

ATOMIC THEORY

History of the Atom Timeline

Democritus 460 BC
and Dalton 1803 AD



Thomson
1897



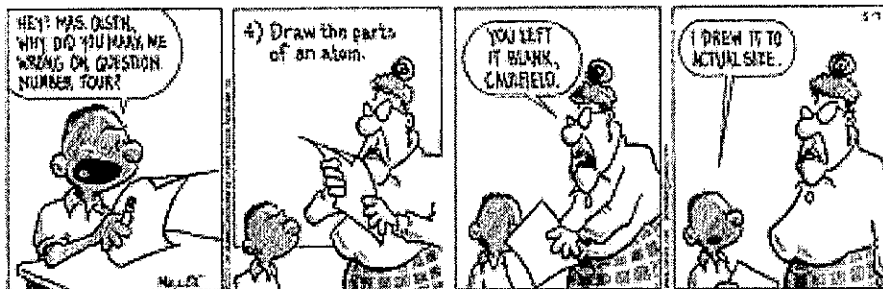
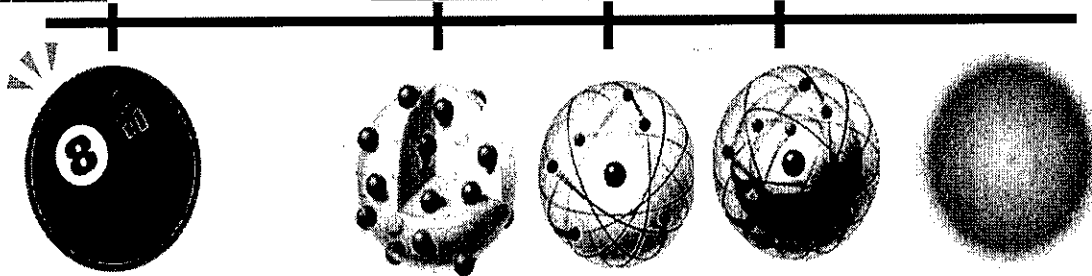
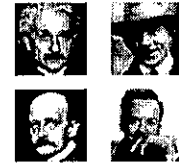
Rutherford
1912



Bohr
1913



Modern
Quantum
Cloud Model
post 1930



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.

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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

- ★ 30 g of carbon are burned in the open air where there is plenty of oxygen. 110 g of carbon oxide forms.
- ★ How much oxygen does this mean the carbon combined with? \_\_\_\_\_

#### Experiment 2

- ★ 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? \_\_\_\_\_

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

# 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? \_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

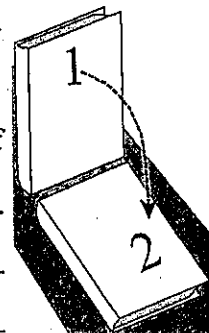
3. Can the book stand between "position 1" and "position 2"? Explain. \_\_\_\_\_

\_\_\_\_\_

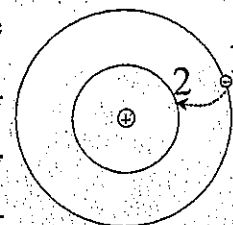
\_\_\_\_\_

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? \_\_\_\_\_



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? \_\_\_\_\_

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? \_\_\_\_\_

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

## Understanding the Rutherford Model



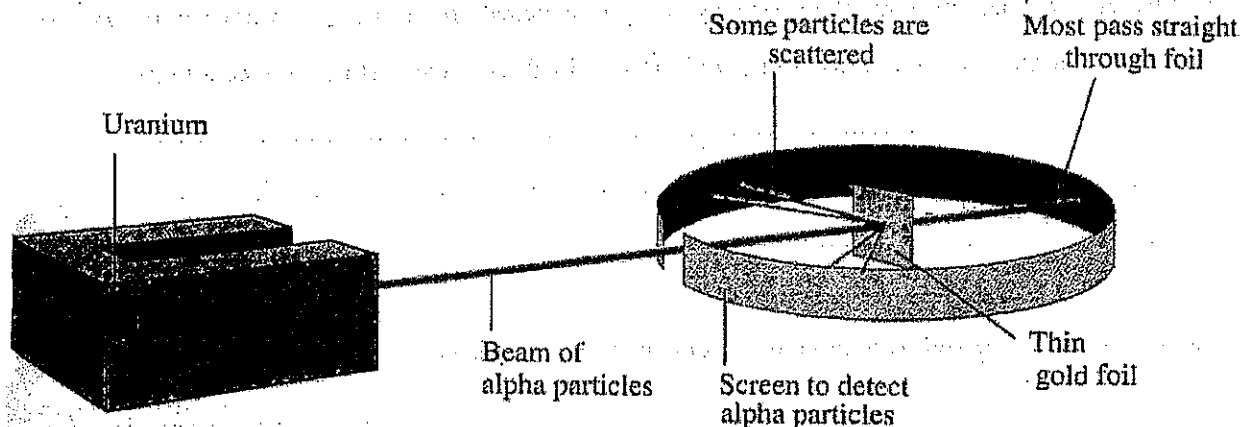
Ernest Rutherford performed an experiment in 1911 that helped him develop the solar system model of the atom. He probed the inside of the atom using small, positively charged particles called alpha particles. Based on his observations, he suggested that the atom is mostly empty space with a small, positively 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? \_\_\_\_\_

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

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?] \_\_\_\_\_

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

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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 therefore 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.

