UNIT 7

Changes in Matter

Chapter 19

Molecules and Compounds

Introduction to Chapter 19

Pure elements are made up of one type of atom. Compounds are made up of molecules which consist of more than one type of atom. Why is it that most of the substances found on earth are compounds? Why do atoms usually associate with other atoms instead of existing alone? In this chapter, you will explore why atoms form chemical bonds to make molecules and compounds.

Investigations for Chapter 19

19.1	**Bonding and Molecules**	*Why do atoms form chemical bonds?*

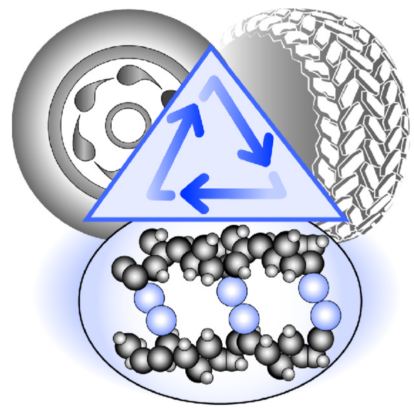
In this Investigation, you will build models of atoms and discover one of the fundamental ideas in chemistry: How electrons are involved in the formation of chemical bonds.

19.2	**Chemical Formulas**	*Why do atoms combine in certain ratios?*

In this Investigation, you will discover how the arrangement of electrons in atoms is related to groups on the periodic table. You will also learn why atoms form chemical bonds with other atoms in certain ratios.

19.3	**Comparing Molecules**	*How can you determine the chemical formula of a compound?*

Atoms combine in whole number ratios to form chemical compounds. In fact, the same two elements may form several different compounds by combining in different ratios. Chemical formulas show the ratios in which elements combine to form a compound. In this Investigation, you will use nuts and bolts to illustrate the meaning of chemical formulas.



Learning Goals

In this chapter, you will:

- ✔ Relate the chemical behavior of an element, including bonding, to its placement on the periodic table.
- ✔ Identify how elements form chemical bonds and the role of electrons in bonding.
- ✔ Predict the chemical formulas of compounds made up of two different elements.
- ✔ Write chemical formulas for compounds made up of many different types of elements.
- ✔ Calculate the formula mass of a compound and compare different compounds based on their formula masses.
- ✔ Identify the environmental and economic impact of recycling plastics.

Vocabulary

Avogadro number	diatomic molecule	ion	polymer
chemical bond	energy level	monoatomic ion	react
chemical formula	formula mass	octet	subscript
covalent bond	ionic bond	polyatomic ion	valence electron

19.1 Bonding and Molecules

Most of the matter around you and inside of you is in the form of compounds. For example, your body is about 80 percent water. You learned in the last unit that water, H_2O, is made up of hydrogen and oxygen atoms combined in a 2:1 ratio. If a substance is made of a pure element, like an iron nail, chances are (with the exception of the noble gases) it will eventually react with another element or compound to become something else. Why does iron rust? Why is the Statue of Liberty green, even though it is made of copper? The answer is fairly simple: Most atoms are unstable unless they are combined with other atoms. In this section, you will learn how, and *why*, atoms combine with other atoms to form molecules. Molecules are made up of more than one atom. When atoms combine to make molecules, they form chemical bonds.

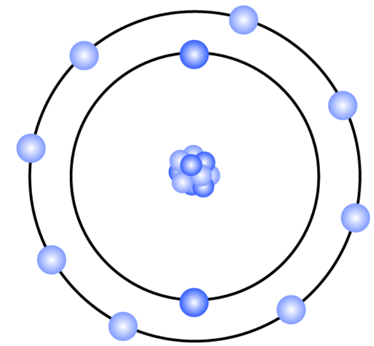


Figure 19.1: *A neon atom is chemically stable because it has a complete octet, or eight valence electrons.*

Why do atoms form chemical bonds?

The outer electrons are involved in bonding

Electrons in atoms are found in energy levels surrounding the nucleus in the form of an electron cloud. The higher the energy level, the more energy is required in order for an electron to occupy that part of the electron cloud. The outermost region of the electron cloud contains the valence electrons and is called the *valence shell*. The maximum number of valence electrons that an atom can have is *eight*. The exception to this rule is the first energy level, which only holds *two* electrons. Valence electrons are the ones involved in forming chemical bonds.

Stable atoms have eight valence electrons

Stable atoms have eight valence electrons. When an atom has eight valence electrons, it is said to have an octet of electrons. Figure 19.1 shows neon with a complete octet. In order to achieve this octet, atoms will lose, gain, or share electrons. An atom with a complete octet is chemically *stable*. An atom with an incomplete octet, like sodium (figure 19.2), is chemically *unstable*. Atoms form bonds with other atoms by either sharing them, or transferring them in order to complete their octet and become stable. This is known as the octet rule.

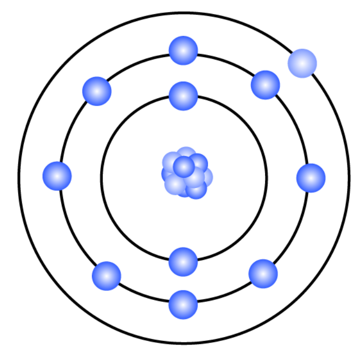


Figure 19.2: *A sodium atom is chemically unstable because it has only one valence electron.*

Exceptions to the octet rule

Look at a periodic table on page 321. Which atoms do you think are an exception to the octet rule? Remember, the first energy level only needs two electrons, not eight. Hydrogen, with only one electron, needs only one more to fill its valence shell. Helium, with two electrons, has a full valence shell (figure 19.3). This means that helium is chemically stable and does not bond with other atoms.

Stable atoms have full valence shells

What about lithium? It has three electrons. This means that its first shell is full but there is one extra electron in the second shell. Would it be easier for lithium (figure 19.4) to gain seven electrons to fill the second shell—or to lose one electron? You probably would guess that it is easier to lose one electron that gain seven. You would be correct in your guess, for lithium loses one electron when it bonds with other atoms. Table 19.1 shows the number of valence electrons and the number needed to complete the octet for the first 18 elements.

Table 19.1: ***Elements, number of valence electrons, and number needed to complete the octet***

element	valence electrons	number needed	element	valence electrons	number needed
H	1	1	Ne	8	0
He	2	0	Na	1	7
Li	1	7	Mg	2	6
Be	2	6	Al	3	5
B	3	5	Si	4	4
C	4	4	P	5	3
N	5	3	S	6	2
O	6	2	Cl	7	1
F	7	1	Ar	8	0

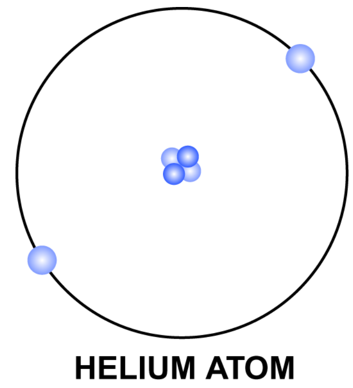


Figure 19.3: *Helium atoms have only two electrons, both of which are in the outermost level. Helium is an exception to the octet rule.*

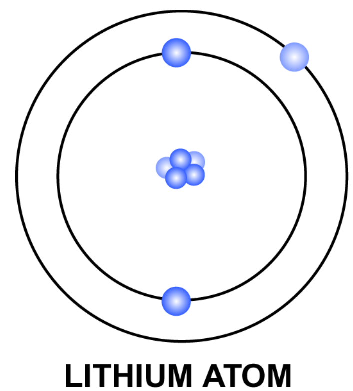


Figure 19.4: *Lithium atoms have three electrons. Since the first energy level only holds two electrons, lithium has one valence electron. If lithium loses that electron, it will have a full valence shell with two electrons.*

Using the periodic table to determine valence electrons

Do you remember how the periodic table is organized? With the exception of the transition metals, the *column* of the table tells you how many valence electrons each element has. For example, the atoms of the elements in column 1 have only *one* valence electron. Elements in column 2 have *two* valence electrons. Next, we skip to column 13 headed by boron. Atoms in this column have three valence electrons. Columns 14 through 18 have *four*, *five*, *six*, *seven*, and *eight* valence electrons, respectively.

