a chemical change. Words that commonly indicate a chemical change include *flammability*, *corrosiveness*, *reactivity*, and *oxidation*. Household examples include iron rusting, milk souring, grass decomposing, and an egg frying. In each case, something with a new chemical composition is formed.

Physical properties of substances describe unchanging composition characteristics, while chemical properties describe chemical behaviors with other substances that result in composition (atom-linking arrangement) changes.

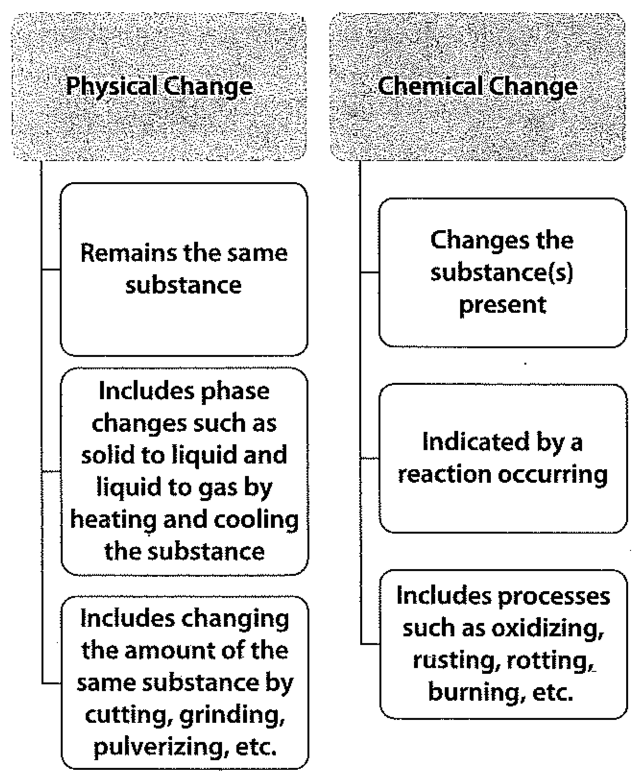


Figure 1.2

EXERCISE
1·3

Indicate which of the changes below are chemical. Check that the ones you don't mark are physical changes.

1. Wood rots ______________________________
2. Ice melts ______________________________
3. A cake bakes ______________________________
4. Paper is cut ______________________________
5. A plant dies ______________________________
6. Silver is made into a thin wire ______________________________
7. A car is dented ______________________________
8. Water is heated and turned into steam ______________________________
9. Alcohol evaporates ______________________________
10. A battery charges ______________________________
11. A piece of cake is split in two ______________________________
12. A piece of metal rusts ______________________________
13. Green leaves turn orange in fall ______________________________
14. A letter is opened ______________________________
15. Paint dries ______________________________

Density

Another physical property is density. *Density* is the ratio of the amount of matter present (mass) to the amount of space that matter occupies (volume). In other words, density is the ratio of mass to volume of a substance. For any substance (elements and compounds), the ratio remains constant at a given set of conditions.

Standard densities are reported at 1 atmosphere of pressure and a temperature of 25°C. The density of a substance can be used as an identifying property. Someone selling 24-karat (pure) gold could be easily checked, since the density of 24-karat gold is 19.3 g/cm^3, the density of 22-karat gold is 17.7 g/cm^3, and the density of 18-karat gold only 15.5 g/cm^3 (see Table 1.1).

Table 1.1 Density of Substances at 20°C

Substance	Density in g/cm^3	Substance	Density in g/mL
Gold	19.3	Acetone	0.79
Iron	7.9	Ethanol	0.772
Mercury	13.6	Gasoline	0.70
Silver	10.5	Hexane	0.658
Sodium chloride	2.2	Water	0.998
Zinc	7.14		

Identifying an unknown metal can be easily accomplished using the property of density. Measuring a mass of 62.5 g and a volume of 6.00 cm^3 for a metal gives a ratio of 10.4 to 1. The metal is most likely silver.

EXERCISE

1·4

Answer the following questions rounding the answer to three total digits. A discussion about proper rounding will be presented in Chapter 2.

1. A gold ring was purchased from a street vendor for a ridiculously low price. The buyer has brought the ring to you to confirm its metal content. You have found the mass of the ring to be 23.46 g and its volume to be 2.25 cm^3. What do you tell the buyer?

 ____

 ____

2. In order to make coins that have less mass but will still fit into vending machines, a country is looking at different metals to substitute for the ones that they presently use. Unfortunately, a technician forgot to include the identities of the metals with the test results. What metals were tested if the following table represents the results?

SAMPLE	MASS	VOLUME	IDENTITY
Metal 1	55.33 g	7.75 cm^3	____________
Metal 2	61.5 g	4.52 cm^3	____________
Metal 3	19.4 g	2.45 cm^3	____________

3. Determine the volume that 45.8 g of carbon tetrachloride will occupy if it has a density of 1.60 g/mL.

4. Determine the mass of zinc that will occupy 18.6 cm^3 if it has a density of 7.14 g/cm^3.

5. What is the density of lead if a piece of lead 1.00 cm by 20.0 cm by 30.0 cm has a mass of 6,800 g?

6. A sheet of aluminum has a mass of 27.0 g and measurements of 10.0 cm by 1.00 cm by 1.00 cm. Calculate the density of aluminum.

7. What is the density of a saltwater solution that has a mass of 17.84 g and a volume of 15.00 mL?

8. A piece of wood has a mass of 27.0 kg and a volume of 15.4 dm^3. What is its density, and will it sink or float in water?

9. At 20°C, what mass of ethanol will occupy 250 mL?

10. Platinum has a density of 21.410 g/cm^3. What volume of platinum has a mass of 50.0 g?

Graphical analysis of density

When comparing measured values for different-sized samples of a material, the plot of volume versus mass will result in a straight line. The slope of this line represents the density. Given the data in Table 1.2, a graph such as the one in Figure 1.3 can be made. By making a best-fit line through the points, the ratio of 1.0 to 1.0 is obtained. This means that no matter how the mass changes, the volume will change by the same ratio. If the mass is halved, the volume will also halve. If the mass is tripled, the volume will triple. This relationship is called a *direct relationship*, and since the slope always has the same value, this confirms that density is an intensive property.

Table 1.2 Density of Water at 25°C

Mass (in g)	Volume (in mL)
1.0	1.0
3.5	3.5
5.0	5.0
8.0	8.0
12.0	12.0
18.0	18.0
22.0	22.0
30.0	30.0
43.0	43.0

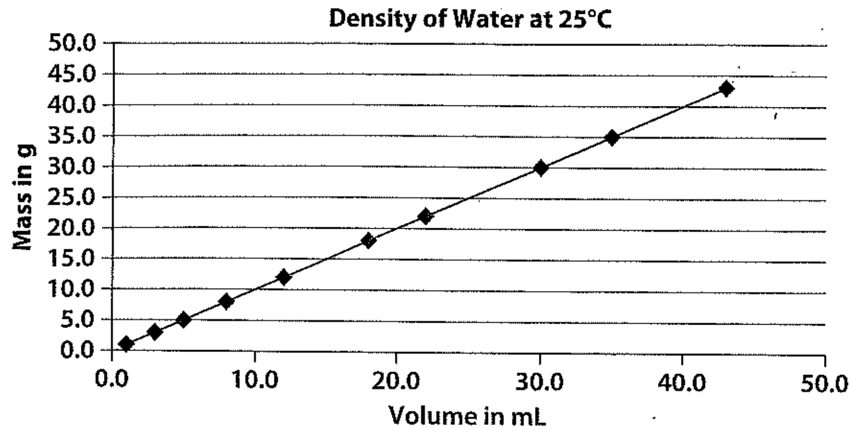
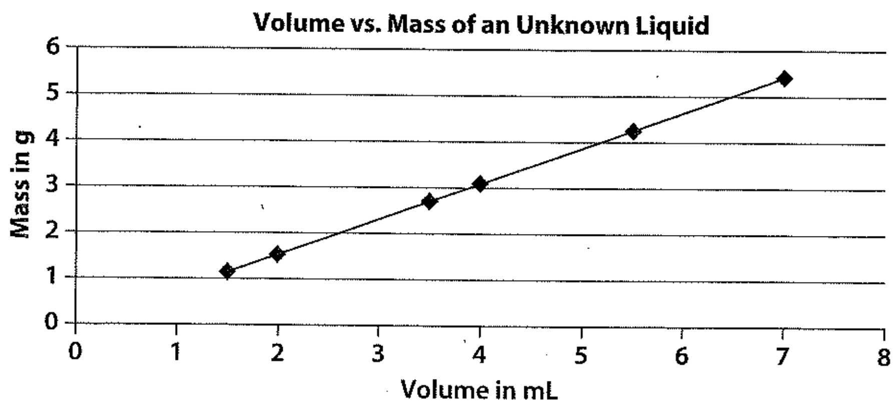


Figure 1.3

EXERCISE

1·5

Use the following graph and Table 1.1 on page 5 to answer questions 1 and 2.



1. The density of the unknown liquid is ______________________________.
2. The identity of the unknown liquid is ______________________________.

Use the information in the following table to answer questions 3 through 5.

Mass in g	Volume in mL
1.10	2.42
3.60	7.92
4.10	9.02
5.30	11.7
5.70	12.6

3. Using the data from the table, draw a graph of the volume versus the mass. Draw a best-fit line through the data and determine the slope of the line.

4. What is the density of the unknown object? ______________________

5. If it does not dissolve, will this substance sink beneath or float on the liquid in problems 1 and 2 of this exercise? ______________________

EXERCISE
1·6

Classify each of the following substances by placing a check mark in the appropriate column.

SUBSTANCE	ELEMENT	COMPOUND	HOMOGENOUS MIXTURE	HETEROGENEOUS MIXTURE
1. Grape jelly	______	______	______	______
2. Neon gas	______	______	______	______
3. Trail mix	______	______	______	______
4. Water	______	______	______	______
5. Beef stew	______	______	______	______

EXERCISE
1·7

Classify each of the following properties by placing check marks in the appropriate columns.

PROPERTY	PHYSICAL	CHEMICAL	INTENSIVE	EXTENSIVE
1. Boiling point	______	______	______	______
2. Height	______	______	______	______
3. Reactivity with water	______	______	______	______
4. Blue color	______	______	______	______

5. Toxicity	________	________	________	________
6. Density	________	________	________	________

EXERCISE

1·8

Indicate whether the following changes are physical, chemical, or both by placing a check mark in the appropriate column.

CHANGE	PHYSICAL	CHEMICAL	BOTH
1. An incandescent bulb burns out	________	________	________
2. A spoon is bent	________	________	________
3. Bread molds	________	________	________
4. Soda goes flat	________	________	________
5. Sweat evaporates	________	________	________
6. A lightbulb lights up	________	________	________
7. Copper is electroplated	________	________	________
8. Food is digested	________	________	________
9. An egg is hard-boiled	________	________	________
10. A propane barbecue is ignited	________	________	________

EXERCISE

1·9

Calculate the missing numeric value(s), rounding the answer to two total digits, and if not given, determine the identity of the unknown substance. All substances are at 20°C.

SUBSTANCE	MASS	VOLUME	DENSITY	IDENTITY
1	________	1.1 cm^3	19.3 g/cm^3	________
2	9.95 g	12.6 mL	________	________
3	1.056 g	________	________	sodium chloride
4	________	4.39 cm^3	________	iron
5	25 g	________	0.658 g/mL	________

Measurement, mathematical notations, and conversions

·2·

The metric system

The *International System of Units* is the system of science. This system uses the metric system. There are standard units you will need to practice working with in order to be comfortable: these include units for *mass* (grams), *length* (meters), and *volume* (liters). Some of the common prefixes in chemistry that may appear in front of the standard unit name include *deci-* (0.1), *centi-* (0.01), and *milli-* (0.001). Science most frequently uses *scientific notation* for expressing numerical values. Often a chemist will need to change between units, so understanding the size of a unit and its representation in scientific notation is important (see Table 2.1).

Table 2.1 Metric Prefixes

Prefix and symbol	Common word	Numeric notation	Scientific notation
tera-, T	trillion	1,000,000,000,000	1×10^{12}
giga-, G	billion	1,000,000,000	1×10^{9}
mega-, M	million	1,000,000	1×10^{6}
kilo-, k	thousand	1,000	1×10^{3}
hecto-, h	hundred	100	1×10^{2}
deca-, da	ten	10	1×10^{1}
deci-, d	tenth	0.1	1×10^{-1}
centi-, c	hundredth	0.01	1×10^{-2}
milli-, m	thousandth	0.001	1×10^{-3}
micro-, µ	millionth	0.000001	1×10^{-6}
nano-, n	billionth	0.000000001	1×10^{-9}
pico-, p	trillionth	0.000000000001	1×10^{-12}

Knowing the size of a unit will be important in dimensional analysis, which involves changing from one unit to another. Knowing there are 100 centimeters in 1 meter gives two possible ratios to use:

$$\frac{1 \text{ m}}{100 \text{ cm}} \text{ and } \frac{100 \text{ cm}}{1 \text{ m}}$$

How are measurements written? Measured values should always include units, and often scientific notation will be used. This notation will include only place values that were measured. How to make actual measurements will be covered in a later section, as will practicing conversions.

Any numeral 1 to 9 in a measurement is a measured value; the question is the 0—was it measured or is it a placeholder? Some statements of measurement, such as 120 g, are unclear, since the 0 is just a placeholder if only the hundreds and tens values were measured. If the placeholder 0 were left off, the measurement would be 12 and not 120 g. Just looking at the number, one does not know if the 0 was actually measured. This is where scientific notation is advantageous; this 0 would not be written in scientific notation if it is only a placeholder. We will address this concept in more detail in the section on significant figures.

To write in scientific notation, only one digit can be placed in front of the decimal point. Other measured digits are behind the decimal point. A measurement of 120 g would then be 1.2×10^2 g—or 1.20×10^2 g if the 0 were indeed part of the measurement. This way of expressing values is very clear as to the accuracy of the measurement. To determine the nonscientific notation measurement of 291 g, the number 2.91 is multiplied by 100, or 10^2. This then would have been written in scientific notation as 2.91×10^2 g. A measurement of 1.35×10^3 g could be 1,350 g, but it could also be 1,352 g or 1,348 g—we don't know for sure. On the other hand, 1.350×10^3 g is *definitely* 1,350 g. Small values are treated the same way, but with negative exponents: 0.0055 g is 5.5×10^{-3} g.

Generally, the exponents of 1 (10^1 and 10^{-1}) are not used, though this can lead to some uncertainty. Is the value 10 g exactly 10 g or is it only approximately 10 g? Using 1.0×10^1 would be more certain but is not common practice. Another way to express that value is 10. g (notice the decimal point after the 0).

EXERCISE 2·1

Express the following measured values in scientific notation.