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.

### Calculating the Number of Neutrons

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

- ☛  $N = A - Z$
- ☛  ${}^A_ZX$  is the symbol for an isotope of element X with atomic number Z and atomic mass number A
- ☛ For the element  ${}^{23}_{11}\text{Na}$ , the number of neutrons (N) is determined as follows:  
 $N = 23 - 11 = 12$

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

## 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     |                   |
|                     |                        |             |                   |                    |         |        | $^{16}_8\text{O}$ |
|                     | 2-8-8-2                |             |                   |                    |         |        |                   |
|                     |                        |             |                   |                    | sulfur  |        |                   |
| 18                  |                        |             |                   |                    |         |        |                   |
|                     |                        | 9           |                   |                    |         |        |                   |
|                     |                        |             |                   |                    |         |        | $^4_2\text{He}$   |
|                     | 2-8-9-2                |             |                   |                    |         |        |                   |



6 • Structure of the Atom  
The Subatomic Particles  
(1 of 8)

| <u>Name</u> | <u>Symbol</u>  | <u>Mass</u>        | <u>Charge</u> | <u>Location</u>                              |
|-------------|----------------|--------------------|---------------|----------------------------------------------|
| protons     | p <sup>+</sup> | 1 u                | 1+            | part of the nucleus                          |
| neutron     | n <sup>o</sup> | 1 u                | 0             | part of the nucleus                          |
| electron    | e <sup>-</sup> | $\frac{1}{1837}$ u | 1-            | normally at large distances from the nucleus |

J.J. Thompson is given credit for discovering **electrons** using a Crookes tube and testing many different gases. Cathode rays were found to be beams of electrons.

Cavendish is given credit for the discovery of the **neutron**.

6 • Structure of the Atom  
Terms I-- Atomic Structure  
(2 of 8)

|                          |                                                                                                                        |
|--------------------------|------------------------------------------------------------------------------------------------------------------------|
| <b>atoms</b>             | the smallest particle of an element . It consists of a central nucleus and electron clouds outside the nucleus.        |
| <b>nucleus subatomic</b> | the dense central portion of an atom. smaller than an atom. The proton, neutron, and electron are subatomic particles. |
| <b>net charge</b>        | the difference in the positive charge due to protons and the negative charge due to electrons in an atom.              |
| <b>nucleons</b>          | the particles that make up the nucleus.                                                                                |

6 • Structure of the Atom  
Terms II-- Atomic Structure  
(3 of 8)

|                          |                                                                                                                                          |
|--------------------------|------------------------------------------------------------------------------------------------------------------------------------------|
| <b>atomic number</b>     | the number of protons in an atom. This number determines the identity of an element.                                                     |
| <b>mass number</b>       | the number of protons + neutrons                                                                                                         |
| <b>isotopes</b>          | atoms with the same number of protons, but different numbers of neutrons. Atoms with the same atomic number, but different mass numbers. |
| <b>isotopic notation</b> | shorthand notation for a nucleus that shows the mass #, atomic # and the symbol. U-238 would be ${}_{92}^{238}\text{U}$                  |

6 • Structure of the Atom  
Calculating Atomic Mass  
(4 of 8)

Any **real** sample of an element contains more than one naturally occurring **isotope**. For instance, **boron**

| isotope                        | abundance | mass # | isotopic mass   |
|--------------------------------|-----------|--------|-----------------|
| boron-10 ${}_{5}^{10}\text{B}$ | 19.78%    | 10     | mass = 10.013 u |
| boron-11 ${}_{5}^{11}\text{B}$ | 80.22%    | 11     | mass = 11.009 u |

The atomic mass is the weighted average of the isotopes.

$$\text{at. mass} = \frac{(19.78\%)(10.013\text{u}) + (80.22\%)(11.0009\text{u})}{100} \text{ or}$$

$$\text{at. mass} = (0.1978)(10.013\text{u}) + (0.8022)(11.0009\text{u}) = 10.81 \text{ u}$$

**6 • Structure of the Atom**  
**Determining Numbers of Protons, Neutrons,**  
**and Electrons from the Isotopic Notation**  
(5 of 8)

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Consider the following symbol:  ${}_{16}^{33}\text{S}^{2-}$

The 16 is the **atomic number** which is the number of **protons**.

The 33 is the **mass number** which is the mass of one of the isotopes. This mass is due to the **protons and neutrons**.

The number of **neutrons** is the **mass number - the atomic number**.  $33 - 16 = 17$  neutrons.

Since the charge is 2-, there are **2 more electrons than protons**. In this case, there are 18 electrons.

