Ionic Bonds and Formulas

Richard Parsons, (RichardP)

To access a customizable version of this book, as well as other interactive content, visit www.ck12.org

AUTHOR
Richard Parsons, (RichardP)

EDITOR
Shonna Robinson, (ShonnaR)

CONTRIBUTOR
Jonathan Edge, (JonathanE)

CK-12 Foundation is a non-profit organization with a mission to reduce the cost of textbook materials for the K-12 market both in the U.S. and worldwide. Using an open-content, web-based collaborative model termed the **FlexBook®**, CK-12 intends to pioneer the generation and distribution of high-quality educational content that will serve both as core text as well as provide an adaptive environment for learning, powered through the **FlexBook Platform®**.

Printed: October 20, 2013

flexbook
next generation textbooks

CHAPTER 1 Ionic Bonds and Formulas

CHAPTER OUTLINE

1.1 Ions and Ion Formation

Lesson Objectives

The student will:

- explain why atoms form ions.
- identify the atoms most likely to form positive ions and the atoms most likely to form negative ions.
- given the symbol of a main group element, indicate the most likely number of electrons the atom will gain or lose.
- predict the charge on ions from the electron affinity, ionization energies, and electron configuration of the atom.
- describe what polyatomic ions are.
- given the formula of a polyatomic ion, name it, and vice versa.

Vocabulary

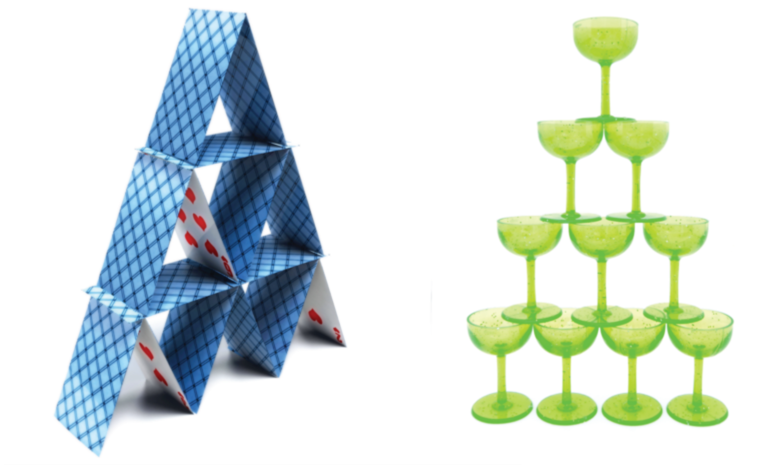
- polyatomic ion

Introduction

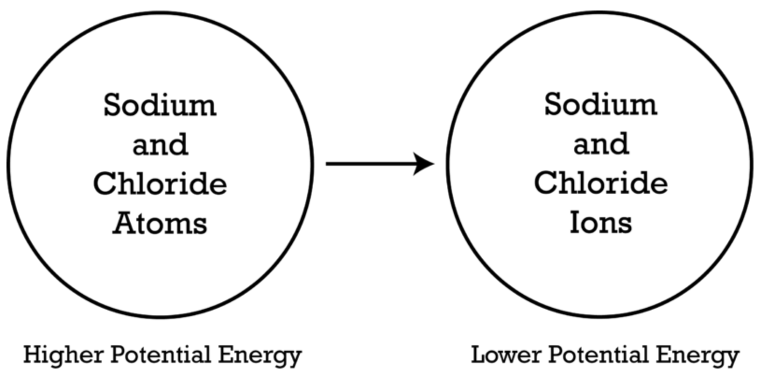
Before students begin the study of chemistry, they might think that the most stable form for an element is that of a neutral atom. As it happens, that particular idea is not true. There are approximately 190,000,000,000,000,000 kilotons of sodium in the earth, yet almost none of that is in the form of sodium atoms. Sodium reacts readily with oxygen in the air and explosively with water, so it must be stored under kerosene or mineral oil to keep it away from air and water. Essentially all of the sodium on earth that exists in its elemental form is man-made.

If those 1.9×10^{17} kilotons of sodium are not in the form of atoms, in what form are they? Virtually all the sodium on Earth is in the form of sodium ions, Na^+. The oceans of the earth contain a large amount of sodium ions in the form of dissolved salt, many minerals have sodium ions as one component, and animal life forms require a certain amount of sodium ions in their systems to regulate blood and bodily fluids, facilitate nerve function, and aid in metabolism. If sodium ions and not sodium atoms can be readily found in nature, it seems reasonable to suggest that ions are chemically more stable than atoms. By *chemically stable*, we mean less likely to undergo chemical change.

One of the major tendencies that causes change to occur in chemistry (and other sciences as well) is the tendency for matter to alter its condition in order to achieve lower potential energy. You can place objects in positions of higher potential energy, such as by stretching a rubber band or pushing the south poles of two magnets together, but if you want them to remain that way, you must hold them there. If you release the objects, they will move toward lower potential energy.



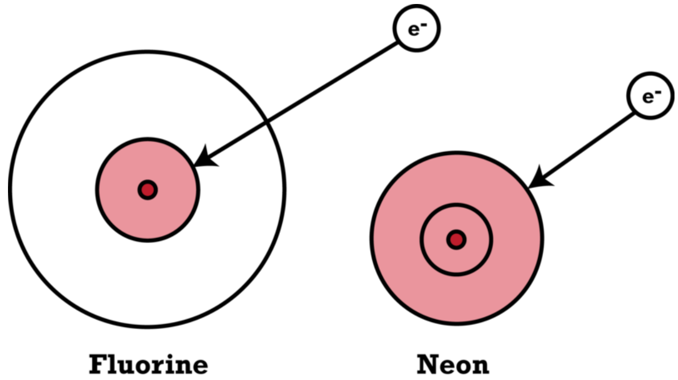
As another example, you can build a house of playing cards or a pyramid of champagne glasses that will remain balanced (like the ones pictured above), provided no one wiggles the table. If someone does wiggle the table, the structures will fall to lower potential energy. In the case of atoms and molecules, the particles themselves have constant random motion. For atoms and molecules, this molecular motion is like constantly shaking the table.



Comparing a system that contains sodium atoms and chlorine atoms to a system that contains sodium ions and chloride ions, we find that the system containing the ions has lower potential energy. This is due to the random motion of the atoms and molecules, which causes collisions between the particles. These collisions are adequate to initiate the change to lower potential energy.

Ion Formation

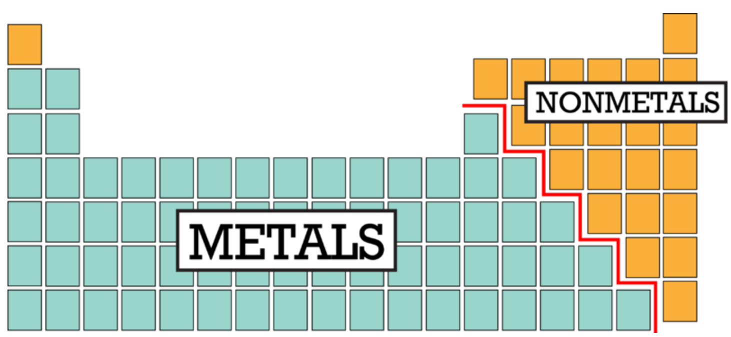
Recall that an atom becomes an ion when it gains or loses electrons. Cations are positively charged ions that form when an atom loses electrons, and anions are negatively charged ions that form when an atom gains electrons. Ionization energies and electron affinities control which atoms gain electrons, which atoms lose electrons, and how many electrons an atom gains or loses. At this point, you should already know the general trends of ionization energy and electron affinity in the periodic table (refer to the chapter "Chemical Periodicity" for more details about these trends).



