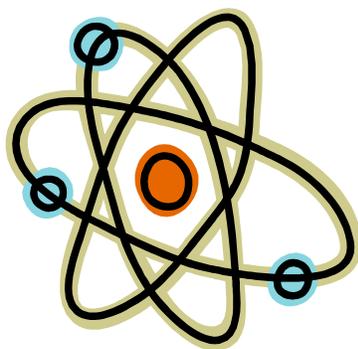


Chemistry A

Atomic Structure



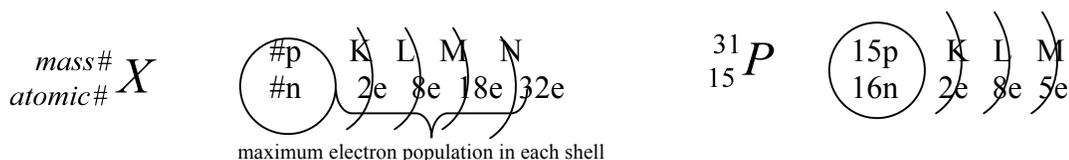
Worksheet #1: Atomic Energy LevelsParts of an Atom Review:

1. Think back to the Periodic Table Packet to complete the following table:

Particle Name	Charge	Mass	Location in Atom	What tells you how many are in an atom?
PROTON				
ELECTRON				
NEUTRON				

Electrons in the Atom:

The electrons are distributed in energy levels or "shells" which Bohr named with capital letters starting with K for the lowest energy level and proceeding alphabetically as you move to higher levels. Since electrons are negative and repel each other there is a limit to the number of electrons that can occupy a given energy region--**Bohr found the maximum population of an energy level to be $2n^2$ where n is the number of energy level.** Thus the K "shell" or $n=1$ energy level can hold a maximum of 2 electrons, the L "shell" or $n=2$ level can hold a maximum of 8 electrons, etc. For the first eighteen elements when one energy level is filled the electrons enter the next higher level; for later elements there is some overlapping of energy levels and the order of filling is more complicated. A simple diagram or representation of the atom called a **Bohr diagram tells the number of protons and neutrons in the nucleus and shows the number of electrons occupying each energy level.** A Bohr diagram for phosphorus (atomic number 15, mass number 31) would be:



2. Name the first four energy levels (shells) and tell the maximum number of electrons that fit in each shell.

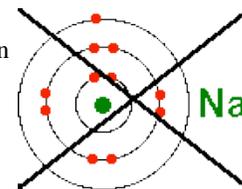
What is the general equation used to determine the number of electrons in a shell? _____

3. Draw Bohr diagrams for elements with atomic numbers 1-10. (i.e. element 1 is H, 2 is He)

H	He	Li	Be	B
C	N	O	F	Ne

Worksheet #2: Electron Orbitals

The Heisenberg Uncertainty Principle states that it is impossible to know both the location and the speed of an electron. This means the Bohr Model of an atom that looks like a solar system is very inaccurate. Bohr Diagrams are useful for counting electrons and illustrating different energy levels, but they are worthless as pictures of what atoms really look like.



The Bohr Model of the atom uses circles to show where the electrons are. Electrons do not really move around the nucleus in perfect circles. They do live in regions of space called orbitals.

An orbital is the region in space where an electron is most likely to be found. Orbitals are organized by the shell (energy level) that they belong to. At left you see what is called the 1s orbital. This ball-shaped orbital gives a better idea of what a hydrogen atom actually looks like. The dots in the picture represent possible positions of the one electron the hydrogen has. The nucleus (too small to see) is in the center of the picture. Notice that there are more dots near the center of the sphere than near the edges. This shows that the electron is most likely to be near the nucleus. But that does not mean it is impossible to find the electron very far away from the nucleus.

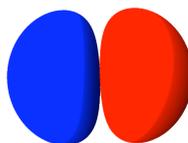
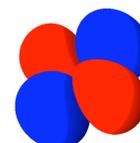
The 1s Orbital



The 2s orbital of an atom is shaped just like the 1s orbital. The difference is that it is larger and has a higher energy than the 1s orbital. The orbitals of an electron have a variety of shapes. The shapes fall into certain categories that are known by letters. **The s orbitals are all spherical and there is only one s orbital in each shell. The p orbitals all look like barbells made out of balloons and there are always three p orbitals per shell. Each p orbital points in a direction perpendicular to the others (p_x , p_y , p_z). Most of the d orbitals look like various combinations of four balloons. There are always five d orbitals in a shell that contains d orbitals.**

