Calculations In Chemistry

Modules 5-7

A Note to the Student

The focus of these lessons is to provide methods to help you solve *calculations* in first-year chemistry. This is only one part of a course in chemistry, but it is often the most challenging.

Problem Notebook: The purchase of a spiral *problem notebook* is suggested as a place to write your work when solving the problems in these lessons.

Choosing a Calculator: As you do problems in these lessons (and assigned homework) that require a calculator, use the *same* calculator that you will be allowed to use during quizzes and tests. Calculators have many different labels and placements of keys. It is advisable to practice the rules and keys for a calculator before quizzes and tests.

Many courses will *not* allow the use of a graphing calculator or other types of calculators with extensive memory during tests. If a type of calculator is specified for your course, buy *two* if possible. When one becomes broken or lost, you will have a familiar backup if the bookstore is sold out later in the term.

If no type of calculator is specified for your course, any inexpensive calculator with a $\boxed{1/x}$ *or* $\boxed{x^{-1}}$, $\boxed{y^x}$ or $\boxed{\wedge}$, $\boxed{\log}$ *or* $\boxed{10^x}$, *and* $\boxed{\ln}$ functions will be sufficient for most calculations in introductory chemistry courses.

When to Do the Lessons: You will receive the maximum benefit from these lessons by completing each topic *before* it is addressed in your class.

Where to Start and Lesson Sequence: The order of these lessons may not always match the order in which topics are covered in your course. If you are using these modules as part of a course, you should do the lessons in the order in which they are assigned by your instructor. If you are using these lessons on your own to assist with a course, begin by

- Determining the *topics* that will be covered on your *next* graded assignment: problem set, quiz, or test.
- Find that topic in the Table of Contents.
- Download the modules that precede and include the topics.
- Find the *prerequisite* lessons for the topic, listed at the beginning of the module or lesson. Print the needed lessons. Do the prerequisites, then the topics related to your next graded assignments.
- Follow the instructions on "How to Use These Lessons" on page 1.

If you begin these lessons after the start of your course, when time permits, review prior topics in these lessons as needed, starting with Module 1. You will need all of these introductory modules for later topics -- and for your final exam.

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If we did not make a complete analysis of the elements of the problem, we should obtain an equation not homogeneous, and, a fortiori, we should not be able to form the equations which express … more complex cases.

. . . every undetermined magnitude or constant has one dimension proper to itself, and the terms of one and the same equation could not be compared if they had not the same exponent of dimensions.

-- Joseph Fourier, *The Analytical Theory of Heat* (1822)

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Module 5 – Word Problems

Prerequisite: Complete Modules 2 and 4 before starting Module 5.

Timing: Begin Module 5 as soon as you are assigned word-problem calculations.

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Introduction

This module includes terms and procedures that we will use to simplify problem solving for the remainder of the course. Be sure to complete all parts of Lessons 5A to 5E.

In this module you will learn to identify *given* quantities and equalities in word problems. You will then be able to solve nearly all of the initial problems assigned in chemistry with the same conversion method used in Module 4. In addition, you will be asked to *organize* your data before you solve. Most students report that by using this structured approach, they have a better understanding of the steps to take to solve science calculations.

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Lesson 5A: Answer Units -- Single or Ratio?

Types of Units

In these lessons, we will divide the units of measurements into three types.

- **Single units** have one kind of base unit in the numerator, but no denominator. Examples include meters, cubic centimeters, grams, and hours.
- **Ratio units** have one kind of base unit in the numerator and one kind in the denominator. Examples include meters/second and g/cm^3.
- **Complex units** are all other units, such as 1/sec *or* (kg·meters^2)/sec^2.

Most of the calculations encountered initially in chemistry involve single units and ratios, but not complex units. Rules for single units will be covered in this module. Distinctions between single and ratio units will be covered in Module 11. Rules for complex units will be addressed in Module 17.

Rule #1: First, Write the WANTED *Unit*

To solve word problems,

Begin by writing "WANTED: **?** " and the *unit* of the answer.

The *first* time you read a word problem, look *only* for the *unit* of the answer.

Example: For the following question,

Q. At an average speed of 25 miles/hour, how many hours will it take to go 450 miles?

Begin by writing:

* * * * *

WANTED: **? hours**

Writing the answer unit first is essential to

- help you choose the correct *given* to start your conversions,
- prompt you to write DATA that you will need to solve, and
- tell you when to stop conversions and do the math.

Rules for Answer Units

When writing the WANTED unit, it is important to distinguish between single units and ratio units.

1. An answer unit is a *ratio* unit if a problem asks you to find

   a. "unit X *over one* unit Y," or

   b. " unit X **/**unit Y " or " unit X•unit Y^{-1} " or

   c. " unit X *per* unit Y " where there is no number after *per*.

   All of those expressions are equivalent. All are ways to represent ratio units.

   Example: $\frac{\text{grams}}{\text{mL}}$, also written grams/mL *or* g • mL^{-1}, is a ratio unit.

   For an answer unit, if there is no number in the bottom unit or after the word *per*, the number *one* is understood, and the WANTED unit is a ratio unit.

   Example: "Find the speed in miles/hour (or miles per hour)" is equivalent to "find the miles traveled per *one* hour."

   A ratio unit means something per ONE something.

2. An answer unit is a *single* unit if it has a one kind of base unit in the numerator (top term) but no denominator.

   Example: If a problem asks you to find miles, or cm^3, or dollars, a single unit is WANTED.

3. If a problem asks for a "unit per *more than one* other unit," it WANTS a *single* unit.

   Example: If a problem asks for "grams per 100 milliliters," or the "miles traveled in 27 hours," it is asking for a single unit.

   A ratio unit must be something per *one* something.

Writing Answer Units

1. If you WANT a *ratio* unit, write the unit as a *fraction* with a top and a bottom.

   Example: If an answer unit in a problem is miles/hour, to start:

   Write: WANTED: **? $\frac{\textbf{miles}}{\textbf{hour}}$ =**

   Do *not* write: WANTED: ? miles/hour *or* ? mph

The slash mark (/), which is read as "per" or "over," is an easy way to *type* ratios and conversion factors. However, when solving with conversions, *writing* ratio answer units with a clear numerator and denominator will help in arranging conversions to solve.

2. If a problem WANTS a single unit, write the WANTED unit without a denominator.

   WANTED: **? miles =** or WANTED: **? mL =**

   Single units have a *one* as a denominator and are written without a denominator.

Practice

Cover the answers below with a sticky note or cover sheet. Then, for each problem, write "WANTED: **?** " and the unit that the problem is asking you to find, using the rules above. After that WANTED unit, write an equal sign.

Do not finish the problem. Write only the WANTED unit.

1. If 1.12 liters of a gas at STP has a mass of 3.55 grams, what is the molar mass of the gas in grams/mole?
2. At an average speed of 25 miles/hour, how many minutes will it take to go 15 miles?
3. If a car travels 270 miles in 6 hours, what is its average speed?
4. A student needs 420 special postage stamps. The stamps are sold with 6 stamps on a sheet, each stamp booklet has 3 sheets, and the cost is $14.40 per booklet. How much is the cost of all of the stamps?
5. How much is the cost per stamp in problem 4?

ANSWERS

1. Write **WANTED**: **? $\frac{\textbf{grams}}{\textbf{mole}}$ =** This is a ratio unit. Any unit that is in the form "unit X **/** unit Y" is a ratio unit.

2. Write **WANTED**: **? minutes =**

   This problem is asking for a single unit. *If* the problem asked for minutes per one mile, that would be a ratio unit, but minutes per *15* miles is asking for a single unit.

3. In this problem, no unit is specified. However, since the data are in miles and hours, the easiest measure of speed is miles per hour, written

   WANTED: ? $\frac{\textbf{miles}}{\textbf{hour}}$ = which is a familiar unit of speed. This problem is asking for a ratio unit.

4. **WANTED: ? $ =** *or* **WANTED: ? dollars =** The answer unit is a single unit.

5. **WANTED: ? $/stamp =** *or* **? cents/stamp =** The cost per *one* stamp is a ratio unit.

* * * * *

Lesson 5B: Mining The DATA

The method we will use to *simplify* problems is to divide solving into three parts.

WANTED:

DATA:

SOLVE:

This method will break complex problems into pieces. You will always know what steps to take to solve a problem because we will solve all problems with the same three steps.

Rules for DATA

To solve word problems, get rid of the words.

By translating words into numbers, units, and labels, you can solve most of the initial word problems in chemistry by chaining conversions, as you did in Module 4. To translate the words, write in the DATA section on your paper every *number* you encounter as you read the problem, followed by its *unit* and a *label* that describes the quantity being measured.

In the initial problems of chemistry, it is important to distinguish numbers and units that are parts of equalities from those that are not. To do so, we need to learn the many ways that quantities that are *equal* or *equivalent* can be expressed in words and symbols.

Rules for Listing DATA in Word Problems

1. Read the problem. Write "WANTED: ?" followed by the WANTED unit and an = sign.
2. On the next line down, write "DATA:"
3. Read the problem a second time.
   - Each time you find a number, *stop*. Write the number on a line under "DATA:"
   - After the number, write its *unit* plus a *label* that helps to identify the number.
   - Decide if that number, unit, and label is *paired* with another number, unit, and label as part of an equality**.**
4. In the DATA section, *write* each number and unit in the problem in an *equality*
   a. If you see *per* or **/** (a slash). Write p*er* or **/** in DATA as an equal sign (=).
      - If a number is shown after *per* or /, write the number in the equality.

        Example: If you read " $8 *per* 3 lb. " write in the DATA: "$8 **=** 3 lb."
      - If *no* number is shown after *per* or /, write *per* as " =**1** "

        Example: If you see "25 km/hour," write "25 km = **1** hour"
      - Treat *unit x* • *unit y*$^{-1}$ the same as *unit x* **/** *unit y*.

        Example: If you see " 75 g • mL^{-1} " write " 75 g = **1** mL "

b. If the same quantity is measured using two different units.

Examples: If a problem says, "0.0350 moles of gas has a volume of 440 mL,"

write in your DATA: "0.0350 moles of gas = 440 mL"

If a problem says a bottle is labeled "2 liters (67.6 fluid ounces),"

write: "2 liters = 67.6 fluid ounces "

In both cases, the *same* physical quantity is being measured in two different units.

c. Any time two measurements are taken of the same process.

If a problem says, "burning 0.25 grams of candle wax releases 1700 calories of energy," write in your DATA section,

"0.25 grams candle wax = 1700 calories of energy"

Both sides are measures of what happened as this candle burned.

5. Watch for words such as *each* and *every* that mean *one*. *One* is a number, and you want *all* numbers in your DATA table.

   If you read, "*Each* student was given 2 sodas, " *write* " 1 student = 2 sodas"

6. Continue until *all* of the *numbers* in the problem are written in your DATA.

7. Note that when writing the WANTED unit, you write "per one" as a ratio unit and "per more than one" as a single unit.

   In the DATA, however, "per one" and "per more than one" are written in the same way: as an equality.

Practice

1. For each phrase below, write the equality that you will add to your DATA. On each side of the equal sign, include a number and a unit. After each unit, if two different entities are being measured in the problem, add additional words that identify what is being measured by that number and unit. After every few, check your answers.

   a. The car was traveling at a speed of 55 miles/hour.

   b. A bottle of designer water is labeled 0.50 liters (16.9 fluid ounces).

   c. Every student was given 19 pages of homework.

   d. To melt 36 grams of ice required 2,880 calories of heat.

   e. The molar mass is 18.0 grams $H_2O \bullet$ mole H_2O^{-1}.

   f. The dosage of the aspirin is 2.5 mg per kg of body mass.

   g. If 0.24 grams of NaOH are dissolved to make 250 mL of solution, what is the concentration of the solution?

2. For Problems 1-4 in the **Practice** for Lesson 5A, write DATA: and then list the data *equalities* that are supplied in the problem.

ANSWERS

Terms that are equal may always be written in the reverse order. If there are two different entities in a problem, attach labels to the units that identify which entity the number and unit are measuring.

1a. 55 miles = 1 hour (Rule 4a) 1b. 0.50 liters = 16.9 fluid ounces (Rule 4b)

1c. 1 student = 19 pages (Rule 5) 1d. 36 grams ice = 2,880 calories heat (Rule 4c: Equivalent)

1e. 18.0 grams H_2O = 1 mole H_2O (Rule 4b) 1f. 2.5 mg aspirin = 1 kg of body mass (Rule 4a)

1g. 0.24 g NaOH = 250 mL of soln. (Rule 4b)

2.	Problem 1.	DATA:	1.12 L gas STP = 3.55 g	(2 measures of same gas)
	Problem 2.	DATA:	25 miles = 1 hour	(Write / as = 1)
	Problem 3.	DATA:	270 miles = 6 hours	(2 measures of same trip)
	Problem 4.	DATA:	6 stamps = 1 sheet	
			1 booklet = 3 sheets	
			$14.40 = 1 booklet	

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Lesson 5C: Solving For Single Units

The Law of Dimensional Homogeneity

By the law of dimensional homogeneity, the *units* on both sides of an *equality* must, at the *end* of the calculation, be the same. One implication of this law is: to find a WANTED single unit, a single unit amount must be supplied in the data. Using this law, we will simplify problem solving by starting single-unit calculations with an equality:

? WANTED single unit = # *given* single unit

then convert the *given* to the WANTED *unit*.

DATA Formats If a Single Unit is WANTED

If a problem WANTS a *single* unit, *one* number and unit in the DATA is likely to be

- either a number and its unit that is not paired in an equality with other measurements, *or*
- a number and its unit that is paired with the WANTED unit in the format

  "? unit WANTED = # unit *given*"

We will define the *given* as the term written to the right of the equal sign: the starting point for the terms that we will multiply to solve conversion calculations.

If a problem WANTS a *single*-unit amount, by the laws of science and algebra, at least *one* item of DATA must be a single-unit amount. In problems that can be solved using conversions, often *one* measurement will be a single unit, and the rest of the DATA will be equalities.

If a single unit is WANTED, *watch* for one item of data that is a single unit amount. In the DATA, write the single number, unit, and label on a line by itself.

It is a good practice to circle that single unit amount in the DATA, since it will be the *given* number and unit that is used to *start* your conversions.

Variations on the above rules will apply when DATA includes two amounts that are equivalent in a problem. We address these cases in Module 11. However, for the problems you are initially assigned in first-year chemistry, the rules above will most often apply.

To SOLVE

After listing the DATA provided in a problem, below the DATA, write SOLVE. Then, *if* you WANT a single unit, write the WANTED and *given* measurements in the format of the *single-unit starting template*.

$$? \text{ unit WANTED} = \# \text{ and } \textbf{\textit{unit}} \textit{ given} \bullet \frac{}{\textbf{\textit{unit}} \textit{ given}}$$

The *given* measurement that is written after the = sign will be the circled single unit listed in the DATA.

To convert to the WANTED unit, use the equalities in the DATA (and other fundamental equalities, such as metric prefix definitions, if needed).

Summary: The 3-Step Method to *Simplify* Problem Solving

1. **WANTED:**

   When reading a problem for the first time, ask *one* question: what will be the *unit* of the answer? Then, write "WANTED: ***?***", the *unit* the problem is asking for, and a *label* that describes what the unit is measuring. Then add an = sign.

   Write WANTED ratio units as $\frac{x}{y}$ fractions and single units as single units.

2. **DATA:**

   Read the problem a second time.

   - Every time you encounter a *number*, under DATA write the number and its unit. Add a label after the unit if possible, identifying what is being measured.
   - Then see if that number and unit are equal to another number and unit.

   If a problem WANTS a single unit, most often *one* measurement will be a single unit and the rest will be equalities. Circle the *single* unit in the DATA.

3. **SOLVE:**

   Start each calculation with an *equality*: ? WANTED unit = # *given* unit.

   If you WANT a single unit, substitute the WANTED and *given* into this format.

   $$? \text{ unit WANTED} = \# \text{ and } \textbf{unit} \textit{ given} \bullet \frac{}{\textbf{unit} \textit{ given}}$$

   Then, using equalities, convert to the WANTED unit.

Solve the following problem in your notebook using the 3-step method above.

Q. If a car's speed is 55 miles/hr., how many minutes are needed to travel 85 miles?

* * * * * (* * * mean cover the answer below, write *your* answer, then check it.)

Your paper should look like this.

WANTED: **?** minutes =

DATA: 55 miles = 1 hour

85 miles

SOLVE: $\textbf{? minutes} = \textbf{85 miles} \cdot \frac{1\text{ hour}}{55\text{ miles}} \cdot \frac{60\text{ min.}}{1\text{ hour}} = \boxed{93\text{ minutes}}$

You can solve *simple* problems without listing WANTED, DATA, SOLVE, but this 3-part method works for *all* problems. It works especially well for the complex problems that soon you will encounter. By using the same three steps for every problem, you will know what to do to solve *all* problems. That's the goal.

Practice

Many science problems are constructed in the following format.

"Equality, equality," then, "**?** WANTED unit = a *given* number and unit."

The problems below are in that format. Using the rules above, solve on these pages *or* by writing the WANTED, DATA, SOLVE sections in your notebook.

If you get stuck, read part of the answer at the end of this lesson, adjust your work, and try again. Do problems 1 and 3, and problem 2 if you need more practice.

Problem 1

If 2.2 pounds = 1 kg, what is the mass in grams of 12 pounds?

WANTED: ? (Write the **unit** you are looking for.)

DATA: (Write every number and unit in the problem here. If solving for a single unit, often *one* number and unit is unpaired, and the rest are in equalities, Circle the unpaired single unit.)

SOLVE: (Start with " **?** unit **WANTED** = # and unit ***given*** • $\frac{\qquad\qquad}{\text{unit } given}$ "

?

* * * * *

Problem 2

If there are 1.6 km/mile, and one mile is 5,280 feet, how many feet are in 0.500 km?

WANTED: ?

DATA:

SOLVE:

?

* * * * *

Problem 3

If there are 3 floogles per 10 schmoos, 5 floogles/mole, and 3 moles have a mass of 25 gnarfs, how many gnarfs are in 4.2 schmoos? (Assume the whole numbers are exact.)

WANTED:

DATA:

SOLVE:

ANSWERS

1. WANTED: ? g =

   DATA: 2.2 pounds = 1 kg

   12 pounds

   SOLVE:

   $$? \text{ g} = 12 \cancel{\text{pounds}} \cdot \frac{1 \cancel{\text{kg}}}{2.2 \cancel{\text{pounds}}} \cdot \frac{10^3 \text{ g}}{1 \cancel{\text{kg}}} = \frac{12 \cdot 10^3}{2.2} \text{ g} = 5.5 \times 10^3 \text{ g}$$

   A single unit is WANTED, and the DATA has one single unit.

   Note that the SOLVE step begins with "how many grams equal 12 pounds?"

   Fundamental conversions such as kilograms to grams need not be written in your DATA section, but they will often be needed to solve. Be certain that you have mastered the metric system fundamentals.

2. WANTED: ? feet =

   DATA: 1.6 km **=** 1 mile

   1 mile **=** 5,280 feet

   0.500 km

   SOLVE:

   $$? \textbf{ feet} = 0.500 \cancel{\text{km}} \cdot \frac{1 \cancel{\text{mile}}}{1.6 \cancel{\text{km}}} \cdot \frac{5{,}280 \textbf{ feet}}{1 \cancel{\text{mile}}} = \frac{0.500 \cdot 5280}{1.6} \text{ feet} = = 1{,}650 \textbf{ feet}$$

3. WANTED: ? gnarfs =

   DATA: 3 floogles = 10 schmoos

   5 floogles = 1 mole

   3 moles = 25 gnarfs

   4.2 schmoos

   SOLVE:

   At the SOLVE step, first state the question, "how many gnarfs equal 4.2 schmoos?"

   Then add the first conversion, set up to cancel your *given* unit.

   $$\textbf{? gnarfs} = 4.2 \text{ schmoos} \cdot \frac{\qquad\qquad}{\text{schmoos}}$$

   Since only one equality in the DATA contains schmoos, use it to complete the conversion.

   $$\textbf{? gnarfs} = 4.2 \cancel{\text{schmoos}} \cdot \frac{3 \text{ floogles}}{10 \cancel{\text{schmoos}}}$$

   On the right, you now have floogles. On the left, you WANT gnarfs, so you must get **rid** of floogles. In the next conversion, put floogles where it will **cancel**.

   $$\textbf{? gnarfs} = 4.2 \cancel{\text{schmoos}} \cdot \frac{3 \text{ floogles}}{10 \cancel{\text{schmoos}}} \cdot \frac{\qquad\qquad}{\text{floogles}}$$

Floogles is in *two* conversion factors in the DATA, but one of them takes us back to schmoos, so let's use the other.

$$\textbf{? gnarfs} = 4.2 \text{ schmoos} \bullet \frac{3 \cancel{\text{floogles}}}{10 \text{ schmoos}} \bullet \frac{1 \text{ mole}}{5 \cancel{\text{floogles}}}$$

Moles must be gotten rid of, but moles has a known relationship with the *answer* unit. Convert from moles to the answer unit. Since, after unit cancellation, the answer unit is now where you WANT it, stop conversions and do the arithmetic.

$$\textbf{? gnarfs} = 4.2 \text{ schmoos} \bullet \frac{3 \cancel{\text{floogles}}}{10 \text{ schmoos}} \bullet \frac{1 \cancel{\text{mole}}}{5 \cancel{\text{floogles}}} \bullet \frac{25 \textbf{ gnarfs}}{3 \cancel{\text{moles}}} = \frac{4.2 \bullet \cancel{3} \bullet 25}{10 \bullet 5 \bullet \cancel{3}} \textbf{ gn.} = \textbf{2.1 gnarfs}$$

* * * * *

Lesson 5D: Finding the *Given*

Ratio Unit *Givens*

In chemistry, the initial quantitative topics generally involve solving for single units, so that will be our initial focus as well. Conversion factors may also be used to solve for ratio units, as we did in Lesson 4E.

However, we will defer most of the rules to use conversions to solve for ratio units until Lesson 11B, when ratio units will be needed to solve for the concentration of chemical solutions. If you need to solve word problems that have ratio-unit answers, now or at any later point, Lesson 11B may be done at any time after completing this lesson.

Single-Unit *Givens*

When solving for single units, the *given* quantity is not always clear. For example,

Q. A student needs special postage stamps. The stamps are sold 6 per sheet, each stamp booklet has 3 sheets, 420 stamps are needed, and the cost is $43.20 per 5 booklets. What is the cost of the stamps?

Among all those numbers, which is the *given* needed as the first term when you SOLVE?

For a single-unit answer, finding the *given* is often a process of elimination. If all of the numbers and units are paired into equalities except one, that one is your *given*.

In your notebook, write the WANTED and DATA sections for the stamps problem above (don't SOLVE yet). Then check your work below.

* * * * *

Answer: Your paper should look like this.

WANTED: **?** $ = *or* **?** dollars = (you could also solve in cents)

DATA: 1 sheet **=** 6 stamps

3 sheets = 1 booklet

420 stamps

$43.20 **=** 5 booklets

Since you are looking for a *single* unit, dollars, your data has one number and unit that did not pair up in an equality: 420 stamps. That is your *given*.

To SOLVE, the rule is

> If you WANT a single unit, *start* with a single unit as your *given*.

Apply the above rule, assume all of these numbers are exact, and SOLVE the problem.

* * * * *

Answer

SOLVE: If you WANT a single unit, start with the single-unit starting template.

$$? \$ = 420 \text{ stamps} \cdot \frac{\qquad}{\text{stamps}}$$

Putting the *given* unit where it *must* be to cancel in the next conversion will help you to pick the DATA for and arrange the DATA in the next conversion.

If you needed that hint, adjust your work and finish.