An atom's attraction for adding electrons is related to how close the new electron can approach the nucleus of the atom. In the case of fluorine (electron configuration $1s^2 2s^2 2p^5$), the first energy level is full but the second one is not full. This allows an approaching electron to penetrate the second energy level and approach the first energy level and the nucleus. In the case of neon, both the first energy level and the second energy levels are full. This means that an approaching electron cannot penetrate either energy level. Looking at these situations sketched in the figure above, it is apparent that the approaching electron can get much closer to the nucleus of fluorine than it can with neon. Neon, in fact, has zero electron affinity. In comparison, the electron affinity of fluorine is -328 kJ/mole.

Spontaneous changes occur when accompanied by a decrease in potential energy. Without the decrease in potential energy, there is no reason for the activity to occur. When fluorine takes on an extra electron, it releases energy and moves toward lower potential energy. If neon took on an extra electron, there would be no decrease in potential energy, which is why neon does not spontaneously attract additional electrons. In comparison, the electron affinity of sodium is $+52.8$ kJ/mole. This means energy must be put in to force a sodium atom to accept an extra electron. Forcing sodium to take on an extra electron is not a spontaneous change because it requires an increase in potential energy.

Metals and Nonmetals



Metals, the atoms found on the left side of the table, have low ionization energies and low electron affinities. Therefore, they will lose electrons fairly readily, but they tend not to gain electrons. The atoms designated as nonmetals, the ones on the right side of the table, have high ionization energies and high electron affinities. Thus, they will not lose electrons, but they will gain electrons. The noble gases have high ionization energies and low electron affinities, so they will neither gain nor lose electrons. The noble gases were called inert gases (because they wouldn't react with anything) until 1962, when Neil Bartlett used very high temperature and pressure to force xenon and fluorine to combine. With a few exceptions, metals tend to lose electrons and become cations, while nonmetals tend to gain electrons and become anions. Noble gases tend to do neither.

In many cases, all that is needed to transfer one or more electrons from a metallic atom to a nonmetallic one is for the atoms bump into each other during their normal random motion. This collision at room temperature is sufficient to remove an electron from an atom with low ionization energy, and that electron will immediately be absorbed by an atom with high electron affinity. Adding the electron to the nonmetal causes a release of energy to the surroundings. The energy release that occurs by adding this electron to an atom with high electron affinity is greater than the energy release that would occur if this electron returned to the atom from which it came. Hence, this electron transfer is accompanied by a lowering of potential energy. This complete transfer of electrons produces positive and negative ions, which then stick together due to electrostatic attraction.

Numbers of Electrons Gained or Lost

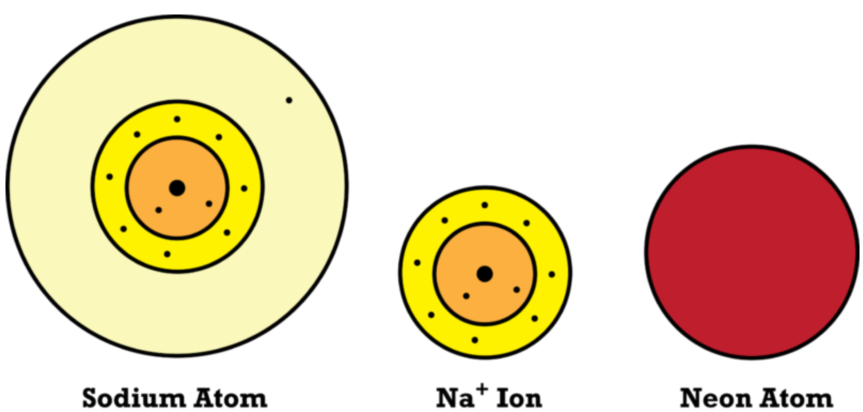
So far, we have been considering the ionization energy of atoms when one electron is removed. It is possible to continue removing electrons after the first one is gone. When a second electron is removed, the energy required is called the second ionization energy. The energy required to remove a third electron is called the third ionization energy, and so on. **Table** 1.1 shows the first four ionization energies for the atoms sodium, magnesium, and aluminum. As a reminder, the electron configurations for these atoms are:

Na: $1s^2 2s^2 2p^6 3s^1$
Mg: $1s^2 2s^2 2p^6 3s^2$
Al: $1s^2 2s^2 2p^6 3s^2 3p^1$

TABLE 1.1: The first four ionization energies of selected atoms

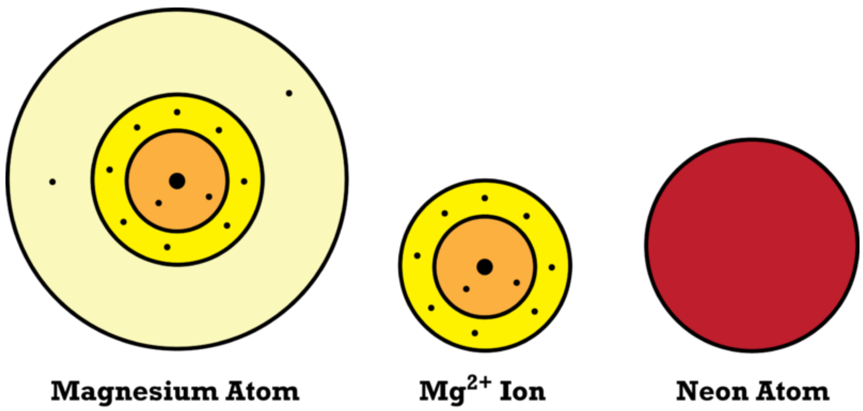
Atom	1st Ionization Energy (kJ/mole)	2nd Ionization Energy (kJ/mole)	3rd Ionization Energy (kJ/mole)	4th Ionization Energy (kJ/mole)
Na	496	4562	6912	9643
Mg	738	1450	7732	10,540
Al	578	1816	2745	11,577

In the chapter "Chemical Periodicity," we learned that $IE_1 < IE_2 < IE_3 < IE_4$. If we examine the size that the ionization energy increases, however, and use that information along with the electron configurations and the type of ion formed, we can gain new insight. For each atom, there is one increase in ionization energies where the next ionization energy is at least four times the previous one. In the case of sodium, this very large jump in ionization energy occurs between the first and second ionization energy. For magnesium, the huge jump occurs between the second and third ionization energies, and for aluminum, it is between the third and fourth ionization energies. If we combine this information with the fact that sodium only forms a $+1$ ion, magnesium only forms a $+2$ ion, and aluminum only forms a $+3$ ion, we have a consistency in our observations that allows us to suggest an explanation.



