

### 3.3: Ions

#### Learning Objectives

- Define the two types of ions.
- Use Lewis diagrams to illustrate ion formation.

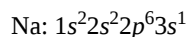
Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while some atoms lack only one or two electrons to have an octet. In cases where an atom has three or fewer valence electrons, the atom may lose those valence electrons quite easily until what remains is a lower shell that contains an octet. Atoms that lose electrons acquire a positive charge as a result because they are left with fewer negatively charged electrons to balance the positive charges of the protons in the nucleus. Positively charged ions are called cations. Most metals become cations when they make ionic compounds.

Some atoms have nearly eight electrons in their valence shell and can gain additional valence electrons until they have an octet. When these atoms gain electrons, they acquire a negative charge because they now possess more electrons than protons. Negatively charged ions are called anions. Most nonmetals become anions when they make ionic compounds.

The names for positive and negative ions are pronounced CAT-eye-ons (cations) and ANN-eye-ons (anions), respectively.

#### Electron Transfer

We can use electron configurations to illustrate the electron transfer process between sodium atoms and chlorine atoms.



As demonstrated here, a sodium atom (Na) has one valence electron in the third principal energy level. It is likely to achieve an octet in its outermost shell by losing its one valence electron. The cation produced in this way,  $\text{Na}^+$ , is called the sodium ion to distinguish it from the element. The sodium ion,  $\text{Na}^+$ , has the electron configuration with an octet of electrons from the second principal energy level. It is now the same as that of the noble gas neon. The term **isoelectronic** refers to an atom and an ion of a different atom (or two different ions) that have the same electron configuration. The sodium ion is isoelectronic with the neon atom. The equation below illustrates this process.

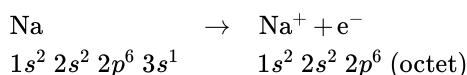


Figure 3.3.1 is a graphical depiction of this process.

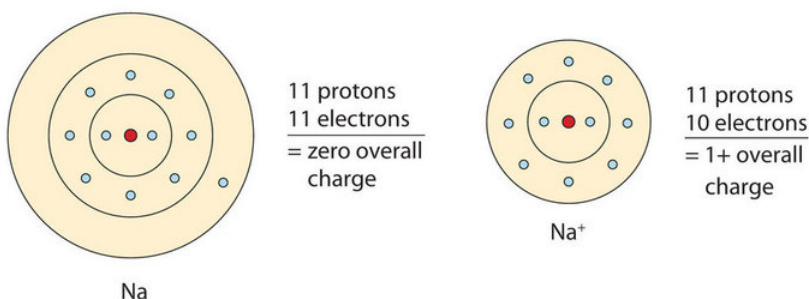


Figure 3.3.1: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge.

Now, let's consider chlorine atom, Cl:  $1s^2 2s^2 2p^6 3s^2 3p^5$

Only one more electron is needed to achieve an octet in chlorine's valence shell. When a chlorine atom gains an electron, its outermost principal energy level achieves an octet. In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again, the octet rule has been satisfied. The resulting anion,  $\text{Cl}^-$ , is called the chloride ion; note the slight change in the suffix (*-ide* instead of *-ine*) to create the name of this anion. This process is illustrated below. (In table salt, this electron comes from the sodium atom.)

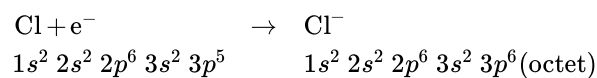


Figure 3.3.2 is a graphical depiction of this process.

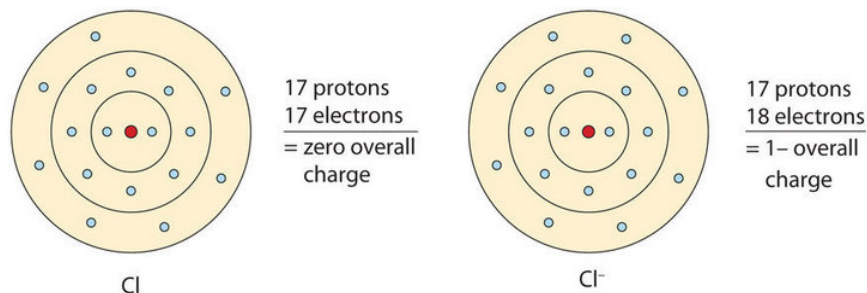


Figure 3.3.2: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1- charge.