* * * * *

$$\mathbf{? \$ =} \ 420 \text{ stamps} \cdot \frac{1 \text{ sheet}}{6 \text{ stamps}} \cdot \frac{1 \text{ booklet}}{3 \text{ sheets}} \cdot \frac{\$ 43.20}{5 \text{ booklets}} = \boxed{\mathbf{\$ 201.60}}$$

Practice

For each problem below, use the WANTED, DATA, SOLVE method. If you get stuck, peek at the answers and try again. Do at least two problems. If you plan on taking physics, be sure to do problem 3.

On each of these, *before* you do the math, double-check each conversion, one at a time, to make sure it is legal.

1. A bottle of drinking water is labeled "12 fluid ounces (355 mL)." What is the mass in centigrams of 0.55 fluid ounces of the H_2O? (Use the definition of one gram).

2. You want to mail a large number of newsletters. The cost is 18.5 cents each at special bulk rates. On the post office scale, the weight of exactly 12 newsletters is 10.2 ounces. The entire mailing weighs 125 lb. There are 16 ounces (oz.) in a pound (lb.).
   a. How many newsletters are being mailed?
   b. What is the cost of the mailing in dollars?

3. If the distance from an antenna on Earth to a geosynchronous communications satellite is 22,300 miles, given that there are 1.61 kilometers per mile, and radio waves travel at the speed of light (3.0×10^8 meters/sec), how many seconds does it take for a signal from the antenna to reach the satellite?

ANSWERS

1. WANTED: **?** cg =

DATA: 12 fl. oz = 355 mL

0.55 fl. oz

1.00 g $H_2O_{(l)}$ = 1 mL $H_2O_{(l)}$ (metric definition of one gram)

SOLVE:

$$? \textbf{ cg} = 0.55 \text{ fl. oz.} \cdot \frac{355 \text{ mL}}{12 \text{ fl. oz}} \cdot \frac{1.00 \text{ g } H_2O_{(l)}}{1 \text{ mL } H_2O_{(l)}} \cdot \frac{1 \text{ cg}}{10^{-2} \text{ g}} = \textbf{1,600 cg}$$

2a. WANTED: **?** newsletters

DATA: 18.5 cents = 1 newsletter

12 exact newsletters = 10.2 ounces

16 oz. = 1 lb. (a definition with infinite *sf*)

125 lb.

SOLVE: $? \text{ newsletters} = 125 \text{ lb.} \cdot \frac{16 \text{ oz.}}{1 \text{ lb.}} \cdot \frac{12 \text{ newsls}}{10.2 \text{ oz.}} = \textbf{2,350 newsletters}$

2b. WANTED: **?** dollars

(Strategy: Since you want a single unit, you can start over from your single *given* unit (125 lb.), repeat the conversions above, then add 2 more.

Or you can start from your single unit answer in Part a, and solve using the two additional conversions.

In problems with multiple parts, to solve for a later part, using an answer from a previous part often saves time.)

DATA: same as for Part a.

SOLVE: $\textbf{? dollars} = 2{,}350 \text{ newsls} \cdot \frac{18.5 \text{ cents}}{1 \text{ newsl}} \cdot \frac{1 \text{ dollar}}{100 \text{ cents}} = \textbf{\$ 435}$

3. WANTED: **?** seconds =

DATA: 22,300 miles

1.61 km = 1 mile

3.0×10^8 meters = 1 sec

SOLVE:

$$? \text{ sec} = 22{,}300 \text{ mi.} \cdot \frac{1.61 \text{ km}}{1 \text{ mile}} \cdot \frac{10^3 \text{ meters}}{1 \text{ km}} \cdot \frac{1 \text{ s}}{3.0 \times 10^8 \text{ m}} = \frac{22{,}300 \cdot 1.61 \cdot 10^3}{3.0 \times 10^8} \text{ sec} = \textbf{0.12 s}$$

(This means that the time up *and* back for the signal is 0.24 seconds. You may have noticed this one-quarter-second delay during some live broadcasts which bounce video signals off satellites but use faster land-lines for audio, or during overseas communications routed through satellites.)

* * * * *

Lesson 5E: Some Chemistry Practice

Listing Conversions and Equalities

Which is the best way to write DATA pairs: as *equalities* or in the *fraction* form as conversion-factor ratios? Mathematically, either form may be used.

In DATA: the equalities

1.61 km = 1 mile	can be listed as	$\frac{1.61\text{ km}}{1\text{ mile}}$, $\frac{3.0 \times 10^8\text{ meters}}{1\text{ sec.}}$
3.0×10^8 meters = 1 sec.		

In these lessons, we will generally write *equalities* in the DATA section. This will emphasize that when solving problems using conversions, you need to focus on the relationship between two quantities. However, listing the data in the fraction format is equally valid. Data may be portrayed both ways in science texts.

Why "Want A Single Unit, Start With A Single Unit?"

Mathematically, the order in which you multiply conversions does not matter. You could solve with your single unit *given* written anywhere on top in your chain of conversions.

However, if you start with a *ratio* as your *given* when solving for a single unit, there is a 50% chance of starting with a ratio that is inverted. If this happens, the units will never cancel correctly, and you would eventually be forced to start the conversions over. *Starting* with the single unit is a method that automatically arranges your conversions "right-side up."

Practice

Let's do some chemistry. The problems below supply the DATA needed for conversion factors. In upcoming modules, you will learn how to write these conversions automatically even when the problem does not supply them. That small amount of additional information is all that you will need to solve most initial chemistry calculations.

You're ready. For problems 1-3, solve two of these problems in your notebook now and one in your next study session. Do include chemical formulas after units. Don't let strange terms like *moles* or *STP* bother you. You've done gnarfs. You can do these.

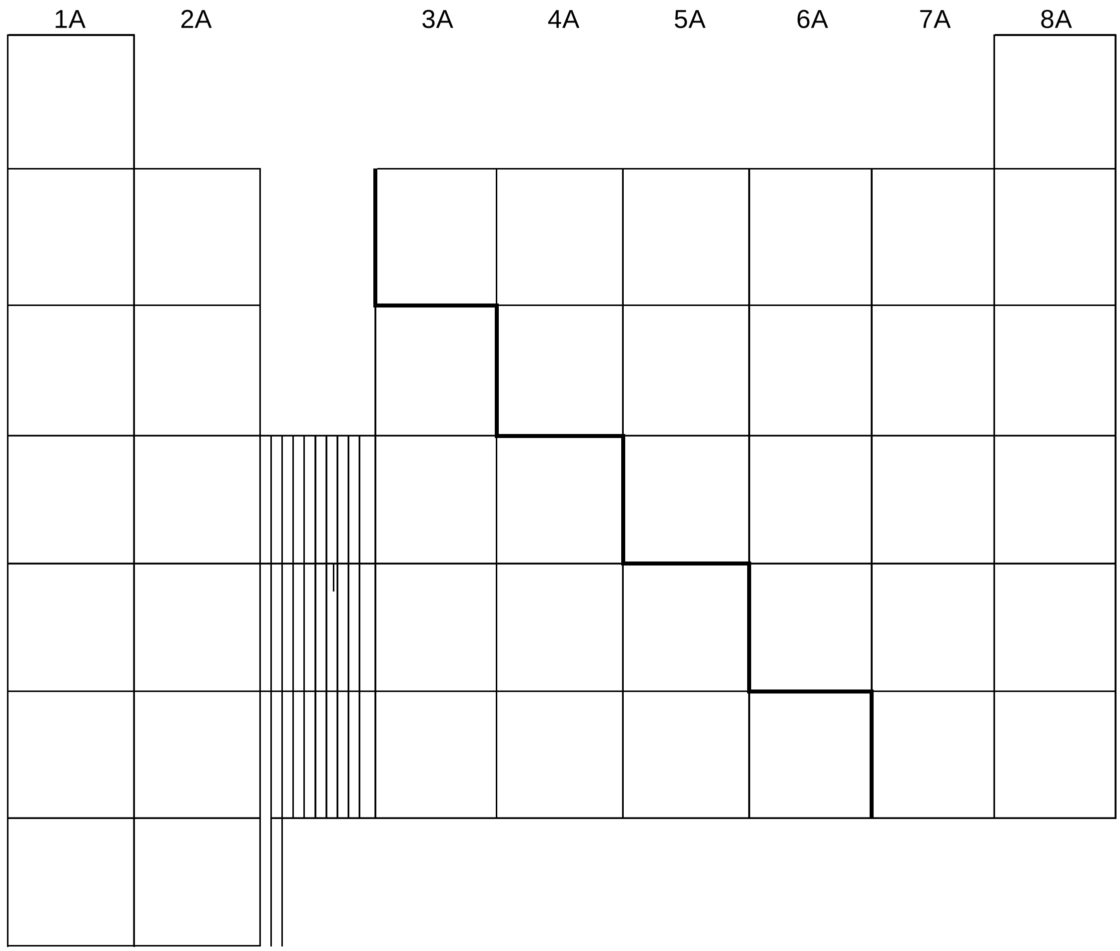
1. Water has a molar mass of 18.0 grams H_2O per mole H_2O. How many moles of H_2O are in 450 milligrams of H_2O?

2. If one mole of all gases has a volume of 22.4 liters at STP, and the molar mass of chlorine gas (Cl_2) is 71.0 grams Cl_2 per mole Cl_2 , what is the volume, in liters, of 28.4 grams of Cl_2 gas at STP ?

3. If 1 mole of H_2SO_4 = 98.1 grams of H_2SO_4 and it takes 2 moles of NaOH per 1 mole of H_2SO_4 for neutralization, how many liters of a solution that is 0.240 mol NaOH/liter is needed to neutralize 58.9 grams of H_2SO_4?

4. On the following table, fill in the names and symbols for the atoms in the first 3 *rows* and the first 2 and last 2 *columns*.

Periodic Table



* * * * *

ANSWERS

1. WANTED: **? moles H_2O =**

   DATA: 18.0 grams H_2O = 1 mole H_2O

   450 mg H_2O

   SOLVE:

$$\textbf{? moles } H_2O = 450 \text{ mg } H_2O \cdot \frac{10^{-3} \text{ g}}{1 \text{ mg}} \cdot \frac{1 \textbf{ mole } H_2O}{18.0 \text{ g } H_2O} = 2.5 \times 10^{-2} \textbf{ moles } H_2O$$

   Write chemistry data in 3 parts: Number, unit, formula. Writing complete labels will make complex problems easier to solve. 450 has 2 *sf*.

2. WANTED: ? L Cl_2 v (always attach chemical formulas, if known, to units)

DATA: 1 mole gas = 22.4 L gas

71.0 g Cl_2 = 1 mole Cl_2

28.4 g Cl_2

SOLVE:

$$\textbf{? L Cl}_{\mathbf{2}} = 28.4\text{ g Cl}_2 \bullet \frac{1\text{ mole Cl}_2}{71.0\text{ g Cl}_2} \bullet \frac{22.4\text{ L Cl}_2}{1\text{ mole Cl}_2} = 8.96\ \textbf{L Cl}_{\mathbf{2}}$$

3. WANTED: ? L NaOH solution

DATA: 1 mole H_2SO_4 = 98.1 grams H_2SO_4

2 moles NaOH = 1 mole H_2SO_4 (assume whole numbers are exact)

0.240 moles NaOH = 1 liter NaOH solution

58.9 grams H_2SO_4

SOLVE:

$$\text{? L NaOH} = 58.9\text{ g H}_2\text{SO}_4 \bullet \frac{1\text{ mole H}_2\text{SO}_4}{98.1\text{ g H}_2\text{SO}_4} \bullet \frac{2\text{ mole NaOH}}{1\text{ mole H}_2\text{SO}_4} \bullet \frac{1\text{ L NaOH soln.}}{0.240\text{ mole NaOH}} = \boxed{\textbf{5.00}\text{ L NaOH soln.}}$$

* * * * *

Lesson 5F: Area and Volume Conversions

Timing: Do this lesson *if* you are assigned area and volume conversions based on taking distance conversions to a power, *or if* you are majoring in science or engineering.

Pretest: If you think you know this topic, try the last problem in the lesson. If you can do that problem, you may skip the lesson.

* * * * *

Area

The rules are

Rule A1. Area, by definition, is distance squared. All units that measure area can be related to distance units squared.

Rule A2. Any unit that measures distance can be used to *define* an *area unit*. The area unit is simply the distance unit squared.

Rule A3. Any equality that relates two distance units can be used as an area conversion by *squaring* both sides of the distance conversion.

Rule A4. In conversions, write “square units” as units^2.

By Rule A2, area units can be any distance unit squared, such as square centimeters, square kilometers, or square miles.

Using Rule A3, we can calculate a conversion factor between any two area units that are distance units squared by starting from the distance to distance equality.

<u>For example</u>: Since $\boxed{1 \text{ mile} = 1.61 \text{ km}}$ is a distance conversion,

and any equality squared on both sides remains true,

$(1 \text{ mile})^2 = (1.61 \text{ km})^2$

$1^2 \text{ mile}^2 = (1.61)^2 \text{ km}^2$

$\boxed{1 \text{ mile}^2 = 2.59 \text{ km}^2}$ which can be used as an area conversion.

Based on the above, you can say that "one *square* mile is equal to 2.59 *square* kilometers."

Note that in squaring an equality, all parts (each *number* and *unit*) must be squared.

When an area conversion based on a distance conversion is needed, the area conversion can be calculated separately, as above. However, the area conversion can also be constructed in or after the *given* as part of your chained conversions.

The logic: any two quantities that are equal can be used as a conversion factor. Since the value of any conversion factor **=** 1, and both sides of an equation can be taken to a power and the equation will still be true, then

$$\text{if } \mathbf{A = B}\text{, then } \frac{A}{B} = 1 \text{ and } \left(\frac{A}{B}\right)^{\mathbf{2}} = 1^{\mathbf{2}} = 1 = \frac{\mathbf{A^2}}{\mathbf{B^2}}$$

Since A^2/B^2 and $(A/B)^2$ both equal 1, both are legal conversion factors.

The general rule is:

> Any distance to distance equality or conversion can be *squared* and used as an area conversion, or *cubed* and used as a volume conversion.

Use that rule to complete this un-finished conversion, solve, then check below.

$$? \text{ miles}^2 = 75 \text{ km}^2 \bullet \left(\frac{1 \text{ mile}}{1.61 \text{ km}}\right)$$

* * * * *

For km^2 in the *given* to cancel and convert to miles^2 on top, *square* the miles-to-km distance conversion. As above, when you square the conversion, be sure to square everything (each number and each unit) inside the parentheses. Adjust your work and finish if needed.

* * * * *

$$? \text{ miles}^2 = 75 \text{ km}^2 \bullet \left(\frac{1 \text{ mile}}{1.61 \text{ km}}\right)^{\mathbf{2}} = 75 \cancel{\text{km}^2} \bullet \frac{1^2 \text{ mile}^2}{(1.61)^2 \cancel{\text{km}^2}} = \frac{75}{2.59} \text{ miles}^2 = \mathbf{29 \text{ miles}^2}$$

The result above means that the *given* 75 square kilometers is equal to 29 square miles.

Practice A

1. If 25.4 mm = 1 inch and 12 inches = 1 foot
   a. ? in. = 1.00 mm
   b. ? in^2 = 1.00 mm^2
   c. ? mm^2 = 2.00 ft^2
2. A standard sheet of notebook paper has dimensions of 8.50 x 11.0 inches.
   a. What is the area of one side of the sheet of paper, in square inches?
   b. Using your *part a* answer and $\boxed{\text{2.54 cm = 1 inch}}$, calculate the area of one side of the sheet of paper in square centimeters.
3. Under the grid system used to survey the American Midwest, a *section*, which is one square mile, is 640 acres. The smallest unit of farm land typically surveyed was a "quarter quarter section" of 40 acres. If 1 mile = 1.61 km, 40.0 acres is how many km^2?

Volume

Volume, by definition, is distance cubed. Note that in each of the following equations used to calculate the volume of solids, measurements of *distance* are multiplied *three* times.

- Volume of a rectangular solid = $l \times w \times h$
- Volume of a cylinder = $\pi r^2 h$
- Volume of a sphere = $4/3\, \pi r^3$

The rules for volume calculations using distance units parallel those for area calculations.

Rule V1. Volume, by definition, is distance cubed. All units that measure volume can be related to distance units cubed.

Rule V2. Any unit that measures distance can be used to *define* a *volume unit*. The volume unit is simply the distance unit cubed.

Rule V3. Any equality that relates two distance units can be used as a volume conversion factor by *cubing* both sides of the distance conversion.

Rule V4. In conversions, write "cubic units" as $units^3$ (cubic meters = m^3)

In chemistry, volume units are used more often than area units. Some key relationships used in distance and volume calculations are

- 1 meter = 10 *deci*meters = 100 *centi*meters, which means
- 1 decimeter = 10 centimeters.

Since volume is distance cubed, and one milliliter is *defined* as one cubic centimeter, we can write metric fundamental rules 4 and 5:

4. $\boxed{1\ cm^3 = 1\ cc\ = 1\ mL}$ and

5. A cube that is 10 cm x 10 cm x 10 cm = 1 dm x 1 dm x 1 dm =

$\boxed{=\ 1{,}000\ cm^3 = 1{,}000\ mL = 1\ L = 1\ dm^3}$ (see Lesson 2A.)

In the English measurement system, volume units include fluid ounces, teaspoons, tablespoons, cups, quarts, and gallons. However, any English distance unit, such as inches, feet, or miles, can also be used to define a volume unit, such as in^3, ft^3, and $miles^3$.

A conversion that can be used to convert between English and metric volume units is the "soda can" equality: 12.0 fluid ounces = 355 mL.

Any distance to distance *equality* can be cubed to serve as a volume conversion.

For example, since $\boxed{1\ \text{foot} \equiv 30.48\ cm}$, $\mathbf{1\ foot^3} \equiv (30.48)^3\ cm^3 = \mathbf{28{,}317\ cm^3}$

and since $\boxed{1\ km\ \equiv 10^3\ m}$, $\mathbf{1\ km^3} \equiv (10^3)^3\ m^3 = \mathbf{10^9\ m^3}$

Note that each number and each unit must be cubed when an equality is cubed.

This general rule applies to both area and volume conversions:

A conversion factor written as a fraction or equality can be taken to *any power* needed in order to cancel units, and the conversion will remain legal (equal to one).

Use that rule to solve this problem.

Q. Lake Erie, the smallest Great Lake, holds an average 485 km^3 of water. What is this volume in cubic miles? (1.61 km = 1 mile).

* * * * *

WANTED: ? $miles^3$ (in calculations, write cubic units as $units^3$.)

DATA: 1.61 km = 1 mile

484 km^3

SOLVE: $? \ miles^3 = 485\ km^3 \bullet \left(\frac{1\ mile}{1.61\ km}\right)$

The above conversion is un-finished. Complete it, solve, then check below.

* * * * *

To get the *given* km^3 to convert to $miles^3$, use the miles-to-km distance conversion, cubed. When cubing the conversion, be sure to cube everything inside the parentheses.

* * * * *

$$\textbf{?}\ miles^3 = 485\ km^3 \bullet \left(\frac{1\ mile}{1.61\ km}\right)^{\mathbf{3}} = 485\ \cancel{km^3} \bullet \frac{1^3\ mi.^3}{(1.61)^3\ \cancel{km^3}} = \frac{485}{4.17}\ mi.^3 = \mathbf{116\ miles^3}$$

To cube 1.61, either multiply 1.61 x 1.61 x 1.61 *or* use the y^x function on your calculator.

Practice B

Use the conversions above. Do at least every other problem now, but save one or two until prior to your test on this material. The more challenging problems are at the end. If you get stuck, read a *part* of the answer, then try again. Be sure to do problem 4.

1. If one mile = 1.61 km, solve: **?** km^3 = 5.00 $miles^3$
2. How many cubic millimeters are in one cubic meter?
3. If 25.4 mm = 1 inch, how many cubic inches are equal to 1.00 cubic millimeters?
4. 0.355 liters
   a. is how many cubic centimeters?
   b. Using 12 in. = 1 foot and 1 in. = 2.54 cm , convert your *part a* answer to cubic feet.
5. ? dm^3 = 67.6 fluid ounces (Finish. Include the soda-can conversion.)
6. The flathead V-twin engine on the 1947 Indian Chief motorcycle has a 74 cubic inch displacement. What is this displacement in cc's? (1 in. = 2.54 cm)
7. Each minute, the flow of water over Niagara Falls averages 1.68×10^5 m^3. What is this flow
   a. In cubic feet? (1 meter = 3.28 feet)
   b. In gallons? (1 gallon = 3.79 liters)
8. Introduced in 1960, the Chevrolet big block engine, when configured with dual four-barrel carburetors and 11.3:1 compression, developed 425 horsepower at 6200 RPM. The cylinders of this hydrocarbon-guzzling behemoth displaced 6.70 L. Immortalized by the Beach Boys, what is this displacement in cubic inches? (1 in. = 2.54 cm)

ANSWERS

Practice A

1\. a. $? \text{ in.} = 1.00 \text{ mm} \cdot \frac{1 \text{ inch}}{25.4 \text{ mm}} = 0.0394 \text{ in.}$

b. $? \text{ in}^2 = 1.00 \text{ mm}^2 \cdot \left(\frac{1 \text{ inch}}{25.4 \text{ mm}}\right)^2 = 1.00 \cancel{\text{mm}^2} \cdot \frac{1^2 \text{ in}^2}{(25.4)^2 \cancel{\text{mm}^2}} = \frac{1}{645} \text{ in}^2 = \mathbf{0.00155\ in^2}$

c. $? \text{ mm}^2 = 2.00 \text{ ft}^2 \cdot \left(\frac{12 \text{ in.}}{1 \text{ ft.}}\right)^2 \cdot \left(\frac{25.4 \text{ mm}}{1 \text{ in}}\right)^2 = 2.00 \text{ ft}^2 \cdot \frac{(12)^2 \text{ in}^2}{1^2 \text{ ft}^2} \cdot \frac{(25.4)^2 \text{ mm}^2}{1^2 \text{ in}^2} = \mathbf{1.86 \times 10^5\ mm^2}$

2\. a. Area = length x width = 8.50 in. *x* 11.0 in. = **93.5 in^2**

b. WANT: **?** cm^2 (a wanted *single* unit)

DATA: 2.54 cm **=** 1 inch (a *ratio*)

93.5 in^2 (a *single* unit. Answers from earlier parts are DATA for later parts)

SOLVE: (if you *want* a single unit, *start* with the *single* unit in the data as your *given*)

$$? \text{ cm}^2 = 93.5 \text{ in}^2 \cdot \left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right)^{\mathbf{2}} = 93.5 \cancel{\text{in}^2} \cdot \frac{(2.54)^2 \text{ cm}^2}{1^2 \cancel{\text{in}^2}} = \mathbf{603 \text{ cm}^2}$$

3. WANTED: ? km^2 (in conversions, use exponents for squared, cubed)

DATA: 1.61 km = 1 mile

1 section = 1 mile^2 = 640 acres (any two equal terms can be used as a conversion)

40.0 acres (the single unit to use as your *given*)

* * * * *

SOLVE: $\mathbf{?}\text{ km}^2 = 40.0 \text{ acres} \cdot \frac{1 \text{ mile}^2}{640 \text{ acres}} \cdot \left(\frac{1.61 \text{ km}}{1 \text{ mile}}\right)^{\mathbf{2}} = \frac{40}{640} \text{ mi}^2 \cdot \frac{2.59 \text{ km}^2}{1 \text{ mi}^2} = \mathbf{0.162 \text{ km}^2}$

Practice B **(**Other conversions than those below can be used if they arrive at the same answer.)

