Ions and Ion Formation

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CONCEPT 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

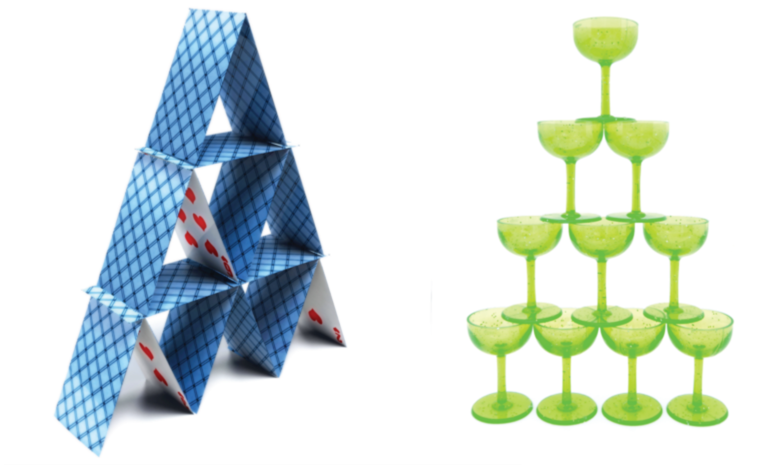
- polyatomic ion

Introduction

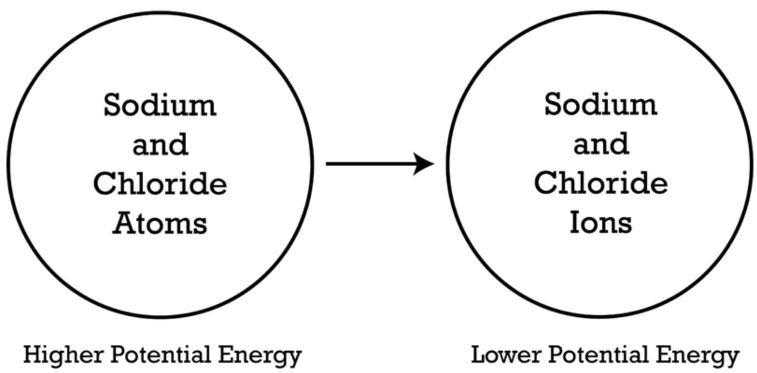
Before students begin the study of chemistry, they might think that the most stable form for an element is that of a neutral atom. As it happens, that particular idea is not true. There are approximately 190,000,000,000,000,000 kilotons of sodium in the earth, yet almost none of that is in the form of 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.



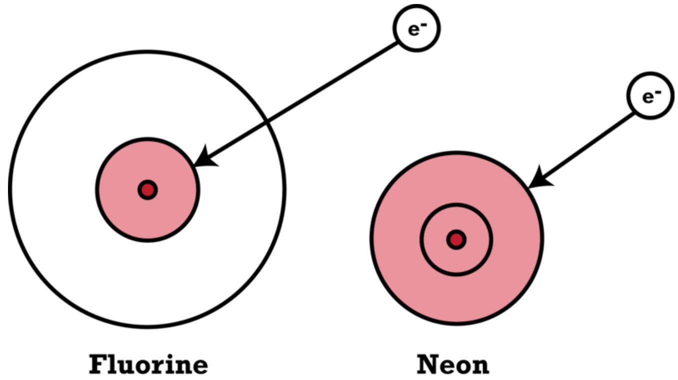
As another example, you can build a house of playing cards or a pyramid of champagne glasses that will remain balanced (like the ones pictured above), provided no one wiggles the table. If someone does wiggle the table, the structures will fall to lower potential energy. In the case of atoms and molecules,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

