Name ________________________ **Chemistry** **___/___/___**

Lewis Dot Diagram/ Ions/ Electron Configuration of Ions

How to Write Lewis Dot Diagrams

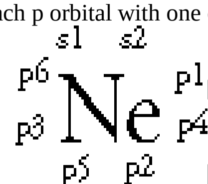
1. Write the symbol for the element you are going to be writing the electron configuration. Let the symbol represent the nucleus of the atom and all of the electrons, except those in the outer energy level.

2. Write the electron configuration for the element. Determine how many electrons are in the outer energy level. (The outer energy level is determined by the highest number before the letters (subshell)). When writing dot diagrams we are only concerned with the electrons in the outermost s and p subshells since d and f subshell electrons are always inside the outermost s and p subshells.

3. Pretend there is an imaginary box around the symbol, leaving you with four sides on which to represent electrons (top, right, left, and bottom). You will use dots to represent the electrons in the outer energy level.

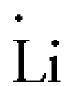
4. There are two rules to remember: (1.) s-orbital electrons are should be drawn on the same side of the symbol. (For convenience, I use the top as the location for my s-orbital electrons and the sides and bottom as my p-orbital electrons. I fill in the p-orbital electron dots in a clockwise manner.) (2.) Electrons will fill an empty orbital before they will pair with other electrons (Hund's Rule). Thus, you should fill each p orbital with one electron until you get to your forth p-orbital electron and then the electrons will pair up.

   I write my Lewis Dot Diagrams in the following order:



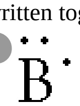
Examples:

Lithium: $1s^2 2s^1$ -- 2 is the outer energy level. Since it is $2s^1$, there is only 1 electron in the outer energy level. Thus the Lewis Dot Diagram should look like:

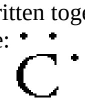


Beryllium: $1s^2 2s^2$ -- 2 is the outer energy level. Since it is $2s^2$, there are 2 electrons in the outer energy level. Remember, both of these electrons in the outer energy level are in the s-orbital. Thus they should be written together and the Lewis Dot Diagram should look like:

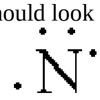
Boron: $1s^2 2s^2 2p^1$ -- 2 is the outer energy level. Since it is $2s^2$ and $2p^1$, there are 3 electrons in the outer energy level. Remember, both of the electrons in the s-orbital have to be written together. The 1 p-orbital electron can go in any empty space. The Lewis Dot Diagram should look like:



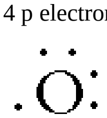
Carbon : $1s^2 2s^2 2p^2$ -- 2 is the outer energy level. Since it is $2s^2$ and $2p^2$, there are 4 electrons in the outer energy level. Remember, both of the electrons in the s-orbital have to be written together. The 2 p-orbital electrons will each fill an empty space before they pair up. The Lewis Dot Diagram should look like:



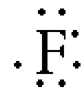
Nitrogen: $1s^2 2s^2 2p^3$ -- -- 2 is the outer energy level. Since it is $2s^2$ and $2p^3$, there are 5 electrons in the outer energy level. Remember, both of the electrons in the s-orbital have to be written together. The 3 p-orbital electrons will each fill an empty space before they pair up. The Lewis Dot Diagram should look like:



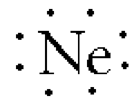
Oxygen: $1s^2 2s^2 2p^4$ -- 2 is the outer energy level. Since it is $2s^2$ and $2p^4$, there are 6 electrons in the outer energy level. Remember, both of the electrons in the s-orbital have to be written together. The 4 p-orbital electrons will each fill an empty space before they pair up. Since there are only 3 empty spaces but 4 p electrons, two of the electrons must double up. The Lewis Dot Diagram should look like:



Fluorine: $1s^2\ 2s^2\ 2p^5$ -- 2 is the outer energy level. Since it is $2s^2$ and $2p^5$, there are 7 electrons in the outer energy level. Remember, both of the electrons in the s-orbital have to be written together. The 5 p-orbital electrons will each fill an empty space before they pair up. Since there are only 3 empty spaces but 5 p electrons, four of the electrons must double up. The Lewis Dot Diagram should look like:



Neon: $1s^2\ 2s^2\ 2p^6$ -- 2 is the outer energy level. Since it is $2s^2$ and $2p^6$, there are 8 electrons in the outer energy level. Remember, both of the electrons in the s-orbital have to be written together. The 6 p-orbital electrons will each fill an empty space before they pair up. Since there are only 3 empty spaces but 6 p electrons, all six of the electrons must double up. The Lewis Dot Diagram should look like:



Taking Dot Diagrams Further - Ions

Group	**1 / IA**	**2 / IIA**	**13 / IIIA**	**14 / IVA**	**15 / VA**	**16 / VIA**	**17 / VIIA**	**18 / VIIIA**
Element	K	Ba	Al	C	P	Te	Br	Ar
# p^+	**19**	**56**	**13**	**6**	**15**	**52**	**35**	**18**
# e^-	**19**	**56**	**13**	**6**	**15**	**52**	**35**	**18**
# e- in outer shell (valence electrons)	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**
Easier to gain/lose e^- and how many	**Lose 1**	**Lose 2**	**Lose 3**	**Gain 4 / Lose 4**	**Gain 3**	**Gain 2**	**Gain 1**	**neither**
After gain/loss, #p^+	**19**	**56**	**13**	**6**	**15**	**52**	**35**	**18**
After gain/loss, #e-	**18**	**54**	**10**	**10/2**	**18**	**54**	**36**	**18**
After gain/loss, atomic charge (oxidation number)	**1+**	**2+**	**3+**	**4+/-**	**3-**	**2-**	**1-**	**0**

When atoms gain or lose electrons, they form **ions**. There are two types of ions, **cations** and **anions**. A **cation** is any atom or group of atoms with a **positive** charge. An **anion** is any atom or group of atoms with a **negative** charge. An ionic charge is written to the upper right on an element's symbol. If an element gains one electron, the charge is written $:F^-$ or F^{1-}. Note: a one does not have to be written, but you may write it if you prefer. However, whenever a number is written, it is written before the charge. For example magnesium would be written Mg^{2+} , aluminum Al^{3+}, and oxygen O^{2-}. Because the noble gases do not gain or lose electrons, they do not form ions and have an overall charge of zero. You do not write a charge with these elements. Neon would simply be written as Ne.

Write the symbol and charge and whether the atom forms an anion or cation in the chart below for each of the elements used above.

Group	**1 / IA**	**2 / IIA**	**13 / IIIA**	**14 / IVA**	**15 / VA**	**16 / VIA**	**17 / VIIA**	**18 / VIIIA**
Element	K	Ba	Al	C	P	Te	Br	Ar
Forms an anion or cation	**Cation**	**Cation**	**Cation**	**Both**	**Anion**	**Anion**	**Anion**	**Neither**
Element with its ionic charge	**K^{1+}**	**Ba^{2+}**	**Al^{3+}**	**$C^{4+/-}$**	**P^{3-}**	**Te^{2-}**	**Br^{1-}**	**Ar**

Above each main block group on your periodic table, write the ionic charge (**oxidation number**) associated with that group.

Electron Configuration of Ions

The only difference between writing a normal electron configuration and writing the electron configuration for an ion is that when writing an electron configuration for an ion, you have to remember to add or subtract electrons from your total. Remember if an ion has a **positive (+) charge (cation)** you must **subtract** electrons from your total. And, if the element is an **anion (negative charg**e), you must **add** electrons to your total. Look at the example below:

Given: Al
Aluminum will lose three electrons. So although a neutral atom of aluminum has 13 electrons, the ion of aluminum, Al^{3+}, has lost three electrons and only has 10. Thus, you should write the electron configuration for 10 electrons.
Answer: Al^{3+} - $1s^2\ 2s^2\ 2p^6$

Given: S
Sulfur will gain two electrons. So although a neutral atom of sulfur has 16 electrons, the ion of sulfur, S^{2-}, has gained two electrons and now has 18. Therefore, you should write the electron configuration for 18 electrons.
Answer: S^{2-} - $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Homework: For each of the following elements, write (1.) Lewis dot diagram (2.) numbers of valence electrons (3.) element symbol with ionic charge (4.) whether the element forms an anion or a cation and the (5.) electron configuration of the ion (you may write the shortcut if you'd like).