1. $? \text{ km}^3 = 5.00 \text{ miles}^3 \cdot \left(\frac{1.61 \text{ km}}{1 \text{ mile}}\right)^{\mathbf{3}} = 5.00 \cancel{\text{mi}^3} \cdot \frac{4.17 \text{ km}^3}{\cancel{1 \text{ mi}^3}} = \mathbf{20.9 \text{ km}^3}$

2. $? \text{ mm}^3 = 1 \text{ meter}^3 \cdot \left(\frac{1 \text{ }\mathbf{m}\text{m}}{10^{\mathbf{-3}} \text{ meter}}\right)^{\mathbf{3}} = 1 \text{ meter}^3 \cdot \frac{1^3 \text{ mm}^3}{10^{-9} \text{ meter}^3} = \mathbf{1 \times 10^9 \text{ mm}^3}$

3. $? \text{ in}^3 = 1.00 \text{ mm}^3 \cdot \left(\frac{1 \text{ inch}}{25.4 \text{ mm}}\right)^{\mathbf{3}} = 1.00 \text{ mm}^3 \cdot \frac{1^3 \text{ in}^3}{(25.4)^3 \text{ mm}^3} = \mathbf{6.10 \times 10^{-5} \text{ in}^3}$

4. a. $? \text{ cm}^3 = 0.355 \text{ L} \cdot \frac{1{,}000 \text{ cm}^3}{1 \text{ L}} = \mathbf{355 \text{ cm}^3}$ (metric fundamentals)

b. $? \text{ ft}^3 = 355 \text{ cm}^3 \cdot \left(\frac{1 \text{ inch}}{2.54 \text{ cm}}\right)^{\mathbf{3}} \cdot \left(\frac{1 \text{ foot}}{12 \text{ in}}\right)^{\mathbf{3}} = 355 \text{ cm}^3 \cdot \frac{1^3 \text{ in}^3}{(2.54)^3 \text{ cm}^3} \cdot \frac{1^3 \text{ ft}^3}{(12)^3 \text{ in}^3} = \mathbf{0.0125 \text{ ft}^3}$

5. $? \text{ dm}^3 = 67.6 \text{ fl. oz.} \cdot \frac{355 \text{ mL}}{12.0 \text{ fl oz.}} \cdot \frac{10^{-3} \text{ L}}{1 \text{ mL}} \cdot \frac{1 \text{ dm}^3}{1 \text{ L}} = \mathbf{2.00 \text{ dm}^3}$

6. $? \text{ cc's} = ? \text{ cm}^3 = 74 \text{ in}^3 \cdot \left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right)^{\mathbf{3}} = 74 \text{ in}^3 \cdot \frac{(2.54)^3 \text{ cm}^3}{1^3 \text{ in}^3} = 1{,}200 \text{ cm}^3 = \mathbf{1{,}200 \text{ cc's}}$

7a. $? \text{ ft}^3 = 1.68 \times 10^5 \text{ m}^3 \cdot \left(\frac{3.28 \text{ ft}}{1 \text{ meter}}\right)^{\mathbf{3}} = 1.68 \times 10^5 \text{ m}^3 \cdot \frac{(3.28)^3 \text{ ft}^3}{(1)^3 \text{ m}^3} = \mathbf{5.93 \times 10^6 \text{ ft}^3}$

7b. Hint: 1 m = 10 dm , 1 dm^3 = 1 liter

* * * * *

$? \text{ gallons} = 1.68 \times 10^5 \text{ m}^3 \cdot \left(\frac{10 \text{ dm}}{1 \text{ meter}}\right)^{\mathbf{3}} \cdot \frac{\mathbf{1 \text{ L}}}{\mathbf{1 \text{ dm}^3}} \cdot \frac{1 \text{ gal}}{3.79 \text{ L}} = \frac{1.68}{3.79} \times 10^8 \text{ gal.} = \mathbf{4.43 \times 10^7 \text{ gallons}}$

8. WANTED: ? in^3 displacement

DATA: 6.70 L displacement

1 inch = 2.54 cm (metric-English bridge)

* * * * *

Strategy: This problem includes numbers you don't need. Since a displacement is wanted, start with a displacement as your *given*, then head for the *cm* needed in the metric part of the metric/English bridge conversion.

* * * * *

SOLVE: $? \text{ in}^3 = 6.70 \text{ L} \cdot \frac{1{,}000 \text{ cm}^3}{1 \text{ L}} \cdot \left(\frac{1 \text{ in}}{2.54 \text{ cm}}\right)^3 = 6{,}700 \text{ cm}^3 \cdot \frac{1 \text{ in}^3}{(2.54)^3 \text{ cm}^3} = \mathbf{409 \text{ in}^3}$

* * * * *

Lesson 5G: Density and Solving Equations

Timing: This lesson should be done *if* you are assigned textbook problems on the *density* of substances that are in the shape of geometric objects such as spheres, cylinders, or rectangular solids.

Pretest: If you think you know this topic, try the last problem in the lesson. If you can do that problem, you may skip the lesson.

* * * * *

Solving Problems Using Mathematical Equations

Calculations in chemistry can generally be solved using conversions, mathematical equations, or both.

Conversions can be used for problems in which all of the relationships can be expressed as two quantities that are equal or equivalent. Equations are required for more complex relationships. In these lessons, when we study gas laws and energy, we will discuss in detail the circumstances in which equations must be used.

Many problems can be solved with either conversions or equations. Conversion methods usually involve less memorization, less algebra, and fewer steps. For most of the early topics in first-year chemistry courses, conversions are the easier way to solve.

An exception is problems involving the density of substances that are in geometric shapes. To calculate volumes, these problems require mathematical equations. (In these lessons, we will call mathematical formulas *equations*, and reserve the term *formula* for chemical formulas.)

Volumes for regular geometric shapes are calculated using equations, including

- Volume of a cube = $(\text{side})^3$
- Volume of a rectangular solid = $l \times w \times h$
- Volume of a cylinder = $\pi r^2 h$
- Volume of a sphere = $4/3\, \pi r^3$

Density is defined as mass per unit of *volume*. In equation form: $\boxed{D = m/V}$.

Because density is the *ratio* between mass and volume, it can be used as a conversion factor. Some calculations involving density may be solved using either conversions or the density equation, but in many density problems, equations are required to calculate the volume of a geometric shapes such as a cylinder or a sphere. If an equation is used for one part, by using the D = m/V equation for the other part, the same equation-solving *method* can be used to solve both parts of the problem.

In a density problem that requires a geometric volume calculation, both the density equation and the geometric volume equations include *volume* as one of the terms. If we can solve for volume in one equation, we can use that volume to solve for quantities in the other equation.

In general, if a problem involves *two* equations linked by a common quantity, a useful method to solve is to

- list the equations and DATA for the two equations in separate columns.
- Find the value of the *linked* quantity in the column with one missing variable instead of two (usually the column that does *not* include the WANTED quantity), then
- Add the value of the linked quantity to the other column and solve for the WANTED quantity.

Let us learn this method by example.

Q. If aluminum (Al) has a density of 2.7 g/cm^3, and a 10.8 gram Al cylinder has a diameter of 0.60 cm, what is the height of the cylinder? ($V_{cylinder} = \pi r^2h$)

Do the following steps in your notebook.

1. First, read the problem and write the *answer* unit. WANTED = ? *unit* and label.
2. To use conversions, at this point we would list the problem's numbers and units, most of them in equalities. However, *if* you see a mathematical *equation* is needed to solve the problem, write that *equation* in your DATA instead, and draw a $\boxed{\text{box}}$ around it. Then, under the equation, list each *symbol* in the equation, followed by an = sign.
3. If *two* equations are needed to solve the problem, write and $\boxed{\text{box}}$ the two equations in two separate columns. Under each equation, write each symbol in that equation.
4. Usually, one symbol will be the same in both equations. Circle that *linked* symbol in the DATA in both columns. That symbol will have the same *value* in both columns.

Finish those steps and then check your answer below.

* * * * *

At this point, your paper should look like this.

WANTED:	**?** cm height Al cylinder =	
DATA:	$\boxed{V_{cylinder} = \pi r^2 h}$	$\boxed{\textbf{D}\text{ensity} = \textbf{m}\text{ass}/\textbf{V}\text{olume}}$
	V =	D =
	r =	m =
	h =	V =

Next, do the following steps.

5. Write "= **?** WANTED" after the symbol that is WANTED in the problem.
6. Transfer the problem data to the DATA table. After each symbol in the DATA, write the number and unit in the problem that corresponds to that symbol. Use the *units* of the numbers to match up the symbols: grams is mass, mL or cm^3 is volume, etc.
7. After any remaining symbol that does *not* have DATA in the problem, write a **?**.

After you have finished those steps, check your answer below.

* * * * *

Your DATA table should look like this.

DATA:	$\boxed{V_{cylinder} = \pi r^2 h}$	$\boxed{\textbf{D}\text{ensity} = \textbf{m}\text{ass}/\textbf{V}\text{olume}}$
	V = **?**	D = 2.7 g/cm^3
	r = 1/2 diameter = 0.30 cm	m = 10.8 grams
	h = **? WANTED**	V = **?**

8. A fundamental rule of algebra: if you know values for all of the symbols in a mathematical equation except one , you can solve for that missing symbol. If you are missing values for two symbols, you cannot solve for those values directly.

   In the above data, column 1 has two missing values, and column 2 has one. At this point, you can solve for the missing value only in column 2.

   In a problem involving two relationships, usually you will need to solve *first* for the common, linked symbol in the column with*out* the WANTED symbol. Then, use that answer to solve for the WANTED symbol in the other column.

9. When solving an *equation*, solve in symbols before you plug in numbers. In algebra, symbols move faster than numbers with units.

   Solve for the *missing* column 2 data, and then check your answer below.

   * * * * *

   SOLVE: (In column 2, D = m/V ; and we want V. Solve the D equation for V in symbols, then plug in the numbers for those symbols from the DATA.)

   $\boxed{\mathbf{D = m/V}}$

   WANTED = $\mathbf{V = \frac{m}{D}} = \frac{10.8 \text{ g}}{2.7 \text{ g/cm}^3} = 4.0 \text{ cm}^3$

   (In the unit cancellation, 1/(1/X) = X. See Lesson 17C.)

10. Put this solved answer in the DATA. Since the problem is about one specific cylinder, the volume of that cylinder must be the same in both columns. Write your calculated volume in *both* columns.
11. Now solve the equation that contains the WANTED symbol for the WANTED symbol. First solve using the symbols, then plug in the numbers and their units.

    EQUATION: $\boxed{V_{cyl.} = \pi r^2 h}$; so

WANTED = height = $\mathbf{h} = \frac{\mathbf{V}}{\pi\,\mathbf{r^2}} = \frac{4.0\text{ cm}^3}{\pi\,(0.30\text{ cm})^2} = \frac{4.0\text{ cm}^3}{\pi\,(0.090\text{ cm}^2)} = \boxed{\textbf{14 cm height}}$

SUMMARY: Steps for Solving With Equations

1. First write WANTED = ? and the unit you are looking for.
2. When you see that you need a mathematical equation to solve, under DATA, write and $\boxed{\text{box}}$ the equation.
3. If you need two equations, write them in separate columns.
4. Under each equation, list each symbol in that equation.
5. Write "**?** WANTED" after the WANTED symbol in the problem.
6. After each symbol, write numbers and units from the problem. Use the units to match the numbers and units with the appropriate symbol.
7. Label remaining symbols without DATA with a **?**
8. Circle symbols for variables that are the same in both equations.
9. Solve equations in symbols before plugging in numbers.
10. Solve for ? in the column with *one ?* first.
11. Write that answer in the DATA for both columns, then solve for the WANTED symbol.

Flashcards: Using the table below, cover the answer column, then put a check by the questions in the left column you can answer quickly and automatically. For the others, make flashcards using the method in Lesson 2C.

One-way cards (with notch at top right): Back Side -- Answers

Density =	Mass/Volume
Volume of a cube =	$(\text{side})^3$
Volume of a sphere =	$4/3\,\pi\,r^3$
Volume of a cylinder =	$\pi\,r^2\,h$

Practice: Practice any needed flashcards above, then try two of the problems below. Save one problem for your next study session.

Use the steps for solving with equations above. Answers are at the end of this lesson. If you get stuck, read a part of the answer, and then try again.

1. What is the density of liquid water?
2. If the density of lead is 11.3 grams per cubic centimeter, what is the mass of a ball of lead that is 9.0 cm in diameter?
3. A gold American Eagle $50 coin has a diameter of 3.26 cm and mass of 36.7 grams. Assuming that the coin is in the approximate shape of a cylinder and is made of gold alloy (density = 15.5 g/mL), find the height of the cylinder (the thickness of the coin).

4. If a solid copper cube with the length on a side of 1.80 cm has a mass of 52.1 grams, what is the density of the copper, in grams per cubic centimeter?

ANSWERS

1. **WANTED**: mass/volume ratio for liquid water. Hint: What's the definition of one gram?

* * * * *

1.00 g (mass) liquid water = 1 mL (volume) , so mass/volume = 1.00 g / 1 mL = **1.00 g/mL**

2. **WANTED**: ? grams lead

DATA:

$\boxed{V_{sphere} = 4/3\pi\ r^3}$	$\boxed{\textbf{D}\text{ensity} = \textbf{m}\text{ass}/\textbf{V}\text{olume}}$
V = **?**	D = 11.3 g/cm^3
r = 1/2 diameter = 4.5 cm	m = **? WANTED**
	V = **?**

Strategy: First solve for the ? in the column with **one ?**. Then use that answer to solve for the variable that is WANTED in the other column.

SOLVE: Column 1 has one **?**, and column 2 has two. Solve column one first.

$$\textbf{?} = \textbf{V} = 4/3\ \pi\ r^3 = 4/3\ \pi\ (4.5\text{ cm})^3 = \mathbf{382\ cm^3}$$

In problems that solve in steps, carry an extra *sf* until the final step.

Add this answer to the *volume* DATA in *both* columns. Then solve the Column 2 equation for the WANTED mass. First solve in symbols, then plug in the numbers.

If needed, adjust your work, then finish.

* * * * *

$\boxed{D = \mathbf{m}/V}$ and **m**ass is WANTED,

$$\text{WANTED} = \mathbf{m} = D \bullet V = 11.3\ \frac{g}{cm^3} \bullet 382\ cm^3 = \mathbf{4.3 \times 10^3\ grams}\quad (2\ sf)$$

Units must be included and must cancel to give the WANTED unit.

Use the *sf* in the original data to determine the *sf* in the final answer.

You can also solve the column 2 data for grams using conversion factors.

$$?\ g = 382\ cm^3 \bullet \frac{11.3\ g}{1\ cm^3} = 4.3 \times 10^3\ g$$

3. (Hint: You will need 1 mL = 1 cm^3)

* * * * *

WANTED: **? cm** height of gold cylinder (thickness of coin)

DATA:

$\boxed{V_{cylinder} = \pi\ r^2\ h}$	$\boxed{D = \textbf{m}\text{ass}/\textbf{V}\text{olume}}$
V = **?**	D = 15.5 g/mL
r = 1/2 diameter = 1.63 cm	m = 36.7 grams

h = **? WANTED** V = **?**

Strategy: First complete the column with one **?**, then use that answer to solve for the variable WANTED in the other column. Column 1 has two **?** and column 2 has one.

SOLVE: $\boxed{D = m/V}$;

$$\textbf{WANTED} = V = \frac{m}{D} = \frac{36.7\text{ g}}{15.5\text{ g/mL}} = \mathbf{2.36\mathit{8}\text{ mL}}$$ (Carry extra sig fig until end)

(For help with the unit cancellation in equations, see Lesson 17C.)

Fill in that **V**olume in both columns. Then solve the equation that contains the WANTED symbol, first in symbols, and then with numbers.

EQUATION: $\boxed{V = \pi r^2 h}$

$$\textbf{WANTED} = \text{height} = \mathbf{h} = \frac{\mathbf{V}}{\pi\,\mathbf{r^2}} = \frac{2.36\mathit{8}\text{ mL}}{\pi\,(1.63\text{ cm})^2} = \frac{2.368\ \mathbf{cm^3}}{8.347\ \mathbf{cm^2}} = \boxed{\mathbf{0.284\text{ cm}}}$$

Note carefully the unit cancellation above. By changing mL to cm^3 (they are identical), the base units are *consistent*. They then cancel properly.

A *height* of a cylinder, or *thickness* of a coin, must be in *distance* units such as cm.

Your work must include unit s, and your answers must include correct units to be correct.

4. WANTED: $\frac{\textbf{? grams}}{cm^3}$ copper cube =

DATA: 52.1 grams copper

Side of cube = 1.80 cm

Strategy: This one is tricky because you are not told that you need to calculate volume. Note, however, that you WANT grams per *cubic* cm. You are given grams and cm. In density problems, be on the lookout for a volume calculation.

The equation for the volume of a cube is $V_{cube} = (\text{side})^3$.

If you needed that hint, adjust your work and try the question again.

* * * * *

DATA: $\boxed{V_{cube} = (\text{Side})^3}$ $\boxed{D = \mathbf{m}\text{ass}/\mathbf{V}\text{olume}}$

V = **?** D = **? WANTED**

side = 1.80 cm m = 52.1 g copper

V = **?**

SOLVE: First solve the column with *one* **?** then put that answer in both columns.

$$\text{Volume of cube} = (\text{side})^3 = (1.80\text{ cm})^3 = \boxed{5.83\mathit{2}\ \mathbf{cm^3}}$$

Now solve for the WANTED symbol in the other equation.

$$D = \textbf{? WANTED} = \frac{\text{mass}}{\text{volume}} = \frac{52.1\text{ g Cu}}{5.83\mathit{2}\text{ cm}^3} = 8.93\ \frac{\text{g Cu}}{\text{cm}^3}$$

* * * * *

Summary: Word Problems

1. To solve word problems, *get rid of the words*.
2. Organize your work into 3 parts: WANTED, DATA, and SOLVE.
3. First, under WANTED, write the unit you are looking for. As a part of the unit, include a label that describes what the unit is measuring.
4. If a *ratio* unit is WANTED, write the unit as a fraction with a top and a bottom.
5. Under DATA, to solve with conversions,
   - write every number in the problem. Attach the units to the numbers. If the problem involves more than one substance, add a label to the unit and number that identifies which substance is being measured.
   - If numbers and units are paired with other numbers and units, write those DATA terms in an equality.
   - Write *per* or a slash (**/**) in the data as **=** . Treat "• unit$^{-\#}$" as "**/** unit$^{\#}$".
     If no number is given after the *per* or /, write **= 1** .
   - Write as equalities two different measurements of the same entity, or any units and labels that are equivalent or mathematically related in the problem.
6. To SOLVE, *start* each calculation with an *equality*:

   ? WANTED unit = **#** *given* unit.

   If you WANT a single unit, start with a single number and unit as your *given*, use the format of the *single-unit starting template*

   $$? \text{ unit WANTED} = \# \text{ and } \textbf{\textit{unit}}\ \textit{given} \bullet \frac{\qquad\qquad}{\textbf{\textit{unit}}\ \textit{given}}$$

   and chain conversions to solve.
7. Any distance to distance equality or conversion can be *squared* and used as an area conversion, or *cubed* and used as a volume conversion.
8. For problems that require mathematical equations to solve,
   - write and $\boxed{\text{box}}$ the equations in your DATA.
   - List each symbol in the equation below the equation.
   - Match the data in the problem to the symbols.
   - Solve in symbols before plugging in numbers.
9. For problems requiring two equations to solve, solve the two equations separately. Solve for the linked variable in the non-WANTED column first. Use that answer as DATA to solve for the WANTED symbol in the other column.

# # # # #

Module 6 – Atoms, Ions, and Periodicity

Prerequisites: None. This module may be started at any point.

Each lesson in this module has a pretest. If you pass the pretest, you may skip the lesson. Module 6 covers fundamentals. Depending on your background, you may be able to skip several lessons or complete them very quickly.

To do this module, you will need an alphabetical list of the atoms (at the end of these lessons) and a periodic table that closely resembles the type of table you will be allowed to consult during quizzes and tests in your course.

* * * * *

Lesson 6A: Atoms

Pretest: Using a list of atoms or a periodic table, try problem 6 at the end of this lesson. If you find problem 6 easy, you may skip to Lesson 6B.

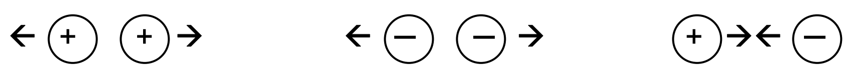
* * * * *

Terms and Definitions

The precise definition for some of the fundamental particles in chemistry is a matter of occasional debate, but following simplified and somewhat arbitrary definitions will provide us with a starting point for discussing atoms.

1. **Matter**. Chemistry is primarily concerned with the measurement and description of the properties of matter and energy. Matter is anything that has mass and volume. In planetary environments, nearly all matter is composed of extremely small particles called atoms. A substance's identity depends on the atoms that make up the substance and their arrangement in space.
2. **Electrical charges.** Some particles have a property known as electric charge.

   There are two types of charges, positive and negative. Particles with like electrical charges repel. Unlike charges attract.

3. **Atoms.** In these lessons, we will define an atom as a particle with a single nucleus, plus the electrons that surround the nucleus.

   There are 91 different kinds of atoms that are found in the Earth's crust. More than 20 additional atoms have been synthesized by scientists using nuclear reactions. All of the millions of different substances on earth are consist of only about 100 different kinds of atoms. It is how the atoms are grouped and arranged in space that results in so many different substances.

   A list of the atoms is found at the end of these lessons. Each atom is represented by a one- or two-letter **symbol**. The first letter of the symbol is always capitalized. The second letter, if any, is always lower case.

4. **Atomic structure.** Atoms can be described as combinations of three **subatomic particles:** protons, neutrons, and electrons.

   a. **Protons (symbol p^+)**

      ➢ Each proton has a **1+** electrical charge (one unit of positive charge).

      o Protons have a mass of about 1.0 amu (**atomic mass units**).

      o Protons are found in the center of the atom, called the nucleus. The nucleus is extremely small and occupies very little volume in the atom.

      ➢ The number of protons in an atom is defined as the **atomic number** (symbol Z) of the atom.

      ➢ The number of protons determines the ***name*** (and thus the symbol) of the atom.

      ➢ The number of protons in an atom is never changed by *chemical* reactions.

   b. **Neutrons (symbol n^0)**

      o A neutron has an electrical charge of zero.

      o A neutron has about the same mass as a proton, 1.0 amu.

      o Neutrons are located in the nucleus of an atom, along with the protons.

      o Neutrons are thought to act as the glue of the nucleus: the particles that keep the repelling protons from flying apart.

      o Neutrons, like protons, are never gained or lost in chemical reactions.

      o The neutrons in an atom in most cases have very little influence on the chemical behavior of the atom.

   c. **Electrons (symbol e^-)**

      ➢ Each electron has a **1—** electrical charge : one unit of negative charge, equal in magnitude but opposite the proton's charge.

      o Electrons have very little mass, weighing about 1/1837th amu.

      o Electrons are found outside the nucleus of an atom, in regions of space called **orbitals**.

      o Nearly all of the *volume* of an atom is due to the space occupied by the electrons around the nucleus.

      ➢ Electrons are the *only* subatomic particles that can be gained or lost during chemical reactions.

   Each of the above points will be addressed at various times in your course. For this lesson, the items above identified by the ➢ symbol must be *memorized*.

5. **Neutral atoms.** If an atom has an equal number of protons and electrons, the balance between positive and negative charges gives the atom a *net* charge of zero. The charges are said to "cancel" to produce an overall **electrically neutral** atom.

Practice A

Commit to memory the points in Section 4 above labeled with a ➤ . Then, for the problems below, consult the alphabetical list of atoms at the end of these lessons. Apply the rules listed above from memory. Check answers at the end of the lesson.