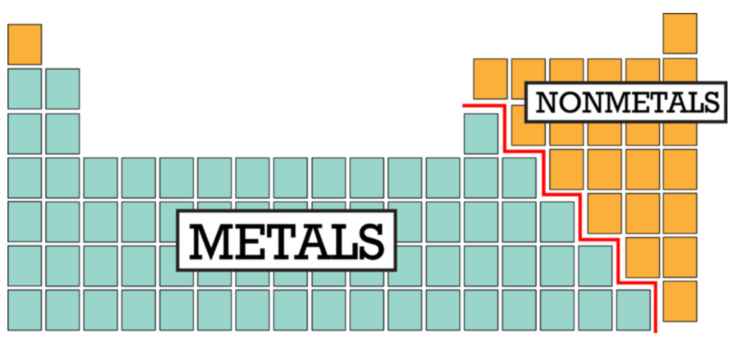
Recall that an atom becomes an ion when it gains or loses electrons. Cations are positively charged ions that form when an atom loses electrons, and anions are negatively charged ions that form when an atom gains electrons. Ionization energies and electron affinities control which atoms gain electrons, 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

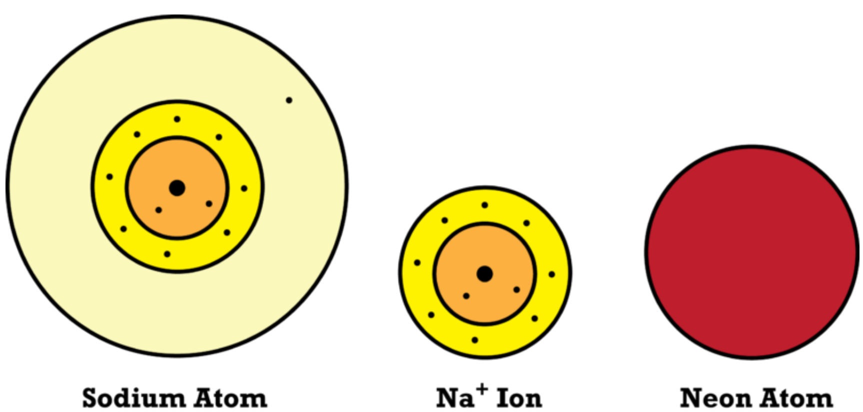
So far, we have been considering the ionization energy of atoms when one electron is removed. It is possible to continue removing electrons after the first one is gone. When a second electron is removed, the energy required is called the second ionization energy. The energy required to remove a third 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

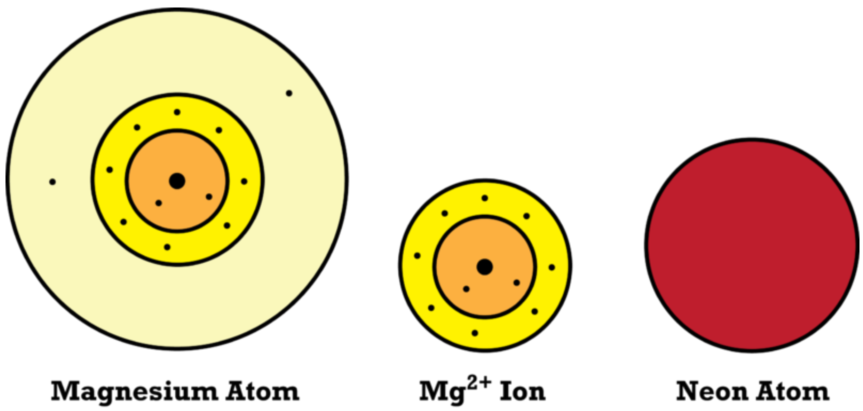
Atom	1st Ionization Energy (kJ/mole)	2nd Ionization Energy (kJ/mole)	3rd Ionization Energy (kJ/mole)	4th Ionization Energy (kJ/mole)
Na	496	4562	6912	9643
Mg	738	1450	7732	10,540
Al	578	1816	2745	11,577

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.

