Chemistry: Form WS2.1.1A

ATOMS

Name ______________________________

Date ________________ Period ______

The Dalton Model

In 1803, John Dalton proposed the theory that all modern chemistry is based on. It is called the Atomic Theory of Matter. He stated the following:

- ☆ Matter is made of small particles called atoms.
- ☆ Atoms are indestructible. They cannot be created or destroyed during chemical or physical changes.
- ☆ Atoms of an element are identical. They have the same mass.
- ☆ Atoms of different elements have different masses.
- ☆ Compounds are formed by combining atoms of different elements.

The idea of atoms was not new, even in 1803. It was proposed centuries earlier, around 460 BC, by Democritus, a Greek philosopher. Dalton's ideas were different than those of Democritus because they were supported by quantitative data.

* * *

Following is an example of the type of data Dalton analyzed. Examine the data. Then explain in your own words what the data might mean.

Sample Data for Dalton's Experiment

Experiment 1	Experiment 2
★ 30 g of carbon are burned in the open air where there is plenty of oxygen. 110 g of carbon oxide forms.	★ 30 g of carbon are burned in a closed container under conditions of low oxygen. 70 g of carbon oxide forms.
★ How much oxygen does this mean the carbon combined with? ____________________	★ How much oxygen does this mean the carbon combined with? ____________________

1. What is the ratio of the masses of the oxygen combined with carbon in *Experiment 1* and *Experiment 2*? _

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2. Is the ratio of masses of combined oxygen in *Experiment 1* and *Experiment 2* an integer or an irrational number? ____

3. How might this data be explained? ____________________________________

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Name ____________________

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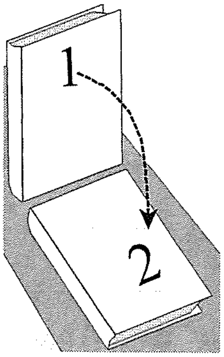
Understanding the Bohr Model

When atoms absorb energy, the electrons speed up and move further away from the nucleus. Eventually they lose this extra energy as light. Neils Bohr had a perplexing problem. He noticed that the light given off by the electrons of excited atoms never produced a full spectrum. Instead, the light consisted of bright lines of various frequencies with none of the in-between frequencies represented. He also had a complex equation into which he could substitute simple numbers, integers such as 1, 2, or 3, and the equation predicted the frequencies of the bright lines. The problem was, he did not know what these integers represented. The whole mystery unraveled, however, when Bohr developed a model of the atom with circular pathways for the electron. These pathways were at fixed distances from the nucleus. Electrons could be found only in these circular pathways. If an electron absorbed enough energy, it could jump up to another level, but it could never be found between levels. Inevitably, the electron lost energy and fell back down to a previous level, giving off the extra energy as a specific frequency of light. Bohr's mystery numbers represented the energy levels of the electrons.

Refer to the introduction above and your knowledge of chemistry and the world to answer the questions below.



1. The drawing to the right shows a book in two possible positions on a table. In which position does the book have more energy? ____________________

2. What will probably happen to the book if it is left in "position 1" for an extended period of time? Why? ____________________

3. Can the book stand between "position 1" and "position 2"? Explain. ____________________

4. If the book falls from "position 1" to "position 2," over and over, how does the energy change compare from one time to the next? ____________________

5. The drawing to the right shows an electron in "position 1." What will happen to the electron over time? What is the evidence for this? ____________________

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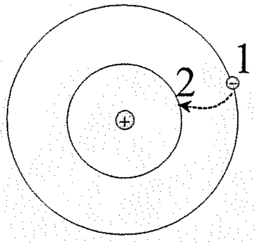

6. According to Bohr, can the electron referred to in the previous question be found between "position 1" and "position 2?" What is Bohr's evidence? ____________________

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7. When the electron moves from "position 1" to "position 2," how does the energy change compare from one time to the next? ____________________

8. Barium always gives a green flame test. Why? ____________________

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9. What is Bohr's model of the atom? How does the evidence support his model? ____________________

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Chemistry: Form WS2.2.1A

ATOMS

Name ______________________

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Understanding the Rutherford Model

Ernest Rutherford performed an experiment in 1911 that helped him develop the solar system model of the atom. He probed the inside of the atom using small, positively charged particles called alpha particles. Based on his observations, he suggested that the atom is mostly empty space with a small, positively charged center and negatively charged electrons revolving around the outside like planets around the sun. This is the image that most people have of an atom, but how did it get that way? By answering a series of questions below and analyzing Rutherford's experiment, you will find out.