1. Write the symbols for these atoms.

   a. Carbon b. Oxygen c. Osmium d. Tungsten

2. Name the atoms represented by these symbols.

   a. K b. F c. Fe d. Pb

3. Assume each atom below is electrically neutral. Fill in the blanks.

Atom Name	Symbol	Protons	Electrons	Atomic Number
Sodium				
	N			
		6		
			82	
				9

More Terms and Definitions

6. **Chemical reactions** cannot create or destroy atoms, nor change an atom from one kind to another. However, during a chemical reaction, how atoms are bonded and arranged changes, and this alters the identity and the behaviors of the substances involved in the reaction.
7. **Physical changes.** When a substance undergoes a **physical change**, it does not change its identity. Melting ice to water is a physical change, because both ice and liquid water are composed of particles that internally have the same atoms in the same geometry. A physical change is not considered to be a chemical reaction.
8. **Ions.** During chemical reactions, the number of protons and neutrons in an atom never changes, but atoms can gain or lose one or more electrons. Any particle (atom or group of bonded atoms) that does not have an equal number of protons and electrons is termed an **ion**, which is a particle with a net electrical charge.

   - Neutral particles that *lose electrons* become **positive ions**.
   - Neutral particles that *gain electrons* become **negative ions.**

   The symbol or chemical formula for a particle that is not electrically neutral places the value of the net charge as a superscript to the right of the symbol.

   An ion is *not* the same as a neutral particle with the same atom or atoms. The ion has a different number of electrons and different chemical behavior.

Examples of atoms and ions

a. All atoms with a single nucleus containing 16 protons are examples of **sulfur** (symbol **S**).

   An atom of sulfur in its elemental state has 16 protons and 16 electrons. The positive and negative charges balance to give a net charge of zero. The symbol for the neutral sulfur atom is written as **S**. The symbol $\mathbf{S^0}$ may also be written to emphasize that the sulfur atom has a neutral charge.

   In substances, some sulfur atoms may be found that contain 16 protons and 18 electrons. This atom is an *ion* of sulfur. Although the 16 protons cancel the charge of 16 electrons, the two un-cancelled electrons result in an overall charge of **–2**. The symbol for this particle is $\mathbf{S^{2-}}$ and it is named a **sulfide ion**.

b. All atoms with **19** protons are named **potassium** (symbol **K**).

   In nature, potassium is always found with 18 electrons. The 18 electrons balance the charge of 18 protons. This leaves one positive charge un-cancelled, so the ion has a net charge of **+1**. This ion form of potassium is symbolized as $\mathbf{K^+}$. For the charges on ions, if no number after the sign is shown, a **1** is understood.

c. All atoms with 88 protons are named **radium** (symbol **Ra**).

   Ra^{2+} ions must have 2 more positive protons than negative electrons. Because all radium atoms must have 88 protons, an Ra^{2+} ion must have how many electrons?

   * * * * *

   86 electrons.

Practice B

For the problems below, use the alphabetical list of atoms at the end of these lessons.

1. Calcium has atomic number 20.
   a. A neutral Ca atom has how many protons? How many electrons?
   b. How many protons and electrons are found in a Ca^{2+} ion?

2. In their nucleus, during chemical reactions, atoms always keep a constant number of

   ______________s, which have a positive charge. Atoms take on a charge and become

   ions by gaining or losing ______________s, which have a ______________charge.

3. In terms of subatomic particles, an atom that is a positive ion will always have more

   ________________ than __________________.

4. For these symbols, write the atom names from memory.

   a. Sr = ______________ b. Si = ______________ c. P = ______________

5. For these atom names, write their symbols from memory.

   a. Bromine = ______ b. Boron = _______ c. Barium = ______________

6. For the particles composed of single atoms at the right, fill in the blanks.

Symbol	Protons	Electrons
O		
O^{2-}		
Mg^{2+}		
I^{-}		
	79	79
	1	0
	35	36
	34	36
	13	10

ANSWERS

Part A

1a. **C** 1b. **O** 1c. **Os** 1d. **W** 2a. **Potassium** 2b. **Fluorine** 2c. **Iron** 2d. **Lead**

3.

Atom Name	Symbol	Protons	Electrons	Atomic Number
sodium	**Na**	**11**	**11**	**11**
nitrogen	N	**7**	**7**	**7**
carbon	**C**	6	**6**	**6**
lead	**Pb**	**82**	82	**82**
fluorine	**F**	**9**	**9**	9

Part B

1. a. **20 protons, 20 electrons.** b. **20 protons, 18 electrons**
2. In their nucleus, during chemical reactions, atoms always keep a constant number of **PROTONS** , which have a positive charge. Atoms take on a charge, to become ions , by gaining or losing **ELECTRONS** , which have a **NEGATIVE** charge.
3. In terms of sub-atomic particles, an atom that is a positive ion will always have more **PROTONS** than **ELECTRONS** .

4a. Sr = **Strontium** b. Si = **Silicon** c. P = **Phosphorus**

5a. Bromine = **Br** b. Boron = **B** c. Barium = **Ba**

6.

Symbol	Protons	Electrons
O	**8**	**8**
O^{2-}	**8**	**10**
Mg^{2+}	**12**	**10**
I^{-}	**53**	**54**
Au	79	79
H^{+}	1	0
Br^{-}	35	36
Se^{2-}	34	36
Al^{3+}	13	10

* * * * *

Lesson 6B: The Nucleus, Isotopes, and Atomic Mass

Pretest: If you think you know this topic, try 2-3 parts of each practice set. If you can do those correctly, skip the lesson.

* * * * *

The Nucleus

At the center of an atom is the nucleus. The nucleus contains all of the protons and neutrons in the atom.

The nucleus is very small, with a diameter that is roughly 100,000 times smaller than the effective diameter of most atoms, yet the nucleus contains all of the atom's positive charge, and nearly all of its mass.

Because the nucleus contains nearly all of the atom's mass in a tiny volume, it is extremely dense. Outside of the nucleus, nearly all of the volume of an atom is occupied by its electrons. Because electrons have low mass but occupy a large volume compared to the nucleus, the region occupied by the electrons has a very low density. In terms of mass, an atom is mostly empty space.

However, an electron has a charge that is equal in magnitude (though opposite) to that of the much more massive proton. It is the *charges* of the subatomic particles inside an atom, rather than their masses, that play the major role in determining an atom's size and chemical behavior.

Types of Nuclei

Only certain combinations of protons and neutrons form a nucleus that is stable. In a nuclear reaction (such as radioactive decay or in a nuclear reactor), if a combination of protons and neutrons is formed that is unstable, the nucleus will decay.

The combinations of protons and neutrons found in nuclei can be divided into three types.

- **Stable:** Stable nuclei are combinations of protons and neutrons that do not change in a planetary environment such as Earth over many billions of years.
- **Radioactive:** Radioactive nuclei are *somewhat* stable. Once formed, they can exist for a time on Earth (from a few seconds to several billion years), but they fall apart (**decay**) at a constant, characteristic rate.
- **Unstable:** In nuclear reactions, if combinations of protons and neutrons form that are unstable, they decay within a few seconds.

Nuclei that exist in the earth's crust include all of the stable nuclei plus some radioactive nuclei.

All atoms that have between one and 82 protons (except technetium with 43 protons) have at least one stable nucleus. Atoms with 83 to 92 protons can be found in the earth's crust, but all are radioactive. Atoms with 93 or more protons exist on earth only when they are created in nuclear reactions (such as in nuclear reactors).

Radioactive atoms comprise a very small percentage of the matter found on earth. Over 99.99% of the earth's atoms have nuclei that are stable. The nuclei in those stable atoms have not changed since the atoms came together to form the earth billions of years ago.

Terminology

Protons and neutrons are termed the **nucleons** because they are found in the nucleus. A combination of a certain number of protons and neutrons is called a **nuclide**. A group of nuclides that have the same number of protons (so they are all the same atom) but differing numbers of neutrons are called the **isotopes** of the atom.

Stable Nuclei

Some atoms have only one stable nuclide; other atoms have as many as 10 stable isotopes.

Example: All atoms with 17 protons are called chlorine. Only two chlorine nuclei are stable: those with

- 17 protons and **18** neutrons, and
- 17 protons and **20** neutrons.

Nuclei that have 17 protons and *other* numbers of neutrons can be made in nuclear reactions, but in all of those combinations, within a few seconds, the nucleus decays by emitting a particle from the nucleus.

Nuclide Symbols

Each nuclide can be assigned a **mass number** which is the *sum* of its number of protons and neutrons.

$$\boxed{\textbf{Mass Number} \text{ of a nucleus } = \mathbf{p^{+} + n^{0}} = \textbf{Protons + Neutrons}}$$

Example: A nucleus with **2** $\mathbf{p^{+}}$ and **2** $\mathbf{n^{0}}$ is helium with a mass number of **4**.

A nuclide can be identified in two ways,

- either by its number of protons and number of neutrons,
- or by its **nuclide symbol** (also termed its **isotope symbol**).

A nuclide symbol has two required parts: the *atom symbol* and the *mass number*. The mass number is written as a superscript in front of the atom symbol.

For example, the two stable isotopes of chlorine can be represented as

- 17 protons + 18 neutrons *or as* $\mathbf{^{35}Cl}$ (a nuclide named chlorine-35); and
- 17 protons + 20 neutrons *or as* $\mathbf{^{37}Cl}$ (named chlorine-37).

Knowing one representation for the composition of a nucleus, you need to be able to write the other.

Using a table of atoms, atom symbols, and atomic numbers that can be found at the end of these lessons, try these questions.

Q1. A nuclide with 6 protons and 8 neutrons would have what nuclide symbol?

* * * * *

A1. Atoms with 6 protons are always named carbon, symbol **C**. The mass number of this nuclide is 6 protons + 8 neutrons = **14** . This isotope of carbon, used in "radiocarbon dating," is carbon-14, and its symbol is written $\mathbf{^{14}C}$ **.**

Try another.

Q2. How many protons and neutrons would be found in $\mathbf{^{20}Ne}$?

* * * * *

A2. All atoms called neon contain 10 protons. The mass number 20 is the total number of protons *plus* neutrons, so neon-20 contains **10 protons and 10 neutrons.**

Nuclide symbols may also include the **nuclear charge** of the particle written in front of and below the atom symbol. Including both the mass number and the nuclear charge is called the **A-Z notation** for a nuclide, illustrated at the right. **A** is the symbol for mass number and **Z** is the symbol for nuclear charge. In atoms, Z is also the number of *protons* in the nucleus.

$$\boxed{^{37}_{17}\mathrm{Cl}}$$

Nuclide symbols can also be used to identify sub-atomic particles (particles smaller than atoms), and in those cases the nuclear charge may be zero or negative. Including the Z values is helpful when balancing nuclear reactions (a future topic), but the Z values are not required to identify an atom, since the Z repeats what the symbol identifies: the number of protons in the nucleus of the atom.

Practice A: Consult a table of atoms or periodic table to fill in the blanks below.

1.

Protons	Neutrons	Atomic Number	Mass Number	Nuclide Symbol	Nuclide Name
	6	6			
7	7				
					Iodine-131
				^{235}U	
		2	4		

2. Which nuclides in Problem 1 must be radioactive? Why?

The Mass of Nuclides

The mass of a single nuclide is usually measured in **atomic mass units**, abbreviated **amu**. For use in conversions, one amu = 1.66 x 10^{-24} grams.

Protons and neutrons have essentially the same mass, and both are much heavier than electrons. The mass of

- a proton is **1.0** amu,
- a neutron is **1.0** amu, and
- an electron is 1/1837th of an amu.

Based on those masses, you might expect that the mass of a $^{35}\mathbf{Cl}$ atom would be just over 35.0 amu, since it is composed of 17 protons, 18 neutrons, and 18 electrons. In fact, for neutral atoms of $^{35}\mathbf{Cl}$, the actual mass is 34.97 amu, slightly *lighter* than the combined mass of its protons, neutrons, and electrons.

Why do the masses of the three subatomic particles *not* add exactly to the mass of the atom? When protons and neutrons combine to form nuclei, a small amount of mass is either converted to, or created from, energy. This change is the relationship postulated by Einstein:

Energy gained or lost = mass lost or gained times the speed of light squared
Which in equation form is written: $\mathbf{E = mc^2}$

In nuclear reactions, if a small amount of mass is lost, a very large amount of energy is created. In forming nuclei, however, because the gain or loss in mass is relatively small, the mass of a nuclide or atom in amu's will *approximately* (but not exactly) *equal* its mass number.

The sum of the mass numbers of a nuclide *roughly* equals its mass in amu .

The Average Mass of Atoms (Atomic Mass)

In chemical reactions, the isotopes of an atom behave the same. In addition, for *most* atoms (those not formed by radioactive decay), one kind of atom may have several stable isotopes, but in visible-sized samples of that atom found in substances on earth, the percentage of each isotope will be the same.

For these reasons, when dealing with visible amounts of most atoms, atoms not formed by radioactive decay on earth (over 99.9% of all matter on earth) will have the same *average mass* in any matter found on earth.

This average mass of an atom, called its **atomic mass,** can be calculated from the **weighted average** of the mass of its isotopes.

For example, in all samples of chlorine, the ratio of the nuclides is very close to 3 nuclides with a mass of 35.0 amu for every one with a mass of 37.0 amu. The average mass is therefore the average of

(35.0 + 35.0 + 35.0 + 37.0) amu . Find the average mass of these particles.

* * * * *

Average = (Sum of values)/(number of values) =

* * * * *

= (35.0 + 35.0 + 35.0 + 37.0) amu/4 = 142/4 = **35.5 amu**

Precise Calculation of Weighted Averages

For most atoms, the characteristic atomic mass cannot be calculated precisely using the method above, because the ratios between the isotopes are not as simple and exact as the "3 to one" ratio that is very close to true for chlorine. However, for all atoms, if we know a precise mass of the isotopes and the relative abundance of the isotopes (the percent that is each isotope in samples of the atoms), we calculate a precise atomic mass using this general formula for the weighted average:

$$\boxed{(1.00)(\text{average value for mixture}) = (\text{fraction})_1(\text{value})_1 + (\text{fraction})_2(\text{value})_2 + \ldots} \quad (1)$$

In this equation, the **fraction** for a component in a mixture can be calculated by dividing, for any uniform sample of a mixture, the number of particles that are the component by the total number of particles in the mixture. In mathematical terms, this means

- $\boxed{\text{fraction = part/total}}$, which will be a number less than one (**0.**XXX…);
- the sum of the fractions in the mixture must add up to 1.00 .

If a *percentage* is known, the fraction is simply the percentage divided by 100.

For example, if the percentage is 20%, the fraction is 0.20 .

(Dividing by 100 moves the decimal to the left, twice.)

Another way to write equation (1) above is

$$\boxed{(1.00)(\text{average value for mixture}) = \sum (\text{fraction for component})(\text{value of component})} \quad (2)$$

where the symbol $\sum$ represents a summation.

Precise Calculation of Atomic Mass

Since the atomic mass of an atom is the weighted average of the atomic masses of its isotopes, the equation for atomic mass can be written as

atomic mass of atom = $\sum$ (isotope fraction)(isotope mass)

or as

atomic mass = (isotope fraction)$_1$(isotope mass)$_1$ + (isotope fraction)$_2$(isotope mass)$_2$ + **…**

Let's apply this formula to chlorine atoms again, but this time using measurements that are more precise than in the example above.

Q. All samples of chlorine atoms in the earth's crust contain

- **75.78%** atoms that have ^{35}**Cl** nuclei with a mass of 34.97 amu; and
- **24.22%** atoms that have ^{37}**Cl** nuclei with a mass of 36.97 amu.

a. What fraction of chlorine atoms are ^{37}Cl ?

b. Calculate the atomic mass of chlorine atoms.

* * * * *

a. The fraction is the percentage divided by 100. ^{37}Cl fraction = 0.2422

Try part b.

* * * * *

The atomic mass of an atom is a weighted average. Substituting into the equation,

atomic mass Cl = (0.7578)(34.97 amu) + (__________)(____________) = _________ amu

Fill in the blanks to complete the calculation above, then check your answer by looking up the atomic mass of chlorine either online or in the table of atoms inside the cover of your chemistry text.

* * * * *

atomic mass Cl = (0.7578)(34.97 amu) + (**0.2422**)(**36.97 amu**) = ___________ amu

* * * * *

= 26.5$\underline{0}$*0* amu + 8.9$\underline{4}$*6* amu = 35.4$\underline{5}$*4* = 35.45 amu = average mass for a chlorine atom

(*SF*: carry extra sf until the final step; when adding, round to highest *place* with doubt.)

No single atom of chlorine will have this average mass, but in visible amounts of substances containing chlorine, the chlorine atoms have this *average* mass. Use of this average mass (atomic mass) will simplify calculations involving mass.

The numeric value for the atomic mass in amu that is found in tables is also the mass of the atom in "grams per mole." The number 6.022 x 10^{23}, called one mole, is a value that will simplify the math when converting between grams of a substance and its number of particles.

Practice B

1. Silver has two stable isotopes: ^{107}Ag (106.91 amu) and ^{109}Ag (108.90 amu). Assuming that 51.8% of naturally occurring silver is silver-107,
   a. calculate the atomic mass of Ag.
   b. Compare your answer to the value listed for silver in the table at the end of these lessons.
   c. What would be the atomic mass of Ag in grams per mole?

Isotopes and Chemistry

The rules and the reactions for "standard chemistry" are very different from those of *nuclear* chemistry. For example,

- chemical reactions can release substantial amounts of energy, such as seen in the burning of fuels or in conventional explosives. Nuclear reactions, however, can involve *much* larger amounts of energy, as in stars or nuclear weapons.
- An important rule in chemical reactions is that atoms can neither be created nor destroyed. In nuclear reactions, atoms are often created and destroyed.

Because the rules are very different, a clear distinction must be made between *chemistry* and *nuclear chemistry*. By convention, it is assumed that the rules that are cited as part of "chemistry" refer to processes that do *not* involve changes in nuclei (unless *nuclear* chemistry is specified). Processes that change the composition of the nucleus are termed *nuclear* reactions which by definition are *not* chemical reactions.

The good news is that, except for experiments in nuclear chemistry, because all isotopes of an atom nearly always have the same chemical behavior, and because in visible amounts of substances, the atoms of a given atom have the same average mass, we can ignore the fact that atoms have isotopes as we investigate nearly all *chemical* reactions and processes.

We will return to the differences among isotopes when we consider nuclear chemistry, which includes reactions such as radioactive decay, fission, and fusion.

Practice C

Fill in the blanks below.

Protons	Neutrons	Electrons	Atomic Number	Mass Number	Nuclide Symbol	Ion Symbol
	144	88	90			
	148					Pu^{2+}
		78			^{206}Pb	
	0					H^{+}

					^{3}H	H^{-}
		36			^{90}Sr	
11		10		23		
15	16	18				

ANSWERS

Practice A

1.

Protons	Neutrons	Atomic Number	Mass Number	Nuclide Symbol	Nuclide Name
6	6	6	**12**	^{12}C	**Carbon-12**
7	7	**7**	**14**	^{14}N	**Nitrogen-14**
53	**78**	**53**	**131**	^{131}I	Iodine-131
92	**143**	**92**	**235**	^{235}U	**Uranium-235**
2	**2**	2	4	^{4}He	**Helium-4**

2. Uranium must be radioactive, because no nuclei with more than 82 protons are stable.

Practice B

1a. Since there are only two Ag isotopes, ^{109}Ag must be 48.2%.

(0.518)(106.91 amu) + (0.482)(108.90 amu) = (55.3*8* + 52.4*9*) amu = 107.8*7* = **107.9 amu**

1b. It should match. 1c. **107.9 g/mole** (value for amu = value for g/mole)

Practice C

Protons	Neutrons	Electrons	Atomic Number	Mass Number	Nuclide Symbol	Ion Symbol
90	144	88	90		^{234}Th	Th^{2+}
94	148	**92**	**94**	**242**	^{242}Pu	Pu^{2+}
82	**124**	78	**82**	**206**	^{206}Pb	Pb^{4+}
1	0	**0**	**1**	**1**	^{1}H	H^{+}
1	**2**	**2**	**1**	**3**	^{3}H	H^{-}
38	**52**	36	**38**	**90**	^{90}Sr	Sr^{2+}
11	**12**	10	**11**	23	^{23}Na	Na^{+}
15	16	18	**15**	**31**	^{31}P	P^{3-}

Lesson 6C: Elements, Compounds, and Formulas

Prerequisites: Lesson 6A.

Pretest: Use the list of atoms on the last page of these lessons or in a textbook. With a perfect pretest score, skip to Lesson 6D. Answers are at the end of the lesson.

1. In this list:

   A. H_2O B. Cl_2 C. Au D. S_8 E. CO_2 F. Co G. H_2SO_4

   a. Which formulas represent elements?

   b. Which formulas represent a substance without ionic or covalent bonds?

   c. Which formulas represent substances that are diatomic?

2. Write the number of oxygen atoms present in each of these compounds.

   a. $Co(OH)_2$ b. CH_3COOH c. $Al_2(SO_4)_3$

3. Write the total number of atoms in each of the compounds in question 2.

* * * * *

Substances

The definitions below are general and highly simplified, but they will give us a starting point for discussing how atoms may combine to yield different substances.

1. A **substance** contains *one* kind of chemical particle. In a substance, all of the neutral units have the same number and kind of atoms, chemically bonded in the same manner and geometry. **Chemical formulas** can be used to represent a substance. A **mixture** is a combination of two or more substances.

   Substances have *characteristic* properties: their melting points, color, and densities are some of the properties that will be the same for a substance no matter what steps are taken to form the substance. These properties can help in identifying the substance.

2. In a substance, if the smallest particles that can be separated from each other relatively easily are neutral particles with two or more atoms, the particles are called **molecules**. If a substance consists of charged particles that can separate from each other when dissolved in water, the separated particles are called ions, and the smallest electrically neutral combination of ions is called a **formula unit.**

3. **Elements** are substances that contain only one kind of atom. Each atom has an **elemental state:** a substance that contains only that atom and is in the form that is the most stable at room temperature and pressure.

   The basic particles for some elements, termed the **monatomic elements,** are individual atoms. The chemical formulas for monatomic elements are written as one instance of the atom's formula, reflecting the fact that the basic unit is a single atom.

   <u>For example</u>, the basic particles of the **noble gases** (helium, neon, argon, krypton, xenon, and radon) are single atoms. Therefore, the formulas for these elements are written as **He** for helium, **Ne** for neon, etc.

Other elements are found in our environment as neutral particles consisting of two or more atoms chemically bonded to form a new larger unit.

Ne, Cl_2, and S_8 are all formulas for elements because they are substances that contain only one kind of atom, and those formulas represent the most stable form in which a collection of those atoms will exist in at normal room temperature and pressure.

4. **Bonds** are forces that hold particles together. Molecules of the **diatomic elements** consist of two atoms (*di-* means two), and their chemical formulas reflect the fact that each unit contains 2 atoms. In chemical formulas, a **subscript** written after a symbol represents the number of that kind of atom or ion that is bonded within the particle.

   For example, the elemental forms of oxygen, nitrogen, and chlorine are all diatomic. The chemical formula for chlorine is Cl_2, nitrogen is N_2, and oxygen is O_2.

   Polyatomic elements are neutral molecules that contain 2 or more atoms, but only one kind of atom.

   For example, the elemental formula for sulfur is S_8, indicating that it exists as eight atoms bonded together.

   Over 70% of the elements found in the earth's crust are metals. Metals have a more complex nature than simple monatomic or polyatomic elements, but metal formulas are represented by single atoms, such as **Ag** for silver, and **Al** for aluminum.

   Some elements have multiple forms that are stable at room temperature. The elemental forms of carbon, for example, include graphite, diamond, and fullerenes (carbon molecules shaped like soccer balls). Each of these has very different bonding and properties, but all are composed entirely of carbon atoms. As an element, carbon is usually represented by the simplified monatomic formula **C**.

5. A **compound** is a substance that consists of two or more different atoms chemically bonded together to form new substance. While there are just over 100 elements, there are millions of known compounds. In a given compound, the ratio of the atoms is always the same, which is reflected in their formulas. H_2O, NaCl, and H_2SO_4 are all formulas for compounds, because they contain two or more different atoms. Compounds can be classified as either ionic or covalent, depending on the type of bonding present.