With two oppositely charged ions, there is an electrostatic attraction between them because opposite charges attract. The resulting combination is the compound sodium chloride. Notice that there are no leftover electrons. The number of electrons lost by the sodium atom (one) equals the number of electrons gained by the chlorine atom (one), so the compound is electrically neutral. In macroscopic samples of sodium chloride, there are billions and billions of sodium and chloride ions, although there is always the same number of cations and anions.

### ✓ Example 3.3.1

Write the electron configuration of aluminum atom ( $Z=13$ ). How many electrons must Al lose/gain to achieve octet? Write the formula of the resulting ion and its electron configuration.

#### Solution

The electron configuration of Al atom is  $1s^2 2s^2 2p^6 3s^2 3p^1$ . The second shell has octet ( $2s^2 2p^6$ ) while the valence shell has 3 electrons ( $3s^2 3p^1$ ). Al can achieve octet by losing the 3 valence electrons. The resulting cation is  $\text{Al}^{3+}$  with electron configuration,  $1s^2 2s^2 2p^6$ .

### Exercise

Write the electron configuration of oxygen atom ( $Z=8$ ). How many electrons must O lose/gain to achieve octet? Write the formula of the resulting ion and its electron configuration.

#### Answer

The electron configuration of O atom is  $1s^2 2s^2 2p^4$ . The second shell has six electrons ( $2s^2 2p^4$ ) and needs two electrons to achieve octet. Oxygen will gain 2 electrons. The resulting anion is  $\text{O}^{2-}$  with electron configuration,  $1s^2 2s^2 2p^6$ .

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge. On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1- charge. Figure 3.3.3 shows how the charge on many ions can be predicted by the location of an element on the periodic table. Note the convention of first writing the number and then the sign on a multiply charged ion. The barium cation is written  $\text{Ba}^{2+}$ , not  $\text{Ba}^{+2}$ .

1A							8A
H <sup>+</sup>	2A		3A	4A	5A	6A	7A
Li <sup>+</sup>					N <sup>3-</sup>	O <sup>2-</sup>	F <sup>-</sup>
Na <sup>+</sup>	Mg <sup>2+</sup>		Al <sup>3+</sup>		P <sup>3-</sup>	S <sup>2-</sup>	Cl <sup>-</sup>
K <sup>+</sup>	Ca <sup>2+</sup>					Se <sup>2-</sup>	Br <sup>-</sup>
Rb <sup>+</sup>	Sr <sup>2+</sup>						I <sup>-</sup>


Figure 3.3.3: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

### ✓ Example 3.3.2

Which of these ions is *not likely* to form?

- Mg<sup>+</sup>
- K<sup>+</sup>

#### Solution

(a) Mg is in Group 2A and has two valence electrons. It achieves octet by losing two electrons to form Mg<sup>2+</sup> cation. Losing only one electron to form Mg<sup>+</sup> does not make an octet, hence, Mg<sup>+</sup> is not likely to form.

### Exercise

Which of these ions is *not likely* to form?

- S<sup>3-</sup>
- N<sup>3-</sup>

#### Answer

(a) S is in Group 6A and has six valence electrons. It achieves octet by gaining two electrons to form S<sup>2-</sup> anion. Gaining three electrons to form S<sup>3-</sup> does not make it octet, hence, S<sup>3-</sup> is not likely to form.

## Lewis Diagrams

Chemists use simple diagrams to show an atom's valence electrons and how they transfer. These diagrams have two advantages over the electron shell diagrams. First, they show only valence electrons. Second, instead of having a circle around the chemical symbol to represent the electron shell, they have up to eight dots around the symbol; each dot represents a valence electron. These dots are arranged to the right and left and above and below the symbol, **with no more than two dots on a side**. For example, the representation for sodium is as follows:



and the representation for chlorine is as follows:



For the above diagrams, it does not matter what sides the dots are placed on in Lewis diagrams as long as each side has a maximum of two dots.