Partial Periodic Table

Number of valence electrons in parentheses

Transition metals - groups 3 - 12
(Variable number of valence electrons)

(1) 1	(2) 2	3	4	5	6	7	8	9	10	11	12	(3) 13	(4) 14	(5) 15	(6) 16	(7) 17	(8) 18
H 1																	He 2
Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54

How do you show valence electrons in a diagram?

Valence electrons are often represented using *dot diagrams*. This system was developed in 1916 by G.N. Lewis, an American chemist. The symbol of the element in the diagram represents the nucleus of an atom and all of its electrons except for the valence electrons. The number of dots placed around the symbol of the element is equal to the number of valence electrons. The arrangement of the dots has no special significance and does not show the actual location of the electrons around the nucleus of the atom. Dots are shown in pairs around the four sides of the symbol as a reminder that electrons occur in pairs in the valence shell. Electrons begin to pair up only when no more single spaces are left. This is why the first four electrons are shown as single dots around the symbol, as in the diagram for carbon in figure 19.5.

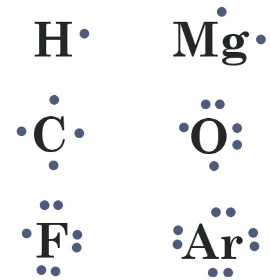


Figure 19.5: *Dot diagrams show the numbers of valence electrons.*

Types of chemical bonds

Chemical bonds result in molecules with different properties

Sodium is a soft, silvery metal so reactive that it must be stored so it does not come into contact with the air. Chlorine exists as a yellow-green gas that is very poisonous. When atoms of these two elements chemically bond, they become the white crystals that you use to make your food taste better: sodium chloride, also known as table salt. When chemical bonds form between atoms, the molecules formed are very different from the original elements they are made out of. What is the "glue" that helps hold atoms together to form so many different compounds? To answer this question, we must study the *types* of chemical bonds.

Ionic bonds

Recall that atoms will gain, lose, or share electrons in order to gain eight valence electrons in their outermost shell, that being the most stable configuration. **Ionic bonds** are formed when atoms gain or lose electrons. Sodium has one valence electron in its third energy level. If sodium loses that electron, its second energy level becomes full and stable with eight electrons. Chlorine has seven valence electrons. If chlorine gains only one electron, its valence shell will be full and stable. Do you think these two atoms are likely to bond?

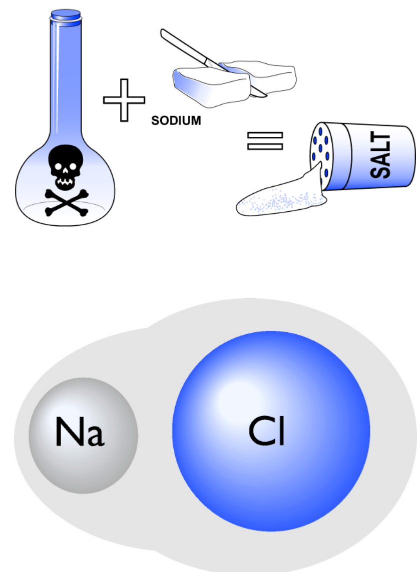


Figure 19.6: *Sodium and chlorine form an ionic bond to make sodium chloride (table salt).*

Ionization

In the last unit, you learned that all atoms are electrically neutral because they have the same number of protons and electrons. When atoms gain or lose electrons, they become **ions**, or atoms that have an electrical charge.

A neutral sodium atom has 11 positively charged protons and 11 negatively charged electrons. When sodium loses one electron to become more stable, it has 11 protons (+) and 10 electrons (-) and becomes an ion with a charge of +1. This is because it now has one more positive charge than negative charges (figure 19.7).

Figure 19.7: *When sodium loses an electron, it becomes an ion with a +1 charge.*

A neutral chlorine atom has 17 protons and 17 neutrons. When chlorine gains one electron to complete its stable octet, it has 17 protons (+) and 18 electrons (-) and becomes an ion with a charge of -1. This is because it has gained one negative charge (figure 19.8).

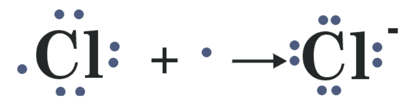


Figure 19.8: *When chlorine gains an electron, it becomes an ion with a -1 charge.*

Opposites attract

Because the sodium ion has a positive charge and the chlorine ion has a negative charge, the two atoms become attracted to each other and form an ionic bond. Recall that opposite charges attract. When sodium, with its +1 charge, comes into contact with chlorine, with its -1 charge, they become electrically neutral as long as they are together. This is because +1 and -1 cancel each other out. This also explains why sodium and chlorine combine in a 1:1 ratio to make sodium chloride (figure 19.9).

Figure 19.9: *Sodium and chlorine form an ionic bond. The compound sodium chloride is electrically neutral as long as the two ions stay together.*

Covalent bonds

Most atoms *share* electrons to in order to gain a stable octet. When electrons are shared between two atoms, a **covalent bond** is formed. Covalent bonds can form between two different types of atoms, or between two or more atoms of the same type. For example, chlorine, with seven valence electrons, sometimes shares an electron with another chlorine atom (figure 19.10). With this configuration, both atoms can share an electron through a covalent bond to become more stable. Many elemental gases in our atmosphere exist in pairs of covalently bonded atoms. The gases nitrogen (N_2), oxygen (O_2) and hydrogen (H_2) are a few examples. We call these covalently bonded atoms of the same type **diatomic molecules** (see Table 20.1, "Elements that exist as diatomic molecules," on page 350).

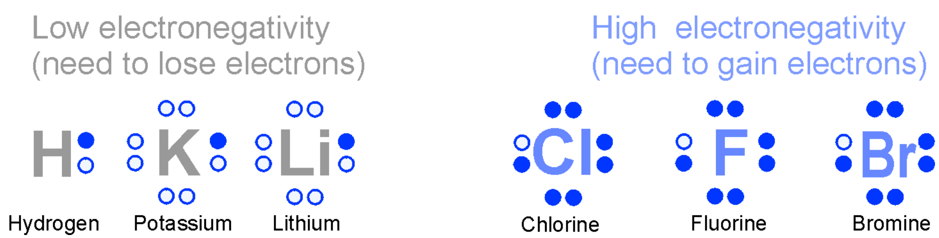
Figure 19.10: *Two chlorine atoms share the pair of electrons between them to form a covalent bond.*

How can you tell whether a bond is ionic or covalent?

Ionic bonds are formed by the attraction of two oppositely charged particles, while covalent bonds are formed when atoms share electrons. Which pairs of atoms are more likely to form ionic bonds? Which are more likely to form covalent bonds? Elements can be classified as *metals, metalloids,* and *nonmetals*. The periodic table inside the cover of this book shows these classifications. Generally, bonds between a metal and a nonmetal tend to be ionic in character while bonds between two nonmetals can be classified as covalent. However, electron pairs are sometimes shared *unequally* in covalent bonds. The attraction an atom has for the shared pair of electrons in a covalent bond is called an atom's **electronegativity**.