The diagram above shows the electron distributions for a sodium atom and a Na^+ ion. For the sodium atom, the first two energy levels are full, and the third energy level contains only a single electron. When we remove the first electron from a sodium atom, we are removing the electron in the third energy level because it is the furthest from the nucleus and thus has the lowest ionization energy. When that electron is removed, the third energy level is no longer available for electron removal. The sodium ion that remains has the same electron configuration as a neon atom. Although this Na^+ ion and a neon atom will have the same electron configuration, the Na^+ ion has a greater ionization energy than neon does because the sodium ion has one more proton in the nucleus. The sodium ion will also be slightly smaller than a neon atom (as indicated by the image above). When you have removed all the electrons in the outer energy level of an atom, the value of the next ionization energy will increase greatly because the next electron must be removed from a lower energy level.

Let's consider the same picture for magnesium.



The magnesium atom has two electrons in the outermost energy level. When those two are removed, the resulting Mg^{2+} ion has the same electron configuration as neon does, but it is smaller than neon because the magnesium ion has two more protons in the nucleus. The first two ionization energies for magnesium are relatively small, but the third ionization is five times as large as the second. As a result, a magnesium atom can lose the first two electrons relatively easily, but it does not lose a third.

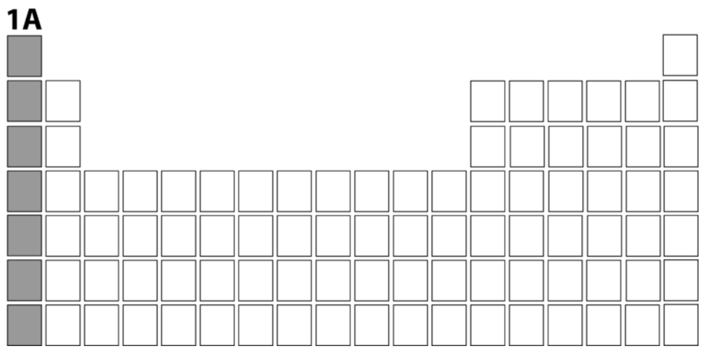
The huge jump in ionization energies is so consistent that we can identify the family of an unknown atom just by considering its ionization energies. If we had an unknown atom whose ionization energies were $IE_1 = 500 \text{ kJ/mol}$, $IE_2 = 1000 \text{ kJ/mol}$, $IE_3 = 2000 \text{ kJ/mol}$, and $IE_4 = 12,000 \text{ kJ/mole}$, we would immediately identify this atom as a member of family 3A. The large jump occurs between the 3rd and 4th ionization energies, so we know that only the first three electrons can be easily removed from this atom.

The logic for the formation of anions is very similar to that for cations. A fluorine atom, for example, has a high electron affinity and an available space for one electron in its outer energy level. When a fluorine atom takes on an electron, the potential energy of the fluorine ion is less than the potential energy of a fluorine atom. The fluoride ion that is formed has the same electron configuration as neon does, but it will be slightly larger than a neon atom because it has one less proton in the nucleus. As a result, the energy levels will not be pulled in as tightly. The electron affinity of a fluoride ion is essentially zero; the potential energy does not lower if another electron is added, so fluorine will take on only one extra electron.

An oxygen atom has a high electron affinity and has two spaces available for electrons in its outermost energy level. When oxygen takes on one electron, the potential energy of the system is lowered and energy is given off, but this oxygen ion has not filled its outer energy level; therefore, another electron can penetrate that electron shell. The oxygen ion (O^-) can accept another electron to produce the O^{2-} ion. This ion has the same electron configuration as neon does, and it will require an input of energy to force this ion to accept another electron.

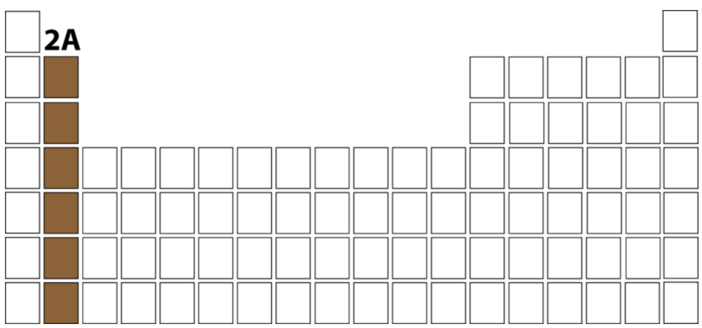
Some Common Ions

All the metals in family 1A (shown in the figure below) have electron configurations ending with a single *s* electron in the outer energy level. For that reason, all members of the 1A family will tend to lose only one electron when ionized. The entire family forms +1 ions: Li^+, Na^+, K^+, Rb^+, Cs^+, and Fr^+. Note that although hydrogen (H) is in this same column, it is not considered to be a metal. There are times when hydrogen acts like a metal and forms +1 ions, but most of the time it bonds with other atoms as a nonmetal. In other words, hydrogen doesn't easily fit into any chemical family.

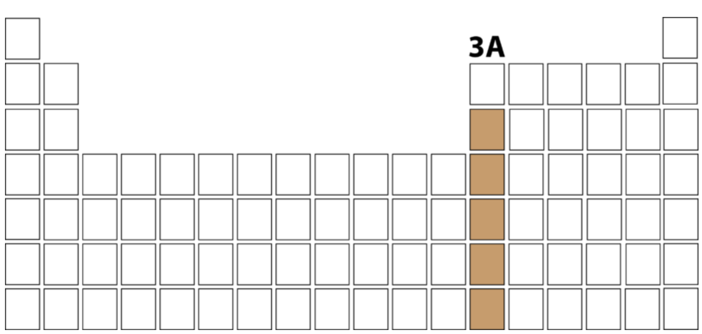


The metals in family 2A (shown in the figure below) all have electron configurations ending with two *s* electrons in the outermost energy level. This entire family will form +2 ions: Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Ra^{2+}.

All members of family 2A form ions with 2+ charge.



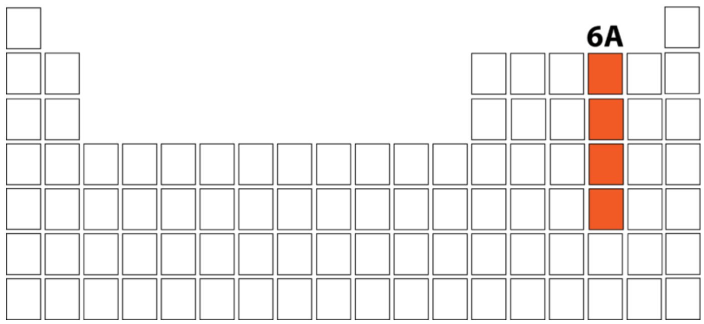
Family 3A members (shown in the figure below) have electron configurations ending in s^2p^1. When these atoms form ions, they will almost always form 3+ ions: Al^{3+}, Ga^{3+}, In^{3+}, and Tl^{3+}. Notice that boron is omitted from this list. This is because boron falls on the nonmetal side of the metal/nonmetal dividing line. Boron generally doesn't lose all of of its valence electrons during chemical reactions.



Family 4A is almost evenly divided into metals and nonmetals. The larger atoms in the family (germanium, tin, and lead) are metals. Since these atoms have electron configurations that end in s^2p^2, they are expected to form ions with charges of +4. All three of the atoms do form such ions (Ge^{4+}, Sn^{4+}, and Pb^{4+}), but tin and lead also have the ability to also form +2 ions. You will learn later in this chapter that some atoms have the ability to form ions of different charges, and the reasons for this will be examined later.