1. 750 g (measured to the tens place) ______
2. 94,632 mL ______
3. 0.010 kg (measured to the thousandths place) ______
4. 0.0058 L ______
5. 200,000 mg (measured to the thousands place) ______
6. 802.0 g ______
7. 5,050.0 mL ______
8. 220. mL ______
9. 0.00101 g ______
10. 0.002 g ______

Express the following values as numbers.

11. 5.1×10^{-4} cm ______
12. 2.8×10^3 g ______

13. 5.2101×10^2 mL

14. 6.33×10^{-5} km

15. 4.60×10^3 g

16. 4×10^{-4} cm^3

17. 1.600×10^2 g

18. 2.834×10^6 g

19. 3.240×10^{-4} mL

20. 2.705×10^{-1} g

Accuracy in measurements

When a measurement is going to be made, it is important to consider the device being used. No device is perfect; therefore, one should consider the accuracy of the devices being considered. The one that can give a measurement closest to the true value is the best. When reading a device such as a beaker or graduated cylinder, look at the marks present. (See Figures 2.1, 2.2, 2.3, and 2.4 for examples of these devices.) A measurement reads by place value the units marked, and then an estimate is made for the next smaller place value. The unit markings you can use are in metric units of hundredths, tenths, ones, tens, hundreds, and so on.

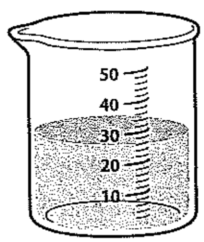


Figure 2.1

If the markings are at 10 mL, 20 mL, and 30 mL, and so on, the place value you can read is the tens. The next smaller place value is the ones, so a measurement made can be recorded only to the ones place. The measurement in Figure 2.1 would be 26 mL, not 26.0 mL.

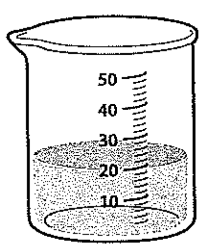


Figure 2.2

Water and aqueous solutions exhibit a downward curvature in the liquid surface, called a *meniscus*, when in a glass container. If the bottom of the meniscus is right on the 20 mL marking, as in Figure 2.2, then a decimal is put after the 0 (20. mL) or the number is written in scientific notation (2.0×10^1 mL). Why is this important? It is difficult otherwise to determine from the written record of a measurement whether the value represented by a 0 was actually measured.

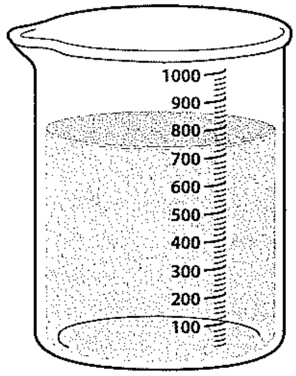
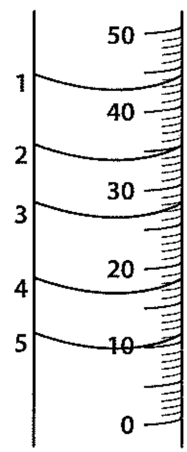


Figure 2.3

What if the beaker is marked at 100 mL, 200 mL, and so on, as in Figure 2.3? The next place value would be the tens. A measurement of 740 mL would be possible. Here the 0 is a placeholder only and does not represent a value that was measured. There would not be a decimal point after the 0, and if written in scientific notation, the number would be 7.4×10^2 mL.

EXERCISE
2·2

Using the figure provided, record the measurement at each marked location.

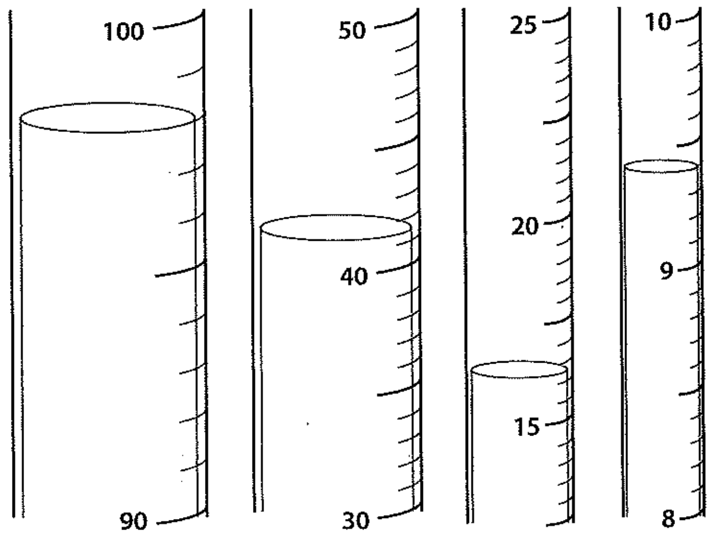


1. ______________________________
2. ______________________________
3. ______________________________
4. ______________________________
5. ______________________________

EXERCISE

2·3

Using the figures below (note the different scales), record the volume of each substance being measured. The unit used on each device is milliliters.



1. ____________________
2. ____________________
3. ____________________
4. ____________________

Significant figures

The digits measured in a measurement and expressed as its value are called significant figures ("sig figs"). One of the ways scientists communicate their results to other scientists is by using significant figures. For instance, if you polled 10 people and 9 liked pepperoni on their pizza, you could say 90% of people prefer pepperoni. What if you asked 100 people, 1,000 people, or 10,000 people? Does 90% really tell you anything unless you know about sample size? Significant figures would distinguish them as 90.%, 90.0%, 90.00%, and 90.000%. Significant figures also affect the way mathematical operations are done. The number of significant figures reported is important when doing mathematical operations. Before doing mathematical operations, you must understand how many digits in each value are actually significant. Counting sig figs requires practice and knowledge of measurements. Here are some rules to help:

- The numerals 1 through 9 are *always* significant.

  1,247 g has four sig figs.

- Zeros that begin numbers are *never* significant. They are only placeholders.

  0.021 g has two sig figs.

- Zeros that are between the numerals 1 through 9 are *always* significant.

  102 g has three sig figs.

- Zeros that trail a number may or may not have been measured—it depends on the presence of a decimal point! If a decimal point *is present*, the trailing 0 was measured.

  370 g has two sig figs. No decimal point is present, so the 0 is a placeholder. A better way to display this value would be 3.7×10^2 g.

  370. g has three sig figs. The decimal point indicates that the 0 was measured.

  0.0250 g has three sig figs. The leading 0s are placeholders, but the trailing 0 was measured.

So why is this important? When reporting experimental data, scientists communicate the details of their work through significant figures.

EXERCISE 2·4

Determine the number of significant figures in each of the following numbers.

1. 110 m ______
2. 633 g ______
3. 900 mL ______
4. 0.250 g ______
5. 8.08 m ______
6. 0.0007446 km ______
7. 5,000,000 g ______
8. 40. L ______
9. 0.007310 kg ______
10. 0.227568 Mm ______
11. 0.157 kg ______
12. 2,500 m ______
13. 0.0250 cm^3 ______
14. 2.680 g ______
15. 44.07 cm ______
16. 50.00 g ______
17. 21 mL ______
18. 8.28 km ______
19. 6.0001 g ______
20. 0.01 g ______

Calculations with significant figures

When making calculations with significant figures, there are additional rules to follow. Two rules apply to the basic mathematical operations of addition, subtraction, multiplication, and division.

- When adding or subtracting two or more measurements, round the answer to the least-accurate decimal place measured. It is also vital that the values be expressed in the same units.

  35.74 mL − 2.4 mL = 33.3 mL

  When this difference is set up as

  35.74 mL
  −2.4 mL

  you can see that the last digit in 35.74 has no digit under it. Since one number was measured to the hundredths place and the other to the tenths place, the least-accurate decimal place is the tenths.

> If the measurements being subtracted or added are given in scientific notation, it is better to first convert the values to the same scientific expression. For example, 13.578×10^2 g − 6.2355×10^{-1} g = 13.572 g can be rewritten as 13.578×10^2 g − 0.0062355×10^2 g to make the placement of the last significant figure more apparent.

- When multiplying or dividing measurements, count the number of sig figs in the numbers and round the answer to the smallest number of sig figs.

  2.50 cm × 1.2 cm = 3.0 cm^2, not 3.00 cm^2, since 2.50 cm has three significant figures and 1.2 cm has two significant figures.

- Exact numbers used in a conversion *do not* count in sig figs (more on this in a couple sections).

  1 kg = 1,000 g

- When a number needs to be rounded, look at only the first digit past the last significant figure. If it is a 5 or higher, round up the last significant figure; if it is a 4 or lower, leave the last significant figure as it is written.

  27.56 g rounded to three significant figures is 27.6 g; 27.46 g rounded to two significant figures is 27 g.

> Sometimes a combination of operations can occur. In that case, the subtraction and addition are completed and rounded first before the multiplication and division are done. An example of this is solving for density using the water-displacement method.

Density is the ratio of the mass to the volume of an object. Assume water is placed in a 100. mL graduated cylinder with a measurement of 62.7 mL before a solid object is submerged in the water. The final volume of the water and the object is 79.2 mL. The object has a mass of 12.875 g. What is the density of the object? First, we must do the subtraction. Both values are measured to the tenths place, so the final answer is rounded to the tenths place: 79.2 mL − 62.7 mL = 16.5 mL.

Next, the division step can be done. A five-digit mass is being divided by a three-digit volume, so the answer must be rounded to three sig figs.

$$\frac{12.875\ \text{g}}{16.5\ \text{mL}} = 0.780\ \text{g/mL}$$

If the problem involves multiple operations but is all multiplication and division, round only once at the end instead of after each operation! When rounding numbers, look *only* at the first digit after the last significant figure. 32.45 g rounded to two digits is 32 g; 32.54 g rounded to two digits is 33 g.

EXERCISE

2·5

Complete the following calculations, expressing your answers with the correct number of significant figures.

1. 5.5 m + 0.781 m = __________
2. 907 mL − 65.2 mL = __________
3. 30.75 cm − 9 cm + 15.3 cm = __________
4. 1240 mg + 38.6 mg − 471.22 mg = __________
5. 9.307 m × 1.5 m = __________
6. 253.14 cm^3 ÷ 7.58 cm^2 = __________
7. $\frac{0.584\ \text{kg}}{4.1\ \text{m} \times 2.36\ \text{m} \times 0.075\ \text{m}} =$ __________
8. $\frac{7.83\ \text{m} - 1\ \text{m}}{7.83\ \text{m}} =$ __________
9. 18.5 cm + 8 cm + 4.73 cm = __________
10. 35.0 m/200 s = __________
11. 39.0 g ÷ 25.82 mL = __________
12. (25.43 cm − 12.0 cm) × 14.37 cm = __________
13. 3.15 dm × 4.0 dm = __________
14. 350.0 m − 200 m = __________
15. 16.53 cm − 6.3 cm = __________
16. 3.00 m × 5.2 m = __________
17. 13.563 L − 11 L = __________
18. 15.54 g + 12 g = __________
19. 0.075 g ÷ 0.003 cm^3 = __________
20. 132 g + 10.00 g + 9.6 g = __________

Accuracy vs. precision

When looking at the devices in Figure 2.4, look at the measurements—if calibrated correctly, which device should give a measurement closer to the real or true value?

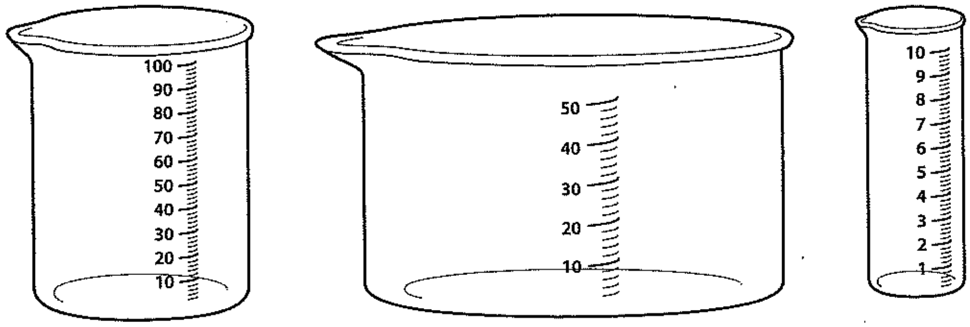


Figure 2.4

When comparing measurements, the one closest to the true value is said to be the most accurate. How does this compare to precision? Precision is how close a group of repeated measurements are to one another.

Look at the following measurements taken by different lab groups measuring the mass of the same material. The true value is 3.00 g.

	Group 1	Group 2	Group 3
Measurement 1	2.99 g	2.88 g	2.99 g
Measurement 2	2.50 g	2.87 g	2.98 g
Measurement 3	2.65 g	2.89 g	3.01 g
Average	2.71 g	2.88 g	2.99 g

Group 1's measurements lack both accuracy and precision. The measurements are not close one another, differing from the highest to lowest (called the range) by ±0.49, and are not close to the true value, with an average of 2.71 g.

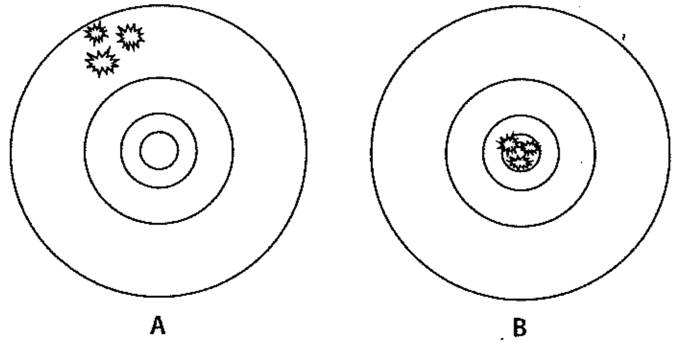
Group 2's measurements have a precision range of ±0.02 (the range from the highest measurement to the lowest measurement)—the measurements are close to one another—but although more accurate than group 1, with an average of 2.88 g, they are not as accurate as group 3, with an average of 2.99 g.

Inaccuracy is often due to a device not being calibrated. If you get on a bathroom scale and it reads 112 pounds repeatedly, but your true weight is 125 pounds, the scale needs to be recalibrated. You would be getting measurements close to one another, so high precision, but not close to the true value of your weight. Group 3's measurements have a precision of ±0.03—not quite as good as group 2's measurements, but they are more accurate, since their measurements are closer to the real value.

EXERCISE
2·6

1. Two students went into the lab and measured the volume of water in the same graduated cylinder. Student 1 recorded the following measurements: 45.2 mL, 42.2 mL, and 43.3 mL. Student 2 recorded the following measurements: 43.9 mL, 43.5 mL, and 43.8 mL. The teacher had actually placed 43.7 mL of water in the graduated cylinder. Which student's measurements were more accurate? Which student's measurements were more precise?