Atoms in column 17 have high electronegativity

For example, in a bond between hydrogen and chlorine, that electron pair is pulled toward the chlorine nucleus. This is because chlorine has very high electronegativity. Atoms in column 17 of the periodic table tend to have very high electronegativity. *Based on what you have learned about valence electrons and stability of atoms, why do you think this is true?* If you suppose that it is because these atoms only need one more electron to complete their octet, you are correct! Atoms with high electronegativity tend to want to *gain* electrons to complete their octet.



Atoms in columns 1 and 2 have low electronegativity

Conversely, atoms in the first two columns of the periodic table tend to have the lowest electronegativity. This is because they want to *get rid* of the electrons in their highest level so that their next level has a full octet. Bonds between atoms with opposite electronegativities tend to be *ionic*.

Linus Pauling

Linus Pauling developed a system for assigning electronegativity values for each element. This is just one of his many accomplishments. He is the only person to have won two unshared Nobel prizes, for chemistry in 1954, and for peace in 1962. His passion was to warn the public about the dangers of nuclear weapons, but he was equally dedicated to chemistry as it helps humanity.

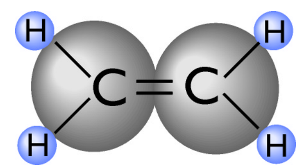


Figure 19.11: *The ethylene molecule is the building block, or subunit, of synthetic plastics. That is why plastics are often referred to as "polyethylenes."*

Environmental Issue: Paper or Plastic?

What is plastic?

Plastics are **polymers**. You may already know that the prefix *poly-* means "many" and the suffix *-mer* means "unit." A polymer is a large molecule that is composed of repeating smaller molecules. The building block or subunit of synthetic plastics is a molecule called ethylene (figure 19.11). Paper is made out of a natural polymer called cellulose. Cellulose, the most abundant polymer on Earth, is made out of many subunits of glucose molecules. The difference between a natural polymer like cellulose, and the man-made polymer that is used to make a bag or a soda bottle is that cellulose is easily digested by microorganisms. In contrast, synthetic plastics are not easily broken down. For this reason, when you throw a plastic cup away, there isn't much chance it will decompose quickly or at all.

Why can't microorganisms digest plastic?

In order for microorganisms to be able to break down a plastic molecule, they must have access to an exposed end or side branch of the molecule. Because synthetic plastics are such long chains of carbon surrounded by hydrogens, there are no places for microorganisms to begin "biting" on the molecule. Since most plastics we use are man-made, microorganisms are not able to consume them.

Biodegradable plastics

One way to approach the plastics problem is to make them *biodegradable*. This means that microbes such as bacteria and fungi can "eat" the plastic. Making biodegradable plastics involves creating exposed ends on the molecules so microbes can get a start. This is done by inserting a food item that microbes readily eat into into a plastic. For instance, starch can be inserted in the polyethylene molecule. Once microbes have eaten the starch, two ends of polyethylene are exposed. Many plastic grocery bags contain starch.

Recycling plastics

You may be familiar with the recycling symbols on the bottoms of plastic bottles. Those symbols allow you to sort the different plastics that make up each kind of plastic. Choosing the kind of plastic that is used for a certain product is a careful decision. Think about the wide variety of plastic containers (and don't forget their lids) that are used for certain products. In order to recycle plastic, you need to melt it so that it can be remolded into new containers or extruded into a kind of fabric that is used for sweatshirts.

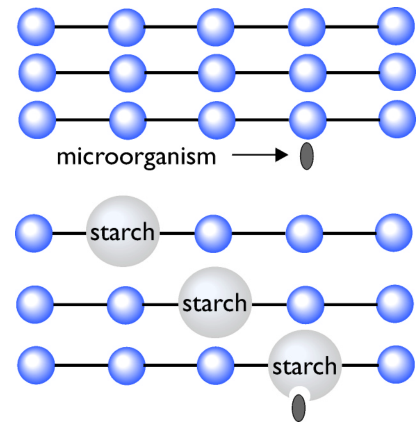


Figure 19.12: *Inserting starch molecules into the polyethylene chain provides a place for microorganisms to begin breaking it down.*

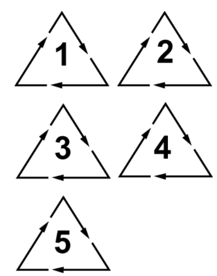


Figure 19.13: *Recycling symbols found on plastic products tell you the type of plastic and are used in sorting the plastics for recycling. Can you find these symbols on products you use every day?*

19.2 Chemical Formulas

In the previous section, you learned how and why atoms form chemical bonds with one another. You also know that atoms combine in certain ratios with other atoms. These ratios determine the **chemical formula** for a compound. In this section, you will learn how to write the chemical formulas for compounds. You will also learn how to name compounds based on their chemical formulas.

Chemical formulas and oxidation numbers

What is the chemical formula for sodium chloride?

A sodium atom will form an *ionic bond* with a chlorine atom to make a molecule of sodium chloride. Because sodium chloride is a compound made out of ions, it is called an **ionic compound**. The chemical formula for sodium chloride is NaCl. This formula indicates that every molecule of sodium chloride contains one atom of sodium and one atom of chlorine, a 1:1 ratio.

Why do sodium and chlorine combine in a 1:1 ratio? When sodium loses one electron, it becomes an ion with a charge of +1. When chlorine gains an electron, it becomes an ion with a charge of -1. When these two ions combine to form an ionic bond, the net electrical charge is zero. This is because (+1) + (-1) = 0.

All compounds have an electrical charge of zero; that is, they are neutral.

Oxidation numbers

A sodium atom always ionizes to become Na+ (a charge of +1) when it combines with other atoms to make a compound. Therefore, we say that *sodium has an oxidation number of 1+*. An **oxidation number** indicates how many electrons are lost, gained, or shared when bonding occurs. Notice that the convention for writing oxidation numbers is the opposite of the convention for writing the charge. When writing the oxidation number, the positive (or negative) symbol is written *after* the number, not *before* it.

What is chlorine's oxidation number? If you think it is 1-, you are right. This is because chlorine gains one electron, one negative charge, when it bonds with other atoms. Figure 19.15 shows the oxidation numbers for some of the elements.