Like family 4A, the elements of family 5A are also divided into metals and nonmetals. The smaller atoms in this family behave as nonmetals, and the larger atoms behave as metals. Since bismuth and arsenic both have electron configurations that end with s^2p^3, they form +5 ions.

Most of the elements in family 6A (shown in figure below) are nonmetals that have electron configurations ending with s^2p^4. These atoms generally have enough electron affinity to attract two more electrons to fill their outermost energy level. They form −2 ions: O^{-2}, S^{-2}, Se^{-2}, and Te^{-2}.



Family 7A are all nonmetals with high electron affinities and electron configurations that end with s^2p^5. When these atoms form ions, they form −1 ions: F^-, Cl^-, Br^-, and I^-. Family 8A, of course, is made up of the noble gases, which have no tendency to either gain or lose electrons.

Polyatomic Ions

Thus far, we have been dealing with ions made from single atoms. Such ions are called monatomic ions. There are also **polyatomic ions**, which are composed of a group of covalently bonded atoms that behave as if they were a single ion. Almost all the common polyatomic ions are negative ions. The only common positive polyatomic ion is ammonium, NH_4^+. The name and formula of ammonium ion is similar to ammonia (NH_3), but it is *not* ammonia, and you should not confuse the two. The following is a list of common polyatomic ions that you should be familiar with.

- Ammonium ion, NH_4^+
- Acetate ion, $C_2H_3O_2^-$
- Carbonate ion, CO_3^{2-}
- Chromate ion, CrO_4^{2-}
- Dichromate ion, $Cr_2O_7^{2-}$
- Hydroxide ion, OH^-
- Nitrate ion, NO_3^-
- Phosphate ion, PO_4^{3-}
- Sulfate ion, SO_4^{2-}
- Sulfite ion, SO_3^{2-}

You should know these well enough so that when someone says the name of a polyatomic ion, you can respond with the formula and charge, and if someone shows you the formula and charge, you can respond with the name.

Lesson Summary

- Ions are atoms or groups of atoms that carry electrical charge.
- A negative ion is called an anion, and a positive ion is called a cation.
- Atoms with low ionization energy and low electron affinity (metals) tend to lose electrons and become positive ions.
- Atoms with high ionization energy and high electron affinity (nonmetals) tend to gain electrons and become negative ions.
- Atoms with high ionization energy and low electron affinity (noble gases) tend to neither gain nor lose electrons.
- Atoms that tend to lose electrons will generally lose all the electrons in their outermost energy level.
- Atoms that tend to gain electrons will gain enough electrons to completely fill the *s* and *p* orbitals in their outermost energy level.
- Polyatomic ions are ions composed of a group of atoms that are covalently bonded and behave as if they were a single ion.

Review Questions

1. Define an ion.
2. In general, how does the ionization energy of metal compare to the ionization energy of a nonmetal?
3. Will an iron atom form a positive or negative ion? Why?
4. Will a bromine atom form a positive or negative ion? Why?
5. Which is larger, a fluorine atom or a fluoride ion?
6. How is the number of valence electrons of a metal atom related to the charge on the ion the metal will form?
7. How is the number of valence electrons of a nonmetal related to the charge on the ion the nonmetal will form?
8. If carbon were to behave like a metal and give up electrons, how many electrons would it give up?
9. How many electrons are in a typical sodium ion?
10. Explain why chlorine is a small atom that tends to take on an extra electron, but argon is an even smaller atom that does not tend to take on electrons.
11. If an atom had the following successive ionization energies, to which family would it belong? Why did you choose this family?

$$\text{1st ionization energy} = 75 \text{ kJ/mol}$$
$$\text{2nd ionization energy} = 125 \text{ kJ/mol}$$
$$\text{3rd ionization energy} = 1225 \text{ kJ/mol}$$
$$\text{4th ionization energy} = 1750 \text{ kJ/mol}$$

1.2 Ionic Compounds

Lesson Objectives

The student will:

- describe how atoms form an ionic bond.
- state, in terms of energy, why atoms form ionic bonds.
- state the octet rule.
- give a brief description of a lattice structure.
- identify distinctive properties of ionic compounds.

Vocabulary

- electrostatic attraction
- ionic bond
- lattice structure
- octet rule

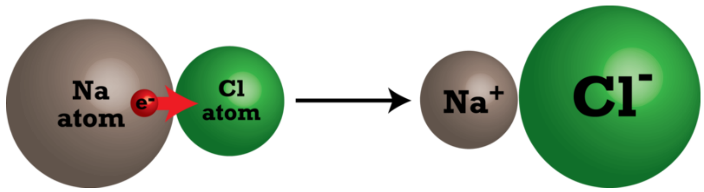
Introduction

Collisions between atoms that tend to lose electrons (metals) and atoms that tend to gain electrons (nonmetals) are usually sufficient enough to remove the electrons from the metal atom and add them to the nonmetal atom. This transfer of electrons forms positive and negative ions, which in turn attract each other due to opposite charges. The compounds formed by this electrostatic attraction are said to be ionically bonded.

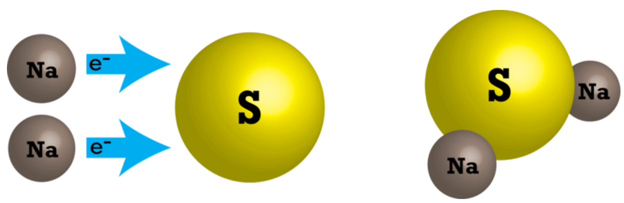
Ionic Bonding

When an atom with a low ionization energy encounters an atom with high electron affinity, it is possible for an electron transfer to occur. The ionization of the metal atom requires an input of energy. This energy input is often accomplished simply by the collision of atoms due to particle motion. Once electrons have been removed from the metal atoms, the electrons are taken on by the nonmetal atoms and energy is released. The energy released provides sufficient energy for the reaction to continue. In some cases, a reaction of this sort must be heated in order to start the reaction, but once the reaction begins, the reaction itself provides enough energy to continue.

The process of transferring an electron from a sodium atom to a chlorine atom, as shown in the sketch below, produces oppositely charged ions, which then stick together because of **electrostatic attraction**. Electrostatic attraction is the attraction between opposite charges. The electrostatic attraction between oppositely charged ions is called an **ionic bond.** Notice in the sketch above that the sodium atom is larger than the chlorine atom before the collision, but after the electron transfer, the sodium ion is now smaller than the chloride ion. Recall that the sodium ion is smaller than a neon atom because it has one more proton in the nucleus than neon does, yet they both have the same electron configuration. The chloride ion is larger than an argon atom because while it has the same electron configuration as argon, it has one less proton in the nucleus than argon. The sodium ion now has high ionization energy and low electron affinity (just like a noble gas) so there is no reason for any further changes. The same is true for the chloride ion. These ions are chemically more stable than the atoms are.



If we had been examining sodium and sulfur atoms, the transfer process would be only slightly different. Sodium atoms have a single electron in their outermost energy level and therefore can lose only one electron. Sulfur atoms, however, require two electrons to complete their outer energy level. In such a case, two sodium atoms would be required to collide with one sulfur atom, as illustrated in the diagram below. Each sodium atom would contribute one electron for a total of two electrons, and the sulfur atom would take on both electrons. The two Na atoms would become Na^+ ions, and the sulfur atom would become a S^{2-} ion. Electrostatic attractions would cause all three ions to stick together.