6. The basic particles for **covalent compounds** (also known as molecular compounds) are molecules. Molecules are held together by **covalent bonds**. In a covalent bond, electrons are shared between two neighboring atoms. Covalent bonds can be single bonds (involving 2 shared electrons), double bonds (4 shared electrons), or triple bonds (6 shared electrons). Covalent bonds hold atoms at predictable angles within the molecule.

7. **Molecular formulas** use atomic symbols and subscripts to represent the number and kind of atoms covalently bonded together to form a single molecule.

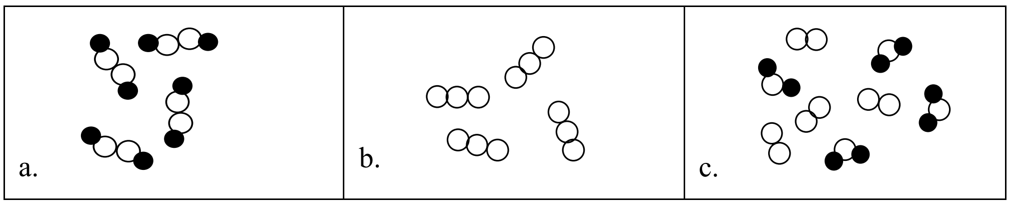
   - Water is composed of molecules that each consist of two hydrogen atoms and one oxygen atom, represented by the molecular formula $\mathbf{H_2O}$.

In chemical formulas, when there is no subscript is written after a symbol, the subscript is understood to be *one*.

- Carbon dioxide is composed of molecules that each consist of two oxygen atoms and one carbon atom. Its molecular formula is written as $\mathbf{CO_2}$.

Practice A: Use the atoms table at the end of these lessons or in a textbook to answer these questions. Answers are at the end of the lesson.

1. Identify each sample sketched below as an *element*, *compound*, or *mixture*. Different atoms are indicated by different shades, and individual particles are separated for clarity.



a. ________________ b.__________________ c. _______________

2. Label the following formulas as representing elements or compounds.

a. Ne b. H_2O c. NaCl d. S_8 e. $C_6H_{12}O_6$

3. Which of these formulas contain chemical bonds?

a. H_2 b. CO c. NH_3 d. He

4. In problems 2 and 3, which formulas represent

a. Diatomic elements? b. Monatomic elements? c. 4 atoms?

More Vocabulary

8. **Structural formulas** can be used to represent chemical particles that are held together by covalent bonds. These formulas show each of the atoms present along with information about their positions within the particle.

O — H
/
H

At the left is a structural formula for water. It shows that the oxygen atom is found in the middle of the molecule, and that water has two directional covalent bonds and a *bent* shape.

The structural formula for carbon dioxide, CO_2 , is **O=C=O** . Carbon dioxide has two double bonds, and the molecule is linear in shape with the carbon atom in the middle.

We generally write structural formulas when knowing the *shape* of the molecule is important, but we write the more compact molecular formulas when it is not.

9. Often, chemical formulas are written as a mixture of different types of formulas. For example,

   - ethyl alcohol can be written as CH_3CH_2OH or as C_2H_6O. The shorter formula, however, is also the molecular formula of dimethyl ether, which is usually written CH_3OCH_3 to show that the O is found in the middle in the ether, rather than toward one end as in the alcohol.

     Ethyl alcohol and dimethyl ether have the same number and kind of atoms, but the differing atomic arrangements give the molecules very different properties. To predict chemical behavior, we often need to know a formula with structural information. In such cases, we write the longer formulas like those above.

10. **Ionic compounds** are substances consisting of a collection of positive and negative **ions** (particles with a net electrical charge). Ions can be **monatomic** (single atoms with an unequal number of protons and electrons) or **polyatomic** (a group of covalently bonded atoms that have an unequal number of protons and electrons). An **ionic bond** is the electrostatic attraction between oppositely charged ions.

11. I**onic formulas** use atomic symbols and numbers to represent the ratio and kind of ions present in an ionic compound. The ions in an ionic compound are always present in a ratio that guarantees overall electrical neutrality.

    A **formula unit** is defined as the smallest combination of ions for which the sum of the electrical charges is zero. Parentheses are used to indicate more than 1 polyatomic ion. Chemical formulas for ionic compounds show the atom ratios in a single neutral formula unit.

    - Table salt consists of a 1:1 ratio of positively charged sodium ions (formula Na^+) and negatively charged chloride ions (Cl^-). The formal name of table salt is sodium chloride, and its ionic formula is written as **NaCl**. The formula unit NaCl contains 2 ions.

    - Calcium phosphate is an ionic compound composed of three monatomic Ca^{+2} ions for every two polyatomic PO_4^{-3} ions. The ionic formula is $\mathbf{Ca_3(PO_4)_2}$, and 1 formula unit represents a total of 5 ions.

    - Copper(II) nitrate is an ionic compound composed of one monatomic Cu^{+2} ion for every two polyatomic NO_3^- ions. The ionic formula is written as $\mathbf{Cu(NO_3)_2}$. Writing the formula as CuN_2O_6 shows the atom ratios, but indicating that the compound contains two NO_3^- groups better conveys the structure and behavior of this compound.

12. Be careful to *write* formulas so that you can distinguish between upper- and lower-case letter combinations such as CS and Cs, Co and CO, NO and No.

    - $Co(OH)_2$ has 1 cobalt atom, 2 oxygen atoms, and 2 hydrogen atoms.

    - CH_3COOH has 2 carbon, 4 hydrogen, and 2 oxygen atoms.

<u>To summarize</u>, although molecules of covalent substances and formula units of ionic compounds have different types of bonds, all compound formulas refer to a single, overall electrically neutral unit of a substance.

<u>Practice B:</u> Use the table of atoms at the end of these lessons or in a textbook to answer these questions.

1. Write the number of oxygen atoms in each of these compounds.

   a. $Al(OH)_3$ b. C_2H_5COOH c. $Co_3(PO_4)_2$

2. Write the total number of atoms for each of the compounds in question 1.

3. Try every third one of these, then check your answers. Need more practice? Do a few more. Name each atom, and write the total number of those atoms, in each of the following chemical formulas.

   a. HCOOH b. $CoSO_4$ c. $No(NO_3)_3$

   d. $Ca_3(PO_4)_2$ e. $(NH_4)_3PO_4$ f. $Pb(C_2H_5)_4$

4. If you need additional practice, redo the pretest at the beginning of Lesson 6C.

ANSWERS

<u>Pretest</u>: 1a. B, C, D, F 1b. C, F 1c. B 2a. 2 2b. 2 2c. 12 3a. 5 3b. 8 3c. 17

<u>Practice A:</u> 1. a. Compound b. Element c. Mixture

2. 2a and 2d are elements because they have one kind of atom. The rest are compounds because they have more than one kind of atom.
3. 3a, 3b, and 3c have bonds, because they have more than one atom. It takes bonds to hold two or more atoms together in particles.
4. 3a, H_2, is the only diatomic element. 5. 2a and 3d are the only monatomic elements.
6. 3d, NH_3 is the only formula with 4 atoms.

<u>Practice B</u>

1a. 3 oxygen atoms 1b. 2 1c. 8 2a. 7 total atoms 2b. 11 2c. 13

3. a. 2 hydrogen
1 carbon
2 oxygen

   b. 1 cobalt
1 sulfur
4 oxygen

   c. 1 nobelium
3 nitrogen
9 oxygen

d.	3 calcium 2 phosphorus 8 oxygen	e.	3 nitrogen 12 hydrogen 1 phosphorus 4 oxygen	f.	1 lead 8 carbon 20 hydrogen

* * * * *

Lesson 6D: The Periodic Table

Pretest: If you think you know this topic, try the last letter of each numbered question at the end of this lesson. If you get those right, you may skip this lesson.

* * * * *

Patterns of Chemical Behavior

Learning the behavior of over 100 different atoms would be a formidable task. Fortunately, the atoms can be organized into **families**. The behavior of one atom in a family will help to predict the chemical behavior other atoms in the family.

The grouping of atoms into families results in the **periodic table.** To build the table, the atoms are arranged in *rows* across (also called **periods**) in order of the number of protons in each atom. This order usually, but not always, matches the order of the increasing atomic mass of the atoms.

At certain points, the chemical properties of the atoms begin to repeat, somewhat like the octaves on a musical scale.

In the periodic table, under most graphic designs, when a noble gas atom is reached, it marks the end of a row. The next atom, with one more proton, starts a new row of the table. This convention places the atoms into vertical **columns** (called **families** or **groups**) with the noble gases in the last column on the right.

Within each *column,* the atoms tend to have similar chemical behavior.

Some Families in the Periodic Table

The noble gases (He, Ne, Ar, Kr, Xe, Rn) are monatomic (composed of single atoms) as elements. Though they can be liquefied by lowering temperature and/or increasing pressure, in their elemental state at room temperature and pressure, all are gases.

These atoms are termed noble because they are chemically "content" with their status as single atoms: these atoms rarely bond with other atoms or each other.

Although the noble gases take part in very few chemical reactions, they are important in predicting chemical behavior. Other atoms tend to react in ways that give them the same electron configuration as the nearest noble gas. The outer electrons of atoms are often said to react to attain a "cloak of nobility."

The alkali metals (Li, Na, K, Rb, Cs, Fr) are in **column one** (also called **group 1A**) of the periodic table, at the far left. As elements, all are soft, shiny metals that tend to react with many substances, including the water vapor present in air.

In chemical reactions, neutral alkali metal atoms tend to *lose* an electron to become a *+1 ion*. This ion has the same number of electrons as the noble gas that has one fewer protons.

Once an alkali metal atom forms a +1 ion, it becomes quite stable. Most chemical reactions do not change its +1 charge.

<u>The halogens</u> (F, Cl, Br, I, and At) are in **column 7** (group 7A) just to the left of the noble gas column. As neutral elements at room temperature, halogen atoms are stable only when the atoms bond to form diatomic molecules; F_2, Cl_2, Br_2, I_2, and At_2.

Like alkali metals, the halogens are very reactive. In reactions, neutral halogen atoms tend to *gain* one electron to become a **halide ion** with a **1—** charge. This ion has the same number of electrons as the noble gas just to the right in the periodic table.

Halogen atoms can also share electrons with neutral atoms of other nonmetals. Shared electrons result in a covalent bond. Including the shared electrons, each neutral halogen atom will tend to be surrounded by the same number of electrons as the nearest noble gas.

<u>Hydrogen</u> is often placed in column one of the table, and the reactions of hydrogen are often like those of the alkali metals. However, other hydrogen reactions are like those of the halogens. Hydrogen is probably best portrayed as a unique family of one that can have characteristics of both alkali metals and halogens.

<u>The main group atoms</u> are those found in the *tall* columns, termed either groups 1, 2, and 13 to 18, or groups 1A, 2A, and 3A-8A, depending on the version of the periodic table that you are using.

<u>The transition metals</u> are in the "middle dip" of the periodic table, in groups 3-12 or the "B" groups. There are 10 atoms in each row of the transition metals.

<u>The inner transition atoms</u> include the 14 **lanthanides** (or **rare earth metals)** in the 6th row and the and 14 **actinides** in the 7th row. These atoms are usually listed below the rest of the periodic table in order to display a table that fits easily on a chart or page.

* * * * *

<u>Predicting Behavior</u>

The following table summarizes the *general characteristics* of the atoms in the columns of the periodic table. The positions of the column numbers, family names, and likely ion charges should be memorized.

Group	1A	2A	3B → 2B or 3 → 12	3A	4A	5A	6A	7A	8A
Family Name	Alkali Metals		Transition Metals					Halogens	Noble Gases
Monatomic ion charge	1 +	2 +		3 + (or 1+)		3 —	2 —	1 —	None

<u>For example:</u> Cesium (Cs) is in column *one* of the periodic table. Based on this placement, it can be predicted to

- behave like other alkali metal atoms; and
- exist as a Cs^+ ion in compounds.

Practice A: Use a copy of the periodic table and your memorized knowledge about the table (first learn the rules, then do the practice) to answer these.

1. Describe the location in the periodic table of the
   a. noble gases

   b. alkali metals

   c. halogens

   d. transition metals

2. Add a charge to these symbols to show the ion that a single atom of these elements tends to form.

a. Br	b. Ra	c. Cs	d. In	e. Te
(35)	(88)	(55)	(49)	(52)

Metals, Metalloids, and Nonmetals

The elements in the periodic table can be divided into metals, metalloids (also called semimetals), and nonmetals.

Metalloids

Many periodic tables include a thick line, like a staircase, as shown in the section of the periodic table below. This line separates the metal and nonmetal atoms.

The six atoms bordering the line in **bold** below are the **metalloids**. They have chemical behavior that is *in-between* that of the metals and the nonmetals.

Unless you are allowed to use a periodic table that has the staircase and identifies the metalloids on tests, you should memorize the location of the staircase and the 6 metalloids.

If you memorize how the staircase looks at boron (**B**), the rest of the staircase is easy. Remembering "11220" will help with the number of metalloids per row going down the table. (Some textbooks include polonium (Po) as a metalloid, others do not.)

					(H)	He
	B	C	N	O	F	Ne
		Si	P	S	Cl	Ar
		Ge	**As**	Se	Br	Kr
			Sb	**Te**	I	Xe
				(Po)	At	Rn

Nonmetals

At the right are the 18 nonmetals. The nonmetals must be *memorized*:
H, C, N, O, P, S, Se, plus the 5 halogens and 6 noble gases.

Note the shape of their positions. All nonmetals are all to the right of the staircase and to the right of the metalloids. All atoms in the last two columns are nonmetals.

			(H)	He
C	**N**	**O**	F	Ne
	P	**S**	Cl	Ar
		Se	Br	Kr
			I	Xe
			At	Rn

Note also that hydrogen, although it is often shown in column one, is considered to be a *non*metal. Hydrogen has unique properties, but it most often behaves as a nonmetal.

Metals

The metals are all of the atoms to the left of the thick line and the six metalloids. The metals include all of the transition metals, as well as all of the inner transition (rare earth) atoms usually listed below the rest of the chart.

In their electrically neutral, elemental form, metal atoms all behave as metals: all are substances that are shiny and conduct electricity. But neutral metal atoms tend to react to form positive ions in compounds. Metal ions do not look or behave like metals.

Of the over 100 elements, over 75 percent are metals. To learn the atoms that are metals, memorize the 6 metalloids and 18 nonmetals. All of the remaining elements are metals.

Practice B: Use a copy of the periodic table and your memorized knowledge about the columns of the table to answer these.

1. How many atoms are non-metals?
2. Without consulting a periodic table, add the metal/nonmetal dividing line to the portion of the periodic table at the right, then circle the metalloid atoms.

					(H)	He
	B	C	N	O	F	Ne
	Al	Si	P	S	Cl	Ar
Zn	Ga	Ge	As	Se	Br	Kr
Cd	In	Sn	Sb	Te	I	Xe
Hg	Tl	Pb	Bi	Po	At	Rn

ANSWERS

Practice A: 1a. Noble gases -- last column 1b. Alkali metals – column one 1c. Halogens – Group 7A (tall column 7), just before the noble gases. 1d. Transition metals – the 10 columns in the middle dip.

2 a. Br^- b. Ra^{2+} c. Cs^+ d. In^{3+} e. Te^{2-}

Practice B: 1. 18 2. See table in lesson.

Lesson 6E: A Flashcard Review System

Previous Flashcards

At this point, you may have a sizeable stack of flashcards, and we will soon add more. Before going further, let's organize the cards. Try this system.

A. Separate your existing flashcards into 4 stacks.

1-Daily: Those you have not yet "practiced until correct" for 3 days.

2-End of Chapter/Quiz: Those you have done for *more* than 3 days. Run again before your next quiz on this material.

3-Test: Those you have done 4 or more times. Run again before starting the practice problems for your next major test.

4-Final Exam Review: Those you have retired until the final.

B. Add cards with those 4 *labels* to the top of each stack. Rubber-band each stack.

You may want to carry the *daily* pack with you for practice during down time.

Module 6 Flashcards

If you have had a previous course in chemistry, you may recall much of the material in Module 6 after a brief review. Other points may be less familiar, and the material in Module 6 will need to be *firmly* in memory for the rest of the course.

For points that are not firmly in memory, make the flashcards. Use the method in Lesson 2C on the sample cards below: cover the answers, put a check next to those which you can answer correctly and quickly. Make a flashcard if the answer is not automatic.

Run your new cards for several days in a row. Run the two-way cards in both directions. Run the cards again before your next quiz, test, and final exam.

For Lesson 6A

One-way cards (with notch at top right): Back Side -- Answers

Like charges	Repel
Unlike Charges	Attract
The particles in a nucleus =	protons and neutrons
Subatomic particle with lowest mass	electron
Subatomic particles with charge	protons and electrons
Mass of a proton in amu	1.0 amu
Mass of a proton in grams/mole	1.0 grams/mole
Protons minus electrons =	Charge on particle
Number of protons determines	Atom name, symbol, and atomic number
Particles gained and lost in chemical reactions	electrons

Zero charge on an atom means	# protons = # electrons
Negative ions have	More electrons than protons
Subatomic particles with mass of 1.0 amu	protons and neutrons

Two-way cards (with*out* notch):

ion	A particle with electrical charge
Protons plus Neutrons =	Mass Number =

For Lesson 6B

One-way cards (with notch) — Back Side -- Answers

To calculate the average atomic mass of an atom, use	$\sum$ (isotope fraction)(isotope mass)
Same # of p^+, different #'s of n^0	isotopes
Different nuclides with same chemical behavior =	isotopes

Two-way cards (with*out* notch):

1 proton and 1 neutron = ? nuclide symbol	^2H = contains what particles?
1 proton and 0 neutrons = ? nuclide symbol	^1H = contains what particles?
1 proton and 2 neutrons = ? nuclide symbol	^3H = contains what particles?
Protons plus neutrons approximately equals	Mass of nuclide in amu approx. equals

For Lesson 6C

Two-way cards (with*out* notch):

Define a Substance	All particles have same chemical formula
A Mixture	2 or more substances
Molecule	Neutral, independent particles with two or more atoms
Molecular Formula	Shows the atoms inside a neutral particle
Structural Formula	Shows atoms and positions in a particle
Elements	Stable neutral substances with one kind of atom
Compounds	Stable neutral particles with more than one kind of atom
Bonds	Forces holding atoms together

For Lesson 6D

One-way cards (with notch) — Back Side -- Answers

One-way cards (with notch)	Back Side -- Answers
Family that rarely bonds to other atoms	noble gases
Lightest non-metal	Hydrogen (H)
Lightest metalloid	Boron (B)
Number of non-metal elements	18

Two-way cards (with*out* notch):

Position of *alkali metals*	First column, below hydrogen
Position of *halogens*	Next-to-last column
Position of *noble gases*	Last column
Position of *rare earths (lanthanides)*	First row below body of table
Position of *transition metals*	In dip between tall columns 2 and 3
Tend to form —1 ions	Ions formed by *halogen* atoms
Family forms +1 ions	Ions formed by *alkali metals*
Family forms +2 ions	Ions formed by *Column 2 atoms*
Name for halogen atoms with a —1 charge	Halide ions

* * * * *

Lesson 6F: Atoms Project – Part 4

The following frequently encountered atoms have symbols based on their Latin names. Test to see if these are in memory in both directions. If they are not, add them to your cards.

Two-way cards (with*out* notch):

copper	Cu
tin	Sn
mercury	Hg
gold	Au
potassium	K

Two-way cards (with*out* notch):

iron	Fe
lead	Pb
silver	Ag
sodium	Na
antimony	Sb

# # # # #

Module 7 – Writing Names and Formulas

Prerequisites: Complete Module 6 before starting Module 7. The other prior modules are not necessary for Module 7.

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Lesson 7A: Naming Elements and Covalent Compounds

Pretest: If you think you know this topic, try the last letter of each question in Practice A and Practice B. If you get those right, skip the lesson.

* * * * *

Systems for Naming Substances

Chemical substances are identified by both a unique name and a chemical formula. For names and formulas that both identify and differentiate substances, a *system* for writing formulas and names is required.

1. Some compounds have names that are **non-systematic** but familiar: Water (H_2O) and ammonia (NH_3) are examples.
2. Historically, chemical substances have been divided into two broad categories. Compounds containing carbon and hydrogen are studied in **organic chemistry**, which has its own system for naming compounds. All other substances are part of **inorganic chemistry**, which is the focus of most first-year courses.
3. Different types of inorganic substances have different naming systems. We will begin with the rules for naming elements, ions, and binary covalent compounds.

Naming Elements

An element is stable, electrically neutral substance that contains of only one kind of atom. The **name** of an element is simply the name of its **atoms**.

Examples

- The element comprised of neutral atoms with 20 protons is called **calcium**. Calcium is a metal, and the formulas of metals are written as if they are monatomic elements. The formula for the element calcium is therefore written as **Ca**.
- Neutral chlorine atoms, at room temperature, are stable when they exist in diatomic molecules. For the element chlorine, the formula is written $\mathbf{Cl_2}$.
- At room temperature, sulfur atoms tend to form molecules with 8 bonded atoms. The formula for the elemental form of sulfur is written $\mathbf{S_8}$.

Note that for elements, the formula easily distinguishes between monatomic, diatomic, or polyatomic structures, but the name does not. This is only an issue for a few of the elements, but for the millions of chemical compounds, a more systematic **nomenclature** (naming system) is needed.

Compounds

In compounds, there is more than one kind of atom, but all the neutral molecules or formula units have the same atoms and structure. Most compounds can be classified as either **ionic** or **covalent**.

Covalent compounds are molecules containing atoms bonded together by electrons shared between atoms. The attractive forces (bonds) within molecules are strong compared to the attractions between molecules. Compounds that are gases or liquids at room temperature are nearly always covalent compounds.

At room temperature, compounds that are solids may be ionic or covalent, but ionic compounds are always solids. Ionic compounds are composed of an array of ions bonded strongly by electrostatic attraction.

Ionic and covalent compounds have different naming systems. To name a compound we must first identify it as ionic or covalent. To make that distinction, we must identify the types of bonds in the compound.

Types of Bonds

1. In **covalent bonds**, electrons are *shared* between two atoms.
2. In **ionic bonds,** an atom (or group of atoms) has lost one or more electrons (compared to its electrically neutral form), and another atom (or group of atoms) has gained one or more electrons. The loss and gain of electrons results in charged particles (ions). The ions are bonded by the attraction of their opposite charges.
3. The following rules will predict whether bonds are ionic or covalent in *most* cases.
   - A bond between two *non*metal atoms is usually a *covalent* bond.
   - A bond between a *metal* and a *non*metal atom is usually an *ionic* bond.

To identify the type of bond, begin by asking: are both atoms non-metals? If so, the bond is covalent.

The non-metals are shown at the right. Recall that hydrogen is classified as a nonmetal, and that all atoms in the last two columns are nonmetals.

The six noble gases rarely bond. The remaining 12 nonmetal atoms nearly always form covalent bonds when they bond with each other.

			(**H**)	He
C	N	O	F	Ne
	P	S	Cl	Ar
		Se	Br	Kr
			I	Xe
			At	Rn

Is one of the atoms in the bond a metal and the other a non-metal? If so, the bond is ionic.

Using those rules and a periodic table, answer these questions.

Q. Predict whether the following bonds will likely be ionic or covalent.

1. C—H   2. C—Na   3. N—Cl   4. K—Cl

* * * * *

Answers

1. C─H Both are non-metals, so is predicted to be a covalent bond
2. C─Na A non-metal and a metal atom; predict ionic bond
3. N─Cl Both are non-metals; predict covalent bond
4. K─Cl A metal and non-metal; predict ionic bond.

* * * * *

Types of Compounds

1. If a compound contains *all* covalent bonds, it is classified as a **covalent compound.**
2. If a compound has *one* or more ionic bonds, even if it also has many covalent bonds, it will tend to have ionic behavior and is classified as an **ionic compound.**

These rules mean that in most cases,

- a compound with *all non*metal atoms is a *covalent compound*.
- a compound that combines *metal* and *non*metal atoms is an *ionic compound*.