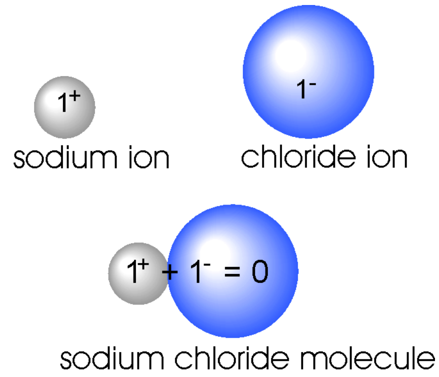


Figure 19.14: *Molecules of all compounds have an electrical charge of zero.*

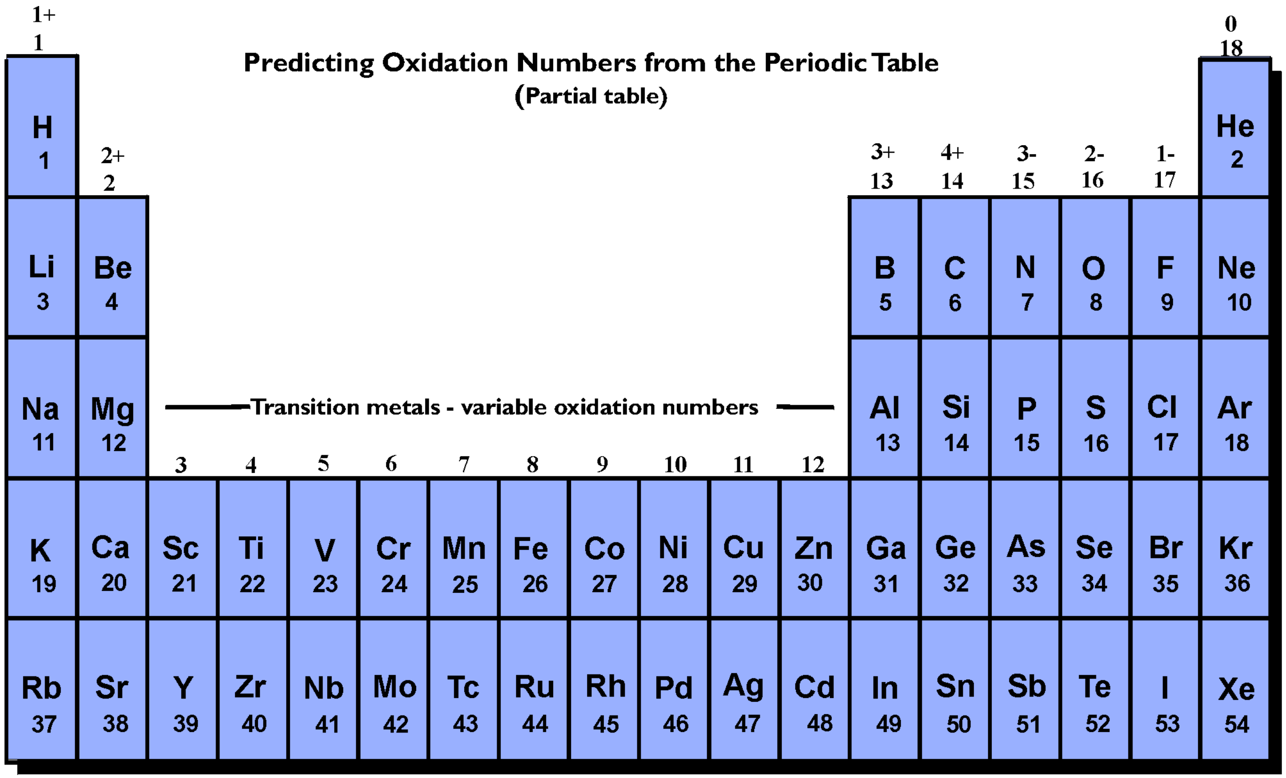
atom	electrons gained or lost	oxidation number
K	loses 1	1+
Mg	loses 2	2+
Al	loses 3	3+
P	gains 3	3-
Se	gains 2	2-
Br	gains 1	1-
Ar	loses 0	0

Figure 19.15: *Oxidation numbers of some common elements.*

Predicting oxidation numbers from the periodic table

In the last section, you learned that you can tell how many valence electrons an element has by its placement on the periodic table. If you can determine how many valence electrons an element has, you can predict its oxidation number. For example, locate beryllium (Be) on the periodic table below. It is in the second column, or Group 2, which means beryllium has two valence electrons. Will beryllium get rid of two electrons, or gain six in order to obtain a stable octet? Of course, it is easier to lose two electrons. When these two electrons are lost, beryllium becomes an ion with a charge of +2. Therefore, the most common oxidation number for beryllium is 2+. In fact, the most common oxidation number for all elements in Group 2 is 2+.

The periodic table below shows the most common oxidation numbers of most of the elements. The elements known as transition metals (in the middle of the table) have variable oxidation numbers.



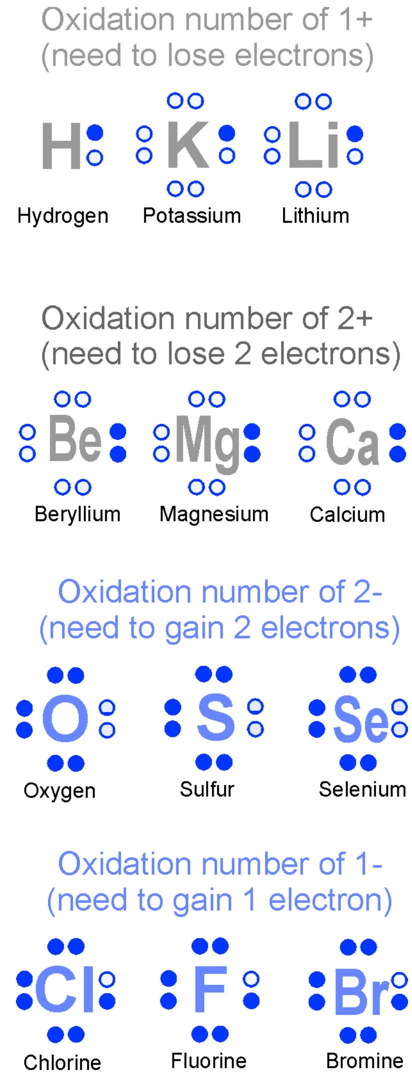


Figure 19.16: *Oxidation number corresponds to the need to gain or lose electrons.*

Writing the chemical formulas of ionic compounds.

Monoatomic ions

Both sodium and chlorine ions are monoatomic ions, that is, ions that consist of a single atom. It's easy to write the chemical formula for compounds made of monatomic ions, if you follow these rules:

1	Write the symbol for the monatomic ion that has a ***positive*** charge first.
2	Write the symbol for the monatomic ion that has a ***negative*** charge second.
3	Add subscripts so that the sum of the positive and negative oxidation numbers is equal to zero—a neutral compound, remember? Note that the subscript tells you how many atoms of that element is in the compound.

Some elements have more than one oxidation number. In this case, roman numerals are used to distinguish the oxidation number. Figure 19.17 shows a few of these elements.

element	oxidation number
copper (I)	Cu^{+}
copper (II)	Cu^{2+}
iron (II)	Fe^{2+}
iron (III)	Fe^{3+}
chromium (II)	Cr^{2+}
chromium (III)	Cr^{3+}
lead (II)	Pb^{2+}
lead (IV)	Pb^{4+}

Figure 19.17: *Elements with variable oxidation numbers.*

Example: Writing a chemical formula

Write the formula for a compound that is made of iron (III) and oxygen.

1. Find the oxidation numbers of each element in the compound.

Iron (III) is a transition metal. The roman numbers indicate that it has an oxidation number of 3+. Its formula is Fe^{3+}.

Oxygen is in group 18 of the periodic table and has an oxidation number of 2-. Its formula is O^{2-}.

2. Determine the ratios of each element and write the chemical formula.

If one iron (III) ion bonds with one oxygen ion, will the compound be neutral? No, since 3+ added to 2- equals 1+. If you have two iron (III) ions for every three oxygen ions, what happens? 2(3+) added to 3(2-) is equal to 0. This means that three iron (III) ions bond with two oxygen ions to get a neutral compound.

The formula for a compound of iron (III) and oxygen is Fe_2O_3.