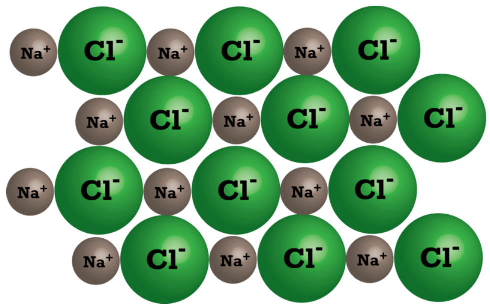
All the valence electrons for the main group elements are in *s* and *p* orbitals. When forming ions, main group metals will lose all of their valence electrons so that the resulting electron configuration will be the same as the previous noble gas. Usually, this means that the ion will have eight valence electrons. (Metals in the second row will form ions that have helium's electron configuration, which contains only two electrons.) Conversely, when nonmetals gain electrons to form anions, the new electron configuration will be the same as the following noble gas. The **octet rule** is an expression of the fact that when main group elements form ions, they tend to achieve a set of 8 valence electrons, which we know is a particularly stable configuration.

Properties of Ionic Compounds

When ionic compounds are formed, we are almost never dealing with just a single positive ion or a single negative ion. When ionic compounds are formed in laboratory conditions, many cations and anions are formed at the same time. The positive and negative ions are not just attracted to a single oppositely charged ion. The ions are attracted to several of the oppositely charged ions. The ions arrange themselves into organized patterns where each ion is surrounded by several ions of the opposite charge.

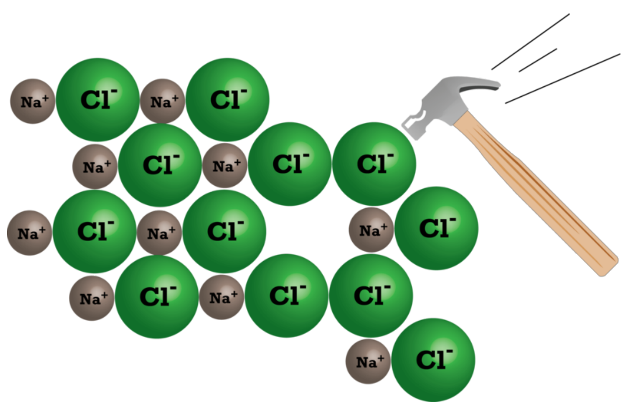
The organized patterns of positive and negative ions are called **lattice structures**. There are a number of different ionic lattice structures. The lattice structure that forms for a particular ionic compound is determined by the relative sizes of the ions and by the charge on the ions. Because ionic compounds form these large lattice structures in the solid phase, they are not referred to as molecules, but rather as lattice structures or crystals.

The image below shows the solid structure of sodium chloride. Only one layer is shown. When layers are placed above and below this one, each sodium ion would be touching six chloride ions – the four surrounding ones, one above, and one below. Each chloride ion will be touching six sodium ions in the same way.



When electrons are transferred from metallic atoms to nonmetallic atoms during the formation of an ionic bond, the electron transfer is permanent. The electrons now belong to the nonmetallic ion. If the ionic lattice structure is taken apart by melting it to a liquid, vaporizing it to a gas, or dissolving it in water, the particles come apart in the form of ions. The electrons that were transferred go with the negative ion when the ions separate. The electrostatic attraction between the oppositely charge ions is quite strong, and therefore ionic compounds have very high melting and boiling points. Sodium chloride (table salt), for example, must be heated to around 800°C to melt and around 1500°C to boil.

If you look again at the image, you can see that negative ions are surrounded by positive ions and vice versa. If part of the lattice is shifted downward, negative ions will then be next to negative ions. Since like charges repel, the structure will break up. As a result, ionic compounds tend to be brittle solids. If you attempt to hammer down on ionic substances, they will shatter. This is very different from metals, which can be hammered into different shapes without the metal atoms separating from each other.



Ionic substances generally dissolve readily in water. When an ionic compound has been melted or dissolved in water, there are ions present that have the ability to move around in the liquid. It is specifically the presence of the mobile ions that allow electric current to be conducted by ionic liquids and ionic solutions. In comparison, non-ionic compounds that are dissolved in water or are in liquid form do not conduct electric current.

The process of gaining or losing electrons completely changes the chemical properties of the substances. The chemical and physical properties of an ionic compound will bear no resemblance to the properties of the elements which formed the ions. For example, sodium is a metal that is shiny, an excellent conductor of electric current, and reacts violently with water. Chlorine is a poisonous gas. When sodium and chlorine are chemically combined to form sodium chloride (table salt), the product has an entirely new set of properties. We could sprinkle sodium chloride on our food, which is not something we would do if we expected it to poison us or to explode when it touches water.

Lesson Summary

- Ionic bonds are the resulting electrostatic attraction holding ions together when electrons transfer from metal atoms to nonmetal atoms.
- The octet rule is an expression of the tendency for atoms to gain or lose the appropriate number of electrons so that the resulting ion has either completely filled or completely empty outer energy levels.
- Ionic compounds form ionic crystal lattices rather than molecules.
- Ionic compounds have very high melting and boiling points.
- Ionic compounds tend to be brittle solids.
- Ionic compounds are generally soluble in water, and the resulting solutions will conduct electricity.
- Ionic compounds have chemical properties that are unrelated to the chemical properties of the elements from which they were formed.

Further Reading / Supplemental Links

This website provides more information about ionic compounds.

- http://misterguch.brinkster.net/ionic.html

This video is a ChemStudy film called "Electric Interactions in Chemistry." The film is somewhat dated but the information is accurate.

- http://www.youtube.com/watch?v=o9TaQLVCFDM

Review Questions

1. What takes place during the formation of an ionic bond?
2. What effect does the transfer of electrons have on the nuclei of the atoms involved?
3. Hydrogen gas and chlorine gas are not acids, but when hydrogen and chlorine combine to form hydrogen chloride, the compound is an acid. How would you explain this?
4. Why do we not refer to molecules of sodium chloride?

1.3 Writing Ionic Formulas

Lesson Objectives

The student will:

- provide the correct formulas for binary ionic compounds.
- provide the correct formulas for compounds containing metals with variable oxidation numbers.
- provide the correct formulas for compounds containing polyatomic ions.

Vocabulary

- empirical formula
- formula unit

Introduction

Ionic compounds do not exist as molecules. In the solid state, ionic compounds are in crystal lattices containing many cations and anions. An ionic formula, like NaCl, is an empirical formula. The **empirical formula** gives the simplest whole number ratio of atoms of each element present in the compound. The formula for sodium chloride merely indicates that it is made of an equal number of sodium and chloride ions. As a result, it is technically incorrect to refer to a molecule of sodium chloride. Instead, one unit of NaCl is called the **formula unit**. A formula unit is one unit of an empirical formula for an ionic compound.