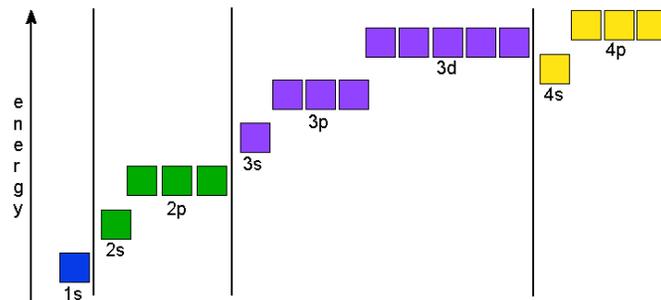


The 2s Orbital

The 2p_x OrbitalThe 3d_{xy} Orbital

The orbitals are organized by their energies. Each shell of the electron structure of an atom is organized by its energy: **the more energy, the farther a shell is from the nucleus. The first shell can only have one kind of orbital: the s orbital. The second shell has both an s orbital and a p orbital. The third shell has s, p, and also d orbitals. The fourth shell has s, p, d, and f orbitals.** After that, as the energy of the shells increases the types of orbitals go by the letter of the alphabet.

Look at the diagram to the right. It shows how the orbitals are organized by their relative energy. You will learn how to determine where the electrons in atoms and ions are placed in such a diagram. You will also learn how to write a shorthand notation for such diagrams.

**Electron Configurations**

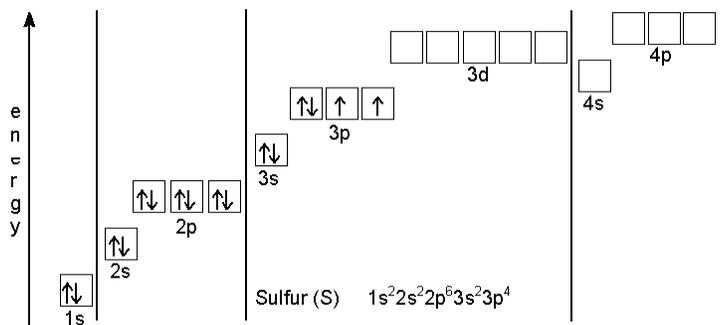
One way to approach the very abstract task of filling electron orbitals is to use an analogy to make it a bit more concrete. Imagine you are the landlord of a very strange apartment building. Your job is to fill the apartments in the building in the most efficient way possible. You are required by the owner of the building to fill the rooms in a certain way. The rules are summarized in the table on the next page.

In the building the different floors are like the different energy levels (or shells) in an atom. The energy levels are numbered starting from one, just like the floors in an apartment. Each room corresponds to one orbital: one box on the diagram above. The rooms can hold to more than two electrons (two people) each. In each room only a man and a woman may be paired together. In the strange world of quantum mechanics there are no same-gender room mates.

Apartment House Rules	Electron Rules
From the Bottom Up: Rooms must be filled from the ground floor up. Fill the one room on the first floor before starting to put new tenants on the second floor. Then fill the s room before the p rooms. At higher floors the order might change a bit.	Aufbau Principle: the electrons fill the available orbitals from lowest energy to highest energy. In the ground state all the electrons are in the lowest possible energy level.
Singles First: the owner of the building wants to have the tenants spread out as much as possible. For that reason singles are placed in rooms before couples. If couples must be placed into a room then all of the other rooms on that floor must already have a single in them.	Hund's Rule: The electrons must be placed into the orbitals in such a way that no pairs are put together unless absolutely necessary. That is, single electrons must be placed into boxes first and then paired up if necessary.
Opposite Gender Only: When two people are placed in a room they must be of opposite genders. No men may room together and no women may room together. This is an arbitrary rule on the part of the owners: we all know this isn't how it works in the real world.	Pauli Exclusion Principle: Electrons come in two varieties based on the direction they are 'spinning'. There is an Up spin and a Down spin. Up and Down spins are always paired together and Up-Up or Down-Down combinations are not allowed. No two electrons can ever be in the same place at the same time.

Here is an example of the rooms in the apartment house having been filled in by the rules. **The element Sulfur has 16 electrons and starting from the bottom up you place two (1 Up and 1 Down) in the 1s orbital.**

Next, electrons are placed in each of the 2p orbitals until there is one in each orbital. Then the rest of the electrons are filled into the 2p orbitals. After that the 3s orbital is filled and then the 3p. Electrons are placed in the boxes of the 3p orbitals one at a time until there is one in each orbital. After that the remaining electron is placed as part of a pair in the first 3p box.