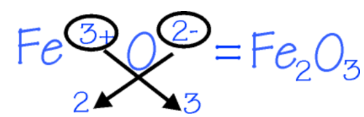


Figure 19.18: *The criss-cross method is a simple way to determine the chemical formula of a compound.*

Ionic compounds made of more than two types of atoms

Not all compounds are made of only two types of atoms

Have you ever taken an antacid for an upset stomach? Many antacids contain calcium carbonate, or $CaCO_3$. How many types of atoms does this compound contain? You are right if you said three: calcium, carbon, and oxygen. Some ionic compounds contain **polyatomic ions**. Polyatomic ions contain more than one type of atom. The prefix *poly* means "many." Table 19.2 lists some common polyatomic ions.

Table 19.2: *Polyatomic ions*

oxidation no.	name	formula
1+	ammonium	NH_4^+
1-	acetate	$C_2H_3O_2^-$
2-	carbonate	CO_3^{2-}
2-	chromate	CrO_4^{2-}
1-	hydrogen carbonate	HCO_3^-
1+	hydronium	H_3O^+
1-	hydroxide	OH^-
1-	nitrate	NO_3^-
2-	peroxide	O_2^{2-}
3-	phosphate	PO_4^{3-}
2-	sulfate	SO_4^{2-}
2-	sulfite	SO_3^{2-}

The positive ion is Ca^{2+}

This is a *monoatomic* ion.

You can determine its oxidation number by looking at the periodic table.

The negative ion is CO_3^{2-}

This is a *polyatomic* ion.

You can determine its oxidation number by looking at the ion chart (Table 19.2).

Figure 19.19: *Which ions does $CaCO_3$ contain?*

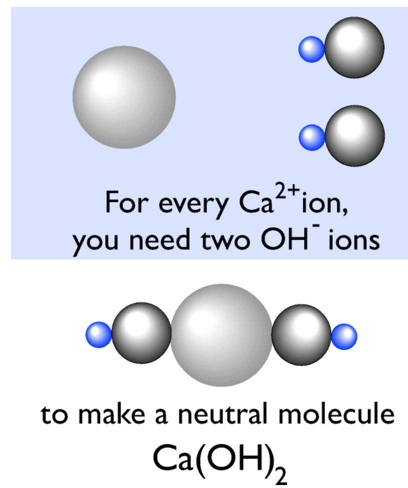


Figure 19.20: *How to write the chemical formula for calcium hydroxide.*

Rules for writing chemical formulas of ionic compounds that contain polyatomic ions

1 Write the chemical formula and oxidation number of the positive ion. If the positive ion is monoatomic, you can find its oxidation number from the periodic table. If the positive ion is polyatomic, use table 19.2 to find the oxidation number of the polyatomic ion.

2 Write the chemical formula and oxidation number for the negative ion. Again, use the periodic table if the negative ion is monoatomic, or table 19.2 if the negative ion is polyatomic.

3 Add the oxidation numbers of the positive and negative ions. Do they add up to zero? If yes, write the formula for the positive ion first and the negative ion second. Do not include the oxidation numbers in the chemical formula. Be sure to write the subscripts!

4 If the oxidation numbers do not add up to zero, figure out how many of each ion you will need so that the oxidation numbers add up to zero. (**Hint:** *Find the least common multiple between the two oxidation numbers. The number that you have to multiply each oxidation number by to equal the least common multiple tells you how many of each ion you need).*

Example: Writing the chemical formula for *aluminum sulfate*

1. Find the formula and charge of the positive ion.

The positive ion is always the first ion in the name. Where can you find the chemical formula for the aluminum ion? Aluminum is monoatomic and its formula is Al. You can find its oxidation number from the periodic table. Al is in group 13 and contains three valence electrons. When it loses those, its charge becomes +3. Therefore, the oxidation number for Al is 3+.

Chemical formula and oxidation number = Al^{3+}

2. Find the formula and charge of the negative ion.

Sulfate is the negative ion. Where can you find the chemical formula and oxidation number for the sulfate ion? Since sulfate is a polyatomic ion, you must consult an ion chart, unless you can remember the formulas and oxidation numbers for all ions!

Chemical formula and oxidation number = SO_4^{2-}

Writing the chemical formula for aluminum sulfate:

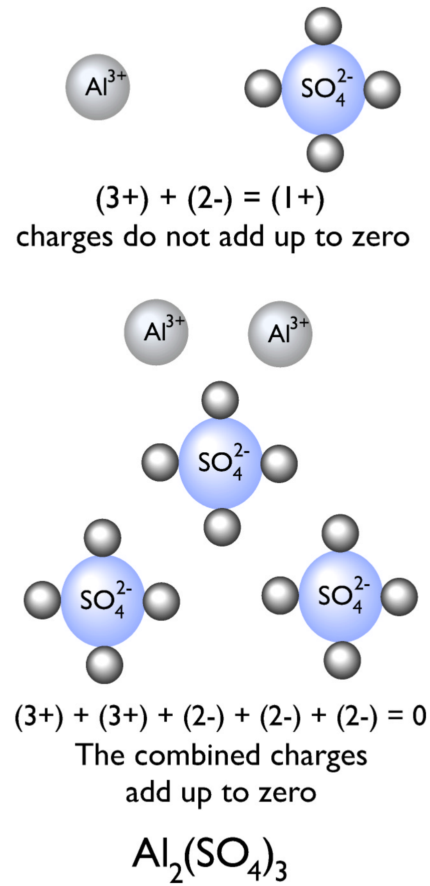


Figure 19.21: *This diagram shows how to determine the chemical formula for aluminum sulfate. How many of each ion does the formula indicate? How many atoms of each element are in one molecule of aluminum sulfate?*

Example, continued

3. Determine how many of each ion are needed so the charges are equal to zero.

The oxidation numbers of (3+) and (2-) add up to (1+), not zero. (3+) + (2-) = (1-)
You need 2 aluminum ions and 3 sulfate ions to make the charges add up to zero.
$(Al^{3+}) \times (2) = (6+)$ $(SO_4^{2-}) \times (3) = (6-)$

4. Write the chemical formula for the compound.

Write the formula, enclosing sulfate in parentheses. Do not change subscripts in the ion.
$Al_2(SO_4)_3$

How do you name ionic compounds?

Compounds with only monoatomic ions

Naming compounds with only monoatomic ions is very simple if you follow these rules:

1	Write the name of the first element in the compound.
2	Write the root name of the second element. For example, ***chlor-*** is the root name of ***chlorine***. Simply subtract the ***-ine*** ending.
3	Add the ending ***-ide*** to the root name. ***Chlor***- becomes ***chloride***.

Compounds that contain polyatomic ions

To name a compound that contains polyatomic ions, follow these steps.

1	Write the name of the positive ion first. Use the periodic table or an ion chart to find its name.
2	Write the name of the negative ion second. Use the periodic table or an ion chart to find its name.

Example: Name BaF_2

1	The first element is barium.
2	The second element is fluorine.
3	The compound's name is barium fluoride.

Figure 19.22: *Example: Naming a binary compound.*

1	The positive ion, Mg^{2+}, is **magnesium**.
2	The negative ion, CO_3^{2-}, is **carbonate**.
3	The name of the compound is **magnesium carbonate**.

Figure 19.23: *How to name a compound with the chemical formula $MgCO_3$.*

Covalent compounds

Compounds that are formed through covalent bonds (shared electrons) are called covalent compounds. Covalent compounds are sometimes referred to as *molecular compounds* because they are sometimes made of many smaller molecules, chemically bonded.

Naming covalent compounds

Covalent compounds that are made of more than two types of elements have their own special naming system that you will learn about in more advanced chemistry courses. Naming covalent compounds that consist of only two elements, often called binary compounds, is fairly straightforward. This naming is very similar to the methods used in naming ionic compounds that contain only two elements described on page 328. However, in this case, the *number* of each type of atom is specified in the name of the compound. Figure 19.24 shows how to name a binary covalent compound.