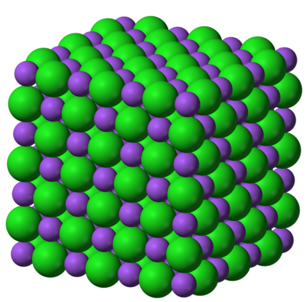


FIGURE 1.1
The three-dimensional crystal lattice structure of sodium chloride.

Sodium sulfide, another ionic compound, has the formula Na_2S. This formula indicates that this compound is made up of twice as many sodium ions as sulfide ions. Na_2S will also form a crystal lattice, but the lattice won't be the same as the NaCl lattice because the Na_2S lattice has to have two sodium ions per sulfide ion.

Predicting Formulas for Ionic Compounds

Determining Ionic Charge

The charge that will be on an ion can be predicted for most of the monatomic ions. Many of these ionic charges can be predicted for entire families of elements. There are a few ions whose charge must simply be memorized, and there are also a few that have the ability to form two or more ions with different charges. In these cases, the exact charge on the ion can only be determined by analyzing the ionic formula of the compound.

All of the elements in family 1A are metals that have the same outer energy level electron configuration, the same number of valence electrons (one), and low first ionization energies. Therefore, these atoms will lose their one valence electron and form ions with a +1 charge. This allows us to predict the ionic charges on all the 1A ions: Li^+, Na^+, K^+, Rb^+, Cs^+, and Fr^+.

Hydrogen is a special case. It has the ability to form a positive ion by losing its single valence electron, just as the 1A metals do. In those cases, hydrogen forms the +1 ion, H^+. In rare cases, hydrogen can also take on one electron to complete its outer energy level. These compounds, such as NaH, KH, and LiH, are called hydrides. Hydrogen also has the ability to form compounds without losing or gaining electrons, which will be discussed in more details in the chapter "Covalent Bonds and Formulas."

All of the elements in family 2A have the same outer energy level electron configuration and the same number of valence electrons (two). Each of these atoms is a metal with low first and second ionization energies. Therefore, these elements will lose both of its valence electrons to form an ion with a +2 charge. The ions formed in family 2A are: Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Ra^{2+}.

There is a slight variation for the elements in family 3A. The line separating metals from nonmetals on the periodic table cuts through family 3A between boron and aluminum. In family 3A, boron, $1s^22s^22p^1$, behaves as a nonmetal due to its higher ionization energy and higher electron affinity. Aluminum, on the other hand, is on the metallic side of the line and behaves as an electron donor. Aluminum, $1s^22s^22p^63s^23p^1$, always loses all three of its valence electrons and forms an Al^{3+} ion. Gallium and indium have the same outer energy level configuration as aluminum, and they also lose all three of their valence electrons to form the +3 ions Ga^{3+} and In^{3+}. Thallium, whose electron configuration ends with $6s^26p^1$, could also form a +3 ion, but for reasons beyond the scope of this book, thallium is more stable as the +1 ion Tl^+.

All the elements in the 6A family have six valence electrons, and they all have high electron affinities. These atoms will, therefore, take on additional electrons to complete the octet of electrons in their outer energy levels. Since each atom will take on two electrons to complete its octet, members of the 6A family will form −2 ions. The ions formed will be: O^{2-}, S^{2-}, Se^{2-}, and Te^{2-}.

Family 7A elements have the outer energy level electron configuration ns^2np^5. These atoms have the highest electron affinities in their periods and will each take on one more electron to complete the octet of electrons in the outer energy levels. Therefore, these atoms will form −1 ions: F^-, Cl^-, Br^-, I^-, and At^-.

Family 8A elements have completely filled outer energy levels. Because of this, it is very energetically unfavorable to either add or remove electrons, and elements found in family 8A do not form ions.

Transition Elements

There are greater variations in the charge found on ions formed from transition elements. Many of the transition elements can form ions with different charges. We will consider some of these elements later in this chapter. There are also some transition elements that only form one ion.

Consider the electron configuration of silver, Ag, is $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^14d^{10}$. When forming an ion, silver loses its single valence electron to produce Ag^+. Note that its electron configuration does not exactly follow the rules that you have been shown for filling orbitals. Electrons have been placed in 4 d orbitals even though the 5s orbital, which is usually lower in energy, is not completely full. Because the 4d and 5s orbitals are so similar in energy, very small perturbations can sometimes make it energetically favorable for the 5s electron to move into a 4d orbital. It happens that a set of half full (5 electrons) or completely full (10 electrons) d orbitals gives a little extra stability to the electron configuration. This also happens with chromium, copper, molybdenum, and gold.

The electron configuration for Zn is $[Ar]4s^2 3d^{10}$. Like main group metals, zinc loses all of its valence electrons when it forms an ion, so it forms a Zn^{2+} ion. Cadmium is similar. The electron configuration for Cd is $[Kr]5s^2 3d^{10}$, so it forms a Cd^{2+} ion.

Writing Basic Ionic Formulas

In writing formulas for binary ionic compounds (binary refers to two elements, *not* two single atoms), the cation is always written first. Chemists use subscripts following the symbol of each element to indicate the number of that element present in the formula. For example, the formula Na_2O indicates that the compound contains two atoms of sodium for every one oxygen. When the subscript for an element is 1, the subscript is omitted. The number of atoms of an element with no indicated subscript is always read as 1. When an ionic compound forms, the number of electrons given off by the cations must be exactly the same as the number of electrons taken on by the anions. Therefore, if calcium, which gives off two electrons, is to be combined with fluorine, which takes on one electron, then one calcium atom must combine with two fluorine atoms. The formula would be CaF_2.

Suppose we wish to write the formula for the compound that forms between aluminum and chlorine. To write the formula, we must first determine the oxidation numbers of the ions that would be formed. We will revisit the concept of oxidation numbers later, but for now, all you need to know is that the oxidation number for an atom in an ionic compound is equal to the charge of the ion it produces.

$$\begin{array}{ccc} 3+ & \quad & 1- \\ \text{Al} & & \text{Cl} \end{array}$$

Then, we determine the simplest whole numbers with which to multiply these charges so they will balance (add to zero). In this case, we would multiply the 3+ by 1 and the 1− by 3.

$$\begin{array}{ccc} 3+ & \quad & 1- \\ \text{Al} & & \text{Cl} \\ (3+)(1) = 3+ & & (1-)(3) = 3- \end{array}$$

You should note that we could multiply the 3+ by 2 and the 1− by 6 to get 6+ and 6−, respectively. These values will also balance, but this is *not* acceptable because empirical formulas, by definition, must have the lowest whole number multipliers. Once we have the *lowest* whole number multipliers, those multipliers become the subscripts for the symbols. The formula for this compound would be $AlCl_3$.

Here's the process for writing the formula for the compound formed between aluminum and sulfur.

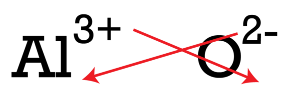
$$\begin{array}{ccc} 3+ & \quad & 2- \\ \text{Al} & & \text{S} \\ (3+)(2) = 6+ & & (2-)(3) = 6- \end{array}$$

Therefore, the formula for this compound would be Al_2S_3.