Answer the questions below based on your knowledge of the world and on the description of Rutherford's Alpha Scattering Experiment which follows.

[1] If you toss a tennis ball at a brick wall, what will happen? ______________________

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[2] If you toss a tennis ball at something that looks like a solid wall, but it is actually smoke, what will happen?

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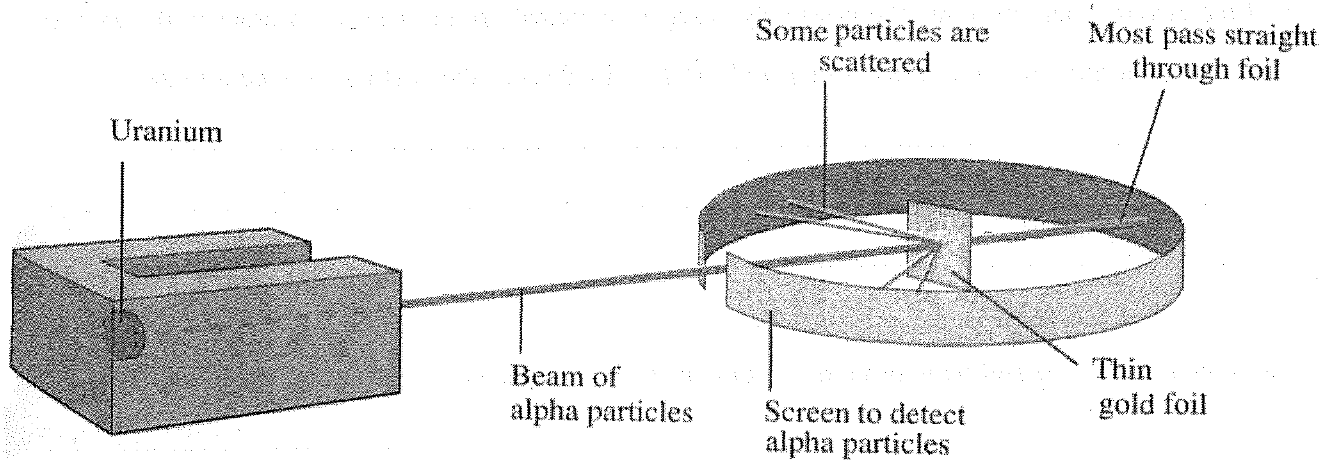
Rutherford performed the following experiment: He aimed a beam of high speed, positively charged particles called alpha particles (similar to our tennis ball) at a piece of solid gold foil (similar to our wall). He set up a special screen all around the foil to help him see where the particles went.



[3] According to Dalton's model, the atom is a solid sphere. What would the alpha particles do when they hit the gold foil if Dalton were correct? ______________________

[4] According to Thomson, the atom is a positively charged cloud with electrons scattered throughout. What would the alpha particles do when they hit the foil if Thomson were correct? ______________________

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[5] When Rutherford performed his experiment, only 1 in 20,000 alpha particles bounced straight back or were deflected greatly. The rest went straight through the gold foil.

a. What does this indicate about the probability of actually hitting anything? ____________________

____

b. What does this indicate about the size of whatever has been hit compared to the size of the gold atoms in the foil? ____

c. Is the atom mostly solid or mostly space? How do you know? ____________________

____

d. Considering the fact that alpha particles are positively charged, what must the charge be on whatever deflected them? ____

e. Based on this evidence, what is in an atom's center? ____________________

____

f. Where might the negatively charged electrons be located? ____________________

____

g. If the electrons and protons (the positively charged particles) are not near each other in the atom, why doesn't their attraction pull them together? [*HINT:* Why doesn't the Earth get pulled into the sun?] ____

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[6] Based on the evidence and on your understanding of the earlier questions, propose a model of the atom. ___

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Chemistry: Form WS2.3.1A

ATOMS

Name ______________________________

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The Number of Neutrons

Atoms are composed of electrons, protons, and neutrons. The relative mass of the electrons is 0 amu, while the relative masses of the protons and neutrons are each 1 amu. The mass of an atom (A) is there fore equal to the sum of the atoms protons (Z) and neutrons (N). The number of protons (Z) is also known as the atomic number. Atomic number and atomic mass are both given on the periodic table. The mass listed on the periodic table for each element is the average mass of the isotopes. When this mass is rounded off, it gives the mass of the most common isotope. The number of neutrons for the most common isotope of an element can be found by rounding off the mass of the element on the periodic table and subtracting the atomic number from it. See the box to the right.