Use the terms accurate *and* precise *to describe the shots fired at the targets in the figure provided.*



2. The shots on target A are ______________________.
3. The shots on target B are ______________________.

Dimensional analysis

Often in chemistry, units need to be changed to other units. This is done through a process called *dimensional analysis*. It is usually just a set of multiplication steps using what are called conversion factors. A *conversion factor* is simply the ratio of two measurements to each other. This was introduced in the section on the metric system, where the ratios $\frac{1\text{ m}}{100\text{ cm}}$ and $\frac{100\text{ cm}}{1\text{ m}}$ both state the relationship between the two units. In dimensional analysis, the conversion factor to use is the one that cancels units. For instance, the simplest conversion would be a one-step problem such as changing 5.89 mL into L.

$$58.9\ \cancel{\text{mL}} \times \frac{1\text{ L}}{1{,}000\ \cancel{\text{mL}}} = 0.0589\text{ L, also written as } 5.89 \times 10^{-2}\text{ L}$$

The two possible ratios are: $\frac{1\text{ L}}{1{,}000\text{ mL}}$ and $\frac{1{,}000\text{ mL}}{1\text{ L}}$. The ratio used has the units set up to cancel so that the correct unit remains.

A two-step problem might be changing 37.0 km/hr into m/s. This requires more steps, with each conversion done in a separate step: km to m and hr to s.

$$\frac{37.0\ \cancel{\text{km}}}{1\ \cancel{\text{hr}}} \times \frac{1{,}000\text{ m}}{1\ \cancel{\text{km}}} \times \frac{1\ \cancel{\text{hr}}}{3{,}600\text{ s}} = 10.3\text{ m/s, also written as } 10.3\text{ ms}^{-1}$$

Conversion problems can range from one step to many steps. This skill is important to master!

EXERCISE

2·7

Write in the blanks the conversion factor(s) needed to complete the following conversions.

1. 2.5 kg to g ____________________
2. 0.783 L to mL ____________________
3. 91.4 cm to m ____________________
4. 6.0 m/min to m/s ____________________
5. 0.846 m/s to km/hr ____________________
6. 0.75 kg to mg ____________________
7. 1,500 nm to km ____________________
8. 0.52 kg to g ____________________
9. 65 cm to km ____________________
10. 1.00 day to s ____________________

EXERCISE

2·8

Using the conversion factor(s) selected above, complete the conversions using dimensional analysis, and record your answers in the blanks provided.

1. 2.5 kg to g ____________________
2. 0.783 L to mL ____________________
3. 91.4 cm to m ____________________
4. 6.0 m/min to m/s ____________________
5. 0.846 m/s to km/hr ____________________
6. 0.75 kg to mg ____________________
7. 1,500 nm to km ____________________
8. 0.52 kg to g ____________________
9. 65 cm to km ____________________
10. 1.00 day to s ____________________

Any relationship can be used as a conversion factor. In Chapter 1 density calculations were covered. Density is a ratio that can be used as a conversion factor! For example, if a substance has a density of 2.50 g/cm^3 and there are 8.50 g of it, what volume would the sample occupy?

$$8.50\ \cancel{g} \times \frac{1\ \text{cm}^3}{2.50\ \cancel{g}} = 3.40\ \text{cm}^3$$

The density ratio was arranged to have units cancel, so the final answer has the units of volume, in this case, cubic centimeters.

EXERCISE

2·9

Using dimensional analysis, solve the following problems.

1. What is the mass of 4.5 cm^3 of gold with a density of 19.3 g/cm^3?

   ____

2. What volume does a 250 g sample of mercury occupy? The density of mercury is 13.6 g/mL.

   ____

3. What volume of hydrochloric acid solution with a density of 1.16 g/mL is needed to obtain a mass of 51.4 g?

   ____

4. If the density of nitric acid is 1.25 g/mL, what is the mass of 23 mL of nitric acid?

   ____

5. The Hope diamond has a mass of 9.1 g. Diamonds have a density of 3.614 g/cm^3. What is the volume of the Hope diamond?

   ____

EXERCISE

2·10

Express the following numbers in scientific notation.

1. 456 m ______________
2. 0.0086 kg ______________
3. 10,000,000 μL ______________
4. 2000 cm ______________
5. 0.0000000075 Tg ______________

EXERCISE

2·11

How many significant figures are in each of the following measurements?

1. 300,000,000 mm ____________________
2. 0.750 L ____________________
3. 800.0 g ____________________

Still stymied by significant figures? Try this little trick. Look at Figure 2.5 and ask yourself these questions:

1. Where is the Pacific Ocean?
2. Where is the Atlantic Ocean?
3. When looking at the number, is the decimal point present or absent?

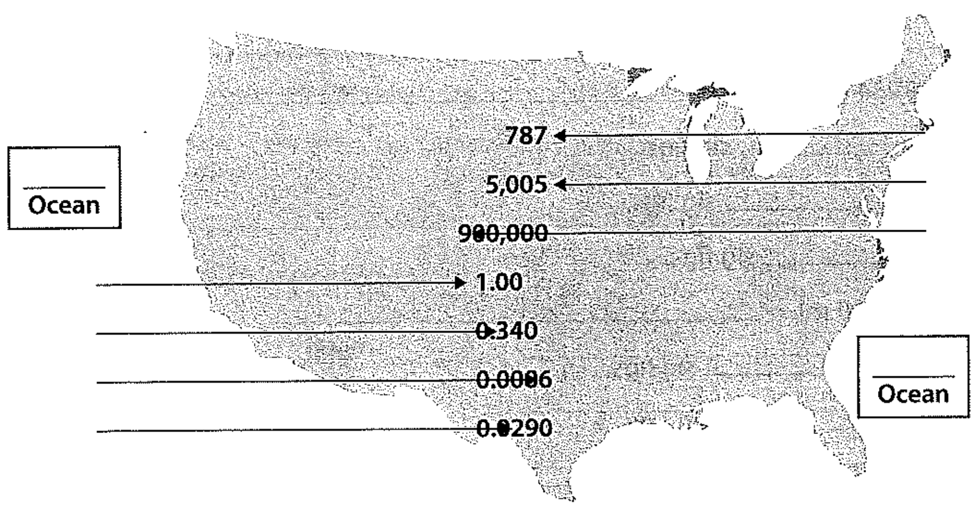


Figure 2.5

Next use the following tips to determine significant numbers:

- If the decimal point is absent (not written), draw an arrow from the Atlantic Ocean until you touch a nonzero number (1, 2, 3, 4, 5, 6, 7, 8, or 9). Any digit that your line does *not* go through is *significant*.

  Since 787 has three digits that you did *not* cross through, it has three significant figures.

  5,005, which has 0s that are trapped between two nonzero digits, has four digits that are not crossed through, so it has four significant figures.

  For the number 900,000, you crossed through five 0s, leaving only the one 9 not marked through—so 900,000 has one significant figure.
- If the decimal point is present (written), draw an arrow from the Pacific Ocean until you touch a nonzero number. Again, any digit that your line does *not* cross through is a significant figure.

  For the number 1.00, the arrow stops at the 1. No digits are crossed through, so 1.00 has three significant figures.

For the number 0.340, only the 0 to the left of the decimal point is crossed through, so 0.340 has three significant figures.

In the number 0.0006, the decimal point is present and the arrow is drawn all the way to the 6. All four 0s are crossed through—only the 6 remains—so 0.0006 has one significant figure.

In the last example, 0.0290, the first two 0s are crossed through when the arrow stops at the 2, leaving the 2, the 9, and the last 0 unmarked. Therefore 0.0290 has three significant figures.

With this little trick, all the different types of 0s have been addressed: trapped, trailing without a decimal point, trailing with a decimal point, and leading. Hopefully this will help you determine which digits are significant in a measurement.

EXERCISE

2·12

Determine how many significant figures are in each of the calculations below. Remember that the number of significant figures in a measurement does not always determine how many significant figures there are in your answer. Next, complete the calculation, rounding the answer so that it has the correct number of significant figures.

1. 45.6 g + 0.234 g + 0.87 g = ______________________

   45.6 g has __________ sig figs.

   0.234 g has __________ sig figs.

   0.87 has __________ sig figs.

2. 1,808 mL − 5.00 mL = ______________________

   1,808 mL has __________ sig figs.

   5.00 mL has __________ sig figs.

3. 0.00685 km × 0.0007 km = ______________________

   0.00685 km has __________ sig figs.

   0.0007 km has __________ sig figs.

4. 9,040,000 m ÷ 35,000 s = ______________________

   9,040,000 m has __________ sig figs.

   35,000 s has __________ sig figs.

5. $\dfrac{5.36\,\text{g}}{(4.5\,\text{mL} - 2.2\,\text{mL})}$ = ______________________

   5.36 g has __________ sig figs.

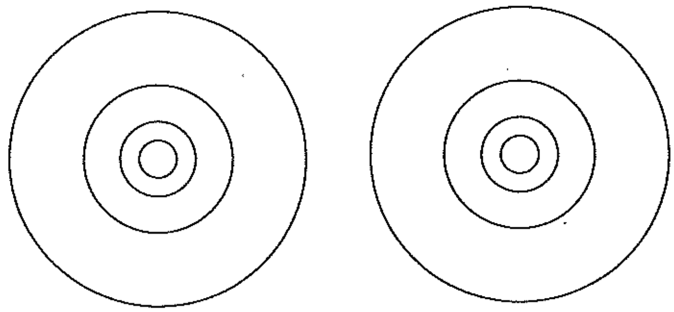
   4.5 mL has __________ sig figs.

   2.2 mL has __________ sig figs.

EXERCISE

2·13

Draw the indicated number of shots on each target that best represents the given description.



A B

1. On target A, draw one shot that is accurate.
2. On target B, draw five shots that are neither accurate nor precise.

EXERCISE

2·14

Complete the following conversions.

1. How many seconds are in 12.0 hr? ____________________
2. How many liters are in 0.50 μ? ____________________
3. How many nanometers are in 5 km? ____________________
4. How many m/s are in 120 km/hr? ____________________
5. How many g/mL are in 0.00486 kg/m^3?
   (For this problem use the conversions
   1 mL = 1 cm^3 and 1,000,000 cm^3 = 1 m^3.) ____________________

Atomic structure and nuclear reactions

3

Atomic structure

Even though there are many tiny subatomic particles, in general the chemist is concerned with the electrons, protons, and neutrons. The concept of the atom and its structure has undergone many changes through time. The term *atom* was introduced by Democritus, who lived between 460 and 370 B.C. He stated atoms were indestructible and indivisible, and differed only in size, shape, and motion. The main philosophers of the time—including Aristotle—disagreed with Democritus. They argued instead there were four elements (earth, air, water, and fire) with four qualities (dryness, hotness, coldness, and moistness), and also two forces: conflict and harmony. The concept of the atom was set aside until John Dalton (1766–1844) reintroduced it. Dalton stated that:

- *Elements* are made of small particles called atoms.
- *Atoms* of the same element are identical, with the same size, mass, and properties.
- *Compounds* are composed of atoms of different elements combined in a set ratio to each other.
- A *chemical reaction* involves rearranging the atoms.

As further experiments were done, ideas about the atom underwent revision. The charged particles were found first—the electron (by J. J. Thomson in 1897) and the proton (by Ernest Rutherford in 1919). The last particle discovered was the neutron (by James Chadwick in 1932). Thomson used cathode-ray tubes to find electron particles deflected by an electric field, and he also found the charge-to-mass ratio of the electron. Robert Millikan in 1909 did an oil-drop experiment and established the elementary electric charge of a single electron. At first Thomson hypothesized these particles were throughout the atom, like plums in pudding (plum-pudding model). In another experiment called the gold-foil experiment, Rutherford was astonished to discover the atom had a *nucleus* composed of a concentrated positive charge, with the electrons outside the nucleus in an area that was mostly empty space. This experiment changed the model of the atom to the current basic structure. Other experiments were done by Neils Bohr, and theories were developed to explain how the electrons were arranged in the atom. The Bohr model of the hydrogen atom proposed that electrons orbit the nucleus. Later, parts of the Bohr model were found to be incorrect and other models were developed, which will be explored in Chapter 4. From these experiments and ideas, the basic structure of the atom has been determined.

The basic particles of the atom found in the nucleus are the *proton* and *neutron*. The *electron* is found outside the nucleus in shaped electron clouds called *orbitals*. We will look closely at electron orbitals in Chapter 4. The mass of a proton and the mass of a neutron are almost equal to each other and are each considered to be approximately 1.0 amu (atomic mass unit). The electron's mass is extremely small in comparison (1/1840 amu).

For any element, the *atomic number* is the number of protons in any atom of that element. It defines the element. How can there be different atoms of the same element? All atoms of an element have the same number of protons, but they can have different numbers of neutrons. (This is like all the members of a family having the same last name but different first names.) These atoms with different numbers of neutrons are called *isotopes* of the element.

Since an atom is not charged, the number of electrons (e^-) in an atom will be equal to the number of protons (p^+). The positive charge of the protons is cancelled by the negative charge of the electrons: $p^+ - e^- = 0$. If the number of protons and the number of electrons are not equal, the particle is called an ion. If the ion is positively charged—due to its having more protons than electrons—it is called a cation. An example is the lithium cation, with three protons and two electrons resulting in a positive charge of 1: $3^+ - 2^- = 1^+$. This is written as Li^+. If the ion is negatively charged, due to its having more electrons than protons, it is called an anion. For instance, the oxide anion has eight protons and 10 electrons, resulting in a charge of 2 negative: $8^+ - 10^- = 2^-$. This is written as O^{2-}.

> Note that when the charge is only 1, the 1 is not included in the symbol; when it is greater than 1, then the integer is included. Charges are always integers—one-half or other fractions of an electron or proton do not exist.

How are the number of neutrons calculated? Each isotope has a *mass number*, an integer which is equal to the number of protons plus the number of neutrons. The mass number (A) minus the atomic number (Z) equals the number of neutrons (n^0): $A - Z = n^0$. For example, uranium-238, also written as ${}^{238}_{92}U$, has a mass number of 238 (the sum of the protons and neutrons is at the top left of the isotope notation) and an atomic number of 92 (the atomic number, which is the number of protons, is found at the bottom left of the isotope notation); the number of neutrons is therefore 238 minus 92, or 146. The isotope notation then sets up nicely as a subtraction problem, with the mass on top and the number of protons below:

$$\begin{array}{r} 238 \\ \underline{-92} \\ =146 \end{array}$$

Sometimes the bottom number—the atomic number—is not given in the notation, since it is understood based on the given element symbol. *All* atoms of uranium (U) have an atomic number of 92. If you know the element, you can find the number of protons by looking for the atomic number on the periodic table. For instance, ${}^{63}Cu$ has a mass number of 63 and has an atomic number of 29, therefore $63 - 29 = 34$ neutrons. The following table has additional examples for reference. Notice that the charge on the ion does *not* change the number of neutrons or protons, only the number of electrons.