The Greek prefixes in figure 19.24 are used in naming binary covalent compounds. If the molecule contains only one atom of the first element, the prefix *mono-* is not used. It is used in the name of the second element in the compound as in the example below:

CO
carbon monoxide

CO_2
carbon dioxide

N_2O_4

dinitrogen tetraoxide

Figure 19.24: *To name a binary covalent compound, specify the number of each type of atom using a Greek prefix. The ending of the name of the second element in the compound is modified by adding the suffix -ide.*

prefix	meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9

Figure 19.25: *Greek prefixes used in naming binary covalent compounds.*

Empirical and molecular formulas

There are two ways to write chemical formulas

The simplest whole-number ratios by which elements combine are written in a form called the empirical formula. The actual number of atoms of each element in the compound is written in a form called the molecular formula. For some compounds, the empirical and molecular formula is the same as is the case with water, H_2O. A molecule of the sugar glucose has a molecular formula of $C_6H_{12}O_6$. To find the empirical formula of glucose, calculate the simplest whole number ratio of the atoms. The empirical formula for glucose is CH_2O.

19.3 Comparing Molecules

If you have ever bought paper, you know that it is sometimes sold in a package of 500 sheets called a *ream*. Do you think someone in a factory counts individual sheets of paper and packages them for sale? Instead of counting individual sheets, the paper is packaged according to *mass*. Knowing the mass of 500 sheets of paper allows the paper to be packaged quickly and efficiently by machines. If the machines that make the paper suddenly started making sheets that were twice as heavy, what would happen to the number of sheets in a package? "Counting" by mass is a useful way to deal with large numbers of objects that are uniform in size—like atoms in an element or molecules in a compound. In this section, you will learn how to quantify atoms and molecules using mass.

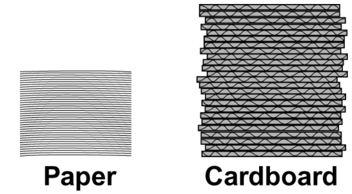


Figure 19.26: *Do 500 sheets of paper have the same mass as 500 sheets of cardboard?*

How do the masses of different molecules compare?

Comparing two different molecules

Does a molecule of *water* (H_2O) have the same mass as a molecule of *calcium carbonate* ($CaCO_3$)? Figure 19.27 shows the molecules' comparative sizes. This question seems difficult to answer at first because molecules are so small that you cannot see them. However, you *can* use what you have learned so far to answer the question. You know that atoms of different elements have different *atomic masses*. You also know that molecules are made of different numbers and types of atoms. Based on this knowledge, you can logically conclude that a molecule of water would have a *different* mass than a molecule of calcium carbonate. You can also conclude that 10 grams of water would have a different *number* of molecules from 10 grams of calcium carbonate.

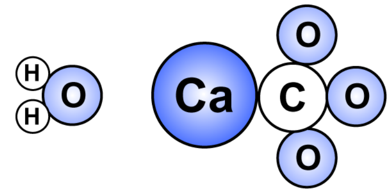


Figure 19.27: *Do you think that a molecule of water has the same mass as a molecule of calcium carbonate?*

Atomic mass units

All atoms are assigned a unit of **relative mass** known as the *atomic mass unit,* or amu. Atomic mass units allow us to compare quantities of matter even though we can't see the molecules and atoms that we want to count or measure.

How is atomic mass determined?

Carbon atoms are used as a standard for determining the atomic mass units for the other elements on the periodic table. One carbon atom is equivalent to 12.01 atomic mass units. Because one hydrogen atom is about 1/12 the mass of one carbon atom (figure 19.28), it is represented as having 1.01 atomic mass units. The actual mass of one atomic mass unit is 1.6606×10^{-24} grams—a very small amount!

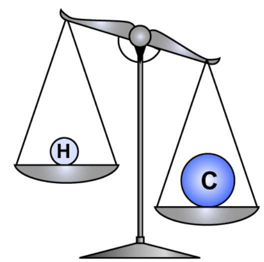


Figure 19.28: *One hydrogen atom is 1/12th the mass of one carbon atom.*

Different objects can be compared by using relative mass

We can use an analogy to explain how the concept of relative mass can be used. Let's say that we have the same number of gumdrops and jawbreakers, and that we will use the variable "x" to represent this number. The sample of x gumdrops has a mass of 100 grams, and the sample of x jawbreakers has a mass of 400 grams. This means that an individual gumdrop has a mass that is 1/4, or 25 percent of, the mass of a jawbreaker. Twenty-five percent can be represented by the number 0.25. This number represents the relative mass of a gumdrop compared with a jawbreaker. Let's call the jawbreaker unit of mass a jmu, for "jawbreaker mass unit." Now, if a single jawbreaker has a mass of 1.0 jmu, then a gumdrop would have a mass of 0.25 jmu. How many jawbreaker mass units would x number of candy bars be if they weighed 800 grams? They would have a mass of 2 jmu.

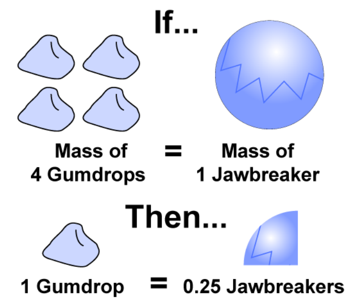


Figure 19.29: *If a single jawbreaker has a mass of 1 jmu (jawbreaker mass unit), what would the mass of 1 gumbdrop be in jmu?*

What does a chemical formula tell you?

Chemical formulas

A chemical formula for a compound gives you three useful pieces of information. First, it tells you which types of atoms and how many of each are present in a compound. Second, it lets you know if polyatomic ions are present. Remember that polyatomic ions are distinct groups of atoms with a collective oxidation number. For example, NO_3 is a polyatomic ion called nitrate with an oxidation number of 1-. As you practice writing chemical formulas, you will start to recognize these ions.

What is formula mass?

Third, a chemical formula allows you to calculate the mass of one molecule of the compound *relative* to the mass of other compounds. Formula mass, like atomic mass, is a way to compare the masses of molecules of different compounds. The formula mass of a compound is determined by adding up the atomic mass units of all of the atoms in the compound as shown in figure 19.30.

Figuring formula mass

The formula for water is H_2O. This means that there are two hydrogen atoms for every one of oxygen in a molecule of water. Using the periodic table, you can see that the atomic mass of hydrogen is 1.007 amu. For our purposes, we will round all atomic masses to the hundredths place. Using 1.01 amu for hydrogen, we can multiply this number by the number of atoms present to determine atomic mass of hydrogen in a water molecule. The atomic mass of oxygen, rounded off, is 16.00. Using this information, the formula mass for water is calculated.

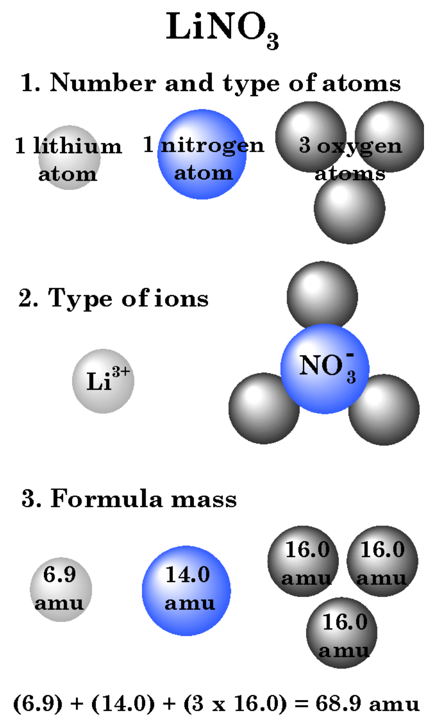


Figure 19.30: *What does a chemical formula tell you?*

How do we make atomic mass units useful to work with?