Calculating the Number of Neutrons

A = atomic mass number
Z = atomic number
N = number of neutrons

☞ N = A – Z

☞ $^{A}_{Z}X$ is the symbol for an isotope of element X with atomic number Z and atomic mass number A

☞ For the element $^{23}_{11}Na$, the number of neutrons (N) is determined as follows:
N = 23 – 11 = 12

Fill in the table below using data from the *Periodic Table.* For each element listed below, look up the mass, round it off, and record the result. Then look up the atomic number, write the symbol for the most common isotope and calculate the number of neutrons.

Element	Mass (A)	Atomic Number (Z)	Isotopic Symbol ($^{A}_{Z}X$)	Number of Neutrons (N)
Oxygen				
Chlorine				
Calcium				
Iron				
Iodine				
Silver				
Bromine				
Copper				
Potassium				
Phosphorus				
Tin				
Zinc				
Sulfur				
Neon				

Chemistry: Form WS2.3.2A

Name ______________________

ATOMS

Date ______________ Period ____

Analyzing the Bohr Atom

Use the information supplied in the table to fill in the remaining blanks in each row.

Number of Electrons	Electron Configuration	Mass Number	Number of Protons	Number of Neutrons	Element	Symbol	Isotope Notation
						B	
	2-7						
13							
					carbon		
			11				
				0			
			19				
						Cl	
							$^{16}_{8}O$
	2-8-8-2						
					sulfur		
18							
		9					
							$^{4}_{2}He$
	2-8-9-2						

Chemistry: Form WS2.3.3B

ATOMS

Name ____________________

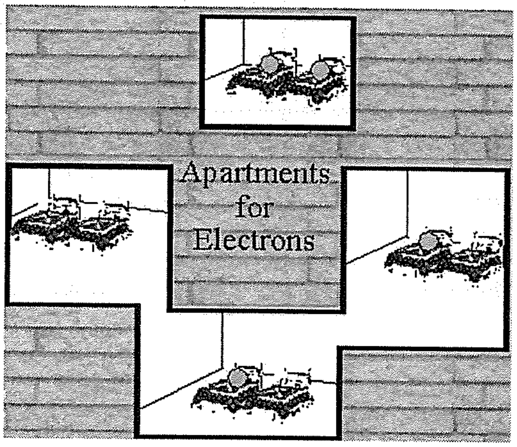
Date ______________ Period ____

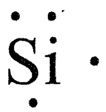
Electron Dot Diagrams, Etc.

Electron dot diagrams are a useful way to show the arrangement of outer electrons of an atom. They show valence electrons as dots at 12 o'clock, 3 o'clock, 6 o'clock, and 9 o'clock, and the rest of the atom, known as the kernel, as a symbol.

It is useful to think of the outer shell as if it contained two apartments, one with one bedroom and the other with three bedrooms. Each bedroom has space for two occupants. Think of electrons as frugal little fellows who do not like to share. Electrons will prefer to move into the cheaper, one bedroom apartment if it is available even if it means sharing a room with another electron. If the cheap apartment is not available, they'll settle for the three bedroom apartment, but they won't share a bedroom until the apartment becomes too crowded to have a choice. As a result, the first space gets two electrons before any of the other spaces get electrons. The next three spaces get the electrons one at a time until pairing becomes necessary.

Silicon, for example, has four valence electrons. As a result, it will have two electrons in one of the clock positions and one electron in each of two of the remaining three.





The dots representing the electrons in silicon are much like the ones occupying the apartments above.

Use the information supplied in the table to fill in the remaining blanks in each row.

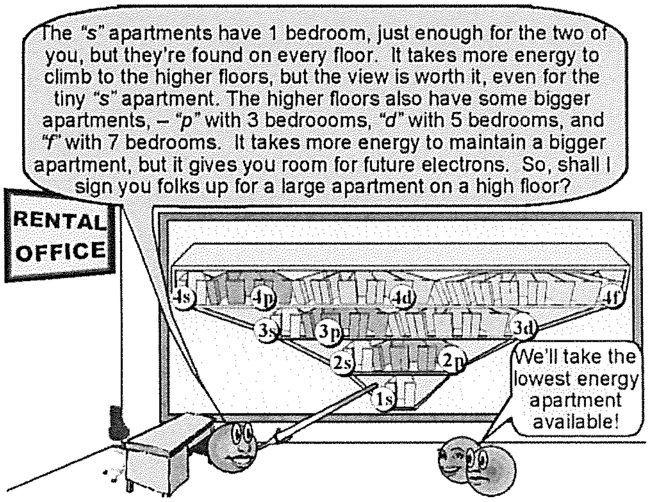
Mass Number	**Atomic Number**	**Isotope Notation**	**Number of Neutrons**	**Bohr Notation**	**Electron Dot**
				2-8-7	
					·Ö:
			2		
20					
		$^{27}_{13}\mathrm{Al}$			
				2-8-18-5	
	20				
					:Är:

Mass Number	Atomic Number	Isotope Notation	Number of Neutrons	Bohr Notation	Electron Dot
31					
	36				
			0		
		$^{14}_{6}\mathrm{C}$			
	21				
					$\overset{\cdot\cdot}{\mathrm{Ba}}$
7					

Chemistry: Form WS2.S2A Name ______________________________

ATOMS Date _________________ Period _____

Location of Electrons



The most common difficulty renting to electrons

Electrons are in regions of the atom known as orbitals. Roughly speaking, they are located in principal energy levels similar to the shells or energy levels of the Bohr model. Each of the energy levels is designated by a quantum number, *n*, from 1 to 7. None of the known elements has atoms with more than 7 principal energy levels. The principal energy level with the lowest energy is 1. The highest is 7. Principal energy levels can be thought of as being subdivided into energy sublevels. The maximum number of sublevels in a principal energy level is *n*, but none of the existing elements use more than 4 sublevels even in principal energy levels 5–7. Sublevels are designated by the letters s, p, d, and f, in increasing order of energy The orbitals are regions within a sublevel where electrons of a given energy are likely to be found. There are a maximum of 2 electrons in an orbital. A useful analogy to help you visualize this is an apartment building. Each floor represents a different principal energy level. Each apartment represents a sublevel. Each bedroom represents an orbital. The electrons are the tenants in the bedrooms. Electrons are most likely to be found in the lowest energy locations available. Knowing this, it is possible to figure out how the electrons are arranged in an atom.

The number of orbitals within a sublevel varies in a predictable pattern. The number of orbitals within a sublevel and the maximum number of electrons is as follows:

Sublevel	**s**	**p**	**d**	**f**
Number of orbitals	1	3	5	7
Maximum Number of Electrons	2	6	10	14

The first energy level has only one sublevel, s; the second energy level has two sublevels, s and p; the third energy level has three sublevels, s, p, and, d; and so on. This results in the pattern shown below.

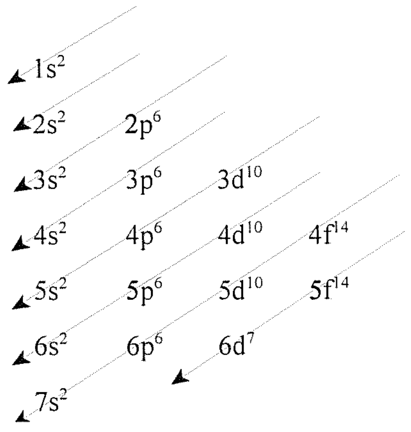
Summary

Principal Quantum Number (n)	**Number of Orbitals (n^2)**	**Orbitals per Sublevel**				**Maximum Number of Electrons ($2n^2$)**
		s	**p**	**d**	**f**	
1	1	1	-	-	-	2
2	4	1	3	-	-	8
3	9	1	3	5	-	18
4	16	1	3	5	7	32

Continue ☞

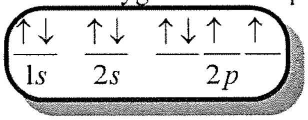
The electrons are arranged according to the following rules:

1. the number of electrons equals the number of protons (atomic number)
2. electrons occupy orbitals in sequence beginning with those of the lowest energy (see diagram to the right)
3. in a given sublevel, a second electron is not added to an orbital until each orbital in the sublevel contains one electron



This results in the order of filling for elements 1 to 109 pictured to the right. Follow each arrow from beginning to end. Then go to the beginning of the next arrow down. When you follow this pattern, you will note that no more than four orbitals are occupied in the outermost principal energy level. This is because, once the p sublevel is filled, the next energy sublevel is always the s in the next principal energy level. Oxygen has 8 protons and 8 electrons. Its electron configuration in sublevel notation is as follows: $1s^2 2s^2 2p^4$. This means there are 2 electrons in the first level and 6 in the second (add the superscripts). As a result the electron arrangement can also be written as follows: 2–6. This is known as Bohr notation.