Element isotope/ ion symbol	**Proton number**	**Neutron number**	**Electron number**	**Atom or ion**
lead-208	82	$208 - 82 = 126$	82	atom
${}^{32}S^{2-}$	16	$32 - 16 = 16$	$16 + x = -2; x = 18$	ion (anion)
${}^{30}Ca^{2+}$	20	$40 - 20 = 20$	$20 + x = 2; x = 18$	ion (cation)

EXERCISE 3·1

Complete the following table using the information given and the periodic table at the end of the book.

PARTICLE	MASS #	ATOMIC #	# OF p^+	# OF n^0	# OF e^-
1. Phosphorus-31	______	______	______	______	______
2. Cobalt-60	______	______	______	______	______
3. ______	14	6	______	______	6
4. ______	______	______	17	20	18
5. Calcium-43	______	______	20	______	______
6. ______	______	94	______	148	94
7. ______	50	24	______	______	24
8. ______	65	29	______	______	27
9. ^{15}N	______	______	______	______	______
10. $^{34}S^{2-}$	______	______	______	______	______

EXERCISE 3·2

Answer the following questions.

1. What is the isotope notation for phosphorus-31? ______
2. What is the isotope notation for cobalt-60? ______
3. What is the isotope notation for the particle in question 3 in Exercise 3-1? ______
4. What is the isotope notation for the particle in question 4 in Exercise 3-1? ______
5. Write the isotope notation for an isotope with a mass of 211 and a Z of 82. ______
6. An isotope has an A of 46 and a Z of 22. Write the notation for this isotope. ______
7. Write the isotope notation of an atom with a Z of 86 and an A of 222. ______
8. Write the isotope notation for magnesium-25. ______
9. What is the isotope notation for iridium-191? ______
10. What is the isotope notation for neon-22? ______

Average mass

All elements have isotopes. Every hydrogen atom has one proton, no matter which isotope of hydrogen it is. Hydrogen-1, or ^{1}H, has one proton and no neutrons; it is often referred to as protium. Hydrogen-2, or ^{2}H, has one proton and one neutron; it is often referred to as deuterium. Hydrogen-3, or ^{3}H, has one proton and two neutrons and is often referred to as tritium.

So why isn't the mass of hydrogen on the periodic table given as 1, 2, or 3? The mass on the table is the average mass of a sample of the element, which takes into account its isotopic mass times the percentage of occurrence (or percent abundance) in the sample, written as a decimal. From the hydrogen mass of 1.008 on the periodic table, we can see that hydrogen-1 is the most abundant isotope, since the average is extremely close to 1 amu! For hydrogen, any sample will contain 99.9885% hydrogen-1, 0.0115% hydrogen-2, and only a trace of hydrogen-3. It is also true that the protons and neutrons do not each have an exact mass of 1, and that when a nucleus is formed some of the mass of the components (protons and neutrons are collectively called "nucleons") is released as energy according to Einstein's famous equation $E = mc^2$. So there are three reasons why actual atomic masses are *not* whole numbers and are not equal to an element's mass number.

For copper there are two stable isotopes in nature: copper-63 and copper-65. Cu-63 occurs at 69.17% and Cu-65 at 30.83% and they add to 100.00%. To find the average mass, the percentages must first be converted to decimal form. Using the formula of mass times the decimal abundance for each isotope and then adding the results together, we get $(63 \times 0.6917) + (65 \times 0.3083) = 63.62$ amu. An atomic mass unit has a mass of 1.9×10^{-23} g. This means the atomic mass found on the periodic table is *not* the mass in grams of an atom! (There will be more about this later, in Chapter 7.) To find the mass in grams of one atom, the mass in amu would be multiplied by the ratio $\dfrac{1.9\times10^{-23}\,\text{g}}{1\ \text{amu}}$.

EXERCISE 3·3

Calculate the atomic mass of each of the following elements using the percent abundances given for each isotope.

1. Magnesium-24 78.99

   Magnesium-25 10.00

   Magnesium-26 11.01

   ____

   ____

2. Bromine-79 50.69

   Bromine-81 49.31

   ____

   ____

The periodic table, periodicity, and periodic trends

•4•

Historical background

The main organizing "tool" for the chemist is the *Periodic Table.* Understanding the construction of the table will allow you to unlock many of the components of chemistry. There have been and continue to be many different versions of the periodic table, however there is a version that has been accepted by the international chemistry community. Dmitri Mendeleev, Henry Moseley, and Glenn Seaborg, along with many others, contributed to the development of the current periodic table. Although the basic structure has not changed recently, the table is still expanding as new elements are synthesized, isolated, and identified.

Mendeleev arranged the known elements of his time by increasing atomic mass, included columns of repeating properties, attributed apparent problems with this structure to incorrect masses, and predicted the properties of four elements: scandium, technetium, gallium, and germanium (see Table 4.1).

Moseley changed the table to an arrangement where the elements were in order by increasing proton number instead of mass, thus fixing the issues in Mendeleev's table, and Seaborg moved the lanthanides and actinides from the body of the table to the rows below the table and started adding new elements in correct locations.

Thus as our knowledge of the elements has expanded, the arrangement of the periodic table has undergone revision. Although the history of chemistry is interesting, the focus of this chapter is understanding what we now know. Does the periodic table look the same worldwide? Yes. Although the names of the elements may differ in different languages, the symbols are universal. Cl represents the same element, whether scientists in the United States call it *chlorine* or those in France call it *chlore.*

The arrangements

The periodic table has columns and rows. The columns are called *families* or *groups*, and the rows are called *periods*. Different variations of the periodic table's column-numbering system exist. Which one should be used? The system used by the *IUPAC*—the International Union of Pure and Applied Chemistry—is the standard. If any questions arise about accepted formats of the periodic table, you should look at these standards; they can be found in a current edition of the *CRC Handbook of Chemistry and Physics.*

Table 4.1 A Rendering of Mendeleev's 1872 Periodic Table

Tabelle II

Reihen	Gruppe I	Gruppe II	Gruppe III	Gruppe IV	Gruppe V	Gruppe VI	Gruppe VII	Gruppe VIII
				RH^4	RH^3	RH^2	RH	—
	R^2O	RO	R^2O^3	RO^2	R^2O^5	RO^3	R^2O^7	RO^4
1	H = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	— = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63
5	(Cu = 63)	Zn = 65	— = 68	— = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	— = 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	J = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	—	—	—	———
9	(—)	—	—	—	—	—	—	———
10	—	—	?Er = 178	?La = 180	Ta = 182	W = 184	—	Os = 195, Ir = 197, Pt = 198, Au = 199
11	(Au = 199)	Hg = 200	Tl = 204	Pb = 207	Bi = 208	—	—	———
12	—	—	—	Th = 231	—	U = 240	—	———

Note: *Gruppe* means "group" and *Reihen* means "row." Also in the table, below the group labels, are the predicted formulas with hydrogen and with oxygen for each group. Notice that the accepted format during that period was to use a superscript instead of a subscript for the number of each atom in the formula. Masses given without symbols indicate elements not known at the time but predicted by Mendeleev. His predictions were later proved correct and identified as scandium, gallium, and germanium.

The periodic table is composed of elements found as solids, liquids, and gases at room temperature, as shown in Figure 4.1. Elements are also classified as metals, metalloids (also called semimetals), and nonmetals, as in Figure 4.2.

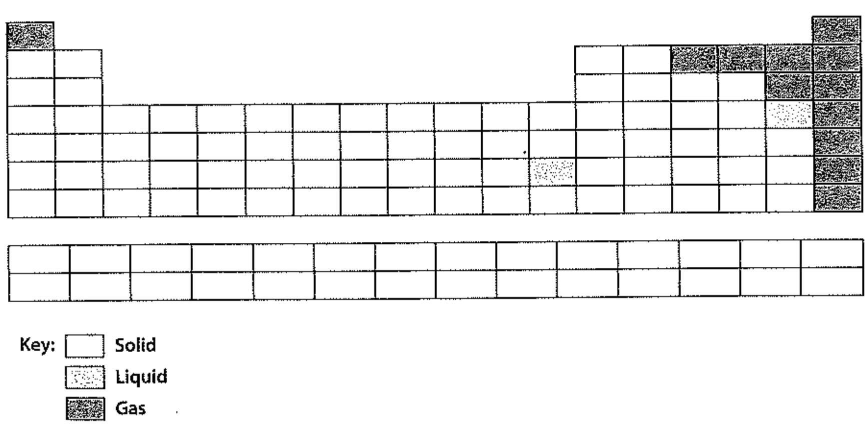


Figure 4.1

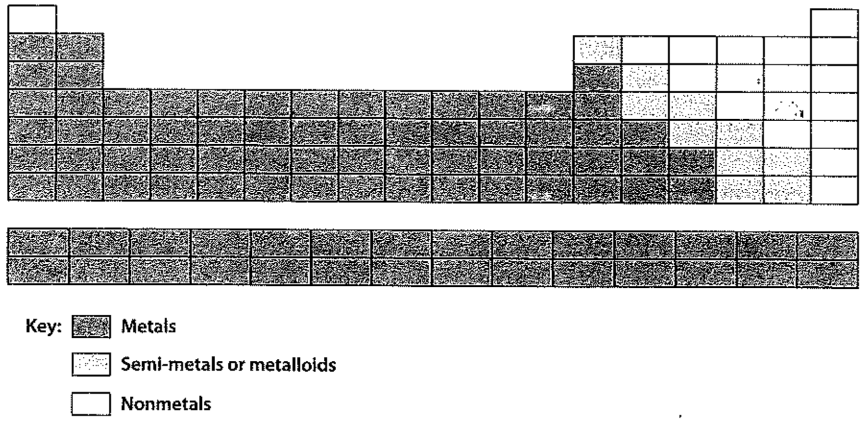


Figure 4.2

EXERCISE
4·1

Complete the following chart using the periodic table at the end of the book.

ELEMENT	METAL, METALLOID, OR NONMETAL	SOLID, LIQUID, OR GAS (AT ROOM TEMP.)	COLUMN #	PERIOD #
1. Oxygen	______	______	______	______
2. Silicon	______	______	______	______
3. Mercury	______	______	______	______
4. ______	______	______	2	6
5. ______	______	______	18	2
6. ______	______	______	17	4
7. Francium	______	______	______	______
8. Radon	______	______	______	______
9. Germanium	______	______	______	______

The arrangements in the table are like a code. They can tell us a lot of things about a substance. Why do elements exist in different states? This can be explained by looking at their structure and the forces holding them together.

The current periodic table has elements arranged by increasing atomic number. Remember, the atomic number represents the number of protons in the element. The atomic number increases by one moving from left to right across the periodic table. Hydrogen is element 1 and has one proton; helium is element 2 and has two protons; lithium is element 3 and has three protons; and so on.

The electrons determine how reactive an element is. The founders of the periodic table arranged the elements based on common characteristics. As more information about the location, function, and movement of electrons was discovered, it was found that how these properties change throughout the table is linked to the electronic structure of the atom.

Each electronic structure has a set of properties that relate to one another. The periods all fill electrons into the same energy levels in four different types of shaped areas, called *orbitals*. The columns fill the same number of electrons into the same type of orbital but at a different average distance from the nucleus, and so at different energy levels. The first column, group 1, has one electron in its outermost orbital, which is an *s* orbital. Column 2—i.e., group 2—has two electrons in the outermost *s* orbital. The arrangement of orbitals and order of filling is important to understand.

Orbital arrangement

The orbitals are designated by the symbols *s*, *p*, *d*, and *f*: *s* has one orbital at each energy level (Figure 4.3), *p* has three orbitals (Figure 4.4), *d* has five orbitals (Figure 4.5), and *f* has seven orbitals (Figure 4.6).

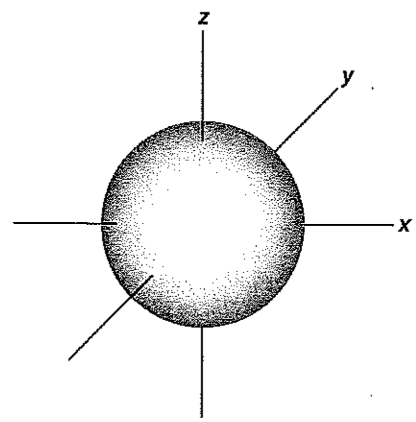


Figure 4.3

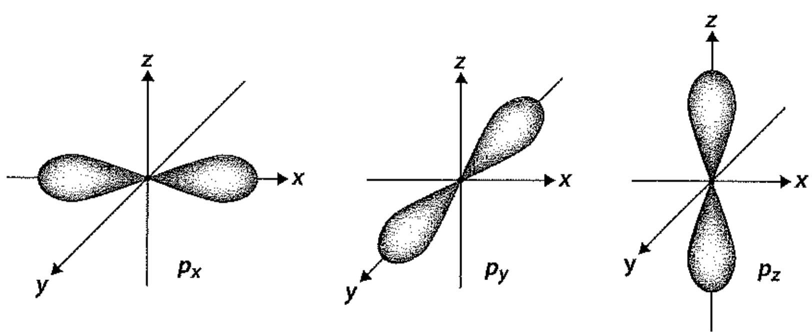


Figure 4.4

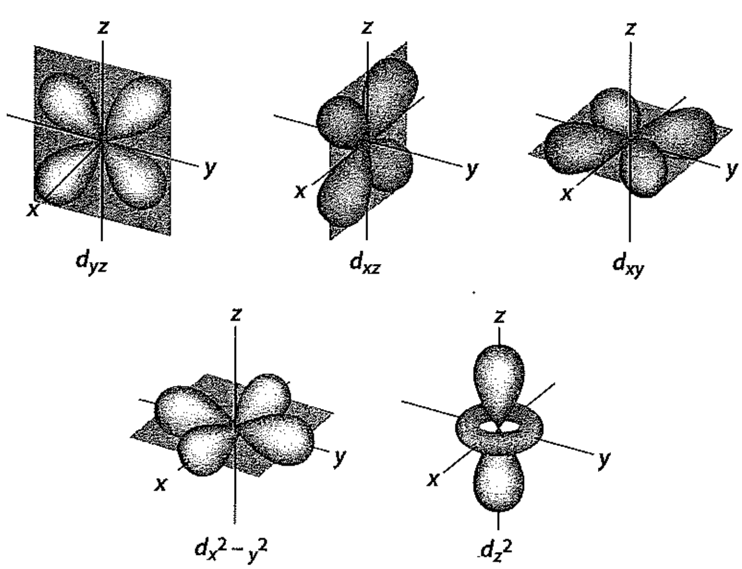


Figure 4.5

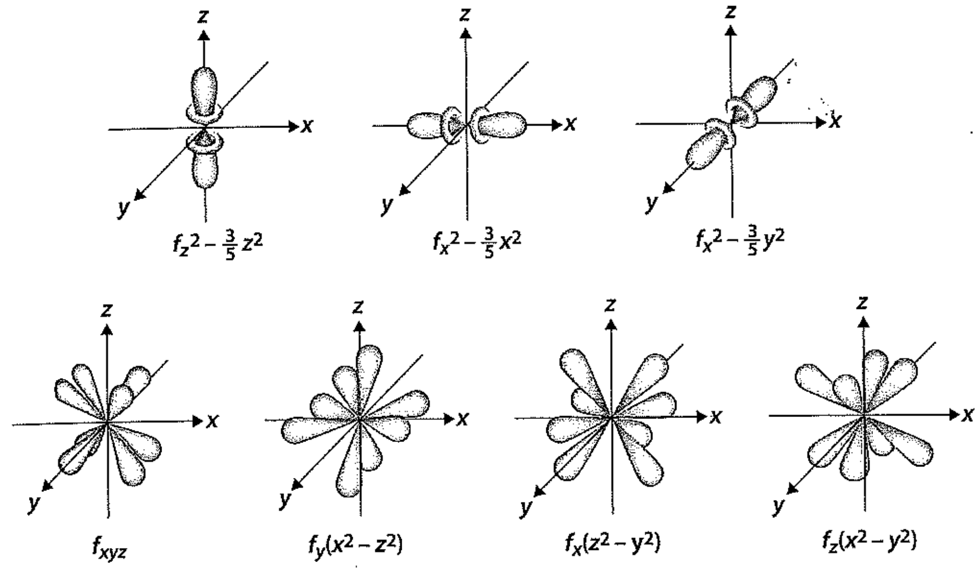


Figure 4.6

Each individual orbital can hold two electrons. Knowing each orbital can contain a specific number of electrons, the one orbital of *s* can contain a maximum of 2 electrons, the three orbitals of *p* can hold 6 electrons, the five orbitals of *d* can hold 10 electrons, and the seven orbitals of *f* can hold 14 electrons. Which orbital is being filled to form an element can be recognized from the periodic table (see Figure 4.7).