Atomic mass units and grams

Working with atomic mass units would be very difficult because each atomic mass unit has a mass of 1/12th the mass of a carbon atom. In order to make atomic mass units useful for conducting, using, and evaluating chemical reactions, it would be convenient to set the value of one amu to equal one gram. One gram is, after all, an amount of matter that we can see! For example, one paper clip has a mass of about 2.5 grams.

How do you relate molecules, atomic mass units, and grams?

If we say one water molecule is equal to 18.02 amu, does it make sense to say that one water molecule is equal to 18.02 grams? No, of course not! For this to make sense, we need to come up with a *number* of molecules or atoms that is easy to work with. From here forward, we will say that 18.02 grams of water contains Avogadro's number of molecules. This number is 6.02×10^{23}—a very, very large number of molecules! Look at the relationships in figure 19.31 to help you understand the Avogadro number.

Avogadro's number

The Avogadro number is the number of atoms in the atomic mass of an element or the number of molecules in the formula mass of a compound when these masses are expressed in grams. One set of 6.02×10^{23} atoms or molecules is also referred to as a mole of that substance. The term mole is used to talk about a number of atoms or molecules just like the term *dozen* is used to talk about quantities of eggs, doughnuts, or cans of soda.

The number, 6.02×10^{23}, was named in honor of Count Amedeo Avogadro (1776 - 1856), an Italian chemist and physicist who first thought of the concept of the molecule. A German physicist actually discovered the Avogadro number nine years after Avogadro's death.

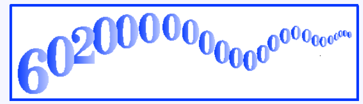
Comparing different compounds

If 6.02×10^{23} water molecules has a mass of 18.02 grams, how much does the same number of molecules of calcium carbonate weigh in grams? If we calculate the mass of the same number of molecules of each substance, we can compare the relative mass of each molecule. An example of how to calculate the formula mass of a compound is provided on the next page.

The formula mass of H_2O is 18.02 amu
18.02 amu = 1 molecule of H_2O
18.02 grams = 6.02×10^{23} molecules of H_2O

Figure 19.31: *The relationship between formula mass, atomic mass units, and grams.*

Science Facts: How large is Avogadro's number?



Imagine that every person on Earth was involved in counting the Avogadro number of atoms. How long do you think it would take? If all 6 billion people counted 3 atoms per second, it would take 1 million years to count 6.02×10^{23} atoms!

Example: Calculating the formula mass of a compound

What is the formula mass of calcium carbonate to the nearest hundredth?

1. Write the chemical formula for the compound.

calcium: Ca^{2+} carbonate: CO_3^{2-}

chemical formula: $CaCO_3$

2. List the atoms, number of each atom, and atomic mass of each atom.

atom	number	atomic mass	total mass (number x atomic mass)
Ca	1	40.08	40.08
C	1	12.01	12.01
O	3	16.00	48.00

3. Add up the values for each type of atom to calculate the formula mass.

$40.08 + 12.01 + 48.00 = 100.09$ amu

The formula mass of calcium carbonate is 100.09 amu.

How do you compare samples of substances?

The Avogadro number of calcium carbonate molecules would have a mass of 100.09 grams. In other words, if you used a balance to weigh 100.09 grams of calcium carbonate, there would be 6.02×10^{23} molecules of calcium carbonate in the sample. Likewise, if you used a balance to weigh 18.02 grams of water, there would be 6.02×10^{23} molecules of water in the weighed sample.

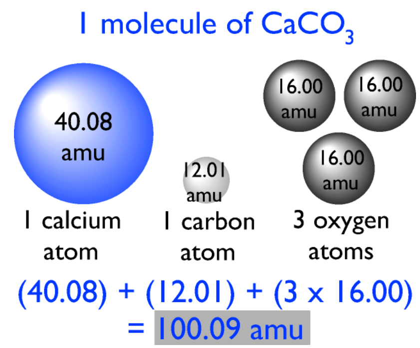


Figure 19.32: *Calculating the formula mass of calcium carbonate.*

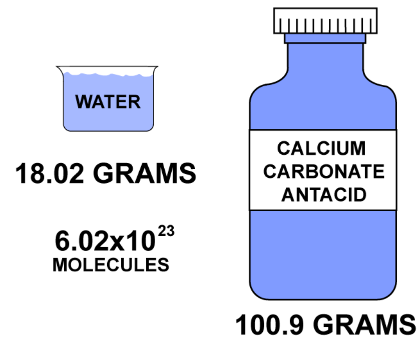


Figure 19.33: *100.09 g of $CaCO_3$ contains the same number of molecules as 18.02 g of H_2O.*

Consumer Chemistry: Hydrates and the Chemical Formulas

Hydrates are ionic compounds that contain precise numbers of water molecules

Have you ever bought a product that contained in the packaging a packet that was labeled: "Silica gel — do not eat"? These packets are often found inside boxes containing electronics equipment—like a DVD player or a stereo receiver. They are found inside shoeboxes, too. What is the purpose of these packets?

The presence of moisture in the packaging of certain products can be a problem. Manufacturers added packets of silica gel to absorb any such water. Ionic compounds like silicon oxide have the ability to incorporate water molecules as part of their structure. Water molecules become chemically bonded to their ions. A hydrate is a compound that has water molecules chemically bonded to its ions. Different compounds can absorb different numbers of molecules, as table 19.3 shows.

Table 19.3: ***Common hydrates***

Name	Formula
Silicon oxide monohydrate	$SiO_2 \cdot H_2O$
Barium chloride dihydrate	$BaCl_2 \cdot 2H_2O$
Calcium nitrate tetrahydrate	$Ca(NO_3)_2 \cdot 4H_2O$
Cobalt chloride hexahydrate	$CoCl_2 \cdot 6H_2O$
Magnesium sulfate heptahydrate	$MgSO_4 \cdot 7H_2O$
Iron (III) nitrate nonahydrate	$Fe(NO_3)_3 \cdot 9H_2O$

prefix	meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9

Figure 19.34: *Greek prefixes.*

Note that the chemical formula of a hydrate shows the ionic compound times a certain number of water molecules. This denotes a ratio of the number of molecules of water absorbed for each molecule of the compound. Note also that the name for each compound is followed by a Greek prefix indicating the number of water molecules and the word "hydrate." Figure 19.34 lists some Greek prefixes and their meanings.

Getting rid of the water molecules

You can remove the water molecules from a hydrate by heating it. When all the water leaves the hydrate through evaporation, it is **anhydrous**, a term that means "without water." Now that you know why packets of silica gel are included with some products, how could you *reuse* a packet of silica gel? How would you know when the packet of silica gel was anhydrous?

How do you calculate the formula mass of a hydrate?

It's easy to calculate the formula mass of a hydrate. First, calculate the formula mass of the ionic compound, then add the formula mass of water times as many molecules of water as are present. The example below shows you how to do this step by step.

Example: Calculating the formula mass of a hydrate

What is the formula mass of $BaCl_2 \times 2H_2O$?