The electron configuration of Sulfur is $1s^2 2s^2 2p^6 3s^2 3p^4$. In that notation $1s^2$ means energy level 1, s-type orbital, 2 e⁻; $2s^2$ means energy level 2, s-type orbital, 2 e⁻; $2p^6$ means energy level 2, p-type orbital, 6 e⁻; $3s^2$ means energy level 3, s-type orbital, 2 e⁻; and, finally, $3p^4$ means energy level 3, p-type orbital, 4 e⁻.

Questions

1. What is the difference between a shell and an electron orbital?
2. How does the Aufbau Principle work? That is, what does it mean to say that orbitals are filled from the bottom up?
3. What is Hund's Rule? That is, how do you go about filling up electron orbitals according to this rule?
4. In the following electron configuration tell what each part means: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$.
5. Why does the 4s orbital get filled with electrons before any electrons can be placed into a 3d orbital?

Worksheet #3: Electron Configurations**“Reading” the Periodic Table**

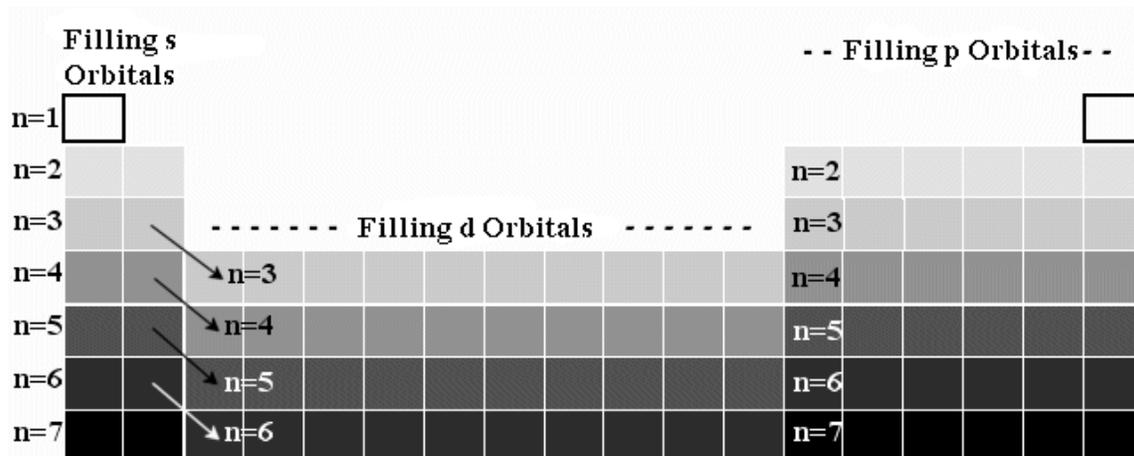
Look at the electron configurations of Oxygen and Sulfur. Notice that in both cases the last orbitals to fill are p orbitals. **Notice also that there are 2 electrons in an s orbital and 4 electrons in p orbital for both Oxygen and Sulfur.**

Oxygen	$1s^2 2s^2 2p^4$
Sulfur	$1s^2 2s^2 2p^6 3s^2 3p^4$

For these reasons Oxygen and Sulfur belong to the *p-block* of elements in the periodic table. In fact, all of the elements from Groups 13 to 18 are part of the *p-block*. All of the elements in this block fill a set of p orbitals last and all of them have their valence electrons in s and p orbitals.

Here you can finally understand the reason that having eight electrons is so important. Each s orbital holds 2 electrons and each p orbital holds 6 electrons. Together they can hold the eight electrons it takes to finish off the orbitals and make the valence shell full.

Other sections of the periodic table are also known by the type of orbital that fills last in that section. **In Groups 1 and 2 the s orbitals are filling. In Groups 3 through 12 the d orbitals are the last to be filled.** Incidentally, one important fact is the the 4s orbital fills before the 3d orbitals start to fill. This is because **the 4s orbital is actually slightly lower in energy than the 3d orbitals.** You will not be much concerned with any orbitals higher than the 4s orbital since there are some interesting but complicated exceptions to the rules when it comes to filling the 3d orbitals.