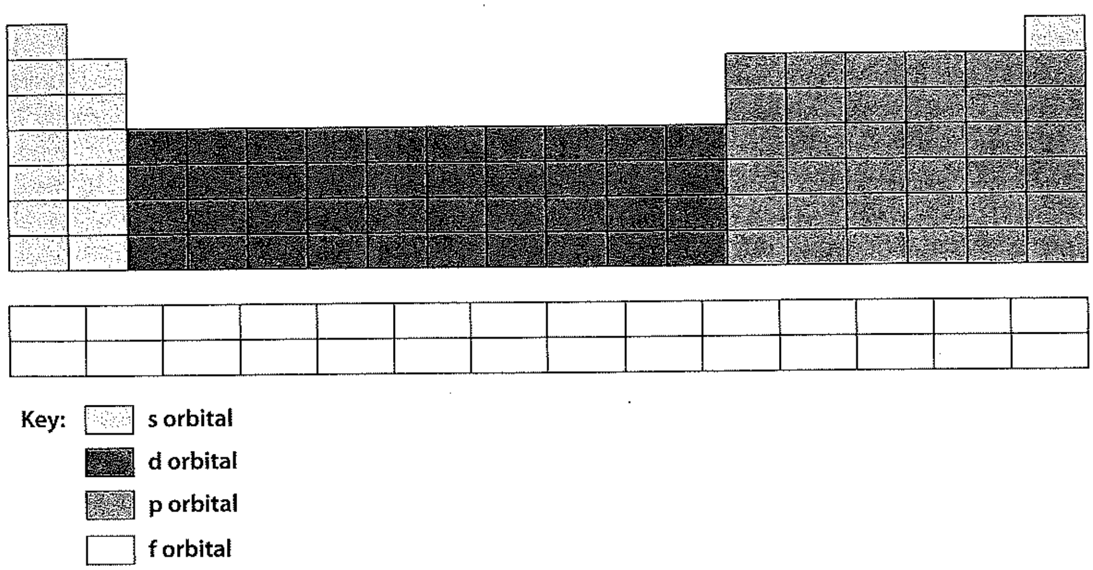


Figure 4.7

Each orbital has a different energy, and they fill in a particular order. If one knows the number of electrons an atom has, the electron configuration can be written. The electron configuration is a notation indicating all the electrons in the atom and where they are located. This notation gives the energy level of the orbital the electron is in, the type of orbital, and the number of electrons. The energy level is indicated by an integer from one to seven, the orbitals each have a letter—*s*, *p*, *d*, or *f*—and the number of electrons varies by the type of orbital, with an upper limit given by this list:

s orbital
- 1 orbital
- 2 electrons (maximum)

p orbital
- 3 orbitals
- 6 electrons (maximum)

d orbital
- 5 orbitals
- 10 electrons (maximum)

f orbital
- 7 orbitals
- 14 electrons (maximum)

The order of filling follows the order shown in Figure 4.8: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, and 7p. You can see that these designations follow the order of increasing atomic number on the periodic table. Note that these designations are *not* included on most standard periodic tables.

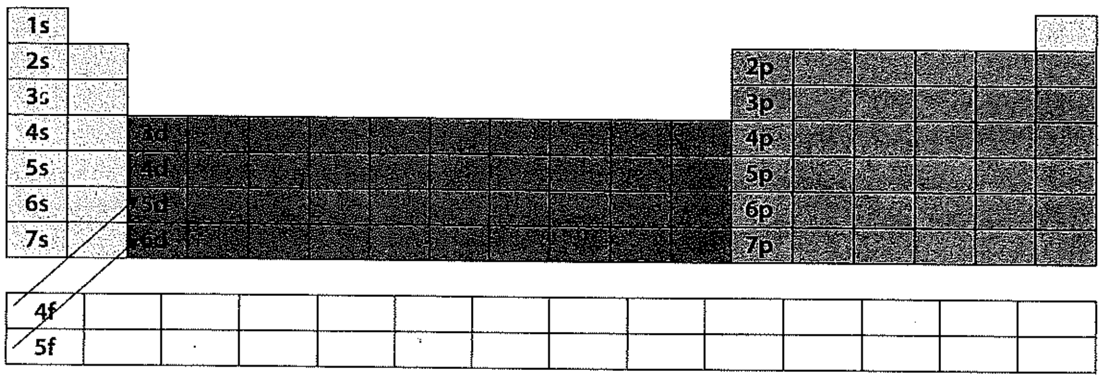


Figure 4.8

Electron configurations

The order that electrons fill orbitals in follows the *Aufbau principle*: Electrons fill orbitals from lowest energy (most stable) to highest energy. For example, magnesium, symbol Mg, has 12 protons and as a neutral atom has 12 electrons. Those electrons are in 1s, 2s, 2p, and 3s. When this is written as $1s^22s^22p^63s^2$, it is called an electron configuration. The superscripts are the number of electrons in the particular type of orbital. The sum of the superscripts will equal the number of electrons in the atom or ion. In this case, 2 + 2 + 6 + 2 = 12!

Atom	Number of electrons	Electron configuration
Calcium	20	$1s^22s^22p^63s^23p^64s^2$
Iron	26	$1s^22s^22p^63s^23p^64s^23d^6$
Tin	50	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^2$

EXERCISE **4·2**

Write electron configurations for the following elements.

1. Carbon ____________________
2. Zinc ____________________
3. Bromine ____________________
4. Cesium ____________________

5. Bismuth ______

6. Potassium ______

7. Manganese ______

8. Iodine ______

9. Cadmium ______

10. Lead ______

11. Francium ______

12. Platinum ______

13. Titanium ______

14. Oxygen ______

15. Beryllium ______

When orbitals are grouped by their first number—called the *principal quantum number*—the order changes to 1s, 2s, 2p, 3s, 3p, 3d, 4s, 4p, 4d, 4f, 5s, 5p, 5d, 5f, etc. By placing the orbitals in this order, it is easier identify what are called the *valence*, or outermost, electrons. Knowing the number of outer electrons will be vital when drawing structures, as only valence electrons are included. It is also important because when an atom forms a positively charged ion by losing electrons, the electrons lost are from the outermost orbital, *not necessarily the last one filled.* In contrast, when electrons are added to form negatively charged ions, they fill the last orbital being filled.

When Mg $1s^22s^22p^63s^2$ forms the ion Mg^{2+}, the two electrons lost are the 3s electrons, leaving $1s^22s^22p^6$. Sulfur (S), element 16, has 16 electrons, resulting in the configuration of $1s^22s^22p^63s^23p^4$. To form sulfur's 2negative ion, S^{2-}, two electrons must be added to the configuration, forming $1s^22s^22p^63s^23p^6$.

EXERCISE

4·3

Write the electron configurations for the following atoms and their corresponding ions.

1. Lithium and Li^+ ______

______

2. Aluminum and Al^{3+} ______

______

3. Lead and Pb^{4+} ______

______

4. Nitrogen and N^{3-} ______

______

5. Chlorine and Cl^- ______

______

6. Oxygen and O^{2-}

7. Sodium and Na^{+}

8. Arsenic and As^{5+}

9. Fluorine and F^{-}

10. Beryllium and Be^{2+}

11. Arsenic and As^{3-}

12. Nickel and Ni^{2+}

13. Cadmium and Cd^{2+}

14. Strontium and Sr^{2+}

15. Iron and Fe^{3+}

Orbital diagrams

Only two electrons can be in an orbital, and they must have opposite spins. This is called the *Pauli exclusion principle*: no two electrons can be in the same place at the same time with the same spin. While the orbitals fill in order of increasing energy in Aufbau's principle, the order of filling the orbitals follows Hund's rule. Hund's rule states when filling the same type of orbital in the same energy level (*p*, *d*, or *f*) that one electron will go into each orbital before the pairing of electrons will occur. For instance, in the three orbitals of *p*, which we can identify as p_x, p_y, and p_z, the first electron will go into p_x, the second into p_y, the third into p_z, and the fourth will go back to p_x and fill the orbital, the fifth will go into p_y, and the last electron into p_z. The defining of orbitals as px, py, and pz is a convention that allows the model to be built regarding the order of orbital filling as presented here. The first three of the six *p* electrons will all have the same type of spin, and the last three will all have the opposite spin of the first three. This is because this arrangement forms the most stable energy; this rule must be understood as configurations are studied. Look at Figure 4.9, the orbital diagram for nitrogen. The seven electrons in nitrogen are arranged in the configuration $1s^22s^22p^3$. The three electrons in the *p* orbitals are split among the three orbitals, following Hund's rule. The three *p* orbitals and their electrons can be written in this case as $2p_x^1 2p_y^1 2p_z^1$. This is a more stable configuration than having $2p_x^2 2p_y^1$ with no $2p_z$ electron.

↑↓	↑↓	↑	↑	↑	
$1s^2$	$2s^2$	$2p_x^1$	$2p_y^1$	$2p_z^1$	for a total of $2p^3$

Figure 4.9

In Figure 4.10, the eight electrons of oxygen are in the configuration $1s^2 2s^2 2p^4$; the fourth electron in the *p* cloud goes back to the first orbital and goes opposite in spin. So the *p* cloud's configuration is written out fully as $2p_x^2 2p_y^1 2p_z^1$.

↑↓	↑↓	↑↓	↑	↑	
$1s^2$	$2s^2$	$2p_x^2$	$2p_y^1$	$2p_z^1$	for a total of $2p^4$

Figure 4.10

EXERCISE

4·4

Draw orbital diagrams for the following elements and ions.

1. Sodium
2. Phosphorus
3. Cobalt
4. Selenium
5. Cadmium
6. Na^+
7. Cd^{2+}
8. O^{2-}
9. P^{3-}
10. Neon
11. Helium
12. Nickel
13. Calcium
14. Ca^{2+}
15. Fe^{3+}

While chemists often describe electrons as "filling orbitals," in fact the electrons *are* the orbitals. An orbital is the probability of an electron being in a certain region around the nucleus. There is nothing there to "fill" until the electron is there! This is just a convenient way for people to describe atoms. So a better description for the subject of the Aufbau principle is "building atoms" rather than "filling orbitals."

Periodic trends

Periodic trends are patterns that can be recognized based on observations of the elements. Two concepts used to describe the reasoning behind different periodic trends are shielding and effective nuclear charge.

Shielding

Orbital diagrams and electron configurations help illustrate the location of electrons to further explain the concepts of shielding and effective nuclear charge. These two concepts will in turn help us understand trends in the periodic table.

Going down a column, the atomic number increases and the outer electron orbitals get farther from the nucleus, as a new principal quantum number must be used.

As an atom adds orbitals, the outer electrons have more inner orbitals between themselves and the pull of all the positive charge from the nucleus. Think of the layers of electrons in between as interference, reducing the effect of the electrical charge of the nucleus. The more layers there are, the more shielding there is. The trend describing how this affects the outer electrons is termed "effective nuclear charge." (See the next section.)

If the atom has a larger radius, this means the outermost electrons are easier to remove, since they are not held as tightly as electrons close to the nucleus (which experience more pull from the positive charge of the nucleus). Distance matters—the coulomb attractive force between opposite charges varies with the distance between the charges according to the formula $1/d^2$!

The greater the period number (principal quantum number), the larger the radius. Magnesium has 12 electrons in the configuration $1s^2 2s^2 2p^6 3s^2$ and calcium has 20 electrons in the configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$. Both configurations end in s^2 electrons, but calcium has an additional energy level and therefore a larger radius. This larger radius exists both because the electrons can only be accommodated at a greater radial distance *and* because there are more inner electrons that provide more shielding. (More periods mean more shielding, thus making it easier to remove outer electrons.)

> The periodic trend is that atoms have larger radii moving down a column. The reason for this trend is that higher energy orbitals with larger radial distances must be used; this increases the shielding moving down a column and reduces the effective nuclear charge (and hence the attraction of the outermost electrons to the nucleus). The result is that the outer electrons are easier to remove moving down a column.

Effective nuclear charge

Moving across a period is different, since the number of outermost energy levels remains the same. Instead of being reduced due to shielding, the effective nuclear charge is increased as the atomic number increases across a period, since there are more protons being added to the nucleus. This greater number of protons results in a larger positive attraction exerted by the nucleus.

This larger attraction results in the atom having a *smaller* radius as you move across a period, due to the *increased* effective nuclear charge (e.g., potassium has 19 electrons in the configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$, and calcium has 20 electrons in the configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$). Both have their outermost electrons in the fourth energy level, but calcium has an additional proton in the nucleus, resulting in a greater effective nuclear charge. This means the calcium atom has a smaller radius than the potassium atom.

The trend is that atoms have smaller radii moving across a period. This is the due to increased effective nuclear charge. The result is that electrons are harder to remove as you move across a period.

EXERCISE 4·5

Choose which atom has the smallest radius from the given group and explain the reasoning behind your choice.