These diagrams are called **Lewis electron dot diagrams**, or simply Lewis diagrams, after Gilbert N. Lewis, the American chemist who introduced them. To write an element's Lewis dot symbol, place the dots representing its valence electrons, one at a time, around the element's chemical symbol. Up to four dots are placed above, below, to the left, and to the right of the symbol (in any order, as long as elements with four or fewer valence electrons have no more than one dot in each position). The next dots, for elements with more than four valence electrons, are again distributed one at a time, each paired with one of the first four. In other words, **place the dots singly on each side before pairing them**. The Lewis electron dot diagram of fluorine, for example, with seven valence electrons, is constructed as follows:

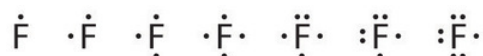
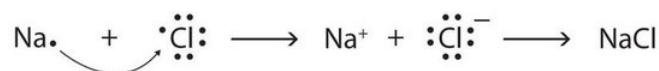


Figure 3.3.4 shows the electron configurations and Lewis diagrams of the elements lithium through neon, which is the entire second period of the periodic table. For the **main group** elements, the **number of valence electrons** is the **same as the group number** listed at the top of the periodic table.



Figure 3.3.4: Lewis Electron Dot Diagrams of the Elements Lithium through Neon

The transfer of electrons can be illustrated easily with Lewis diagrams:



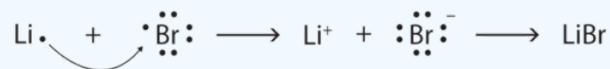
In representing the final formula, the dots are omitted.

### ✓ Example 3.3.3

Starting with lithium and bromine atoms, use Lewis diagrams to show the formation of the ionic compound LiBr.

#### Solution

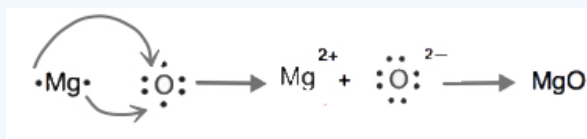
From the periodic table, we see that lithium is in the same column as sodium, so it will have the same valence shell electron configuration. That means that the neutral lithium atom will have the same Lewis diagram that the sodium atom has. Similarly, bromine is in the same column as chlorine, so it will have the same Lewis diagram that chlorine has. Therefore,



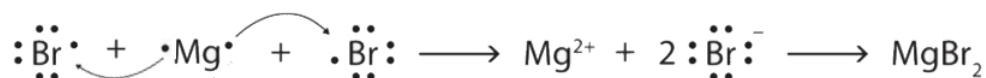
### ? Exercise 3.3.3

Starting with magnesium and oxygen atoms, use Lewis diagrams to show the formation of the ionic compound MgO.

#### Answer



Some ionic compounds have different numbers of cations and anions. In those cases, electron transfer occurs between more than one atom. For example, here is the formation of  $\text{MgBr}_2$ :



Notice that in this example there are two bromide ions ( $1-$  charge) needed for every one magnesium ion ( $2+$  charge) in order for the overall charge of the compound to equal zero. This is called **charge balance**. The number of each type of ion is indicated in the formula by the subscript.

Most of the elements that make ionic compounds form an ion that has a characteristic charge. For example, sodium makes ionic compounds in which the sodium ion always has a  $1+$  charge. Chlorine makes ionic compounds in which the chloride ion always has a  $1-$  charge. Some elements, especially transition metals, can form ions of multiple charges. Figure 3.3.5 shows the characteristic charges for some of these ions. As we saw in Figure 3.3.1, there is a pattern to the charges on many of the main group ions, but there is no simple pattern for transition metal ions (or for the larger main group elements).

1A																					8A
H <sup>+</sup>	2A																				
Li <sup>+</sup>																					
Na <sup>+</sup>	Mg <sup>2+</sup>	3B	4B	5B	6B	7B	8B				1B	2B	3A	4A	5A	6A	7A				
K <sup>+</sup>	Ca <sup>2+</sup>	Sc <sup>3+</sup>	Ti <sup>2+</sup>	V <sup>2+</sup>	Cr <sup>2+</sup>	Mn <sup>2+</sup>	Fe <sup>2+</sup>	Co <sup>2+</sup>	Ni <sup>+</sup>	Cu <sup>+</sup>	Zn <sup>2+</sup>		Al <sup>3+</sup>		P <sup>3-</sup>	S <sup>2-</sup>	Cl <sup>-</sup>				
Rb <sup>+</sup>	Sr <sup>2+</sup>						Fe <sup>3+</sup>	Co <sup>3+</sup>		Cu <sup>2+</sup>						Se <sup>2-</sup>	Br <sup>-</sup>				
Cs <sup>+</sup>	Ba <sup>2+</sup>									Ag <sup>+</sup>	Cd <sup>2+</sup>				Sn <sup>2+</sup>		I <sup>-</sup>				
										Au <sup>+</sup>					Pb <sup>2+</sup>						
										Au <sup>3+</sup>											

Figure 3.3.5: Charges of the Monatomic Ions. Note that some atoms commonly form ions of variable charges.