The above general rules do not cover all types of bonds and compounds, and there are many exceptions. However, they will give us a starting point for both naming compounds and writing formulas that indicate the composition and behavior of compounds.

Q. Using those rules and a periodic table, label these compounds as ionic or covalent.

1. NaCl 2. CH_4 3. Cl_2 4. HCl

* * * * *

Answers

1. NaCl Na is a metal, Cl is non-metal, compound is ionic.
2. CH_4 Both atoms are non-metals; compound is covalent.
3. Cl_2 Both atoms are non-metals; compound is covalent.
4. HCl Both atoms are non-metals; compound is covalent.

Covalent Compounds

The 12 nonmetals that tend to bond are a small percentage of the more than 100 atoms. However, because

- covalent bonds tend to be strong,
- the nonmetal atoms are relatively abundant on our planet, and
- the molecules in living systems are based on a nonmetal (carbon),

a substantial percentage of the compounds studied in chemistry are covalent compounds.

Practice A

For the problems below, use the type of periodic table that you are permitted to view on tests in your course. You should not need to consult the metal versus nonmetal charts found in these lessons, since they should be committed to memory.

1. Label these bonds as ionic or covalent.

   a. Na—I   b. C—Cl   c. S—O   d. Ca—F   e. C—H   f. K—Br

2. Label these compounds as ionic or covalent.

   a. CF_4 b. KCl   c. CaH_2 d. H_2O e. NF_3 f. $NaCH_3O$

3. Which two families of elements are all non-metals?

Naming Binary Covalent Compounds

Binary covalent compounds contain *two* different nonmetals. The naming of binary compounds uses the atom name or the *root* of the atom name. Binary covalent compounds that include *hydrogen* are usually given "common names" such as methane, water, and ammonia, or follow special rules for acid compounds.

For the 11 remaining non-metals that bond, the roots are C=carb-, N=nitr-, O=ox-, F=fluor-, P=phosph-, S=sulf-, C=chlor-, Se=selen-, Br=brom-, I=iod-, and At=astat-. Not all of those roots are "regular," but their use will become intuitive with practice.

For compounds composed of two different *non*metal atoms, the rules for naming are:

1. The name contains two words. The format is *prefix-atomname* then *prefix-root-suffix* .

   Example: The name of N_2Cl_4 is dinitrogen tetrachloride.

2. This rule takes precedence over the rules below. For covalent compounds that contain
   - O atoms, the second word is prefix-*oxide*.
   - H atoms, the compound usually has a name that does not follow these rules.
3. The *first* word contains the name of the *atom* (of the two atom symbols in the formula) that is in a column farther to the *left* in the periodic table. If the two atoms are in the same column, the *lower* atom is named first.
4. The second word contains the *root* of the second atom name, with the suffix *–ide* added.
5. The *number* of atoms of each kind is represented by a Greek prefix**.**

   mono- = 1 atom. (In the first word, *mono-* is left off and assumed if no prefix is given. *Mono-* is included if it applies to the second word.)

di- = 2 atoms	*penta-* = 5 atoms	*octa-* = 8 atoms
tri- = 3 atoms	*hexa-* = 6 atoms	*nona-* = 9 atoms
tetra- = 4 atoms	*hepta-* = 7 atoms	*deca-* = 10 atoms

If an *o* or *a* at the end of a prefix is followed by a first letter of an atom or root that is a vowel, the *o* or *a* in the prefix is *sometimes* omitted (both inclusion and omission of the *o* and *a* are allowed, and you may see such names both ways).

Using a periodic table and the above rules, try the following.

Q1. What is the name of CS_2?

* * * * * (the * * * mean *cover* the answer below, *write* your answer, then check it.)

A1. Carbon is in the column farther to the left in the periodic table, so *carbon* is the atom in the first word. For one atom, the prefix would be *mono-*, but *mono-* is omitted if it applies to the first word. The name's first word is simply *carbon*.

For the root of the second word, sulfur becomes sulf*ide*. Since there are two sulfur atoms, the name of the compound is **carbon disulfide**.

Q2. What is the name of the combination of four fluorine and two nitrogen atoms?

* * * * *

A2. Nitrogen is in the column more to the left in the periodic table, so the first word contains nitrogen. Since there are two nitrogen atoms, add the prefix *di-*. For the second word, the root fluorine becomes fluoride, and the prefix for four atoms is *tetra-*. The name for the compound is **dinitrogen tetrafluoride**.

Flashcards

Cover the answers below, then check those which you can answer correctly and quickly. When done, make flashcards for the others (see the steps in Lesson 2C).

Run the new cards for several days in a row, then add them to the previous flashcards for quiz and test review.

One-way cards (with notch)	Back Side -- Answers
The formula for elemental oxygen	O_2
A bond between a metal and nonmetal is	Usually ionic
A bond between two nonmetals is	Usually covalent
A covalent compound has	Shared electrons and only covalent bonds
An ionic compound has	One or more ionic bonds
A compound with all nonmetal atoms is usually	A covalent compound
Compounds with one or more metal atoms are	Ionic compounds

Two-way cards (with*out* notch):

Formula for ammonia = ?	Name of NH_3 = ?
Formula for carbon monoxide = ?	Name of CO = ?
Formula for dinitrogen tetrachloride = ?	Name of N_2Cl_4 = ?

Practice B

Learn the rules, practice needed flashcards, then try every *other* problem. Wait a day, run the cards again, then try the remaining problems.

1. Write the name for these combinations of nonmetals.

   a. Three chlorine plus one nitrogen. b. One sulfur and six fluorine.

   c. Two sulfurs and one silicon. d. Three chlorine and one iodine.

   e. One oxygen and two chlorines. f. One bromine and one iodine

2. Name these covalent compounds.

   a. SCl_2 b. PI_3 c. SO_2 d. NO

3. Nonmetals often form several stable oxide combinations, including the combinations below. Name that compound!

   a. Five oxygen and two nitrogen b. 10 oxygen and four phosphorus

   c. NO_2 d. N_2O e. SO_3 f. Cl_2O_7

ANSWERS

Practice A

1. a. Na—I **Ionic** b. C—Cl **Covalent** c. S—O **Covalent**
   d. Ca—F **Ionic** e. C—H **Covalent** f. K—Br **Ionic**
2. a. CF_4 **Covalent** b. KCl **Ionic** c. CaH_2 **Ionic**
   d. H_2O **Covalent** e. NF_3 **Covalent** f. CH_3ONa **Ionic**

   (All of the ionic compounds contain a metal atom.) 3. Halogens and noble gases.

Practice B

1. a. Nitrogen is to the left, so it is the first word in the name. When the first word refers to a single atom, the prefix is omitted. For the second word, chlorine becomes chloride, and the prefix tri- is added. The name is **nitrogen trichloride.**

   b. Sulfur hexafluoride. c. Silicon disulfide d. Iodine trichloride (if in same column, name lower first) e. Dichlorine monoxide (oxygen is always last, drop last *o* in mono-) f. Iodine monobromide

2. a. Sulfur dichloride b. Phosphorus triiodide c. Sulfur dioxide d. Nitrogen monoxide
3. a. Dinitrogen pentoxide (or pentaoxide) b. tetraphosphorus decaoxide c. Nitrogen dioxide

   d. Dinitrogen monoxide e. Sulfur trioxide f. Dichlorine heptaoxide (or heptoxide).

* * * * *

Lesson 7B: Naming Ions

Prerequisites: Complete Module 6 and Lesson 7A before starting this lesson.

Pretest: If you think you know this topic, try several problems at the end of this lesson. If you complete them all correctly, you may skip the lesson.

* * * * *

Ions

Ionic compounds are combinations of *ions*: particles with an electrical charge.

In most first-year chemistry courses you will be asked to memorize the names and symbols for more than 50 frequently encountered ions. This task is simplified by the patterns for ion charges that are found in the periodic table. Learning these rules and patterns will help you to speak the language of chemistry.

Categories of Ions

1. All ions are either positive or negative.
   - A positive ion is termed a **cation** (pronounced KAT-eye-un). The charges on positive ions can be 1+, 2+, 3+, or 4+.
   - A negative ion is termed an **anion** (pronounced ANN-eye-un). The charges on negative ions can be 1–, 2–, or 3–.
2. All ions are either **monatomic** or **polyatomic.**
   - A monatomic ion is composed of a single atom.

     Examples of monatomic ions are Na^+, Al^{3+}, Cl^-, and S^{2-}.
   - A polyatomic ion is a particle that has two or more covalently bonded atoms and an overall electric charge.

     Examples of polyatomic ions are OH^-, Hg_2^{2+}, NH_4^+, and SO_4^{2-}.

Ions of Hydrogen

Hydrogen has unique characteristics. It is classified as a nonmetal, and in many of its compounds hydrogen bonds covalently. However, in compounds classified as acids, one or more hydrogens form H^+ ions when the compound is dissolved in water. In addition, when bonded to metal atoms, hydrogen behaves as a hydride ion (H^-).

The Structure and Charge of Metal Ions

More than 70% of the atoms in the periodic table are classified as metal atoms.

- Geologically, in the earth's crust, *most* metals are found as metal *ions*. When metal ions are found in rocks from which the ions can be extracted and converted to metals that have economic value, the rocks are termed **ores**.

  Exceptions to the "metals are found as ions" rule include the coinage metals: copper and silver, which may be found geologically both as ions or in their metallic, elemental form, and gold, which is always found in nature as a metal.

- In *reactions*, neutral metal atoms tend to *lose* electrons to form *positive* ions.
- In compounds that contain both metal and nonmetal atoms, the metal atoms nearly always behave as ions with a *positive* charge. The charge can be 1+, 2+, 3+, or 4+.
- With the exception of mercurous (Hg_2^{2+}) ion, all frequently encountered metal ions are monatomic: the ions are *single* metal atoms that have lost one or more electrons.

  Examples of metal ions are Na^+, Mg^{2+}, Al^{3+}, and Sn^{4+}.

All metals form at least one positive monatomic ion. Some frequently encountered metals form two stable monatomic ions. In many cases, the charge (or possible charges) on a monatomic metal ion can be predicted from the position of the metal in the periodic table.

In first-year chemistry, when you are asked to predict the charge on a monatomic metal atom, you will nearly always be allowed to consult a periodic table. Use a periodic table when learning the following rules for the charges on metal ions.

Metal Ions With One Charge

Metals in the *first two* columns of the periodic table form only *one* stable monatomic ion. The charge on that ion is easy to predict.

- All metals in column *one* (the alkali metals) form only one stable ion: a single atom with a **1+** charge: Li^+, Na^+, K^+, Rb^+, Cs^+, and Fr^+.
- All metals in column *two* form only one stable ion: a single atom with a **2+** charge: Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Ra^{2+}.

The charges on metal ions in the remainder of the periodic table are more difficult to predict. Additional rules for predicting ion charge will be learned when electron configuration is studied in later parts of your course.

In order to solve problems initially, most courses require that the possible charges on certain metals to the right of column 2 in the periodic table be memorized. The rules below will help with that process.

Most metals to the right of the first two columns form two or more stable ions, but some form only one. The following rule should be memorized.

- Metals to the right of the first two columns that form only *one* stable ion include Ag^+, Zn^{2+}, and Al^{3+}.

For help in remembering this group, note the position of these metals in the periodic table.

Naming Metal Ions

How a metal ion is named depends on whether the metal forms only one ion or forms two or more ions.

1. If a metal forms only *one* stable ion, the ion name is the atom name.

   Examples: Na^+ is a sodium ion. Al^{3+} is an aluminum ion.

This rule applies to

- metal ions in columns one and two, plus
- the additional five metal ions listed above, plus
- additional ions that may be studied later in chemistry.

2. For metals that form *two* different positive ions, the **systematic name** (or *modern* name) of the ion is the atom name followed by a roman numeral in parentheses that states the ion's positive charge.

   Examples: Fe^{2+} is named iron(II) and Fe^{3+} is named iron(III) ion.

3. Metals that form *two* different positive ions and were "known to the ancients" also have **common names** for their multiple ions.

   In common names, the lower charged ion uses the Latin root of the atom name plus the suffix *–ous*. The higher-charged ion uses the Latin root plus the suffix *–ic.*

   For metal ions, the systematic (roman numeral) names are preferred, but the common (latin-based) names are often encountered.

Most courses require that the names and symbols for the following ions, and perhaps others, be memorized.

Ion Symbol	Systematic Ion Name	Common Ion Name
Cu^{+}	copper(I)	cuprous
Cu^{2+}	copper(II)	cupric
Fe^{2+}	iron(II)	ferrous
Fe^{3+}	iron(III)	ferric
Sn^{2+}	tin(II)	stannous
Sn^{4+}	tin(IV)	stannic
$Hg_2{}^{2+}$	mercury(I)	mercurous
Hg^{2+}	mercury(II)	mercuric

Lead also forms two ions. Pb^{2+} is named lead(II), and Pb^{4+} is named lead(IV). Lead(II) is encountered in more compounds than lead(IV).

Note the exceptional name and structure of the mercury(I) ion. Mercury(I) is the only frequently encountered metal ion that is polyatomic: It has the structure of a diatomic ion with a 2+ charge. It is given the name mercury(I) matching the format of other metal ions, in part because it behaves in many respects as two loosely bonded +1 ions.

When to Include Roman Numerals In Systematic Names

When naming metal ions, the general *rule* is: Do *not* use roman numerals in systematic names for metal ions that can form only *one* stable ion: ions for atoms in the first two columns, plus Ag^{+}, Zn^{2+}, and Al^{3+}.

However, for the transition metals that form ions, adding the roman numeral, such as using nickel(II) for Ni^{2+}, may be acceptable in your course.

Summary: Metal Ion Rules

- All metal ions are positive. Except for Hg_2^{2+}, all metal atoms are monatomic.
- In column one, all atoms tend to form 1+ ions.
- In column two, all atoms tend to form 2+ ions.
- If a metal forms only one ion, the ion name is the atom name.
- If a metal forms **more** than one ion, the systematic ion name is the atom name followed by a roman numeral in parentheses showing the positive charge of the ion.
- For the metals to the right of column 2, metals that form only one monatomic ion include Ag^+, Zn^{2+}, and Al^{3+}. For naming purposes, assume that other metals form more than one ion and the ( ) is needed in the name.

Flashcards: Using the flashcard steps in Lesson 2C, make cards for any of these that you cannot answer from memory.

One-way cards (with notch)	Back Side -- Answers
cation	A positive ion
anion	A negative ion
Monatomic ion	One atom with a charge
Polyatomic ion	2 or more bonded atoms with an overall charge
All metal ions (except mercurous) are	Monatomic – contain only one atom
The charge on a metal ion is always	Positive
Column one ions have what charge?	+1
Column two ions have what charge?	+2
When is () in an *ion* name needed?	In systematic names, if the metal forms more than one kind of positive ion
In systematic names, which ions do *not* need (roman numerals) to show their charge?	Columns 1 and 2, plus Ag^+, Zn^{2+}, and Al^{3+}

Practice A: Use a periodic table. Memorize the rules, ion symbols, and names in the section above *before* doing the problems. On multi-part questions, save a few parts for your next study session.

1. Add a charge to show the symbol for the stable ion that these atoms form.

   a. Ba b. Al c. Rb d. Na e. Zn f. Ag

2. Write the symbols for these ions.

   a. Cadmium ion b. Lithium ion c. Hydride ion d. Calcium ion

3. Which ions in Problems 1 and 2 are anions?
4. Write the name and symbol for a polyatomic metal ion often encountered.
5. Fill in the blanks.

Ion Symbol	Systematic Ion Name	Common Ion Name
		Stannic
		Cupric
	Iron(III)	
	Copper(I)	
Fe^{2+}		

Monatomic Anions

Nine monatomic anions are often encountered in first-year chemistry. Their names and symbols should be memorized.

- One is H^{-} (hydride).
- Four are halides (the -1 ions of halogens): fluoride, chloride, bromide, and iodide (F^{-}, Cl^{-}, Br^{-}, and I^{-}).
- Two are in tall column 6A: oxide (O^{2-}) and sulfide (S^{2-}).
- Two are in tall column 5A: nitride (N^{3-}), and phosphide (P^{3-}).

For monatomic anions, the name is the root of the atom name followed by *-ide.*

For monatomic ions, the position of the atom in the periodic table predicts the charge.

Group	1A	2A		3A	4A	5A	6A	7A	8A
Family Name	Alkali Metals		Transition Metals			N Family	O Family	Halogens	Noble Gases
Charge on Monatomic ion	1 +	2 +		3 + (or 1+)		3 −	2 −	1 −	None

Polyatomic Ions

A polyatomic ion is a particle that both has two or more atoms held together by covalent bonds and has an overall electrical charge. In polyatomic ions, the total number of protons and electrons in the overall particle is not equal.

An example of a polyatomic ion is the hydroxide ion, OH^{-}. One way to form this ion is to start with a neutral water molecule H—O—H, which has 1+8+1 = 10 protons and 10 balancing electrons, and take away an H^{+} ion (which has one proton and no electrons).

The result is a particle composed of two atoms with a total of 9 protons and 10 electrons. Overall, the particle has a negative charge. The negative charge behaves as if it is attached to the oxygen. A structural formula for the hydroxide ion is

$$H-O^-$$

Polyatomic ions will be considered in more detail when studying the three-dimensional structure of particles. At this point, our interest is the *ratios* in which ions combine. For that purpose, it may help to think of a monatomic ion as a charge that has one atom attached, and a polyatomic ion as a charge with several atoms attached.

Polyatomic Cations

Three polyatomic cations with names and symbols that should be memorized are the NH_4^+ (ammonium), H_3O^+ (hydronium), and Hg_2^{2+} (mercury(I) or mercurous) ions.

Oxyanions

Polyatomic ions with negative charges that contain non-metals and oxygen are termed **oxyanions**. Oxyanions are often part of a *series* of ions that has one *common* atom and the same charge, but different numbers of oxygen atoms.

Example: Nitrate ion = NO_3^-, nitrite ion = NO_2^-

The names and symbols for most oxyanions can be determined from the following rules.

Oxyanion Naming System

1. When an atom has *two* oxyanions that have the same charge, the ion with more oxygens is named root-*ate* , and the ion with one fewer oxygen atoms is root-*ite*.

Example**:** Sulf**ate** is SO_4^{2-} . Sulf**ite** is SO_3^{2-}

2. If an atom has *more* than two oxyanions with the same charge, the
   - *per*–root–*ate* ion has X oxygen atoms:
   - root-*ate* ion has one fewer oxygens;
   - root-*ite* ion has 2 fewer oxygens;
   - *hypo*-root-*ite* ion has 3 fewer oxygens.

Example: Memorize that the ClO_4^- ion is named ***per****chlor**ate*. Then,

- ClO_3^- is chlor**ate**;
- ClO_2^- is chlor**ite**;
- ClO^- is **hypo**chlor**ite**.

A way to simplify naming these ions is to memorize the name and formula for the ion in the series that has the most oxygens, then write out the rest by logic as needed. With practice, this naming process will become automatic.

Learning the Ion Names and Formulas

In most courses, you will be asked to memorize the names and formulas for a list of frequently encountered ions. Even if this is not required, doing so will speed your work and improve your understanding of chemistry.

The following set of flashcards is information that you will rely on heavily for the remainder of your course. You may want to use a unique card color to identify these as the *ion* cards, *or* add the word *ion* for clarity after each ion name.

Your course may not *require* that you know the "latin" names for the metal ions that have more than one possible charge, but learning those names and charges will help you to recall what charges are likely to be found on those metal ions.

Make these "two-way" flashcards following the procedure in Lesson 2C:

- cover the formula, and put a check if you are certain of the formula from the name.
- Then cover the names and put a check if you know the name from the formula.

You will need to be able to translate in *both* directions between the names and the ion formulas. Omit making flashcards for names and formulas that you already know well in both directions.

For a large number of new flashcards, allow yourself several days of practice. In the beginning, writing and saying the answers out loud will speed your progress.

Two-way cards (with*out* notch):

CH_3COO^-	acetate
CN^-	cyanide
OH^-	hydroxide
NO_3^-	nitrate
MnO_4^-	permanganate
CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate
CrO_4^{2-}	chromate
$Cr_2O_7^{2-}$	dichromate
PO_4^{3-}	phosphate
SO_4^{2-}	sulfate
SO_3^{2-}	sulfite
Na^+	sodium ion
K^+	potassium ion

Two-way cards (with*out* notch):

Cu^+	cuprous/copper(I)
Cu^{2+}	cupric/copper(II)
Fe^{2+}	ferrous/iron(II)
Fe^{3+}	ferric/iron(III)
Sn^{2+}	stannous/tin(II)
Sn^{4+}	stannic/tin(IV)
Hg_2^{2+}	mercurous or mercury(I)
Hg^{2+}	mercuric or mercury(II)
O^{2-}	oxide
S^{2-}	sulfide
N^{3-}	nitride
P^{3-}	phosphide
ClO_4^-	perchlorate
ClO_3^-	chlorate

Al^{3+}	aluminum ion
F^{-}	fluoride
Cl^{-}	chloride
Br^{-}	bromide
I^{-}	iodide
Ca^{2+}	calcium ion
Ba^{2+}	barium ion

ClO_2^{-}	chlorite
ClO^{-}	hypochlorite
H^{+}	hydrogen ion
H^{-}	hydride
Mg^{2+}	magnesium ion
NH_4^{+}	ammonium
H_3O^{+}	hydronium

Practice B: Learn the rules and run the flashcards for the ion names and symbols in the section above, *then* try these problems. Work in your notebook. Repeat these again after a few days of flashcard practice.

1. In this chart of ions, from memory, add *charges*, *names*, and ion *formulas*.

Symbol	Ion name
	acetate
CN	
	silver
	hydroxide
Al	
ClO_4	
	nitrate
	sodium
F	

CO_3	
	radium
MnO_4	
CrO_4	
K	
	dichromate
PO_4	
	sulfate
	sulfide
Ba	

2. Circle the **poly**atomic ion symbols in the left column of Problem 1 above.
3. If NO_3^{-} is a nitrate ion, what is the symbol for a nitrite ion?
4. Complete this table for the series of oxyanions containing bromine.

Ion name	Ion Symbol
Per_______________	_______________
___________________________	BrO_3^{-}
Bromite	__________________
Hypo_________________	___________________

5. Write the names for the atoms with these symbols:

   a. Au ______________ b. Ag ______________ c. Hg ___________________

6. Write the symbols for the atoms with these names:

   a. Tin ________ b. Copper ______ c. Iron _________ d. Lead ________

ANSWERS

Practice A

1. a. Ba^{2+} b. Al^{3+} c. Rb^{+} d. Na^{+} e. Zn^{2+} f. Ag^{+}

2. a. Cd^{2+} b. Li^{+} c. H^{-} d. Ca^{2+} 3. Only the hydride ion (H^{-}). 4. Hg_2^{2+}

5.

Ion Symbol	Systematic Ion Name	Common Name
Sn^{4+}	**tin(IV)**	stannic
Cu^{2+}	**copper(II)**	cupric
Fe^{3+}	iron(III)	**ferric**
Cu^{+}	copper(I)	**cuprous**
Fe^{2+}	**iron(II)**	**ferrous**

Practice B

1,2.

Symbol	Ion name
CH_3COO^{-}	acetate
CN^{-}	**cyanide**
Ag^{+}	silver
OH^{-}	hydroxide
Al^{3+}	**aluminum**
ClO_4^{-}	**perchlorate**
NO_3^{-}	nitrate
Na^{+}	sodium
F^{-}	**fluorine**

CO_3^{2-}	**carbonate**
Ra^{2+}	radium
MnO_4^{-}	**permanganate**
CrO_4^{2-}	**chromate**
K^{+}	**potassium**
$Cr_2O_7^{2-}$	dichromate
PO_4^{3-}	**phosphate**
SO_4^{2-}	sulfate
S^{2-}	sulfide
Ba^{2+}	**barium**

3. NO_2^{-}

4.