1. Nitrogen and oxygen

____

2. Potassium and rubidium

____

3. Sulfur, chlorine, and bromine

____

4. Strontium, cesium, and barium

____

5. Rubidium and krypton

____

6. Calcium and gallium

____

7. Fluorine and iodine

____

8. Boron and fluorine

____

9. Carbon, nitrogen, and silicon

____

10. Neon, argon, and krypton

____

Atomic radius versus ionic radius

When ions are formed, what happens to the atom's radius and effective nuclear charge depends on whether the ion is a cation or an anion. Remember that when magnesium $1s^22s^22p^63s^2$ forms the cation Mg^{2+}, the two electrons lost are the 3s electrons, leaving $1s^22s^22p^6$. An entire energy level has been lost, but the number of protons has not changed! The remaining electrons are held more tightly, due to an increased effective nuclear charge based on the reduced number of electrons. In the magnesium atom, 12 protons were pulling on 12 electrons, compared to the 12 protons that are pulling on the 10 electrons in Mg^{2+}.

> The trend is that cations have smaller radii than their corresponding atoms. The reason for this trend is that cations have lost electrons, changing the ratio from an equal number of protons and electrons to one where there are more protons than electrons. This results in an increased effective nuclear charge, making those electrons more tightly bound and harder to remove.

For anions, the opposite is true. For sulfur, symbol S, which has 16 electrons in the configuration $1s^22s^22p^63s^23p^4$, the atom gains two electrons to form S^{2-}, resulting in a configuration of $1s^22s^22p^63s^23p^6$. The number of protons in the nucleus has not changed but two electrons have been gained. In the sulfur atom, 16 protons were pulling on 16 electrons, but in the S^{2-} anion, 16 protons are pulling on 18 electrons. This reduction in effective nuclear charge means the ion has a larger radius than its atom.

> The periodic trend here is that anions have larger radii than their corresponding atoms. The reason for this trend is that anions gain electrons, changing the ratio from an equal number of protons and electrons to one where there are more electrons than protons. This results in a decreased effective nuclear charge, making the outer electrons easier to remove.

EXERCISE

4·6

Write the electron configurations for the following atoms and their corresponding ions. Based on the proton-to-electron ratio, determine which member of the pair has a larger radius, and explain your reasoning.

PARTICLES	CONFIGURATIONS	PROTON-TO-ELECTRON RATIOS
1. F	________	________
F^-	________	________
Which is larger? Explain.	________	
2. Al	________	________
Al^{3+}	________	________
Which is larger? Explain.	________	
3. Fe	________	________
Fe^{2+}	________	________
Which is larger? Explain.	________	
4. As	________	________
As^{3-}	________	________
Which is larger? Explain.	________	

5. N

N^{3-}

Which is larger? Explain.

6. O

O^{2-}

Which is larger? Explain.

7. Na

Na^{+}

Which is larger? Explain.

8. Ga

Ga^{3+}

Which is larger? Explain.

9. Cr

Cr^{6+}

Which is larger? Explain.

10. Ca

Ca^{2+}

Which is larger? Explain.

First ionization energy

The amount of energy required to remove an electron from a gaseous atom in its ground state is defined as its first ionization energy. How much energy is needed depends on how much shielding is present and what effective nuclear charge is exhibited. Remember that shielding increases down a column, resulting in electrons not being held as tightly; these electrons are easier to remove, and thus ionization energy decreases.

> The periodic trend is that first ionization energy decreases down a column. The reason for this trend is that more shielding and greater distance between the nucleus and outer electrons result in the electrons not being held as tightly; therefore it takes less energy to remove the outer electrons.

Going across a period from left to right, ionization energy tends to increase, since effective nuclear charge increases. Why "tends to" instead of just "does"? There are two exceptions to this trend, each due to a different reason.

- Exception 1: Using period 2, moving from beryllium—in Group 2, with a full 2s orbital—to boron, which has a single electron in the higher-energy 2p orbital, it actually takes less energy to remove the single electron in 2p.

- Exception 2: Ionization energy also decreases from Group 15 to Group 16. Look again at Figures 4.9 and 4.10 on page 47. Nitrogen's electrons in the *p* orbitals all have parallel spin, while oxygen has two electrons with opposite spins paired in first *p* orbital and it has two lone electrons with parallel spin in the remaining two *p* orbitals. It takes less energy to remove one of the paired-spin electrons from oxygen.

> For representative elements, ionization energy tends to increase across a period, with slight drops as orbital types change from *s* to *p* and as the three orbitals of *p* move from having one unpaired electron in each orbital to having one orbital with paired electrons. The reason for this trend is that increasing effective nuclear charge across a period results in more energy being needed to remove outer electrons.

To review the trends previously discussed, Figure 4.11 shows the patterns in first ionization energy and atomic radius. Note the directions of the arrows!

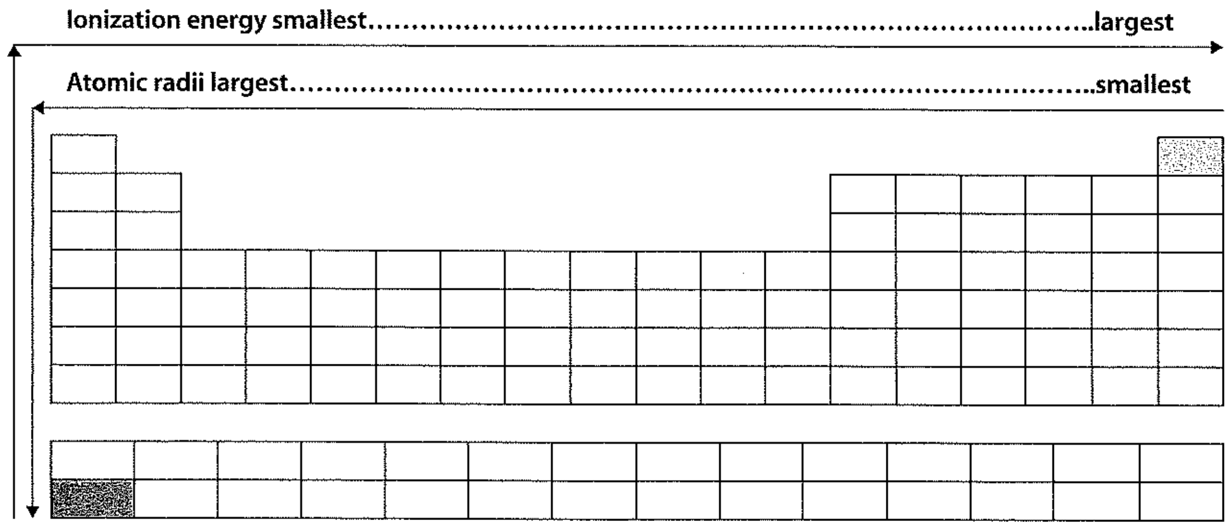


Figure 4.11

EXERCISE 4·7

Choose which atom has the smallest first ionization energy from the given group and explain the reasoning behind your choice. It might be helpful to draw orbital diagrams for the valence electrons of each atom.

1. Carbon and nitrogen

____

2. Potassium and rubidium

____

3. Sulfur, selenium, and bromine

____

4. Rubidium, cesium, and francium

5. Magnesium and aluminum

6. Oxygen and francium

7. Fluorine and nitrogen

8. Helium and neon

9. Boron and gallium

10. Silver and copper

Successive ionization energy

After the energy is added to remove the first electron, adding additional energy can remove additional electrons in sequence. The additional energy to remove each sequential electron is called *successive ionization energy.* The energy required to remove additional electrons is always greater than the first ionization energy, due to the proton-to-electron ratios increasingly favoring the protons as more electrons are removed. Effective nuclear charge increases each time an electron is removed; when an entire energy level is removed, the energy needed increases dramatically as effective nuclear charge on the outer electrons increases dramatically with the smaller radius and the loss of inner shielding. Ionization energies listed in Figure 4.12 are measured in kJ per mole. The concept of a mole will be explained in Chapter 7. The important thing to understand here is that as more electrons are removed, the energy required increases.

	1st electron	2nd electron	3rd electron	4th electron	5th electron
H	1,312				
He	2,373	5,251			
Li	520	7,300	11,815		
Be	899	1,757	14,850	21,005	
B	801	2,430	3,660	25,000	32,820

Figure 4.12

The trend is that successive ionization energies always increase. The reason for this trend is the greater effective nuclear charge that results as electrons are removed, leading to more energy being needed to remove additional electrons.

EXERCISE 4·8

Complete the following chart using the periodic table at the end of the book.

ELEMENT	METAL, METALLOID, OR NONMETAL	SOLID, LIQUID, OR GAS (AT ROOM TEMP.)	COLUMN #	PERIOD #
1. Bromine	______	______	______	______
2. Antimony	______	______	______	______
3. Titanium	______	______	______	______
4. ______	______	______	14	5
5. ______	______	______	17	2
6. ______	______	______	1	1

EXERCISE 4·9

Write electron configurations for the following elements.

1. Krypton ______
2. Technetium ______
3. Selenium ______
4. Barium ______
5. Chlorine ______

EXERCISE 4·10

Write the electron configurations for the following atoms and their corresponding ions.

1. Cadmium and Cd^{2+} ______
______
2. Selenium and Se^{2-} ______
______

EXERCISE
4·11

Draw orbital diagrams for the following elements.

1. Beryllium
2. Fluorine
3. Potassium

EXERCISE
4·12

Complete the following chart by filling in each blank with the best choice from the pair of particles. If the values are equal, write "same" in the blank.

	MORE SHIELDING	GREATER EFFECTIVE NUCLEAR CHARGE	LARGER RADIUS	SMALLER IONIZATION ENERGY
1. Li or K	______	______	______	______
2. Na or Cl	______	______	______	______
3. K or K^+	______	______	______	______
4. Br or Br^-	______	______	______	______

Naming compounds and writing formulas

·5·

The periodic table will be a big help when it comes to naming compounds and writing formulas. When elements from different parts of the table combine, different types of bonds are made, and the bonding determines the naming system to be used. The possible combinations are metals with metals, metals with nonmetals, nonmetals with nonmetals, and two unique groups: hydrogen with a nonmetal, and carbons primarily with hydrogen. This is like having to learn five foreign languages!

Metal-metal combinations

Metal-metal combinations form metallic bonds. They can conduct electricity in all phases, due to electrons' being available to move in a "sea." Metal-metal solutions are called *alloys*. Alloys go by different names, depending on the identity of the metals and the percentage of each metal in the solution. Other metals or elements can be added to change the properties of an alloy: adding carbon to iron makes steel stronger. Bronze is an alloy of copper and tin, brass is composed of copper and zinc, and steel is an alloy of iron and several other elements, including nickel. Since these are not compounds with exact ratios of atoms, we do not write formulas for them. Their names usually relate to historical developments rather than chemical composition.

Metal-nonmetal combinations

Metal-nonmetal combinations form ionic bonds. The metal combines in the form of a *cation* with the nonmetal in the form of an *anion* (or anion groups) to make a formula unit. The metal cation is always written first in the formula and in the compound name, while the anion is always written second in the formula and in the compound name. Sodium chloride, $NaCl$, is an example of a formula unit that contains an ionic bond. One important cation does not contain a metal atom—it is the ammonium ion, NH_4^+, a group of nonmetal atoms combined with a positive charge. As a cation, ammonium is still written first in a formula.

Writing the formula of a substance that contains ionic bonds from the name

Writing the formula for a chemical compound is an important skill in chemistry. There are certain types of information that are needed to accomplish this task.

- Remember, the cation symbol is always written first in the formula, and the anion symbol is always written second.
- To determine the formula, the charge of each ion must be known (see Table 5.1).

Table 5.1 Selected Ions List

Cation name	Cation symbol	Anion name	Anion symbol
aluminum	Al^{3+}	acetate	$C_2H_3O_2^-$
ammonium	NH_4^+	bromide	Br^-
barium	Ba^{2+}	carbonate	CO_3^{2-}
cadmium	Cd^{2+}	chlorate	ClO_3^-
calcium	Ca^{2+}	chloride	Cl^-
cobalt(II)	Co^{2+}	chromate	CrO_4^{2-}
copper(II)	Cu^{2+}	cyanide	CN^-
copper(I)	Cu^+	dichromate	$Cr_2O_7^{2-}$
hydrogen	H^+	fluoride	F^-
iron(II)	Fe^{2+}	hydride	H^-
iron(III)	Fe^{3+}	bicarbonate	HCO_3^-
lead(II)	Pb^{2+}	hydroxide	OH^-
lead(IV)	Pb^{4+}	iodide	I^-
lithium	Li^+	nitrate	NO_3^-
magnesium	Mg^{2+}	nitride	N^{3-}
nickel(II)	Ni^{2+}	nitrite	NO_2^-
potassium	K^+	oxalate	$C_2O_4^{2-}$
rubidium	Rb^+	oxide	O^{2-}
silver	Ag^+	permanganate	MnO_4^-
sodium	Na^+	peroxide	O_2^{2-}
strontium	Sr^{2+}	phosphate	PO_4^{3-}
tin(II)	Sn^{2+}	phosphite	PO_3^{3-}
tin(IV)	Sn^{4+}	phosphide	P^{3-}
zinc	Zn^{2+}	sulfate	SO_4^{2-}
		sulfide	S^{2-}
		sulfite	SO_3^{2-}

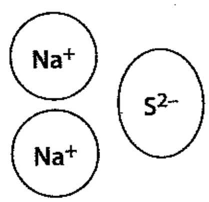
- The ions combine in a ratio so that the sum of the charges is equal to zero. A method called "crisscross and reduce" will help determine how many of each ion is needed. In this method, the charge of the cation becomes the number of anions present, and the charge of the anion becomes the number of cations present. Na^+ combines with S^{2-} to give Na_2S. Drawing pictures also helps us see the number of each ion required (see Figure 5.1). If the numbers in the "crisscross" end up being numbers that can be factored, such as Pb_2S_2, then the numbers need to be reduced to the lowest possible whole numbers: Pb_2S_2 then becomes PbS.

✓ **Sodium sulfide: Na^+ combined with $S^{2-} = Na_2S$. It takes <u>two</u> 1^+ charges to offset <u>one</u> 2^- charge, $(\underline{2} \times 1) + (\underline{1} \times -2) = 0$.**

Crisscross $Na^+ \ S^{2-}$

Na_2S

Drawing



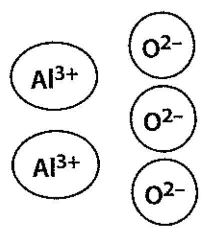
Total 2^+ and $2^- = 0$

✓ **Aluminum oxide: Al^{3+} combined with $O^{2-} = Al_2O_3$. It takes <u>two</u> 3^+ charges to offset <u>three</u> 2^- charges, $(\underline{2} \times 3) + (\underline{3} \times -2) = 0$.**

Crisscross $Al^{3+} \ O^{2-}$

Al_2O_3

Drawing



Total 6^+ and $6^- = 0$

Figure 5.1

- Some ions with more than one possible charge will have a Roman numeral in the name to indicate the charge used. For instance, iron(III) is Fe^{3+}.