### Key Takeaways

- Ions can be positively charged or negatively charged.
- A Lewis diagram is used to show how electrons are transferred to make ions and ionic compounds.

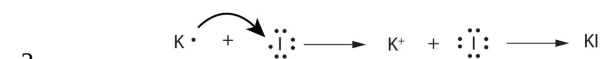
### Exercises

- What are the two types of ions?
- Use Lewis diagrams to illustrate the formation of an ionic compound from a potassium atom and an iodine atom.
- When the following atoms become ions, what charges do they acquire?
  - Li
  - S
  - Ca
  - F
- Identify each as a cation, an anion, or neither.
  - H<sup>+</sup>
  - Cl<sup>-</sup>
  - O<sub>2</sub>
  - Ba<sup>2+</sup>
  - CH<sub>4</sub>
  - CS<sub>2</sub>
- Identify each as a cation, an anion, or neither.
  - NH<sub>3</sub>
  - Br<sup>-</sup>
  - H<sup>-</sup>
  - Hg<sup>2+</sup>
  - CCl<sub>4</sub>
  - SO<sub>3</sub>
- Write the electron configuration for each ion.
  - Li<sup>+</sup>
  - Mg<sup>2+</sup>

- c.  $F^-$
  - d.  $S^{2-}$
7. Write the electron configuration for each ion.
- a.  $Na^+$
  - b.  $Be^{2+}$
  - c.  $Cl^-$
  - d.  $O^{2-}$
8. Draw Lewis diagrams for the ions listed in Exercise 6. Also include Lewis diagrams for the respective neutral atoms as a comparison.
9. Draw Lewis diagrams for the ions listed in Exercise 7. Also include Lewis diagrams for the respective neutral atoms as a comparison.
10. Using Lewis diagrams, show the electron transfer for the formation of  $LiF$ .
11. Using Lewis diagrams, show the electron transfer for the formation of  $MgO$ .
12. Using Lewis diagrams, show the electron transfer for the formation of  $Li_2O$ .
13. Using Lewis diagrams, show the electron transfer for the formation of  $CaF_2$ .
14. What characteristic charge do atoms in the first column of the periodic table have when they become ions?
15. What characteristic charge do atoms in the second column of the periodic table have when they become ions?
16. What characteristic charge do atoms in the third-to-last column of the periodic table have when they become ions?
17. What characteristic charge do atoms in the next-to-last column of the periodic table have when they become ions?

### Answers

1. Cations (positive charged) and anions (negative charged).



- 3.
- a. 1+
  - b. 2-
  - c. 2+
  - d. 1-

- 4.
- a. cation
  - b. anion
  - c. neither
  - d. cation
  - e. neither
  - f. neither

- 5.
- a. neither
  - b. anion
  - c. anion
  - d. cation
  - e. neither
  - f. neither

- 6.
- a.  $1s^2$
  - b.  $1s^2 2s^2 2p^6$

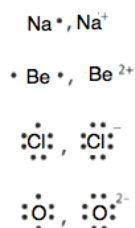
- c.  $1s^2 2s^2 2p^6$   
 d.  $1s^2 2s^2 2p^6 3s^2 3p^6$

7.

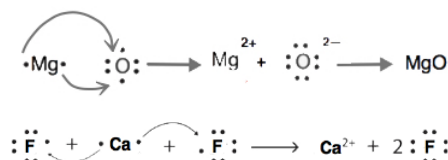
- a.  $1s^2 2s^2 2p^6$   
 b.  $1s^2$   
 c.  $1s^2 2s^2 2p^6 3s^2 3p^6$   
 d.  $1s^2 2s^2 2p^6$

8.

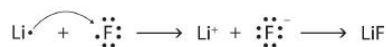
- a.  $\text{Li}^\bullet, \text{Li}^+$   
 b.  $\bullet\text{Mg}^\bullet, \text{Mg}^{2+}$   
 c.  $\text{:}\ddot{\text{F}}\text{:}, \text{:}\ddot{\text{F}}\text{:}^-$   
 d.  $\text{:}\ddot{\text{S}}\text{:}, \text{:}\ddot{\text{S}}\text{:}^{2-}$



9.



10.



11.



13.

12.

14.  $1+$

15.  $2+$

16.  $2-$

17.  $1-$

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