**Abbreviated Electron Configurations**

As we get to larger elements we notice the electron configuration becomes tediously long and repetitive so obviously there must be some way to abbreviate configurations. **The system for abbreviating involves comparing the electron structure of the atom you are considering with that of a noble gas.** Since the noble (or inert) gases have the most stable electron arrangements we use them for a reference. **An element is abbreviated by comparing its structure to the noble gas at the end of the preceding period of the periodic table. You abbreviate the element's configuration by indicating the appropriate noble gas plus all of the additional electrons in the element you are considering.**

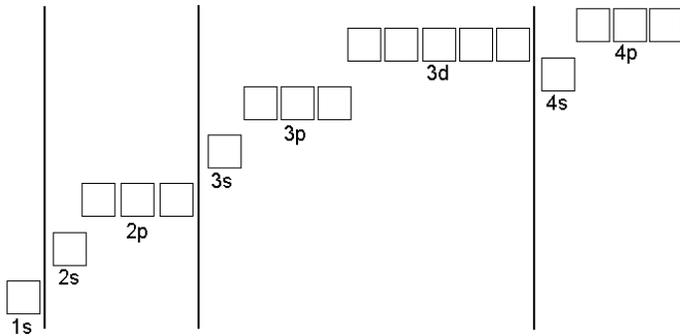
Example: Show the abbreviated configuration for chlorine. Note that the noble gas neon precedes chlorine.

The complete configuration for chlorine is ${}_{17}^{35}\text{Cl } 1s^2 2s^2 2p^6 3s^2 3p^5$ since neon has the configuration ${}_{10}^{20}\text{Ne } 1s^2 2s^2 2p^6$ this sequence can be replaced by [Ne], so the abbreviated configuration for chlorine is ${}_{17}^{35}\text{Cl } [\text{Ne}]3s^2 3p^5$

Questions

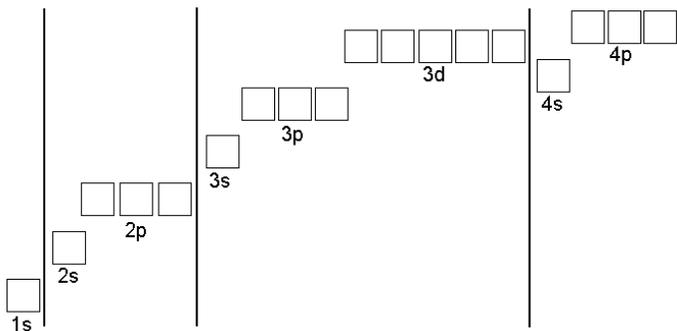
1. Why are Groups 1 and 2 referred to as the *s-block* of the periodic table?
2. Why are Groups 3 through 12 referred to as the *d-block* of the periodic table?

3. Fill the boxes below with the arrow notation for electrons showing the correct ground state electron configuration for the element He.



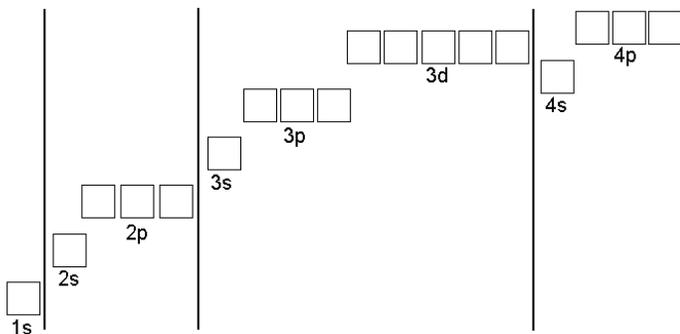
4. Write out the complete electron configuration for He

5. Fill the boxes below with the arrow notation for electrons showing the correct ground state electron configuration for the element Ne.



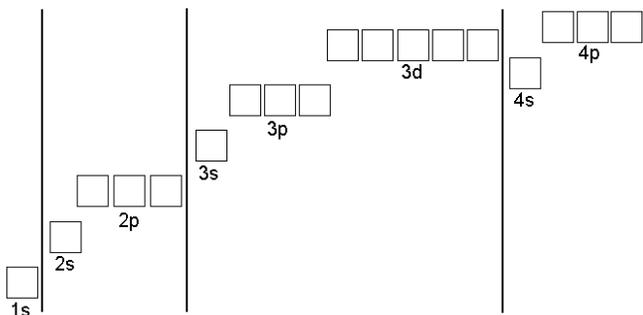
6. Write out the complete electron configuration for Ne

7. Fill the boxes below with the arrow notation for electrons showing the correct ground state electron configuration for the element Ar.