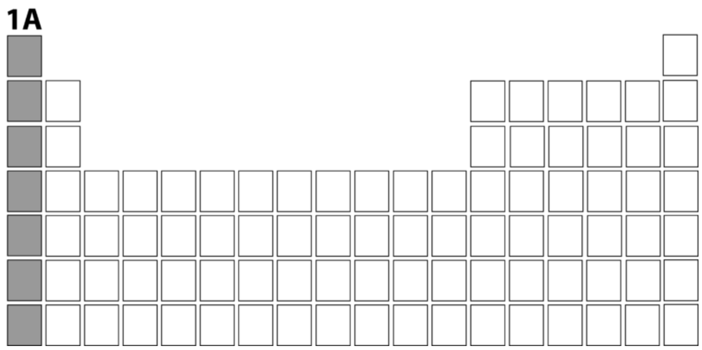
The huge jump in ionization energies is so consistent that we can identify the family of an unknown atom just by considering its ionization energies. If we had an unknown atom whose ionization energies were $IE_1 = 500\ kJ/mol$, $IE_2 = 1000\ kJ/mol$, $IE_3 = 2000\ kJ/mol$, and $IE_4 = 12,000\ kJ/mole$, 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.

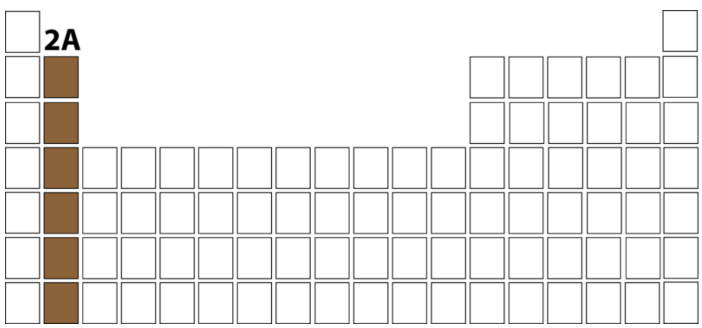
Some Common Ions

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.

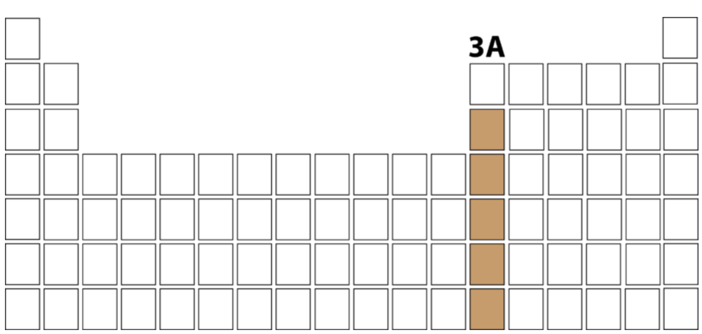


The metals in family 2A (shown in the figure below) all have electron configurations ending with two *s* electrons in the outermost energy level. This entire family will form +2 ions: Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Ra^{2+}.

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



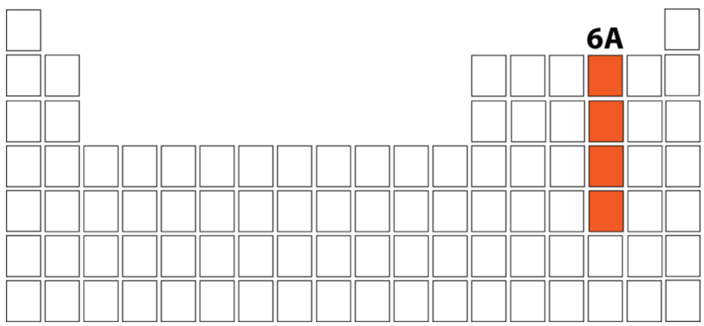
Family 3A members (shown in the figure below) have electron configurations ending in s^2p^1. When these atoms form ions, they will almost always form 3+ ions: Al^{3+}, Ga^{3+}, In^{3+}, and Tl^{3+}. Notice that boron is omitted from this list. This is because boron falls on the nonmetal side of the metal/nonmetal 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?

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