Remember, electrons never pair in an orbital until every orbital in a sublevel has an electron. When they do pair, they spin in opposite directions. This reduces the repulsion between them. The opposite spins of the electrons are represented by up arrows and down arrows. The electron configuration of oxygen can be depicted as follows:



Each horizontal line represents an orbital in a sublevel. Each arrow represents an electron in an orbital. This is called orbital notation.

For each of the elements below, write the sublevel notation, the Bohr notation, and the orbital notation.

Element	Atomic Number	Electron Configuration		
		Sublevel Notation	Bohr Notation	Orbital Notation
H	1			
N	7			
Ca	20			
Al	13			
Cu	29			
C	6			
Ar	18			
Na	11			
S	16			
Ne	10			
P	15			

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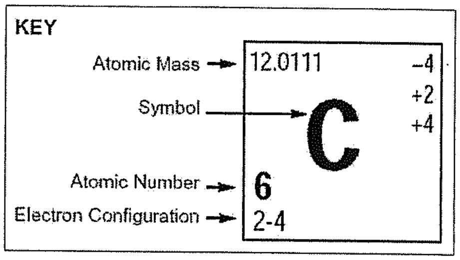
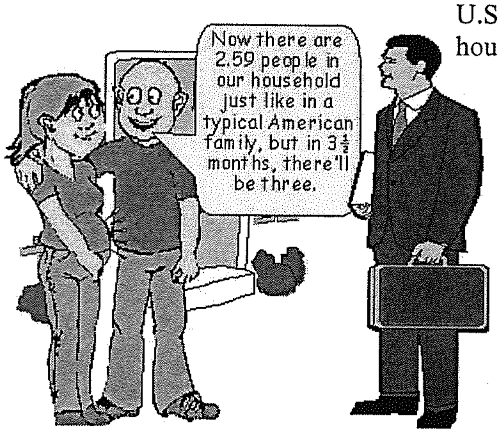
Average Atomic Mass

According to the *Periodic Table,* carbon has a mass of 12.0111 *amu*. The mass of an atom comes from its protons and neutrons. Carbon has 6 protons. Subtracting the number of protons from the mass to get the number of neutrons gives 6.0111. This result is impossible! There are no fractions of neutrons!! Nonsensical numbers such as these exist elsewhere. For example, the U.S. Census reports 2.59 people per household in 2000. You can count on the fact that there is no household with 2.59 people. This number is an average.





Likewise, there is no such thing as a carbon atom with a mass of 12.0111 *amu*. The atomic mass listed on the *Periodic Table* is the average mass of the isotopes of carbon. Carbon has two naturally occurring stable isotopes. The large majority of carbon atoms, 98.89%, are ^{12}C, while only 1.108% are ^{13}C. That is why the average mass is so close to 12.

Average Atomic Mass

$$m_{avg} = p_1 m_1 + p_2 m_2 + \dots p_n m_n$$

m_{avg} – average mass; p_1 – percentage of isotope 1;
m_1 – mass of isotope 1; p_2 – percentage of isotope 2;
m_2 – mass of isotope 2; p_n – percentage of isotope n;
m_n – mass of isotope n; n – the number of isotopes

Example

What is the average mass of chlorine if a sample consists of 77.35% Cl–35 and 22.65% Cl–37?

$$\begin{aligned} m_{avg} &= (0.7735)(35) + (0.2265)(37) \\ &= 27.07 + 8.381 \\ &= 35.45 \end{aligned}$$

The average mass is determined by the procedure illustrated in the box to the right. The mass of each isotope is multiplied by its percentage. Then these products are added to find the average.

Use this procedure to find the average mass for each of the examples below.

1. An element recently discovered in Spring Valley High School called "valleyum" occurs in the following proportions: 81.75% Vm–500 and 18.25% Vm–502
2. A new element, "schoolium," found only in schools occurs in the following proportions: 9.750% Sl–750 and 90.25% Sl–752
3. A new element "newium" has been discovered which occurs in the following proportions: 25.25% Nw–300; 15.25% Nw–301; and 59.50% Nw–302
4. A new element found in the science department office called "departmentium" is found in the following proportions: 11.78% Dt–800; 1.850% Dt–801; and 86.37% Dt–803