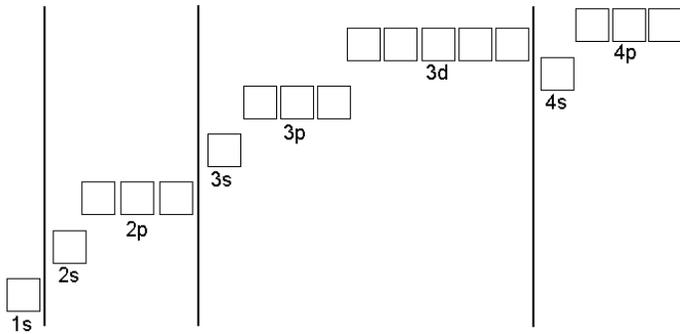
8. Write out the complete electron configuration of Ar

9. Fill the boxes below with the arrow notation for electrons showing the correct ground state electron configuration for the element Na.



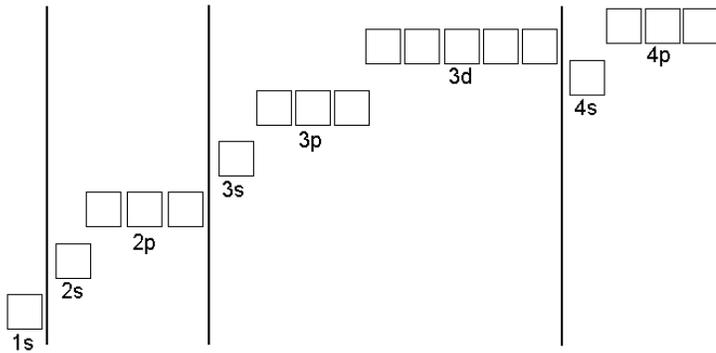
10. Write out the complete electron configuration for Na

11. Fill the boxes below with the arrow notation for electrons showing the correct ground state electron configuration for the element Si.



12. Write out the complete electron configuration for Si

13. Fill the boxes below with the arrow notation for electrons showing the correct ground state electron configuration for the element N.



14. Write out the complete electron configuration for N

15. Write out the abbreviated electron configuration of B.

16. Write out the abbreviated electron configuration of Cl.

17. Write out the abbreviated electron configuration of F.

18. Write out the abbreviated electron configuration of Ca.

19. Write out the abbreviated electron configuration of Kr.

EXTRA CHALLENGES

20. Write out the abbreviated electron configuration of O^{2-} . Notice that this is an anion!

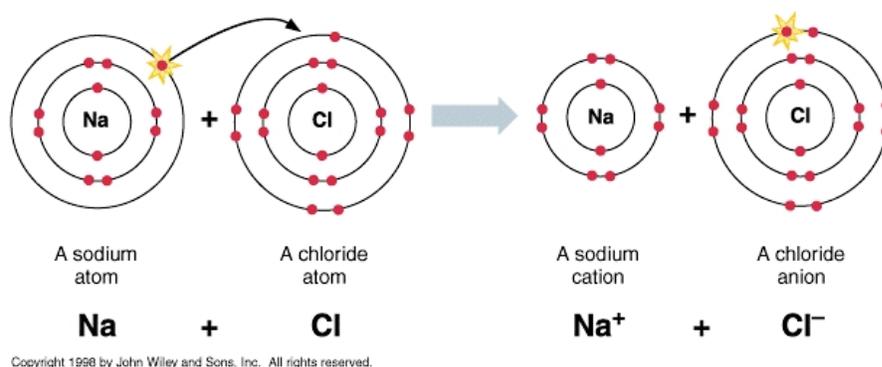
21. Write out the abbreviated electron configuration of Na^+ . Notice that this is a cation!

Worksheet #4- Ions

Chemical reactions involve electrons in the highest energy levels of an atom, those that we call **valence electrons** and describe as being in the outermost shell. As we discuss ions it will be important to remember that the #1 goal of ANY atom is to have a full outer shell. In other words, all atoms want 2 electrons in their 1st shell AND 8 electrons in their 2nd, 3rd, 4th shells. Noble Gases (Group 18) are the happiest atoms because they have achieved this goal. All other groups try for this goal by giving away or stealing electrons.

When a chemical reaction occurs an atom can lose, gain or share valence electrons. Atoms that have gained or lost electrons it is called an **ion**. Ions are not neutral, they have a net (overall) charge. If an atom loses valence electrons it becomes a **positive ion or cation**. If an atom gains valence electrons it becomes a **negative ion or anion**.

Below is a diagram that shows a chemical reaction between sodium (Na) and chlorine (Cl). In this reaction, sodium gives away an electron to chlorine. In other words, chlorine steals an electron from sodium.