Element	Valence Electrons	Lewis Dot Diagram	Symbol with Ionic Charge	Cation/ Anion	# of electrons in Ion	Electron Configuration of Ion
1. rubidium	**1**	**1 dot**	Rb^{1+}	**Cation**	**36**	**[Ar]** $4s^2\ 3d^{10}\ 4p^6$
2. sulfur	**6**	**6 dots**	S^{2-}	**Anion**	**18**	**[Ne]** $3s^2\ 3p^6$
3. krypton	**8**	**8 dots**	**Kr**	**Neither**	**36**	**[Ar]** $4s^2\ 3d^{10}\ 4p^6$
4. iodine	**7**	**7 dots**	I^{1-}	**Anion**	**54**	**[Kr]** $5s^2\ 4d^{10}\ 5p^6$
5. oxygen	**6**	**6 dots**	O^{2-}	**Anion**	**10**	**[He]** $2s^2\ 2p^6$
6. arsenic	**5**	**5 dots**	As^{3-}	**Anion**	**36**	**[Ar]** $4s^2\ 3d^{10}\ 4p^6$
7. calcium	**2**	**2 dots**	Ca^{2+}	**Cation**	**18**	**[Ne]** $3s^2\ 3p^6$
8. francium	**1**	**1 dot**	Fr^{1+}	**Cation**	**86**	**[Xe]** $6s^2\ 4f^{14}\ 5d^{10}\ 6p^6$
9. strontium	**2**	**2 dots**	Sr^{2+}	**Cation**	**36**	**[Ar]** $4s^2\ 3d^{10}\ 4p^6$
10. carbon	**4**	**4 dots**	$C^{4+/-}$	**Both**	**10/2**	**[He]** $2s^2\ 2p^6$ / $1s^2$
11. helium	**2**	**2 dots**	**He**	**Neither**	**2**	$1s^2$
12. chlorine	**7**	**7 dots**	Cl^{1-}	**Anion**	**18**	**[Ne]** $3s^2\ 3p^6$
13. phosphorus	**5**	**5 dots**	P^{3-}	**Anion**	**18**	**[Ne]** $3s^2\ 3p^6$

Write the correct term for each of the following definitions. Not all terms will be used.

actinides, alkali metals, alkaline-earth metals, anion, atomic radius, cation, chalcogens, electron affinity, electronegativity, group, halogens, ion, ionization energy, lanthanides, main-block elements, metal, metalloid, noble gas, nonmetal, octet, octet rule, oxidation number, period, periodic law, periodic table, shielding effect, transition elements, valence electrons

1. **shielding effect**	The reduction in the attraction between the nucleus and its outer electrons due to the blocking effect of inner electrons.
2. **ion**	An atom that gains or loses an electron.
3.**electronegativity**	The tendency for an atom to attract electrons to itself when it is bonded to another atom.
4. **valence electrons**	The electrons in the outer energy level of an atom.
5. **anion**	An atom with a negative charge.
6. **ionization energy**	The amount of energy needed to remove an electron from an atom or ion in the ground state in the gas phase.
7. **atomic radius**	Half the distance between the nuclei of two like atoms located next to one another.
8. **group/family**	The vertical columns on the periodic table.
9. **period**	A horizontal row on the periodic table.
10. **cation**	An atom that loses electrons when it forms an ion.
11. **actinides**	Metallic elements with atomic numbers 90-103 that fill the 5f orbitals.
12. **alkali metals**	Highly reactive metallic elements that form alkaline solutions in water, burn in air, and belong to Group 1(IA) of the periodic table.
13. **alkaline earth**	Reactive, metallic elements that belong to Group 2(IIA) of the periodic table.
14. **oxidation number**	A positive or negative charge of an atom when it forms an ion.
15. **octet**	A stable outer shell of eight electrons, arranged in four orbital pairs.
16. **halogens**	Elements that combine with most metals to form salts and that belong to Group 17(VIIA) of the periodic table.
17. **lanthanides**	Shiny, metallic elements with atomic numbers 58-71 that fill the 4f orbitals.
18. **main-block elements**	Elements that represent the entire range of chemical properties and belong to all A Group elements in the periodic table.
19. **non-metal**	An element that is neither a metal or metalloid (semi-metal).
20. **transition elements**	Metallic elements that have varying properties and belong to Groups 3(IIIB) to 12(IIB) of the periodic table.
21. **periodic law**	Properties of elements tend to change with increasing atomic number in a periodic way.
22. **metalloids**	An element having properties of both metals and nonmetals; also known as semi-metals.