Ion name	**Ion Symbol**
Per**bromate**	$\mathbf{BrO_4^-}$
Bromate	BrO_3^-
Bromite	$\mathbf{BrO_2^-}$
Hypo**bromite**	$\mathbf{BrO^-}$

5a. Au **Gold** b. Ag **Silver** c. Hg **Mercury**

6a. Tin **Sn** b. Copper **Cu** c. Iron **Fe** d. Lead **Pb**

Lesson 7C: Names and Formulas for Ionic Compounds

Pretest: Using a periodic table, if you get these right 100%, you may skip the lesson. Answers are at the end of the lesson.

1. Name $Pb_3(PO_4)_2$ 2. Write formulas for a. tin(IV) chlorate b. radium nitrate

* * * * *

Ionic Compounds: Fundamentals

Ions that have opposite charges attract, and they can combine to form ionic compounds. Ionic compounds must contain both positive and negative ions, and the ions must be present in a ratio that balances the charges and results in electrical neutrality.

The composition of an ionic compound that contains one kind of cation and one kind of anion can be expressed in three ways.

- By a **name**; Example: ammonium phosphate
- As a **solid** formula; Example: $(NH_4)_3PO_4$
- And as balanced, **separated ions**. Example: $3\ NH_4^+ + 1\ PO_4^{3-}$

As a part of solving many chemistry problems, given one type of expression, you will need to write the other two.

Ionic compounds can initially be confusing because their names and solid formulas do not clearly identify the charges on the ions. To solve problems that involve ionic compounds, a key step will be to write out the ions in the *separated-ions* format that shows the number and the formulas of the ions, including their charges.

- To write names and formulas for ionic compounds, the key steps will be:

> - First write the *separated-ions* formula, and then
> - Add *coefficients* that balance the *charges*.

Balancing Separated Ions

It is a fundamental law of our universe that if matter has an electrical charge, it will tend to arrange and/or react in ways that balance and neutralize that charge, so that the total number of positive and negative charges is the same. In the case of charged particles that are ions, the result is this rule:

> In all combinations of ions, whether in solids, melted, or dissolved in water, the total charges on the ions must *balance*: the total number of positive charges must equal the total number of negative charges, so that the overall charge is *zero*.

When ions combine, only *one* ratio will result in electrical neutrality. In problems, you will often need to determine that *ratio*. The way to find that ratio is to write a balanced *separated-ions* formula for the ionic compound. Let's learn to do this with an example.

Q. Find the ratio that balances the charges when S^{2-} and Na^{+} combine.

In your notebook, apply the following steps, then check your answer below.

Step 1. Write the symbols for the two ions in the compound, with their charges, in this format: positive ion symbol **+** negative ion symbol

Step 2. **Coefficients** are numbers written in *front* of ion or particle symbols. In all ion combinations,

(Coefficient *times* charge of cation) must balance (coefficient *times* charge of anion).

In balancing, you *cannot* change the symbol or the charge of an ion.

In balancing, the change that *you* must make is to *write* whole-number *coefficients* in front of the particle symbols that balance the charges.

Step 3. Reduce the coefficients to the *lowest* whole-number ratios.

* * * * *

Answer Step 1. Na^{+} + S^{2-}

Step 2. **2** Na^{+} + **1** S^{2-} This is the *separated*-ions formula.

There *must* be *two* sodium ions for every *one* sulfide ion. Why? For the charges, (2 times 1+ = 2**+**) balances (1 times 2**—** = 2**—**). In ion combinations, the ions are always present in ratios so that the total positive and negative *charges* balance.

Step 3. 2 and 1 are the lowest whole-number ratios.

Only one set of coefficient *ratios* will balance the charges. Those coefficients show the ratios in which the ions must be found in the compound.

Try another. Cover the answer below, then try this question using the steps above.

Q. Add coefficients so that the charges balance: ___ Al^{3+} + ___ SO_4^{2-}

* * * * *

Answer: One way to find the coefficients is to make the coefficient of each ion equal to the *number* of *charges* of the *other* ion.

$$\mathbf{2}\ Al^{3+} + \mathbf{3}\ SO_4^{2-}$$

For these ions, (2 times +3 = +6) balances (3 times —2 = —6). In an ionic compound, the total positives and total negatives must balance.

However, when balancing charge when using this method, you often must adjust the coefficients so that the *final* coefficients are the *lowest* whole-number ratios. Try this problem.

Q. Add proper coefficients: ____ Ba^{2+} + ____ SO_4^{2-}

* * * * *

Answer

If balancing produces a ratio of **2** Ba^{2+} + **2** SO_4^{2-} , write the *final* coefficients as

1 Ba^{2+} + **1** SO_4^{2-}

When balancing separated ions for use in writing solid formulas, you must write coefficients with the *lowest whole-number* ratio that produces electrical neutrality.

Practice A: Add lowest-whole-number coefficients to make these separated ions balanced for charge. Start with the odd numbers; save the evens for your next practice session. After every two, check your answers at the end of the lesson.

1. ____ Na^{+} + ____ Cl^{-}
2. ____ Ca^{2+} + ____ Br^{-}
3. Mg^{2+} + SO_4^{2-}
4. Cl^{-} + Al^{3+}
5. NH_4^{+} + CH_3COO^{-}
6. In^{3+} + CO_3^{2-}
7. Al^{3+} + PO_4^{3-}
8. HPO_4^{2-} + In^{3+}

Writing the Separated Ions from Names

To write the separated ions from the *name* of an ionic compound, follow these steps.

Step 1: The first word in the name is always the positive ion.

Write: positive ion symbol **+** negative ion symbol

Step 2: Add the lowest-whole-number coefficients that balance the charges.

Try those steps on this problem.

Q. Write a balanced separated-ions formula for aluminum carbonate.

* * * * *

Answer: Step 1: Aluminum carbonate → Al^{3+} + CO_3^{2-}

Step 2: Aluminum carbonate → **2** Al^{3+} + **3** CO_3^{2-}

The separated-ions formula shows clearly what the name does not. In aluminum carbonate, there must be 2 aluminum ions for every 3 carbonate ions.

When writing separated ions, write the charges *high*, any subscripts *low*, and the coefficients at the *same* level as the atom symbols.

Practice B

If you have not done so today, run your ion flashcards. Then write balanced *separated-ion* formulas for the ionic compounds below. You may use a periodic table, but otherwise write the ion formulas from memory. Do odds now, evens later. Check answers as you go.

1. Sodium hydroxide →

2. Aluminum chloride →

3. Rubidium sulfite →

4. Ferric nitrate →

5. Lead(II) phosphate →

6. Calcium chlorate →

Writing Solid Formulas From Names

In ionic solid formulas, charges are hidden, but charges must balance. The key to writing a correct solid formula is to write the balanced *separated*-ions first, so that you can see and balance the charges.

To write a *solid* formula from the name of an ionic compound, use these steps.

1. Based on the name, write the *separated* ions. Add lowest whole number coefficients to balance charge. Then, to the right, draw an arrow → .
2. After the →, write the two ion symbols, positive ion first, with a small space between them. Include any *subscripts* that are part of the ion symbol, but *leave out* charges and coefficients.
3. For the ion symbols written after the arrow, **p**ut **p**arentheses () around a **p**olyatomic ion *if* its coefficient in the separated-ions formula is more than 1.
4. Add *subscripts* after each symbol on the right. The subscript will be the same as the coefficient in front of that ion in the *separated-ions* formula.

   Omit subscripts of 1. For polyatomic ions, write the coefficients as subscripts *outside* and *after* the parentheses.

Apply those steps to this example.

Q. Write the solid formula for potassium sulfide.

* * * * *

Answer

1: Write the *separated*-ions formula first. For potassium sulfide: $2\ K^+ + 1\ S^{2-}$

2: Re-write the symbols with*out* coefficients or charges. $2\ K^+ + 1\ S^{2-} \rightarrow$ **K S**

3: Since both K and S ions are monatomic, add no parentheses.

4: The K coefficient becomes a solid formula subscript: $\mathbf{2}\ K^+ + 1\ S^{2-} \rightarrow \mathbf{K_2S}$
The sulfide subscript of one is omitted as understood.
The *solid* formula for potassium sulfide is $\mathbf{K_2S}$.

Try another using the same steps.

Q. Write the solid formula for magnesium phosphate.

* * * * *

<u>Answer</u>

1: Write the balanced separated ions. Magnesium phosphate → **3** Mg^{2+} + **2** PO_4^{3-}

2: Write symbols without coefficients or charges. **3** Mg^{2+} + **2** PO_4^{3-} → **Mg PO_4**

3: Since Mg^{2+} is *mon*atomic (just one atom), it is not placed in parentheses.

Phosphate is *both poly*atomic *and* we need *more than* **1**, so add **( )** . **Mg (PO_4)**

4: The separated coefficient of the Mg ion becomes its solid subscript. **$Mg_3(PO_4)$**

The phosphate ion's separated coefficient becomes its solid subscript. **$Mg_3(PO_4)_2$**

$Mg_3(PO_4)_2$ is the *solid* formula for magnesium phosphate.

Recite the *3-P's rule* until it is committed to memory. When writing ionic-*solid* formulas:

> ➢ ***P*ut *p*aren**theses around ***p*oly**atomic ions -- *if* you need more than one.

Practice C: As you go, check the answers at the end of the lesson. You may want to do half of the lettered parts today and the rest during your next study session.

1. Circle the polyatomic ions.

   a. Na^+ b. NH_4^+ c. CH_3COO^- d. Ca^{2+} e. OH^-

2. When do you need parentheses? Write the rule from memory.

3. Write solid formulas for these ion combinations.

   a. 2 K^+ + 1 CrO_4^{2-} →

   b. 2 NH_4^+ + 1 S^{2-} →

   c. 1 SO_3^{2-} + 1 Sr^{2+} →

4. Balance these separated ions for charge, then write solid formulas.

   a. Cs^+ + N^{3-} →

   b. $Cr_2O_7^{2-}$ + Ca^{2+} →

   c. Sn^{4+} + SO_4^{2-} →

5. From these names, write the separated-ions formula, then the solid formula.

   a. Ammonium sulfite →

   b. Potassium permanganate →

   c. Calcium hypochlorite →

   d. Sodium hydrogen carbonate →

6. Write the solid formula.

   a. Stannous fluoride →

   b. Calcium hydroxide →

   c. Radium acetate →

<u>Writing Separated Ions From Solid Formulas</u>

When placed in water, all ionic solids dissolve to *some* extent. The dissolved ions separate and move about independently in the solution.

This dissolving process can be represented by a chemical equation that has a solid on the left and the separated ions on the right. For example, when solid sodium phosphate dissolves in water, the equation is

$$Na_3PO_{4(s)} \xrightarrow{H_2O} 3\ Na^+_{(aq)} + 1\ PO_4^{3-}{}_{(aq)}$$

The (s) is an abbreviation for the *solid* state. The (aq) is an abbreviation for the **aqueous** state, which means "dissolved in water."

When a compound divides into ions that can move about freely, the reaction is termed **dissociation**. If the reactant is an ionic solid, the ions are already present in the solid: dissolving simply allows the ions to separate, move about, collide, and potentially react with other particles.

Every equation representing ion separation must balance atoms, balance charge, and result in correct formulas for the ions that are actually found in the solution.

In equations for an ionic solid separating into its ions, some subscripts in the solid formula become coefficients in the separated ions, but others do not. In the equation above, the subscript 3 became a coefficient, but the subscripts 1 and 4 did not. To correctly separate solid formulas into ions, you must be able to recognize the ions inside the solid formula. That's why the frequently encountered ion names and formulas must be memorized.

Cover the answer below, try this example, then check the answer for tips that will make this process easier. When needed, read a part of the answer for a hint, then try again.

Q. Write the equation for the ionic solid Cu_2CO_3 separating into its ions.

* * * * *

<u>Answer</u>: Follow these steps in going from a solid formula to separated ions.

Step 1: Decide the *negative* ion's charge and coefficient first.

The first ion in a solid formula is always the positive ion, but many metal ions can have two possible positive charges. Most negative ions only have one likely charge, and that charge is often needed to identify the positive ion's charge, so we usually add the charge to the negative ion first.

In Cu_2CO_3, the negative ion is CO_3, which always has a 2— charge.

This step temporarily splits the solid formula into **Cu_2** and 1 CO_3^{2-} .

Step 2: Decide the positive ion's charge and coefficients.

Given Cu_2 and CO_3^{2-} , the positive ion or ions must include **2** copper atoms *and* must have a total **2+** charge to balance the charge of CO_3^{2-}.

So **Cu_2** , in the separated-ions formula, must be *either* **1** Cu_2^{2+} ***or*** **2** Cu^+ .

Both possibilities balance atoms and charge. Which is correct? Recall that

All *metal* ions are *mon*atomic (except Hg_2^{2+} (mercury(I) ion)).

This means that Cu^+ must be the ion that forms, since Cu_2^{2+} is polyatomic.

Because most metal ions are monatomic, a solid formula with a metal ion will separate

$$\mathbf{M_X}\text{Anion} \rightarrow \mathbf{X}\, M^{+?} + \text{Anion} \qquad (\textit{unless} \text{ the metal ion is } Hg_2^{2+}).$$

You also know that Cu^+ is the copper(I) ion that was previously memorized because it is frequently encountered. Both rules lead us to predict that the equation for ion separation is

$$\mathbf{Cu_2CO_3 \rightarrow 2\, Cu^+ + 1\, CO_3^{2-}}$$

Copper can also be a Cu^{2+} ion, but in the formula above, there is only one carbonate, and carbonate always has a 2— charge. Two Cu^{2+} ions cannot balance the single carbonate.

Step 3: <u>Check</u>: Make sure that the charges balance. Make sure that the number of atoms of each kind is the same on both sides. The equation must also make sense going backwards, from the separated to the solid formula.

Try another.

Q2. Write the equation for the ionic solid **$(NH_4)_2S$** dissolving to form ions.

* * * * *

Answer

- In a solid formula, parentheses are placed around polyatomic ions. When you write the separated ions, a subscript *after* parentheses *always* becomes the polyatomic ion's *coefficient*.

  You would therefore split the formula **$(NH_4)_2S \rightarrow 2\ NH_4 \ + 1\ S$**

- Assign the charges that these ions prefer. $(NH_4)_2S \rightarrow 2\ NH_4^{+} \ + 1\ S^{2-}$
- Check: In the separated formula, do the charges balance?

  Going backwards, do the separated ions combine to give the solid formula?

Keep up your practice, for 15-20 minutes a day, with your *ion* name and formula flashcards (Lesson 7B). Identifying ions without consulting a table will be most helpful in solving the complex problems that lie ahead.

Practice D

If you have not done so today, run your ion flashcards in both directions, then try these. To take advantage of the "spacing effect"(Lesson 2C), do half of the lettered parts below today, and the rest during your next study session.

1. Finish balancing by adding ions, coefficients, and charges.

   a. $PbCO_3 \rightarrow \quad Pb \quad + 1\ CO_3^{2-}$

   b. $Hg_2SO_4 \rightarrow \quad Hg_2 \quad +$

2. Write equations for these ionic solids separating into ions.

   a. $KOH \rightarrow$

   b. $CuCH_3COO \rightarrow$

   c. $Fe_3(PO_4)_2 \rightarrow$

   d. $Ag_2CO_3 \rightarrow$

   e. $NH_4OBr \rightarrow$

   f. $Mg(OH)_2 \rightarrow$

Naming Ionic Compounds

From a solid or a separated-ions formula, writing the *name* is easy.

Step 1: Write the separated-ions formula.

Step 2: Write the *name* of the positive ion in the formula.

Step 3. Write the name of the negative ion.

That's it! In ionic compounds, the name ignores the number of ions inside. Simply name the ions in the compound, with the positive ion named first. Try this problem.

Q. What is the name of K_2CO_3?

* * * * *

Answer

$K_2CO_3 \rightarrow$ 2 $\mathbf{K^+}$ + 1 $\mathbf{CO_3^{2-}}$; the name is potassium carbonate.

With time, you will be able to convert solid formulas to compound names without writing the separated ions, but the only way to develop this accurate intuition is by practice.

Practice E: If you are unsure of an answer, check it before continuing.

1. Return to Practice D and name each compound.
2. In Practice C, Problems 3 and 4, name each compound.
3. Would CBr_4 be named carbon bromide or carbon tetrabromide? Why?
4. Name these ionic and covalent compounds. Try half today and half during your next study session.

a. $CaBr_2$	b. NCl_3	c. NaH	d. $CuCl_2$
e. $RbClO_4$	f. KOI	g. Li_3P	h. PbO
i. NH_4BrO_2	j. SO_2	k. $CaSO_3$	l. P_4S_3

Flashcards: Add these to your collection.

One-way cards (with notch)	Back Side -- Answers
What must be true in all ionic substances?	Total **+** charges = total **—** charges Must be electrically neutral
Numbers you add to balance separated ions	coefficients
To understand ionic compounds:	Write the *separated*-ion formulas
When are parentheses needed in formulas?	In *solid* formulas, put parentheses around polyatomic ions -- *if* you need **>1**
In separated-ion formulas, what do the coefficients tell you?	The ratio in which the ions must be present to balance atoms and charge

* * * * *

Practice F: Combining Ions Worksheet

Fill in the blanks. Complete half of the rows today and the rest during your next study session. Check answers at the end of the lesson.

Ionic Compound NAME	SEPARATED Ions	SOLID Formula
• Name by ion names • Must be two or more words • Put name of + ion first	• Charges must show • Charges must balance • Charges may flow • Coefficients tell ratio of ions	• Positive ion first • Charges balance, but don't show • Put () around polyatomic ions IF you need >1
Sodium chloride	$1\ Na^{+} + 1\ Cl^{-}$	$NaCl$
	$2\ A1^{3+} + 3\ SO_3^{2-}$	$A1_2(SO_3)_3$
Lithium carbonate		
Potassium hydroxide		
	___ Ag^{+} + ___ NO_3^{-}	
	___ NH_4^{+} + ___ SO_4^{2-}	
		$FeBr_2$
		$Fe_2(SO_4)_3$
Cuprous chloride		
Tin(II) fluoride		
	__ $A1^{3+}$ + __ $Cr_2O_7^{2-}$	
		K_2CrO_4
		$CaCO_3$
Aluminum phosphate		

ANSWERS

Pretest: 1. Lead(II) phosphate 2a. $Sn(ClO_3)_4$ 2b. $Ra(NO_3)_2$

Practice A

1. **1** Na^{+} + **1** Cl^{-}
2. **1** Ca^{2+} + **2** Br^{-}
3. **1** Mg^{2+} + **1** SO_4^{2-}
4. **3** Cl^{-} + **1** Al^{3+}
5. **1** NH_4^{+} + **1** CH_3COO^{-}
6. **2** In^{3+} + **3** CO_3^{2-}
7. **1** Al^{3+} + **1** PO_4^{3-}
8. **3** HPO_4^{2-} + **2** In^{3+}

Practice B

1. Sodium hydroxide → **1 Na^{+} + 1 OH^{-}**
2. Aluminum chloride → **1 Al^{3+} + 3 Cl^{-}**
3. Rubidium sulfite → **2 Rb^{+} + 1 SO_3^{2-}**
4. Ferric nitrate → **1 Fe^{3+} + 3 NO_3^{-}**
5. Lead(II) phosphate → **3 Pb^{2+} + 2 PO_4^{3-}**
6. Calcium Chlorate → **1 Ca^{2+} + 2 ClO_3^{-}**

Practice C

1. The polyatomic ions: **b. NH_4^{+}** **c. CH_3COO^{-}** **e. OH^{-}**
2. For ionic solid formulas, put parentheses around polyatomic ions IF you need more than one.

3a. **2** K^{+} + **1** CrO_4^{2-} → **K_2CrO_4**

3b. **2** NH_4^{+} + **1** S^{2-} → **$(NH_4)_2S$**

3c. **1** SO_3^{2-} + **1** Sr^{2+} → **$SrSO_3$**

4a. **3** Cs^{+} + **1** N^{3-} → **Cs_3N**

4b. **1** $Cr_2O_7^{2-}$ + **1** Ca^{2+} → **$CaCr_2O_7$**

4c. **1** Sn^{4+} + **2** SO_4^{2-} → **$Sn(SO_4)_2$**

5a. **2** NH_4^{+} + **1** SO_3^{2-} → **$(NH_4)_2SO_3$**

5b. **1** K^{+} + **1** MnO_4^{-} → **$KMnO_4$**

5c. **1** Ca^{2+} + **2** OCl^{-} → **$Ca(ClO)_2$**

5d. **1** Na^{+} + **1** HCO_3^{-} → **$NaHCO_3$**

6. Write balanced, separated ions first to help with the solid formula.
   a. Stannous fluoride → 1 Sn^{2+} + 2 F^{-} → **SnF_2**
   b. Calcium hydroxide → 1 Ca^{2+} + 2 OH^{-} → **$Ca(OH)_2$**
   c. Radium acetate → 1 Ra^{2+} + 2 CH_3COO^{-} → **$Ra(CH_3COO)_2$**

Practice D and E

1. a. $PbCO_3$ → **1** Pb^{2+} + 1 CO_3^{2-} Part E: **Lead(II) carbonate**
   b. Hg_2SO_4 → 1 Hg_2^{2+} + **1 SO_4^{2-}** **Mercurous sulfate** or **Mercury(I) sulfate**
2. a. KOH → **1 K^{+} + 1 OH^{-}** **Potassium hydroxide**
   b. $CuCH_3COO$ → **1 Cu^{+} + 1 CH_3COO^{-}** **Copper(I) acetate** or **cuprous acetate**
   c. $Fe_3(PO_4)_2$ → **3 Fe^{2+} + 2 PO_4^{3-}** **Iron(II) phosphate** or **ferrous phosphate**
   d. Ag_2CO_3 → **2 Ag^{+} + 1 CO_3^{2-}** **Silver carbonate**
   e. NH_4OBr → **1 NH_4^{+} + 1 BrO^{-}** **Ammonium hypobromite**
   f. $Mg(OH)_2$ → **1 Mg^{2+} + 2 OH^{-}** **Magnesium hydroxide**

E2. C3a. Potassium chromate C3b. Ammonium sulfide C3c. Strontium sulfite

C4a. Cesium nitride C4b. Calcium dichromate C4c. Tin(IV) sulfate *or* stannic sulfate

E3: Carbon tetrabromide. Carbon is a nonmetal, so the compound is covalent (see Lesson 7A). Use *di-, tri-* prefixes in the names of *covalent* compounds. Practice recognizing the symbols of the nonmetals.

E4. a. Calcium bromide b. Nitrogen trichloride c. Sodium hydride

c. Copper(II) chloride or cupric chloride e. Rubidium perchlorate f. Potassium hypoiodite

g. Lithium phosphide h. Lead(II) oxide i. Ammonium bromite j. Sulfur dioxide

k. Calcium sulfite l. Tetraphosphorus trisulfide

Practice F

Ionic Compound NAME	SEPARATED Ions	SOLID Formula
Sodium chloride	1 Na^{+} + 1 $Cl^{—}$	NaCl
Aluminum sulfite	2 $A1^{3+}$ + 3 $SO_3^{2—}$	$A1_2(SO_3)_3$
Lithium carbonate	**2 Li^{+} + $CO_3^{2—}$**	**Li_2CO_3**
Potassium hydroxide	**1 K^{+} + 1 $OH^{—}$**	**KOH**
Silver nitrate	**1** Ag^{+} + **1** $NO_3^{—}$	**$AgNO_3$**
Ammonium sulfate	**2** NH_4^{+} + **1** $SO_4^{2—}$	**$(NH_4)_2SO_4$**
Iron(II) bromide/Ferrous bromide	**1 Fe^{2+} + 2 $Br^{—}$**	$FeBr_2$
Iron(III) sulfate/Ferric sulfate	**2 Fe^{3+} + 3 $SO_4^{2—}$**	$Fe_2(SO_4)_3$
Cuprous chloride	**1 Cu^{+} + 1 $Cl^{—}$**	**CuCl**
Tin(II) fluoride	**1 Sn^{2+} + 2 $F^{—}$**	**SnF_2**
Aluminum dichromate	**2** $A1^{3+}$ **+ 3** $Cr_2O_7^{2—}$	**$Al_2(Cr_2O_7)_3$**
Potassium chromate	**2 K^{+} + $CrO_4^{2—}$**	K_2CrO_4
Calcium carbonate	**1 Ca^{2+} + 1 $CO_3^{2—}$**	$CaCO_3$
Aluminum phosphate	**1 Al^{3+} + 1 $PO_4^{3—}$**	**$AlPO_4$**

Lesson 7D: Naming Acids

Timing: Complete this lesson *if* you are asked to name acids from an acid formula *or* to write a substance formula from an acid name.