Since electrons have a **NEGATIVE** charge, when chlorine gains an electron it becomes negative and when sodium gives away an electron it becomes positive. Let's review:

Cations = Ions with a positive charge.	Anions = Ions with a negative charge.
Cations have lost electrons	Anions have gained electrons
Examples = Na ⁺ , Mg ²⁺ , Al ³⁺	Examples = Cl ⁻ , O ²⁻ , N ³⁻ , C ⁴⁻

Cation**Anion****Forming Cations and Anions**

Atoms that have *less than four* of the possible 8 electrons in their outer shell are most likely to *give away electrons* and become *cations*. This means that groups 1A-3A form cations. Atoms that have *four or more* of the possible 8 electrons in their outer shell are most likely to *steal electrons* and become *anions*. This means groups 4A- 7A form anions. Noble gases do not form cations or anions because they have the 8 electrons that they want.

Oxidation Number

Based on these ideas, we can predict the most likely charge of any atom in groups 1A-7A. An **oxidation number** is the charge assigned to an atom based on the predicted number of electrons gained or lost.

Examples: Li = +1, Mg = +2, Al = +3, C = -4, N = -3, O = -2, Cl = -1. Noble gases do not have oxidation numbers because they do not form cations or anions.

Worksheet #4 Continued**Practice with Ions**

Remember: The charge on an ion will involve a number and a sign. The number will always be the number of electrons involved, e.g., loss of two electrons is +2. The sign will be positive if electrons (negatives) are lost and negative if electrons are gained.

I. Determine the following charges

1. An atom having lost two electrons _____
2. An atom having lost six electrons _____
3. An atom having gained one electron _____
4. An atom having gained three electrons _____
5. An atom having lost five electrons _____
6. An atom having gained two electrons _____
7. An atom having lost one electron _____
8. An atom having gained four electrons _____

II. Determine the charges on the following elements based on their locations on the periodic table.

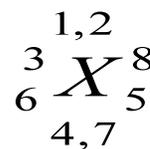
1. Fluorine _____
2. Magnesium _____
3. Aluminum _____
4. Sodium _____
5. Nitrogen _____
6. Zinc _____
7. Hydrogen _____
8. Argon _____
9. Barium _____
10. Sulfur _____

III. Answer the following questions.

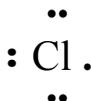
1. Will metals lose or gain electrons?
2. Considering your answer for #1, what type of ion will metals form (positive or negative)?
3. What is the name of a positive ion?
4. What is the name of a negative ion?
5. How many electrons are in a **full** outer shell for period 1 (H and He)?
6. How many electrons are in **full** outer shells for all atoms other than H and He?

Worksheet #5Lewis Dot Diagrams

Earlier you learned that a valence shell is the outermost shell of an atom. The electrons there are involved in chemical reactions. We can draw a Lewis dot (or electron dot) diagram to show the electrons that occupy the valence shell. A generic electron dot diagram:



For example, chlorine has an electron configuration of ${}_{17}^{35}\text{Cl } 1s^2 2s^2 2p^6 3s^2 3p^5$. The highest energy level (shell) is 3. Looking in the 3rd energy level you can see that there are 7 electrons. This indicates that there are 7 valence electrons for chlorine so the electron dot diagram would look like:

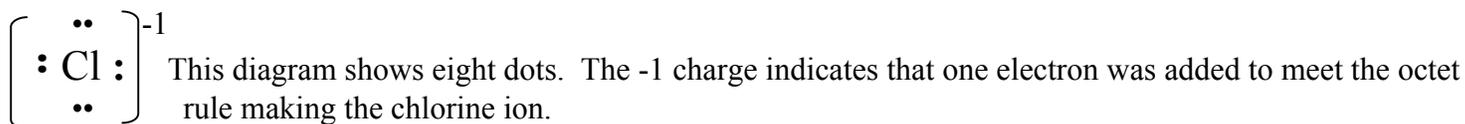
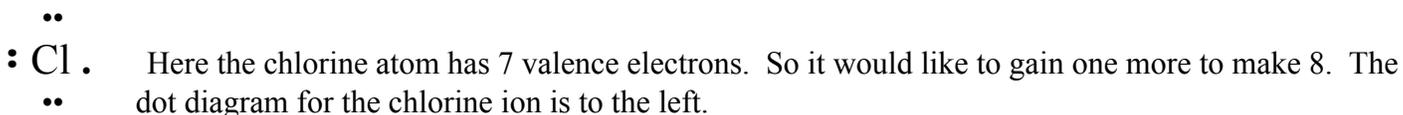