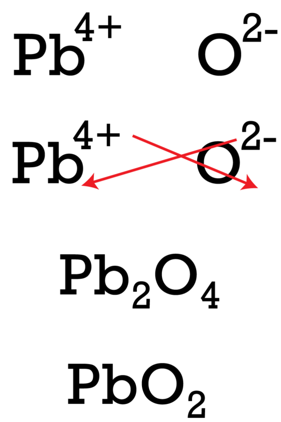
Another method used to write formulas is called the criss-cross method. It is a quick method, but it often produces errors if the user doesn't pay attention to the results. The example below demonstrates the criss-cross method for writing the formula of a compound formed from aluminum and oxygen. In the criss-cross method, the oxidation numbers are placed over the symbols for the elements just as before.

3+	2−
Al	O

In this method, the oxidation numbers are then criss-crossed and used as the subscripts for the other atom (ignoring sign).



This produces the correct formula Al_2O_3 for the compound. Here's an example of a criss-cross error:



If you used the original method of finding the lowest multipliers to balance the charges, you would get the correct formula PbO_2, but the criss-cross method produces the incorrect formula Pb_2O_4. If you use the criss-cross method to generate an ionic formula, it is essential that you check to make sure that the subscripts correspond to the lowest whole number ratio of the atoms involved. Note that this only applies to ionic compounds. When we learn about covalent compounds in the chapter "Covalent Bonds and Formulas," you will see that the formula N_2O_4 describes a different molecule than NO_2, so it would not be reduced to its simplest ratio.

Metals with Variable Oxidation Number

When writing formulas, you are given the oxidation number. When we get to naming ionic compounds, however, it is absolutely vital that you are able to recognize metals that can have more than one oxidation number. A partial list of metals with variable oxidation numbers includes iron, copper, tin, lead, nickel, and gold.

For example, iron can form both Fe^{2+} and Fe^{3+} ions. The electron configuration for neutral Fe is $1s^22s^22p^63s^23p^64s^23d^6$. It is fairly straightforward as to why iron forms the 2+ ion, because it loses all of its valence electrons like other metals do. The electron configuration for the Fe^{2+} ion is $[Ar]3d^6$. Why would the iron ion lose one more electron? Earlier, we mentioned that the *d* orbitals have a slightly lower energy when they are exactly half full or completely full. If this Fe^{2+} ion were to lose one more electron, the 3*d* orbitals would be exactly half full with five electrons. When iron reacts, it will first form Fe^{2+}. However, if the pull on its electrons is particularly strong, it will form Fe^{3+}.

Other examples of metals with variable oxidation states are less intuitive. For example, copper, silver, and gold (a single family of metals) can all lose a single electron to form Cu^+, Ag^+, and Au^+. All subshells in the resulting ions are completely full or empty. However, copper can also form Cu^{2+}, which is actually more stable in many cases. Gold can form Au^{3+}, but Au^{2+} is rarely observed. Silver, as we mentioned earlier, does not commonly lose more than one electron.

The oxidation states available to main group metals are generally easy to predict. However, tin and lead are two exceptions. In addition to losing all four of their valence electrons to make Sn^{4+} and Pb^{4+}, tin and lead will also commonly form Sn^{2+} and Pb^{2+} ions. There are many metals with variable oxidation states, but it is worth memorizing at least the ones mentioned here (Fe, Cu, Au, Sn, Pb).

Polyatomic Ions

Polyatomic ions require additional consideration when you write formulas involving them. Recall from earlier this list of common polyatomic ions:

- Ammonium ion, NH_4^+
- Acetate ion, $C_2H_3O_2^-$
- Carbonate ion, CO_3^{2-}
- Chromate ion, CrO_4^{2-}
- Dichromate ion, $Cr_2O_7^{2-}$
- Hydroxide ion, OH^-
- Nitrate ion, NO_3^-
- Phosphate ion, PO_4^{3-}
- Sulfate ion, SO_4^{2-}
- Sulfite ion, SO_3^{2-}

Suppose we are asked to write the formula for the compound that would form between calcium and the nitrate ion. We begin by putting the charges above the symbols just as before.

$$\begin{array}{cc} 2+ & 1- \\ \mathrm{Ca} & \mathrm{NO_3} \\ (2+)(1) = 2+ & (1-)(2) = 2- \end{array}$$

The multipliers needed to balance these ions are 1 for calcium and 2 for nitrate. We wish to write a formula that tells our readers that there are two nitrate ions in the formula for every calcium ion. When we put the subscript 2 beside the nitrate ion in the same fashion as before, we get something strange – $CaNO_{32}$. With this formula, we are indicating 32 oxygen atoms, which is wrong. The solution to this problem is to put parentheses around the nitrate ion before the subscript is added. Therefore, the correct formula is $Ca(NO_3)_2$. Similarly, calcium phosphate would be $Ca_3(PO_4)_2$. If a polyatomic ion does not need a subscript other than an omitted 1, then the parentheses are not needed. Although including these unnecessary parentheses does not change the meaning of the formula, it may cause the reader to wonder whether a subscript was left off by mistake. Try to avoid using parentheses when they are not needed.

Example:

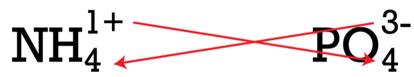
Write the formula for the compound that will form from aluminum and acetate.

$$Al^{3+} \qquad C_2H_3O_2^{1-}$$

The charge on an aluminum ion is +3, and the charge on an acetate ion is −1. Therefore, three acetate ions are required to combine with one aluminum ion. This is also apparent by the criss-cross method. However, we cannot place a subscript of 3 beside the oxygen subscript of 2 without inserting parentheses first. Therefore, the formula will be $Al(C_2H_3O_2)_3$.

Example:

Write the formula for the compound that will form from ammonium and phosphate.



The charge on an ammonium ion is +1 and the charge on a phosphate ion is −3. Therefore, three ammonium ions are required to combine with one phosphate ion. The criss-cross procedure will place a subscript of 3 next to the subscript 4. This can only be carried out if the ammonium ion is first placed in parentheses. Therefore, the proper formula is $(NH_4)_3PO_4$.

Example:

Write the formula for the compound that will form from aluminum and phosphate.

$$Al^{3+} \qquad PO_4{}^{3-}$$

Since the charge on an aluminum ion is +3 and the charge on a phosphate ion is −3, these ions will combine in a one-to-one ratio. In this case, the criss-cross method would produce an incorrect answer. Since it is not necessary to write the subscripts of 1, no parentheses are needed in this formula. Since parentheses are not needed, it is generally considered incorrect to use them. The correct formula is $AlPO_4$.

More Examples:

Magnesium hydroxide ......$Mg(OH)_2$

Sodium carbonate ....Na_2CO_3

Barium acetate ....$Ba(C_2H_3O_2)_2$

Ammonium dichromate ....$(NH_4)_2Cr_2O_7$

Lesson Summary

- The oxidation number for ions of the main group elements can usually be determined by the number of valence electrons.
- Some transition elements have fixed oxidation numbers, while others have variable oxidation numbers.
- Some metals, such as iron, copper, tin, lead, and gold, also have variable oxidation numbers.
- Formulas for ionic compounds contain the lowest whole number ratio of subscripts such that the sum of all positive charges equals the sum of all negative charges.

Further Reading / Supplemental Links

This website provides more details about ionic bonding, including a conceptual simulation of the reaction between sodium and chlorine. The website also discusses covalent bonding, the focus of the chapter "Covalent Bonds and Formulas."