Pretest: If you think you know this topic, try the last two problems on the practice at the end of the lesson. If you get all of those parts right, skip this lesson.

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Acids

An **acid** can be defined as a substance that, when dissolved in water, forms H^+ ions (there are other definitions for acids, but this is a place to start). This dissolving process can be represented by a reaction equation that has a solid, liquid, or gas on the left and the separated ions on the right.

For example, when the covalent gas hydrogen chloride dissolves in water, it forms a solution of hydrochloric acid. The reaction equation is

$$HCl_{(g)} \xrightarrow{H_2O} 1\ H^+_{(aq)} + 1\ Cl^-_{(aq)} \quad (\textit{or}\ \ \mathbf{HCl_{(aq)}})$$

Recall that (aq) is an abbreviation for *aqueous* (dissolved in water). A hydrochloric acid solution is usually represented using the molecular formula $HCl_{(aq)}$, but the separated ions are a more accurate description of the behavior of an acid. The two formulas on the right are equivalent, and we will need both types when *naming* acids.

Acid Nomenclature

Because of the long history of acids in chemistry, the names follow a variety of rules. We can write a long set of rules to cover all cases, but for now it is easier to memorize a few frequently encountered name and formula combinations, then learn a set of rules that generally apply to the remaining cases.

The steps to name acids:

Apply these rules *in order*.

Rule 1: Memorize the names for these acid solutions, by 2-way flashcard if necessary.

$\mathbf{H_2SO_4}_{(aq)}$ is **sulfuric acid**, $\mathbf{H_2SO_3}_{(aq)}$ is **sulfurous acid**, and $\mathbf{H_3PO_4}_{(aq)}$ is **phosphoric acid.**

In addition, $\mathbf{HCN}_{(aq)}$ is **hydrocyanic acid**, and the combination of an H^+ ion and an OH^- ion is…? Water.

Rule 2: Memorize: Four acids that combine one hydrogen and one halogen atom are

HCl = hydro*chlor*ic acid, **HF** = hydro*fluor*ic acid, **HBr** = hydro*brom*ic acid, and

HI = hydroiodic acid.

The next rule will apply to H^+ ions combined with oxoanions: negative ions that contain oxygen. Some oxoanions occur in a series that have the same charge but decreasing numbers of oxygens.

For the *four*-member oxoanion sequence that contain halogen atoms, the naming is

XO_4^{-} XO_3^{-} XO_2^{-} XO^{-} (where X can be the halogen Cl, Br, or I)

Per*halo*ate *halo*ate *halo*ite hypo*halo*ite

Examples: BrO^{-} is hypo**brom***ite* ion, IO_3^{-} is **iod***ate* ion,

Some oxoanion series include just two members.

Examples: NO_3^{-} (nitr**ate**) and NO_2^{-} (nitr**ite**) .

Some oxoanions are not part of a series, such as CO_3^{2-} (carbonate ion).

Rule 3. If an acid contains an H^+ ion and an oxoanion, to name the acid:

a. Write the name of the oxoanion, then cross off the suffix to form the *root* name.
b. If the ion suffix was *–ate* , replace the suffix with *–ic* followed by the word *acid*.
c. If the ion suffix was *–ite* , replace the suffix with *–ous acid*.

Examples:

For the acid $H_2CO_{3(aq)}$

To be neutral, the acid must combine **2** $H^+_{(aq)}$ + 1 $CO_3^{2-}{}_{(aq)}$

(To understand ionic compounds, write the *separated* ions formula.)

The negative ion CO_3^{2-} is named carbon~~ate~~ .

The acid name for $H_2CO_{3(aq)}$ is carbon**ic acid.**

Note that multiple H^+ ions in the acid do not affect the name.

For the acid $HClO_{(aq)}$,

By oxoanion rules, the ion ClO^{-} is named hypochlor~~ite~~ .

The acid name for $HClO_{(aq)}$ is therefore hypochlor**ous acid.**

Q. Apply Rule 3 to name these acid solutions.

a. $HClO_{4(aq)}$ b. $HNO_{2(aq)}$

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a. In the acid $HClO_4$, the negative ion is ClO_4^{-}, named perchlor~~ate~~ .
The name for an $HClO_4$ solution is perchlor**ic acid.**

b. In the acid HNO_2 , the negative ion is NO_2^{-}, named nitr~~ite~~ .
The name for an HNO_2 solution is nitr**ous acid.**

Acid Formulas

In most cases, because the H^+ ion is positive, it is written first in formulas. In compounds that contain carbon and hydrogen (organic compounds), other rules are followed.

For example: the solution consisting of H^+ ion and CH_3COO^{-} ion is named…?

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Acet~~ate~~ → **acetic acid** , contains oxygen and is named by Rule 3 above, but you will see the *formula* written as

CH_3COOH or **CH_3CO_2H** or **$HC_2H_3O_2$** or **HAc** (in which Ac is an abbreviation for acetate ion and is not the atom actinium).

However, most acid formulas write the *acidic* H's in front. We will address additional rules for identifying acid formulas in Module 14.

Practice: As always, it will improve efficiency and effectiveness if you first learn the rules, *then* do the practice, and save a few problems for your next study session.

1. Name these acid solutions. a. HCl b. HIO c. HNO_3 d. H_3PO_4
2. Write molecular formulas representing aqueous solutions of these acids.
   a. Bromous acid b. Sulfurous acid c. Chromic acid
3. Write formulas and names for aqueous solutions containing these ions.
   a. H^+ and MnO_4^- b. H^+ and SO_4^- c. H^+ and IO_3^-
4. The formula for the arsenate ion is AsO_4^{3-}. What is the name and formula for an aqueous solution of an acid composed of H^+ ions and AsO_4^{3-} ions ?

ANSWERS

1a. **Hydrochloric acid** by rule 2. 1b. **Hypoiodous acid** by rule 3 from hypoiodite ion.

1c. **Nitric acid** by rule 3 from nitrate ion. 1d. **Phosphoric acid** by rule 1.

2a. Brom*ous* acid must include brom*ite* ion which is BrO_2^-, so the acid must be **$HBrO_{2(aq)}$** .

2b. Sulfurous acid is memorized as **$H_2SO_{3(aq)}$** .

2c. Chrom*ic* acid must come from chrom*ate* ion which is CrO_4^{2-}, so the acid must be **$H_2CrO_{4(aq)}$** .

3a. The acid's anion is permangan*ate* , so the acid name is **permangan*ic acid***; **$HMnO_{4(aq)}$** .

3b. The neutral molecular formula must be **$H_2SO_{4(aq)}$** which is **sulfuric acid** (Rule 1).

3c. The acid's anion is iod*ate* , so the acid name is **iod*ic acid***; **$HIO_{3(aq)}$** .

4. To be neutral, must be **3** H^+ + **1** AsO_4^{3-} → **$H_3AsO_{4(aq)}$** . Arsen*ate* ion is the anion in arsen***ic acid.***

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Lesson 7E: Review Quiz For Modules 5-7

You may use a calculator and a periodic table. Work on your own paper. State answers to calculations in proper significant figures.

Set a 30-minute limit, then check your answers after the *Summary* that follows.

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1. (See Lesson 5D): If there are 96,500 coulombs per mole of electrons and 1 mole = 6.02 x 10^{23} electrons, what is the charge in coulombs on 100. electrons?
2. (Lesson 5E): One acre is 43,560 square feet. If one foot = 0.3048 meters, 0.250 acres is how many square meters?
3. (Lesson 5F): What is the volume in mL of a metal cylinder that is 5.0 cm in diameter and 2.0 cm long? Use a calculator. $\boxed{V_{cylinder} = \pi r^2 h}$
4. (Lesson 6B): For a particle with atomic number 92 that contains 143 neutrons and 90 electrons, write the nuclide (isotope) symbol and then the symbol for the ion.
5. (Lesson 6B): A particle of the isotope ^{107}Ag is an Ag^+ ion. How many protons, neutrons, and electrons does the particle contain?
6. (Lesson 6B): If an atom has two isotopes with masses of 104.0 amu and 108.0 amu, and 22.0% of the atom in naturally occurring samples is the lighter isotope, what is the atom's atomic mass?
7. (Lesson 6D): Which of these lists contains all non-metals?

   a. C, N, S, Na, O b. H, I, He, P, C c. F, H, Ne, Si, S d. Br, H, Al, N, C

8. (7C): Write the symbols for the ions that are combined to form these compounds.

   a. Ag_2SO_4 b. NaOH c. K_2CrO_4

9. (Lessons 7B-D): Write chemical formulas for these compounds.

   a. Sodium dichromate b. Ammonium phosphate

   c. Aluminum iodate d. Hydroiodic acid

   e. Nitrous acid f. Bromic acid

10. (Lessons 7B-D): Name these compounds.

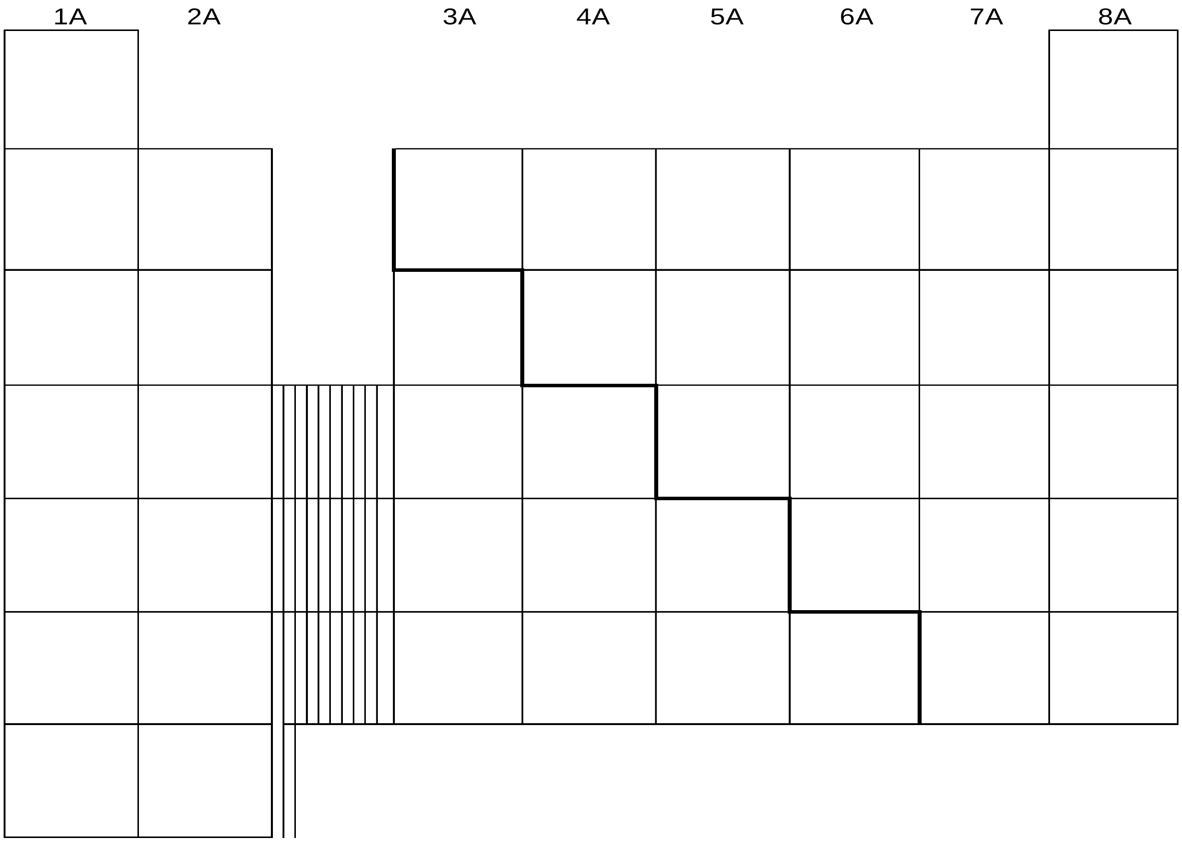
    a. Br_2O_7 b. KClO

    c. $NaHCO_3$ d. $Fe_2(SO_3)_3$

    e. CH_3COOH f. HBrO

11. (Lesson 7B): Which of the compounds in Questions 9 and 10 are covalent?
12. (4F, 6F): On the following table, fill in the names and symbols for the atoms in the first 3 *rows* and the first 2 and last 2 *columns*.

Periodic Table



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Summary: Writing Names and Formulas

1. The name of an element is the name of its atoms.
2. In covalent bonds, electrons are shared. Two nonmetal atoms usually bond with a covalent bond.
3. An ionic bond exists between positive and negative ions. If a metal is bonded to a nonmetal, the bond is generally ionic. The metal is the positive ion.
4. Most compounds with all nonmetal atoms are covalent. Most compounds that have both metal atoms and nonmetal atoms are ionic.
5. If a compound has only covalent bonds, it is covalent. If a compound has *one* or more ionic bonds, it is ionic.
6. Naming binary covalent compounds:
    a. Names have two words.
    b. Compounds with H have many exceptions. Compounds with O end in (prefix)*oxide*. (This rule has precedence.)
    c. The first word contains the name of the atom in the column farther to the left in the periodic table. For two atoms in the same column, the lower one is named first.

   d. The second word contains the root of the second atom name plus a suffix *-ide*.
   e. The number of atoms is shown by a prefix.
      - *Mono-* = 1 atom. (For the first word of the name, *mono* is left off and is assumed if no prefix is given.)
      - *Di-* = 2 atoms, *Tri-* = 3, *Tetra-* = 4, *Penta-* = 5, *Hexa-* = 6, *Hepta-* = 7, *Octa-* = 8.

7. Positive ions are cations (pronounced CAT-eye-uns). Negative ions are anions (pronounced ANN-eye-uns).
8. Metals can lose electrons to form positive ions. Column one atoms tend to form 1+ ions column two atoms tend to form 2+ ions.
9. The name of a metal ion that forms only one ion is the name of the atom.
10. Metals to the right of column two often form two different cations. The name of these ions is
    - the atom name followed by (I, II, III, or IV) stating the positive charge,
    - or a common name consisting of the Latin root plus *–ous* for the lower-charged ion or*–ic* for the higher-charged ion.
11. A polyatomic ion is composed of more than one atom.
12. The name of monatomic anions is the root followed by *-ide.*
13. For oxyanions of a given atom, the *per*–root–*ate*, root-*ate*, root-*ite*, and *hypo*-root-*ite* ions each have the same charge, but one fewer oxygens, respectively.
14. Ionic compounds have positive and negative ions in ratios that guarantee electrical neutrality.
15. To determine the names and formulas for ionic compounds,
    - write the separated-ions formula first, and
    - be certain that all names and formulas are electrically neutral.
16. To balance separated-ions formulas, add coefficients that balance charge. Coefficients are numbers written in front of the ion symbols that show the ratio of the ions in the compound. In balancing, you may not change the symbol or the stated charge of an ion.

    (Coefficient times charge of cation) must balance (coefficient times charge of anion). The overall charge for ionic compounds must equal zero.
17. To write solid formulas for ionic compounds from their names, follow these steps.
    - Write the separated ions with the lowest whole-number coefficient ratios.
    - Write the two ion symbols, positive ion first, with*out* charges, a + sign, or coefficients.
    - *P*ut *p*arentheses ( ) around *p*olyatomic ions IF you need more than one.
    - Make the separated formula coefficients into solid formula subscripts. Omit subscripts of 1.

18. To write separated ions from solid formulas,
    - decide the *negative* ion's charge and coefficients first.
    - Add the *positive* ion's charge based on what balances atoms and charge.
    - Assume that metal atoms are monatomic (except Hg_2^{2+}).
19. To name an ionic compound: name the ions, positive first.
20. To name acid solutions, memorize these:
    - $\mathbf{H_2SO_4}$ = sulfuric acid, $\mathbf{H_2SO_3}$ = sulfurous acid, $\mathbf{H_3PO_4}$ = phosphoric acid.
    - **HCl** = hydrochloric acid, **HF** = hydrofluoric acid, **HBr** = hydrobromic acid, and **HI** = hydroiodic acid.
21. If an acid contains an H^+ ion and an oxoanion, to name the acid:
    a. Write the name of the oxoanion, then cross off the suffix to form the *root* name.
    b. If the ion suffix was *–ate* , replace the suffix with *–ic* followed by the word *acid*.
    c. If the ion suffix was *–ite* , replace the suffix with *–ous acid*.

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ANSWERS – Module 5-7 Review Quiz

Some *partial* solutions are provided below. Your work on calculations should include WANTED, DATA, and SOLVE.

1. **1.60×10^{-17} Coulombs**

   $$? \text{ Coulombs} = 100. \text{ electrons} \cdot \frac{1 \text{ mole of electrons}}{6.02 \times 10^{23} \text{ electrons}} \cdot \frac{96{,}500 \text{ Coulombs}}{1 \text{ mole of electrons}} =$$

2. **$1{,}010\ m^2$** $\quad ?\ \mathbf{m^2} = 0.250 \text{ acres} \cdot \frac{43{,}560 \text{ ft}^2}{1 \text{ acre}} \cdot \left(\frac{0.3048 \text{ m}}{1 \text{ foot}}\right)^2 =$

3. **39 mL** $\quad V_{cylinder} = \pi r^2 h = \pi (2.5 \text{ cm})^2(2.0 \text{ cm}) = 39 \text{ cm}^3 = 39 \text{ mL}$

4. **^{235}U** and **U^{2+}** $\qquad$ 5. **47 protons, 60 neutrons,** and **46 electrons**

6. **107.1 amu** $\quad$ ave. mass = (104.0 g/mol x 0.220) + (108.0 g/mol x 0.780) =

7. **b. H, I, He, P, C** $\qquad$ 8a. **Ag^+** and **SO_4^{2-}** $\qquad$ 8b. **Na^+** and **OH^-**

8c. **K^+** and **CrO_4^{2-}** $\qquad$ 9a. **$Na_2Cr_2O_7$** $\qquad$ 9b. **$(NH_4)_3PO_4$** $\qquad$ 9c. **$Al(IO_3)_3$**

9d. **HI** $\qquad$ 9e. **HNO_2** $\qquad$ 9f. **$HBrO_3$** $\qquad$ 10a. **Dibromine heptoxide (or heptaoxide)**

10b. **Potassium hypochlorite** $\qquad$ 10c. **Sodium hydrogen carbonate** (or **sodium bicarbonate**)

10d. **Iron(III) sulfite** $\qquad$ 10e. **Acetic acid** $\qquad$ 10f. **Hypobromous acid**

11. **Only 10a** $\quad$ Acids contain H^+ *ions*. $\qquad$ 12. See a periodic table.

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NOTE on the Table of Atoms

The atomic masses in this Table of Atoms use fewer significant figures than most similar tables in college textbooks. By "keeping the numbers simple," it is hoped that you will use "mental arithmetic" to do easy numeric cancellations and simplifications before you use a calculator for arithmetic.

Many calculations in these lessons have been set up so that you should not need a calculator at all to solve, if you look for *easy cancellations* first.

After any use of a calculator, use mental arithmetic and simple cancellations to *estimate* the answer, in order to catch errors in calculator use.

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The ATOMS –

The **third** column shows the atomic number: The **protons** in the nucleus of the atom.

The **fourth** column is the molar mass, in **grams/mole**. For radioactive atoms, **( )** is the molar mass of most stable isotope.

Actinium	Ac	89	(227)
Aluminum	Al	13	27.0
Americium	Am	95	(243)
Antimony	Sb	51	121.8
Argon	Ar	18	40.0
Arsenic	As	33	74.9
Astatine	At	84	(210)
Barium	Ba	56	137.3
Berkelium	Bk	97	(247)
Beryllium	Be	4	9.01
Bismuth	Bi	83	209.0
Boron	B	5	10.8
Bromine	Br	35	79.9
Cadmium	Cd	48	112.4
Calcium	Ca	20	40.1
Californium	Cf	98	(249)
Carbon	C	6	12.0
Cerium	Ce	58	140.1
Cesium	Cs	55	132.9
Chlorine	Cl	17	35.5
Chromium	Cr	24	52.0
Cobalt	Co	27	58.9
Copper	Cu	29	63.5
Curium	Cm	96	(247)
Dysprosium	Dy	66	162.5
Erbium	Er	68	167.3
Europium	Eu	63	152.0
Fermium	Fm	100	(253)
Fluorine	F	9	19.0
Francium	Fr	87	(223)
Gadolinium	Gd	64	157.3
Gallium	Ga	31	69.7
Germanium	Ge	32	72.6
Gold	Au	79	197.0
Hafnium	Hf	72	178.5
Helium	He	2	4.00
Holmium	Ho	67	164.9
Hydrogen	H	1	1.008
Indium	In	49	114.8
Iodine	I	53	126.9
Iridium	Ir	77	192.2
Iron	Fe	26	55.8
Krypton	Kr	36	83.8
Lanthanum	La	57	138.9
Lawrencium	Lr	103	(257)
Lead	Pb	82	207.2
Lithium	Li	3	6.94
Lutetium	Lu	71	175.0
Magnesium	Mg	12	24.3
Manganese	Mn	25	54.9
Mendelevium	Md	101	(256)
Mercury	Hg	80	200.6
Molybdenum	Mo	42	95.9
Neodymium	Nd	60	144.2
Neon	Ne	10	20.2
Neptunium	Np	93	(237)
Nickel	Ni	28	58.7
Niobium	Nb	41	92.9
Nitrogen	N	7	14.0
Nobelium	No	102	(253)
Osmium	Os	76	190.2
Oxygen	O	8	16.0
Palladium	Pd	46	106.4
Phosphorus	P	15	31.0
Platinum	Pt	78	195.1
Plutonium	Pu	94	(242)
Polonium	Po	84	(209)
Potassium	K	19	39.1
Praseodymium	Pr	59	140.9
Promethium	Pm	61	(145)
Protactinium	Pa	91	(231)
Radium	Ra	88	(226)
Radon	Rn	86	(222)
Rhenium	Re	75	186.2
Rhodium	Rh	45	102.9
Rubidium	Rb	37	85.5
Ruthenium	Ru	44	101.1
Samarium	Sm	62	150.4
Scandium	Sc	21	45.0
Selenium	Se	34	79.0
Silicon	Si	14	28.1
Silver	Ag	47	107.9
Sodium	Na	11	23.0
Strontium	Sr	38	87.6
Sulfur	S	16	32.1
Tantalum	Ta	73	180.9
Technetium	Tc	43	(98)
Tellurium	Te	52	127.6
Terbium	Tb	65	158.9
Thallium	Tl	81	204.4
Thorium	Th	90	232.0
Thulium	Tm	69	168.9
Tin	Sn	50	118.7
Titanium	Ti	22	47.9
Tungsten	W	74	183.8
Uranium	U	92	238.0
Vanadium	V	23	50.9
Xenon	Xe	54	131.3
Ytterbium	Yb	70	173.0
Yttrium	Y	39	88.9
Zinc	Zn	30	65.4
Zirconium	Zr	40	91.2