1. Write the abbreviated electron configuration and electron dot diagram for the following elements.

hydrogen	beryllium	boron	carbon
fluorine	neon	phosphorus	arsenic

Ion Diagrams

Remember that ions are formed when an atom either gains or loses electrons. When an atom gains an electron it forms a negative ion (anion), when an atom loses an electron it forms a positive ion (cation). The octet rule, a stable electron arrangement is achieved when eight valence electrons are present. For example, the dot diagram for chlorine that was done in the previous example is:



If an atom has less than 4 dots it will give them away until it has zero dots. This means the ion's charge would be positive.

2. Show the ion dot diagram (including the charge) for the ions that each element would form. These are the same elements as in the table above.

hydrogen	beryllium	fluorine	phosphorus
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Worksheet #6Isotopes

Isotopes are atoms of the same element with different masses (or weights) due to different numbers of neutrons in their nuclei. All atoms of the same element must have the same number of protons (and thus the same number of electrons) which is equal to the atomic number. However, atoms of the same element can have different numbers of neutrons and thus a different mass number. The difference in mass does not influence the chemical behavior. Valence electrons are what determine chemical behavior.

Hydrogen has three isotopes. The most abundant isotope is ordinary hydrogen-1. A second stable isotope is hydrogen-2 and the third isotope is the radioactive hydrogen-3. The Bohr diagrams for the isotopes of hydrogen are shown below. Notice that only the mass number and number of neutrons are different.

Hydrogen-1		Hydrogen-2		Hydrogen-3	
1_1H		2_1H		3_1H	

- What is an isotope? _____

- What does the number next to isotopes signify (ie: Carbon 12 vs. Carbon 13)? _____

- How can you tell isotopes apart in lab? _____

- Here are three isotopes of an element: ${}_6^{12}C$ ${}_6^{13}C$ ${}_6^{14}C$
 - The element is: _____
 - The number 6 refers to the _____
 - The numbers 12, 13, and 14 refer to the _____
 - How many protons and neutrons are in the first isotope? _____
 - How many protons and neutrons are in the second isotope? _____
 - How many protons and neutrons are in the third isotope? _____
- Write the symbols for the isotopes of uranium with the following numbers of neutrons:
 - 142 neutrons
 - 143 neutrons
 - 146 neutrons

6. Complete the following chart:

Isotope name	Atomic #	Mass #	# of protons	# of neutrons	# of electrons
Uranium-235					
Uranium-238					
Boron-10					
Boron-11					
Carbon – 12					
Carbon – 14					
Tin – 119					
Tin – 120					
Lithium – 7					
Sodium – 23					

Worksheet #7- Average Atomic Mass

The atomic masses (or atomic weights) on the periodic table are the average of all the isotopes but it is not a straight average. The atomic masses on the periodic table are the averages of all the isotopes based upon abundance.

For example, consider element X:

60% mass 40 amu,

30% mass 42 amu

10% mass 44 amu,

then the periodic table mass would be calculated:

$$.60 \times 40 \text{ amu} = 24.0 \text{ amu}$$

$$.30 \times 42 \text{ amu} = 12.6 \text{ amu}$$

$$.10 \times 44 \text{ amu} = 4.4 \text{ amu}$$

average mass based upon relative abundance

$$= 41.0 \text{ amu}$$

Calculate the average atomic mass for the following isotopes.

Element	Mass of Isotope	Abundance	Atomic Mass Based Upon Relative Abundance
1. Nitrogen			
N-14	14.0031 amu	99.64 %	
N-15	15.0001 amu	0.36 %	
2. Magnesium			
Mg-24	23.9850 amu	78.99 %	
Mg-25	24.9858 amu	10.00 %	
Mg-26	25.9826 amu	11.01 %	
3. Neon			
Ne-20	19.992 amu	90.51%	
Ne-21	20.994 amu	0.27%	
Ne-22	21.991 amu	9.22%	
4. Chromium			
Cr-50	49.9461 amu	4.35 %	
Cr-52	51.9405 amu	83.79%	
Cr-53	52.9407 amu	9.50%	
Cr-54	53.9389 amu	2.36 %	

Worksheet #7 Continued**EXTRA CHALLENGES**

5. Rubidium is a soft, silvery-white metal that has two common isotopes, Rb-85 and Rb-87. If the abundance of Rb-85 is 72.2% and the abundance of Rb-87 is 27.8%, what is the average atomic mass of rubidium?