1. Calculate the formula mass of the ionic compound

The ionic compound is $BaCl_2$. To calculate its formula mass:

1 Ba atom × 137.30 amu = 137.30 amu

2 Cl atoms × 35.45 amu = 70.90 amu

137.30 amu + 70.90 amu = 208.20 amu

2. Calculate formula mass of the water molecules

The formula mass for water is:

2 H atoms × 1.01 = 2.02 amu

1 O atom × 16.00 = 16.00 amu

2.02 + 16.00 = 18.02 amu

There are 2 molecules of water in the hydrate. The total formula mass is:

2 molecules H_2O × 18.02 = 36.04 amu

3. Calculate the formula mass of the hydrate

$BaCl_2 \times 2H_2O$ = 208.20+ 36.04 = 244.24 amu

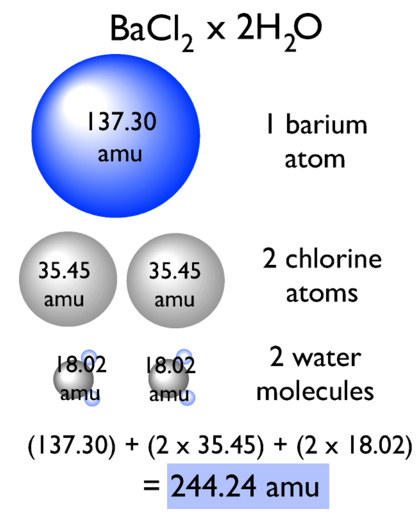


Figure 19.35: *Calculating the formula mass of a hydrate.*

Chapter 19 Review

Vocabulary review

Match the following terms with the correct definition. There is one definition extra in the list that will not match any term.

Set One			
1.	covalent bond	a.	The electrons involved in chemical bonding
2.	ionic bond	b.	Most atoms need eight valence electrons to be stable
3.	octet rule	c.	A bond between atoms in which electrons are lost or gained
4.	valence electrons	d.	A number that represents the number of electrons that are lost or gained in bonding
		e.	A bond between atoms in which electrons are shared

Set Two			
1.	Binary compound	a.	An ion like Na+, K+, or Cl-
2.	Monoatomic ion	b.	Electrons that are involved in bonding
3.	Oxidation number	c.	An ion like CO_3^{2-} or OH^-
4.	Polyatomic ion	d.	A number that indicates how many electrons will be gained or lost during bonding
		e.	A molecule composed of two monatomic ions

Concept review

1. Why do atoms tend to combine with other atoms instead of existing as single atoms?
2. Why are atoms in Group 18 considered to be chemically stable?
3. How can you determine the number of valence electrons by looking at the periodic table?
4. What conditions are met when an atom is chemically stable?
5. What is the major difference between ionic and covalent bonds?
6. Provide one general rule for predicting whether or not a bond will be ionic. (Hint: use the periodic table in your rule.)
7. What are polymers? Give an example of a natural polymer and a synthetic polymer.
8. What is an oxidation number? How can you determine an element's oxidation number by looking at the periodic table?
9. In a chemical formula, what do subscripts tell you?
10. What is the relationship between the formula mass of a compound, the Avogadro number of molecules of that compound, and the mass in grams of the compound?

Problems

1. Fill in the table below.

Element	Atomic number	Valence electrons	Lewis dot diagram
Fluorine			
Oxygen			
Phosphorus			
Carbon			
Beryllium			
Nitrogen			
Sulfur			
Neon			
Silicon			

2. Identify which of the following bonds are ionic or covalent and justify your reasoning.
   a. C-C
   b. Na-Br
   c. C-N
   d. C-O
   e. Ca-Cl

3. Fill in the table below.

Element	Number of valence electrons	Electrons gained or lost during ionization	Oxidation number
Potassium			
Aluminum			
Phosphorus			
Krypton			

4. Which group number on the periodic table is represented by each description?
   a. These atoms form compounds with ions that have an oxidation number of 1^-.
   b. The oxidation state of the atoms in this group is 3^-.
   c. Atoms in this group have four valence electrons in the outermost energy level. The atoms in this group form compounds with ions like H^+, Na^+ and Li^+.
   d. If these ions combined with Al^{3+}, you would need three of them and two aluminum ions in the formula.
   e. Atoms in this group lose two electrons during ionization.

5. Which of the following would be a correct chemical formula for a molecule of N^{3-} and H^+?
   a. HNO_3
   b. H_3N_6
   c. NH_3
   d. NH

6. What is the simplest ratio of carbon to hydrogen to oxygen in a molecule of glucose ($C_6H_{12}O_6$)?
   a. 1:2:1
   b. 6:12:6
   c. 2:4:2
   d. 6:2:6

7. What is the correct name for the compound $NaHCO_3$?
   a. sodium carbonide
   b. sodium hydrogen carbonate
   c. sodium hydrogen carboxide
   d. bicarbonate

8. Which of the following ion pairs would combine in a 1:2 ratio?
   a. NH^{4+} and F^-
   b. Be^{2+} and Cl^-
   c. sodium and hydroxide
   d. hydrogen and phosphate

9. For each of the compounds below, (1)state whether it is an empirical or molecular formula; and (2)if it is a molecular formula, write the empirical formula.
   a. CH_2O
   b. $(CH_2)_2(OH)_2$
   c. $C_7H_5NO_3S$
   d. $C_{10}H_8O_4$

10. Name each of the following binary covalent compounds.
    a. N_4O_6
    b. SiO_2
    c. S_2F_{10}
    d. $SbCl_5$

11. Write the chemical formula for the following compounds. Consult Table 19.2, "Polyatomic ions," on page 329 when needed.
    a. Sodium acetate
    b. Aluminum hydroxide
    c. Magnesium sulfate
    d. Ammonium nitrate
    e. Calcium fluoride

12. Calculate the formula mass for the following household compounds. Use the periodic table on the inside back cover.
    a. Lye drain cleaner, NaOH
    b. Epsom salts, $MgSO_4$
    c. Aspirin, $C_9H_8O_4$
    d. Plant fertilizer, $Ca(H_2PO_4)_2$
    e. Dampness absorber, $CaCl_2 \cdot 6H_2O$

13. Give the scientific name of compounds (a), (b), and (d) in question 10. Consult Table 19.2 on page 329.

Applying your knowledge

1. Many of the atoms on the periodic table have more than one oxidation number. You can figure out the oxidation number for these elements if you know at least one of the oxidation numbers in the compound. You only need to figure out what the oxidation number would be to make the molecule neutrally charged. Fill in the oxidation numbers for each of the following molecules.

Chemical formula of compound	Oxidation number for positive ion	Oxidation number for negative ion
SiO_2		2-
PBr_3		1-
$FeCl_3$		1-
CuF_2		1-
N_2O_3		2^-
P_2O_5		2^-

2. Suppose you are working in the lab with the following compounds: NaCl and Al_2O_3.
   a. Would the same number of molecules of each compound have the same mass? Explain your reasoning.
   b. Explain how you would measure out Avogadro's number of molecules (6.02×10^{23}) of each compound.
   c. Why is there a difference in the mass of the exact same number of molecules of each compound?

3. Find out about recycling plastics in your community. Prepare a pamphlet or bulletin board for your school that provides information on how to recycle plastics. The pamphlet or bulletin board should include practical information about recycling plastics including: how to get involved in community organizations, and what types of plastics are recycled. If your school does not recycle plastics, see if you can form a committee to develop and implement a plan.

4. Examine the household chemicals that are used in your own home. Make a list of the products your family uses, the names of the chemicals in each product and the hazards associated with each chemical. Research environmentally friendly alternatives to some of the products your family uses and prepare a brief presentation for your class.