Some ions are polyatomic, which means that several atoms form a group with a consistent charge, such as the ammonium ion mentioned earlier and the sulfate ion, SO_4^{2-}. If more than one of a polyatomic ion is needed, put the ion in parentheses and the number needed as a subscript after the closing parenthesis (see Figure 5.2).

✓ **Aluminum phosphate: Al^{3+} combined with $PO_4^{3-} = AlPO_4$. It takes <u>one</u> 3^+ charge to offset <u>one</u> 3^- charge, $(\underline{1} \times 3) + (\underline{1} \times -3) = 0$.**

Al^{3+} PO_4^{3-}

$Al_3(PO_4)_3$ and reduce to $AlPO_4$

Al^{3+} PO_4^{3-}

Total 3^+ and $3^- = 0$

✓ **Aluminum sulfate: Al^{3+} combined with $SO_4^{2-} = Al_2(SO_4)_3$. It takes <u>two</u> 3^+ charges to offset <u>three</u> 2^- charges, $(\underline{2} \times 3) + (\underline{3} \times -2) = 0$.**

Al^{3+} SO_4^{2-}

$Al_2(SO_4)_3$

Al^{3+} SO_4^{2-}

SO_4^{2-}

Al^{3+} SO_4^{2-}

Total 6^+ and $6^- = 0$

✓ **Ammonium sulfide: NH_4^+ combined with $S^{2-} = (NH_4)_2S$. It take <u>two</u> 1^+ charges to offset <u>one</u> 2^- charge, $(\underline{2} \times 1) + (\underline{1} \times -2) = 0$.**

NH_4^+ S^{2-}

$(NH_4)_2S$

NH_4^+

S^{2-}

NH_4^+

Total 2^+ and $2^- = 0$

Figure 5.2

EXERCISE

5·1

Using the periodic table at the end of this book and Table 5.1 on page 58, write the formula for each of the following compounds.

1. Potassium oxide ______________________
2. Strontium chloride ______________________
3. Aluminum nitride ______________________
4. Lithium phosphide ______________________
5. Aluminum chlorate ______________________
6. Beryllium phosphite ______________________

7. Calcium sulfate ______________________

8. Sodium nitrate ______________________

9. Aluminum oxalate ______________________

10. Ammonium sulfate ______________________

11. Sodium nitrite ______________________

12. Potassium nitrate ______________________

13. Aluminum sulfide ______________________

14. Iron(II) oxide ______________________

15. Iron(III) oxide ______________________

16. Lead(II) acetate ______________________

17. Gallium chlorate ______________________

18. Ammonium sulfide ______________________

19. Silver oxide ______________________

20. Manganese(IV) oxide ______________________

Writing the name of an ionic compound from the formula

Writing the name of a compound from the formula is the opposite process of writing the formula from the name. To write the name, consider the following points.

- The metal cation name is always written first in the name, and the anion name is written second.
- The metal cation is given the name of the metal without changing the ending.
- If the metal cation has more than one oxidation state, a Roman numeral is put in parentheses after the name to indicate the state used. To know if this is the case, look on a table with oxidation states given, such as Table 5.1 on page 58. Copper(II) nitrate, for example, indicates that the charge on the copper is 2^+. Most of the metals needing a Roman numeral are transition metals.
- The anion has a name ending in *-ide*, *-ate*, or *-ite*. Use Table 5.1 to find which one to use. Usually, single ions end in *-ide* (two exceptions are cyanide, CN^-, and hydroxide, OH^-). Ions containing many atoms, called polyatomic ions, end in *-ate* and *-ite*, such as sulfate, SO_4^{2-}, and sulfite, SO_3^{2-}. If you look closely, you will observe that the number of oxygen atoms in an *-ate* ion is one more than in the corresponding *-ite* ion: sulfate has four oxygen atoms, while sulfite has three. A *per-* in front of an *-ate* ion indicates an additional oxygen atom: chlorate is ClO_3^-, while perchlorate is ClO_4^-. A *hypo-* in front of an *-ite* ion indicates one less oxygen atom: chlorite is ClO_2^-, while hypochlorite, found in bleach, is ClO^-.

Ion name	Number of oxygen atoms
Per___ate	+1
___ate	depends on the ion
___ite	−1
Hypo___ite	−2

Number of atoms in common *-ate* ions			
3		**4**	
Nitrate	NO_3^-	Phosphate	PO_4^{3-}
Carbonate	CO_3^{2-}	Sulfate	SO_4^{2-}
Chlorate	ClO_3^-		
Bromate	BrO_3^-		
Iodate	IO_3^-		

- Even though they are from an old system no longer used by the IUPAC, old names will often be seen on household labels. For instance, Fe^{3+} was called *ferric*. Ferric chloride is now iron(III) chloride, $FeCl_3$. A few examples are listed in the following table:

Ion	Former name	New name
Fe^{3+}	ferric	iron(III)
Fe^{2+}	ferrous	iron(II)
Cu^{2+}	cupric	copper(II)
Cu^+	cuprous	copper(I)
Sn^{4+}	stannic	tin(IV)
Sn^{2+}	stannous	tin(II)

- The first letter of the cation should be capitalized if it starts a sentence; otherwise, the entire name should be lowercase.

Hint: Reversing the crisscross will help identify the charge of the ion.

$Ca(OH)_2$: Ca^{2+} is calcium and OH^- is hydroxide, so the name is calcium hydroxide.

$KMnO_4$: K^+ is potassium and MnO_4^- is permanganate, so the name is potassium permanganate.

$Ni_3(PO_4)_2$: Ni^{2+} is nickel(II) and PO_4^{3-} is phosphate, so the name is nickel(II) phosphate. Always check if a metal ion needs a Roman numeral!

EXERCISE

5·2

Write the names of the following compounds from their formulas.

1. Al_2S_3 ________________
2. K_3PO_4 ________________
3. $CuSO_4$ ________________
4. $Zn(NO_3)_2$ ________________
5. $FeCO_3$ ________________
6. $Sn(BrO)_2$ ________________
7. $PbCl_4$ ________________
8. MnO_2 ________________
9. $Ca(MnO_4)_2$ ________________

10. Cs_2O ______________________

11. AgBr ______________________

12. $CrCl_3$ ______________________

13. K_2CO_3 ______________________

14. FeS ______________________

15. $Al(OH)_3$ ______________________

16. CdS ______________________

17. $FeCl_3$ ______________________

18. $Ca_3(PO_4)_2$ ______________________

19. $Ni(NO_3)_2$ ______________________

20. PbO ______________________

Nonmetal-nonmetal combinations

Nonmetal-nonmetal combinations form covalent bonds. Units of these compounds are called molecules and have their own naming system. This system uses the following prefixes:

Number of atoms	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

Writing the formula of a covalent compound from the name

Writing the formula of a covalent compound requires remembering the prefixes that are associated with the number of atoms found in the compound. The preceding table can be used as a reference.

- The first word in the name has the name of the first element. If only one atom of the first element is present in the formula, no prefix will be added to the element name.
- The second element in the name will *always* have a prefix to indicate the number of atoms of that element that are in the molecule, and the name will *always* end in *-ide*.

Carbon monoxide: The lack of a prefix on carbon indicates that there is one carbon atom, and the prefix *mono-* indicates that there is one oxygen atom, for a formula of CO.

Dinitrogen trioxide: The prefix *di-* indicates that there are two nitrogen atoms, and the prefix *tri-* indicates that there are three oxygen atoms, for a formula of N_2O_3.

EXERCISE

5·3

Write the chemical formula for each of the following compounds.

1. Carbon dioxide ______
2. Sulfur trioxide ______
3. Dihydrogen monoxide ______
4. Tetraphosphorus decoxide ______
5. Disulfur dichloride ______
6. Silicon tetrafluoride ______
7. Ditnitrogen tetroxide ______
8. Nitrogen dioxide ______
9. Dichlorine heptoxide ______
10. Carbon disulfide ______

Writing the name of a covalent compound from the formula

When writing the name of a covalent compound from the formula be sure to follow these pointers.

- The first symbol represents the first element in the name. If there is more than one atom of that element in the formula, add the appropriate prefix in front of the element's name.
- The second symbol represents the second element (and the second word) in the name. No matter how many atoms are present, add the appropriate prefix and end the name in *-ide*. If the prefix ends in *a*, such as *tetra-*, and the element name begins with a vowel, the *a* can be dropped from the name: tetroxide instead of tetraoxide.
- The first letter of the name is capitalized if it starts a sentence; otherwise, the entire name should be lowercase.

SO_2: Only one sulfur atom is listed, so no prefix is added to *sulfur*; two oxygen atoms are listed, indicating the prefix *di-* for *oxide*. The name is sulfur dioxide.

P_4O_6: Four phosphorus atoms are given, so the prefix *tetra-* is added to *phosphorus*; six oxygen atoms are listed, for a prefix of *hexa-* added to *oxide*. The name is tetraphosphorus hexoxide. Note that the *a* is left on *tetra-* but dropped off *hexa-*.

The exception to this type of naming is when the nonmetal starting the compound is a hydrogen ion and the system is in a water-based solution. See the next section on acids.

EXERCISE
5·4

Write the name that corresponds to each of the following chemical formulas.

1. N_2O_5 ____________________
2. CBr_4 ____________________
3. P_2O_3 ____________________
4. PCl_3 ____________________
5. SiO_2 ____________________
6. BF_3 ____________________
7. NO ____________________
8. SF_6 ____________________
9. ClO_2 ____________________
10. CO ____________________

When a hydrogen cation and a nonmetal anion are in solution, this indicates that an acid is present. If the compound is a gas, use the rules for nonmetal-nonmetal combinations found earlier in this chapter. A good practice to use is to first write the name as if it were not an acid, then, if the compound is in aqueous solution (aq), to convert the name to an acid name! An aqueous solution is one where the dissolving medium is water. Acids are a unique form of aqueous solution with a hydrogen cation with the ability to change the pH of a solution.

Some acids have the ability to change the pH more than others. Acids can be classified as strong or weak. Strong acids completely ionize in solution producing hydrogen ions and the anions that in the formula. The most common strong acids include: perchloric, chloric, nitric, sulfuric, hydroiodic, hydrobromic, and hydrochloric. Any other acid is considered weak. Weak acids do not completely ionize in solution. Writing the formulas for these acids will be covered in the next section.

Writing the formula of an acid from the name

When writing the formula of an acid from the name, there are two main types of acids: hydracids and oxoacids. Hydracids contain hydrogen and an anion that does not contain oxygen. Oxoacids contain hydrogen and an anion that does contain oxygen.

- If the name begins with *hydro-*, hydrogen has combined with an anion ending in *-ide*: *hydrochloric* indicates that H^+ has combined with chloride, Cl^-. To determine how many of each atom is present, the sum of the charges must add up to zero. In this case, $1 + -1 = 0$, so the formula is HCl. Again, the crisscross method can be used to determine the formula. Note that HCl can exist as a pure gas, when its name is hydrogen chloride—but in aqueous solution it is named hydrochloric acid. If this seems confusing, it is! You will have to try to remember the distinction. The same applies to HCN (hydrogen cyanide gas, or hydrocyanic acid).

- If the name of the acid does not begin with *hydro-*, H^+ is still present in the formula. This is an indication that the anion used ends in either *-ate* or *-ite*. To know which was used, look at the ending of the acid name; if it ends in *-ic*, the anion used ends in *-ate*, and if it ends in *-ous*, the anion used ends in *-ite*. Again, to determine how many of each part is present, the sum of the charges must equal zero. For instance, the *-ic* ending in *sulfuric acid* indicates that H^+ was combined with sulfate, SO_4^{2-}. It takes two 1^+ charges to offset the one 2^- charge of sulfate, resulting in a formula of H_2SO_4.

Ending of anion name	Indication of H^+ in acid name	Ending of acid name
-ide	Hydro-	-ic
-ate		-ic
-ite		-ous

Hydrofluoric acid: The prefix *hydro-* indicates that H^+ has combined with fluoride, F^-. $(1 \times 1) + (1 \times -1) = 0$, so the formula is HF.

Phosphoric acid: The ending *-ic* indicates that H^+ has combined with phosphate, PO_4^{3-}. $(3 \times 1) + (1 \times -3) = 0$, resulting in a formula of H_3PO_4.

Chlorous acid: The ending *-ous* indicates that H^+ has combined with chlorite, ClO_2^-. $(1 \times 1) + (1 \times -1) = 0$, resulting in the formula $HClO_2$.

EXERCISE **5·5**

Write the formula of each acid listed below.

1. Bromic acid ______________________
2. Periodic acid ______________________
3. Hydroiodic acid ______________________
4. Hydrofluoric acid ______________________
5. Permanganic acid ______________________
6. Sulfurous acid ______________________
7. Carbonic acid ______________________
8. Nitric acid ______________________
9. Acetic acid ______________________
10. Phosphoric acid ______________________

Writing the name of an acid from the formula

Writing the name of an acid from the formula is the opposite of writing the formula from the name. Use the following information to help you acquire this skill.

- The acid formula begins with H, but whether or not the name begins with *hydrogen* depends on what has been combined with the H^+.

- If the name of the anion present ends in *-ide*, the acid name begins with the prefix *hydro-* in front of the anion's root name, and the ending of the anion's name changes to *-ic*. For HCN, the anion present is CN^-, which is cyanide. The *-ide* suffix changes to *-ic* and the prefix *hydro-* is placed in front of the root *cyan*: we write *hydrocyanic acid.*
- If the name of the anion present ends in *-ate*, only the ending of the anion's name changes, to *-ic*. $HClO_4$ has the anion ClO_4^-, which is perchlorate. The *-ate* tells us to write *perchloric acid.*
- If the name of the anion present ends in *-ite*, only the ending of the anion's name changes, to *-ous*. HClO has the anion ClO^-, which is hypochlorite. The *-ite* tells us to write the name as *hypochlorous acid.*

HBr: Br^- is bromide, so the name is hydrobromic acid.

HNO_3: NO_3^- is nitrate, so the name is nitric acid.

HNO_2: NO_2^- is nitrite, so the name is nitrous acid.

EXERCISE
5·6

Write the name of each of the following acids.

1. HI ______
2. H_3PO_4 ______
3. H_2CO_3 ______
4. HFO_2 ______
5. $HBrO_2$ ______
6. HCl ______
7. $H_2C_2O_4$ ______
8. HF ______
9. HNO_2 ______
10. H_2SO_3 ______

Introduction to organic compounds

Organic compounds are molecules made of carbon(s) and saturated with hydrogen. This is only an introduction to organic naming; organic chemistry is its own branch of chemistry. We will look only at *alkanes*, in which the number of hydrogen atoms compared to the number of carbon atoms always follows the formula C_nH_{2n+2}, where n is the number of carbon atoms in the formula. In the compounds that follow this general formula, the carbons link together to form continuous unbranched chains or branched chains.