6. Uranium is used in nuclear reactors and is a rare element on earth. Uranium has three common isotopes. If the abundance of U-234 is 0.01%, the abundance of U-235 is 0.71%, and the abundance of U-238 is 99.28%, what is the average atomic mass of uranium?

7. Titanium has five common isotopes: Ti-46 (8.0%), Ti-47 (7.8%), Ti-48 (73.4%), Ti-49 (5.5%), Ti-50 (5.3%). What is the average atomic mass of titanium?

8. Naturally occurring chlorine that is put in pools is 75.53 percent C-35l (mass = 34.969 amu) and 24.47 percent C-37 (mass = 36.966 amu). Calculate the average atomic mass.

9. Copper used in electric wires comes in two flavors (isotopes): C-63u and Cu-65. Cu-63 has an atomic mass of 62.9298 amu and an abundance of 69.09%. The other isotope, Cu-65, has an abundance of 30.91%. The average atomic mass between these two isotopes is 63.546 amu. Calculate the actual atomic mass of Cu-65.

Atomic Structure Review Sheet

***THIS REVIEW DOES NOT COVER EVERYTHING THAT COULD BE ON THE TEST! Study your packet.

1. Complete the following table:

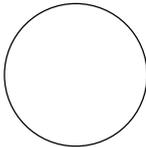
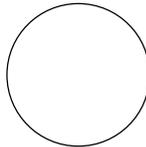
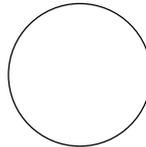
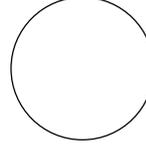
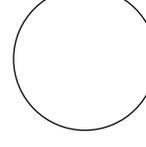
Particle Name	Charge (+)(-)(0)	Mass (1 or 0)	Location in Atom (nucleus or cloud)	What tells you how many are in an atom? (atomic number, mass number...)
PROTON				
ELECTRON				
NEUTRON				

2. Complete the following table:

Element Symbol	Atomic Number	Protons	Neutrons	Electrons	Mass Number
B			6		
	29		35		
Pb					207
	43				99
				28	59

3. Name the first four energy levels (shells) and the maximum number of electrons that fit in each shell.

4. Draw **Bohr** diagrams for the following elements.

lithium	oxygen	neon	magnesium	sulfur
				

6. Name the first three orbitals and the maximum number of electrons that fit in each orbital.

6. For the following elements, write the complete electron configuration and abbreviated configuration.

beryllium	Symbol:	Complete electron configuration:	Abbreviated configuration: []
neon	Symbol:	Complete electron configuration:	Abbreviated configuration: []
aluminum	Symbol:	Complete electron configuration:	Abbreviated configuration: []
potassium	Symbol:	Complete electron configuration:	Abbreviated configuration: []

7. Draw the Lewis dot diagrams for the following **atoms and their ions**.

helium	atom	ion	calcium	atom	ion
nitrogen	atom	ion	bromine	atom	ion
hydrogen	atom	ion	arsenic	atom	ion

9. What is the #1 goal of all atoms?

10. What is an ion? Give an example of a cation AND an anion.

11. What is an isotope? Give an example.

12. Chromium is found to exist in four isotopic forms. Calculate the average atomic mass of chromium based upon the abundance of these isotopes. Show your work!

Cr-50	49.9461 amu	and 4.35 %
Cr-52	51.9405 amu	and 83.79 %
Cr-53	52.9407 amu	and 9.50 %
Cr-54	53.9389 amu	and 2.36 %

13. Silicon is found to exist in three isotopic forms. Calculate the average atomic mass of silicon based upon the abundance of these isotopes. Show your work!

Si-28	27.97693 amu	and 92.23 %
Si-29	28.97649 amu	and 4.67 %
Si-30	29.97376 amu	and 3.10 %

14. Complete each sentence in the blank to the left of each of the following statements.

_____ 1. Electrons that are in the outermost energy level are called _____ electrons.

_____ 2. Isotopes of an element differ in their number of _____.

_____ 3. An atom that has either gained or lost electrons is called a/an _____.

_____ 4. A positively charged atom is called a/an _____.

_____ 5. A negatively charged atom is called a/an _____.

15. What is the difference between a mass number and an average atomic mass? Which is listed on the periodic table of elements?

16. State whether each of the following atoms would be expected to form a cation (C) or an anion (A)

a. Sodium (Na)

c. Calcium (Ca)

b. Fluorine (F)

d. Oxygen (O)