- http://visionlearning/library/module_viewer.php?mid=55

Review Questions

1. Fill in the chart below (**Table** 1.2) by writing formulas for the compounds that might form between the ions in the columns and rows. Some of these compounds don't exist but you can still write formulas for them.

TABLE 1.2:

	Na^+	Ca^{2+}	Fe^{3+}	NH_4^+	Sn^{4+}
NO_3^-					
SO_4^{2-}					
Cl^-					
S^{2-}					
PO_4^{3-}					
OH^-					
$Cr_2O_7^{2-}$					
CO_3^{2-}					

1.4 Naming Ionic Compounds

Lesson Objectives

The student will:

- correctly name binary ionic compounds, compounds containing metals with variable oxidation numbers, and compounds containing polyatomic ions when given the formulas.
- provide chemical formulas for binary ionic compounds, compounds containing metals with variable oxidation numbers, and compounds containing polyatomic ions when given the names.

Introduction

It is necessary for each symbol and each name in chemistry to be completely unique. Using an incorrect substance in a chemistry experiment could have disastrous results, so the names and symbols of elements and compounds must refer to exactly one substance. For beginning students, the system of naming chemicals can seem impossibly complex. This section presents the rules for naming various ionic compounds.

Rules for Naming Ionic Compounds

When an atom gains or loses electrons to form an ion, its name sometimes changes. Main group metals retain their name when forming cations. For example, K^+ is a potassium ion, and Mg^{2+} is a magnesium ion. However, when nonmetallic elements gain electrons to form anions, the ending of their names is changed to "-ide." For example, a fluorine atom gains one electron to become a fluoride ion (F^-), and sulfur gains two electrons to become a sulfide ion (S^{2-}). Polyatomic ions have names that you simply need to memorize. A list of common polyatomic ions was presented earlier in the chapter.

Binary Ionic Compounds

Binary ionic compounds are compounds that contain only two kinds of ions, regardless of how many of each ion is present. To name such compounds, you simply write the name of the cation followed by the name of the anion. Unless you are dealing with a metal that can have multiple oxidation states, there is no need to indicate the relative number of cations and anions, since there is only one possible ratio that will give you a neutral compound.

Examples:

$MgCl_2$ . . . . magnesium chloride

NaBr . . . . sodium bromide

AlF_3 . . . . aluminum fluoride

K_2S . . . . potassium sulfide

CaI_2 .... calcium iodide

Rb_2O .... rubidium oxide

H_3N .... hydrogen nitride

Polyatomic Ions

When naming a compound containing a polyatomic ion, the name of the polyatomic ion does not change regardless of whether it is written first or last in the formula. If the formula contains a positive polyatomic ion and a nonmetal, the ending of the nonmetal is replaced with "-ide." If the compound contains a metal and a polyatomic ion, both the metal and the polyatomic ion are written without any changes to their names.

Examples:

$NaC_2H_3O_2$ .... sodium acetate

$Mg(NO_3)_2$ .... magnesium nitrate

$(NH_4)_2CrO_4$ .... ammonium chromate

$(NH_4)_2S$ .... ammonium sulfide

$Ca(OH)_2$ .... calcium hydroxide

$BaCr_2O_7$ .... barium dichromate

H_3PO_4 .... hydrogen phosphate

Variable Oxidation Number Metals

Metals with variable oxidation numbers may form multiple different compounds with the same nonmetal. Iron, for example, may react with oxygen to form either FeO or Fe_2O_3. These are very different compounds with different properties. When we name these compounds, it is absolutely vital that we clearly distinguish between them. They are both iron oxides, but in FeO, iron has an oxidation number of $+2$, while in Fe_2O_3, it has an oxidation number of $+3$. The rule for naming these compounds is to write the oxidation number of the metal after the name. The oxidation number is written using Roman numerals and is placed in parentheses. For these two examples, the compounds would be named iron(II) oxide and iron(III) oxide. When you see that the compound involves a metal with multiple oxidation numbers, you must determine the oxidation number of the metal from the formula and indicate it using Roman numerals.

In general, main group metal ions have only one common oxidation state, whereas most of the transition metals have more than one. However, there are plenty of exceptions to this guideline. Main group metals that can have more than one oxidation state include tin (Sn^{2+} or Sn^{4+}) and lead (Pb^{2+} or Pb^{4+}). Transition metals with only one common oxidation state include silver (Ag^+), zinc (Zn^{2+}), and cadmium (Cd^{2+}). These should probably be memorized, but when in doubt, include the Roman numerals for transition metals. Do not do this for main group metals that do not have more than one oxidation state. Referring to AgCl as silver(I) chloride is redundant and may be considered wrong. However, copper chloride is definitely incorrect, because it could refer to either CuCl or $CuCl_2$.

Other than the use of Roman numerals to indicate oxidation state, naming these ionic compounds is no different than what we have already seen. For example, consider the formula $CuSO_4$. We know that the sulfate anion has a charge of -2. Therefore, for this to be a neutral compound, copper must have a charge of +2. The name of this compound is copper(II) sulfate.

How about SnS_2? Tin is a variable oxidation number metal. We need a Roman numeral in the name of this compound. The oxidation number of sulfur is -2. Two sulfide ions were necessary to combine with one tin ion. Therefore, the oxidation number of the tin must be $+4$, and the name of this compound is tin(IV) sulfide.

Examples:

PbO . . . . lead(II) oxide

FeI_2 . . . . iron(II) iodide

$Fe_2(SO_4)_3$ . . . . . . iron(III) sulfate

$AuCl_3$ . . . . gold(III) chloride

CuO . . . . copper(II) oxide

PbS_2 . . . . lead(IV) sulfide

The most common error made by students in naming these compounds is to choose the Roman numeral based on the number of atoms of the metal. The Roman numeral in these names is the oxidation number of the metal. For example, in PbS_2, the oxidation state of lead (Pb) is +4, so the Roman numeral following the name lead is IV. Notice that there is no four in the formula. As in previous examples, the empirical formula is always the lowest whole number ratio of the ions involved. Think carefully when you encounter variable oxidation number metals. Make note that the Roman numeral does not appear in the formula but does appear in the name.

Lesson Summary

- Cations have the same name as their parent atom.
- Monatomic anions are named by replacing the end of the parent atom's name with "-ide."
- The names of polyatomic ions do not change.
- Ionic compounds are named by writing the name of the cation followed by the name of the anion.
- When naming compounds that include a metal with more than one common oxidation state, the charge of the metal ion is indicated with Roman numerals in parentheses between the cation and anion.

Review Questions

1. Name the following compounds.
   a. CaF_2
   b. $(NH_4)_2CrO_4$
   c. K_2CO_3
   d. $NaCl$
   e. PbO
   f. $CuSO_4$
   g. $Ca(NO_3)_2$
   h. $Mg(OH)_2$
   i. SnO_2
2. Write the formulas from the names of the following compounds.
   a. Sodium carbonate
   b. Calcium hydroxide
   c. Iron(III) nitrate
   d. Magnesium oxide
   e. Aluminum sulfide
   f. Copper(I) dichromate
   g. Ammonium sulfate
   h. Iron(II) phosphate
   i. Lead(IV) sulfate

1.5 References

1. Benjah-bmm27. 3D structure of sodium chloride. Public domain