Writing the formula of an unbranched alkane from the name

Writing the formula of an organic compound follows a different set of steps and prefixes.

- The number of carbons in the chain is indicated by a prefix. Some of these prefixes are different from those used earlier in the naming of covalent compounds.

Number of carbon atoms	Prefix
1	meth-
2	eth-
3	prop-
4	but-
5	pent-
6	hex-
7	hept-
8	oct-
9	non-
10	dec-

- The *-ane* ending tells us to use the formula $H = 2n + 2$ to determine the number of hydrogen atoms in the formula, where n is the number of carbon atoms.

  Methane: *meth-* indicates one carbon atom, so the number of hydrogen atoms is $2(1) + 2 = 4$. The formula is CH_4.

  Butane: *but-* means four carbon atoms, so the number of hydrogen atoms is $2(4) + 2 = 10$. The formula is C_4H_{10}.

EXERCISE

5·7

Write the formula for each of the following unbranched alkanes.

1. Propane ________________
2. Octane ________________
3. Pentane ________________
4. Decane ________________
5. Nonane ________________
6. Hexane ________________
7. Heptane ________________
8. Methane ________________
9. Ethane ________________
10. Butane ________________

Writing the name of an unbranched alkane from the formula

To write the name of an unbranched alkane from the formula, follow these steps.

- Look at the number of carbons in the formula and use the organic prefix representing that number.
- If the number of hydrogen atoms is equal to $2n + 2$, the name ends in *-ane*.

C_2H_6: Two carbon atoms indicate that the prefix used is *eth-*, and six hydrogen atoms are equal to $2n + 2$, so the name is ethane.

C_3H_8: Three carbon atoms indicate that the prefix used is *prop-*, and eight hydrogen atoms are equal to $2n + 2$, so the name is propane.

EXERCISE 5·8

For each of the following alkane formulas, write the name of the molecule.

1. C_9H_{20} ____________________
2. C_6H_{14} ____________________
3. C_7H_{16} ____________________
4. C_5H_{10} ____________________
5. CH_4 ____________________
6. C_2H_6 ____________________
7. C_3H_8 ____________________
8. C_4H_{10} ____________________
9. C_8H_{18} ____________________
10. $C_{10}H_{22}$ ____________________

Many other organic compounds exist, but naming these could be an entire book in itself! Figure 5.3 may help you determine which naming system to use.

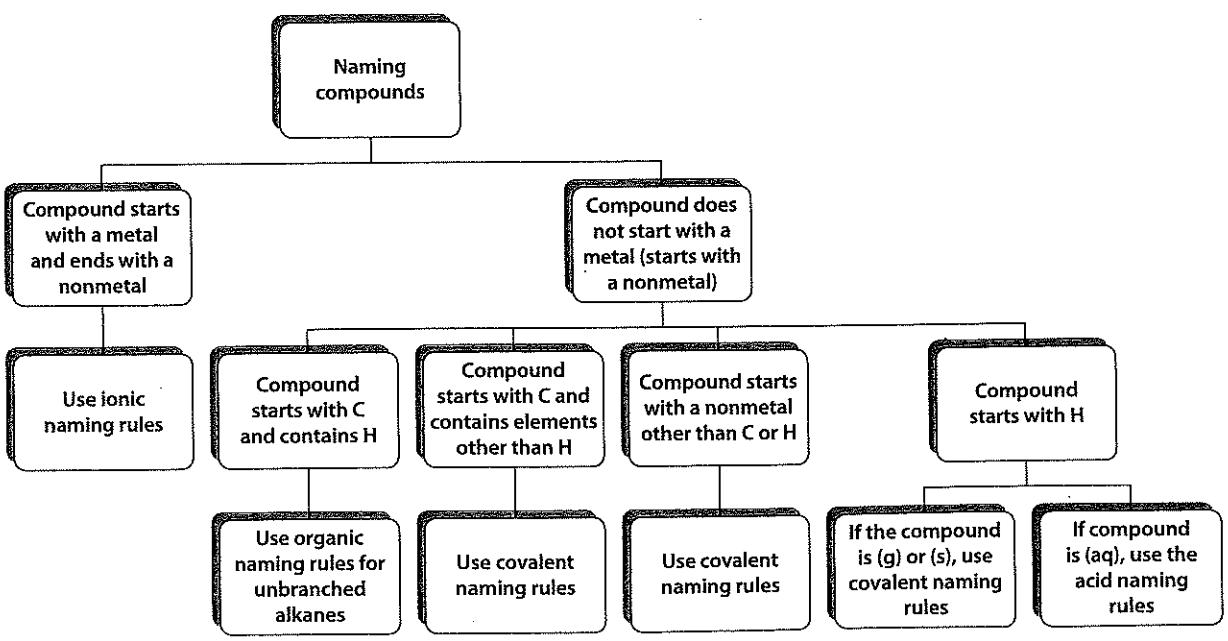


Figure 5.3

EXERCISE

5·9

Look at the formula provided and determine whether the substance listed should be named as an acid (A), a metal-nonmetal combination (M-NM), a nonmetal-nonmetal combination (NM-NM), a molecular compound (M), or none of the choices given (N).

1. K_3PO_4 ______
2. $AuBr_3$ ______
3. CS_2 ______
4. $MgCO_3$ ______
5. HCl(g) ______
6. HI(aq) ______
7. NO_2 ______
8. KF ______
9. $Al(ClO)_3$ ______
10. H_2SO_4(aq) ______
11. CuBr ______
12. $AgNO_2$ ______
13. P_2O_5 ______
14. $(NH_4)_2C_2O_4$ ______
15. NaBr ______

EXERCISE

5·10

Match the formula in the first column to the name in the second column. If there is no matching name for a formula, determine the correct name.

1. K_3PO_4	______	A. Hydrochloric acid
2. $AuBr_3$	______	B. Ammonium oxalate
3. CS_2	______	C. Aluminum chlorate
4. $MgCO_3$	______	D. Phosphorus monofluoride
5. HCl(g)	______	E. Silver nitrite
6. HI(aq)	______	F. Diphosphorus pentoxide
7. NO_2	______	G. Potassium monofluoride
8. KF	______	H. Nitrogen dioxide
9. $Al(ClO)_3$	______	I. Gold(III) bromide
10. H_2SO_4(aq)	______	J. Hydrogen iodide
11. CuBr	______	K. Magnesium carbonate
12. $AgNO_2$	______	L. Potassium phosphate
13. P_2O_5	______	M. Hydrogen chloride
14. $(NH_4)_2C_2O_4$	______	N. Copper(I) bromide
15. NaBr	______	O. None of the choices match

EXERCISE

5·11

Match the name in the first column to the correct formula in the second column.

1. Butane	______	A. CH_4
2. Decane	______	B. C_2H_6
3. Ethane	______	C. C_3H_8
4. Heptane	______	D. C_4H_{10}
5. Hexane	______	E. C_5H_{12}
6. Methane	______	F. C_6H_{14}
7. Nonane	______	G. C_7H_{16}
8. Octane	______	H. C_8H_{18}
9. Pentane	______	I. C_9H_{20}
10. Propane	______	J. $C_{10}H_{22}$

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.

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

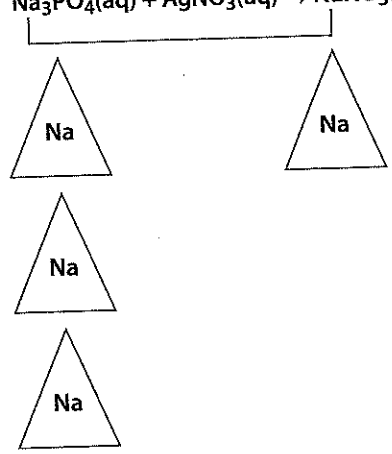


Figure 6.1

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.

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

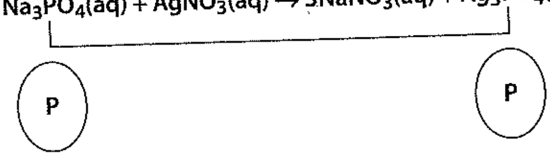


Figure 6.2

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

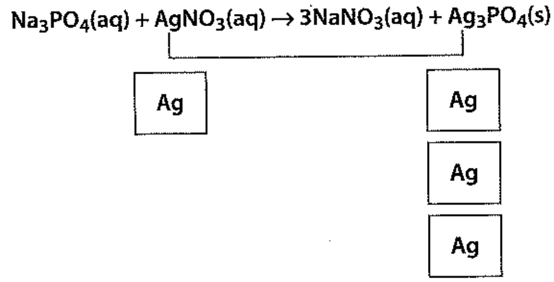


Figure 6.3

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).

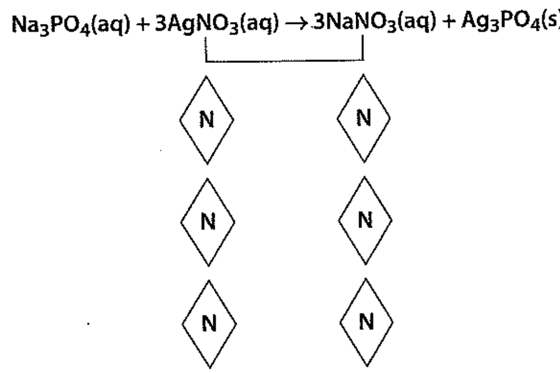


Figure 6.4

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

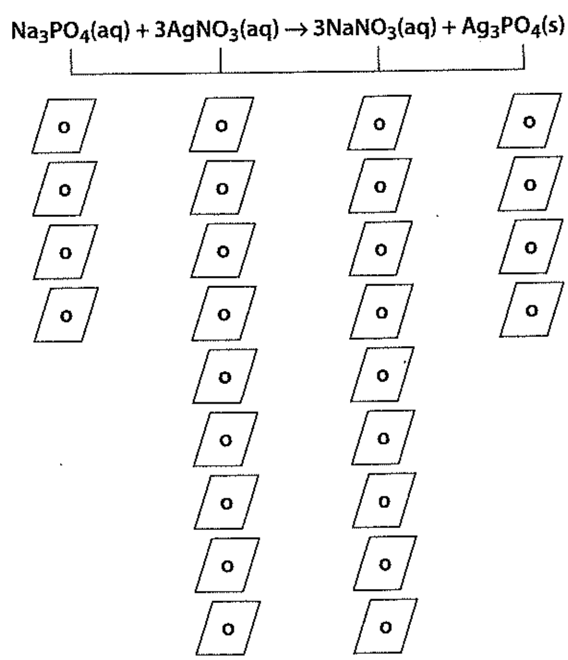


Figure 6.5

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.

4. Calcium metal reacts with oxygen in the air.

5. Hydrogen gas reacts with chlorine gas.

6. Sodium metal reacts with chlorine gas.

7. Sodium oxide solid reacts with water.

8. Silver metal reacts with oxygen in the air.

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

10. Aluminum metal reacts with oxygen in the air.

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 $\rightarrow$ 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.

____

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

____

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.

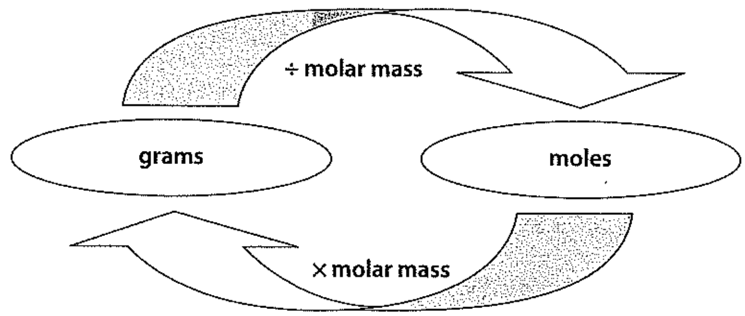


Figure 7.1

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.

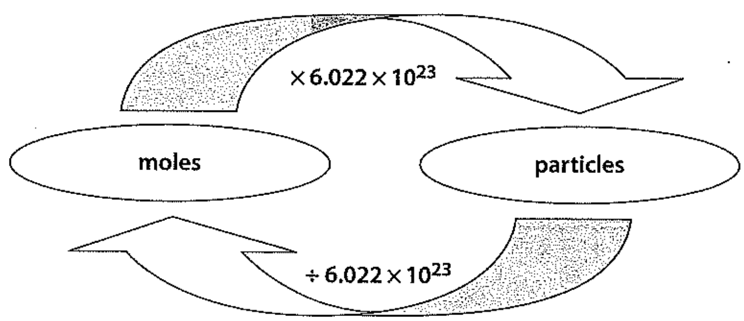


Figure 7.2

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).

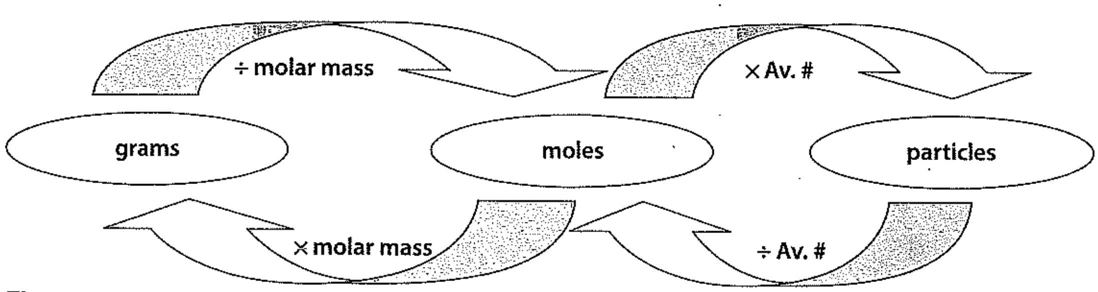


Figure 7.3

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.

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

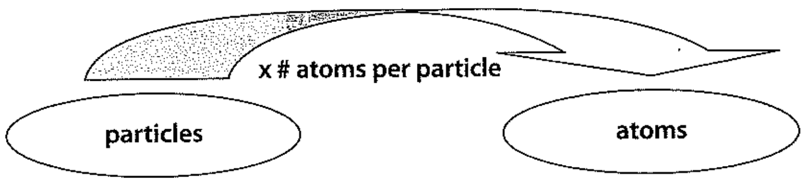


Figure 7.4

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, $\frac{1\ mol\ O_2}{2\ mol\ H_2O}$ 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.

$$37.4 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \times \frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 334 \text{ g } H_2O$$

$$50.7 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{2\text{mol } H_2O}{1 \text{ mol } O_2} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 57.1 \text{ g } H_2O$$

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.

Ui

K

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?

    ______