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**6 • Structure of the Atom**  
**Important People in the Development of the**  
**Atomic Theory**  
(6 of 8)

**Democritus** [atomos]

philosopher who decided that matter was discontinuous

**John Dalton** [billiard-ball model]

experiments with gases... different substances are different combinations of atoms

**J.J. Thomson** [plum-pudding model]

experiments with gas-discharge tubes... atoms have positive and negative parts... the negative electrons are the same from atom to atom

**Ernest Rutherford** [nuclear model/solar system model]

most of the mass of the atom is concentrated in a tiny, positively-charged nucleus

**Niels Bohr** [Bohr Model (quantized e- energy levels/orbits)]

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**6 • Structure of the Atom**  
**Early Experimental Observations That Would**  
**Later Be Explained By The Atomic Theory**  
(7 of 8)

**The Law of Conservation of Mass**

the mass of all the reactant molecules = the mass of all the product molecules

**The Law of Definite Composition**

the percentage composition of any sample of a substance is the same

**The Law of Multiple Proportions**

when two compounds made of the same two elements (such as CO and CO<sub>2</sub>) are broken down to give the same mass of *one* element... the masses of the *other* element will be in simple whole-number ratio.

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**6 • Structure of the Atom**  
**Rutherford's Gold Foil Experiment**  
(8 of 8)

**Ernest Rutherford's classic gold foil experiment** led to the **nuclear model** of the atom.



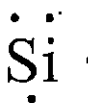
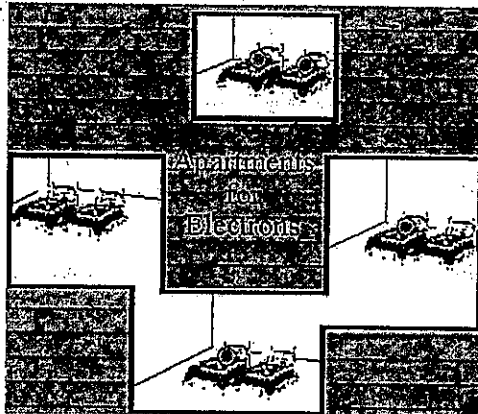
- **the nucleus is tiny** - because most of the alpha's missed the nucleus and went straight through the foil
  - **the nucleus is positively charged** - because the (+) charged alpha was repelled by the (+) charged nucleus
  - **the nucleus is incredibly dense** - because the nucleus was able to bounce back at a very large angle
-

## 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.

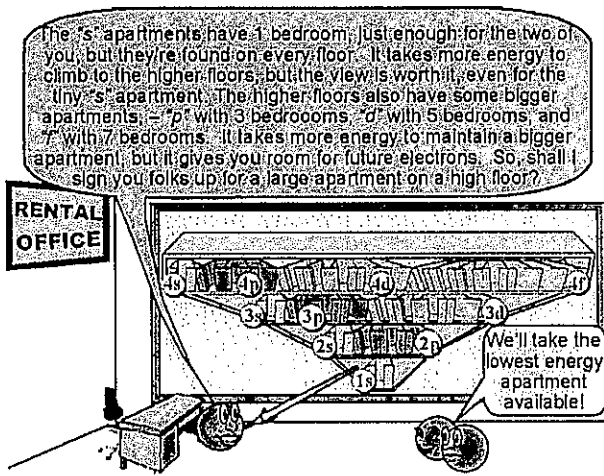
Use the information supplied in the table to fill in the remaining blanks in each row.

| Mass Number | Atomic Number | Isotope Notation        | Number of Neutrons | Bohr Notation | Electron Dot                   |
|-------------|---------------|-------------------------|--------------------|---------------|--------------------------------|
|             |               |                         |                    | 2-8-7         |                                |
|             |               |                         |                    |               | $\cdot \ddot{\text{O}} \cdot$  |
|             |               |                         | 2                  |               |                                |
| 20          |               |                         |                    |               |                                |
|             |               | ${}_{13}^{27}\text{Al}$ |                    |               |                                |
|             |               |                         |                    | 2-8-18-5      |                                |
|             | 20            |                         |                    |               |                                |
|             |               |                         |                    |               | $\cdot \ddot{\text{Ar}} \cdot$ |

| Mass Number | Atomic Number | Isotope Notation  | Number of Neutrons | Bohr Notation | Electron Dot |
|-------------|---------------|-------------------|--------------------|---------------|--------------|
| 31          |               |                   |                    |               |              |
|             | 36            |                   |                    |               |              |
|             |               |                   | 0                  |               |              |
|             |               | $^{14}_6\text{C}$ |                    |               |              |
|             | 21            |                   |                    |               |              |
|             |               |                   |                    |               | Ba           |
| 7           |               |                   |                    |               |              |

## Location of 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 most common difficulty renting to electrons

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

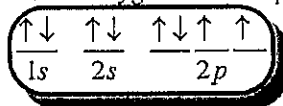
| Principal Quantum Number ( $n$ ) | Number of Orbitals ( $n^2$ ) | Orbitals per Sublevel |   |   |   | Maximum Number of Electrons ( $2n^2$ ) |
|----------------------------------|------------------------------|-----------------------|---|---|---|----------------------------------------|
|                                  |                              | 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                                     |

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:



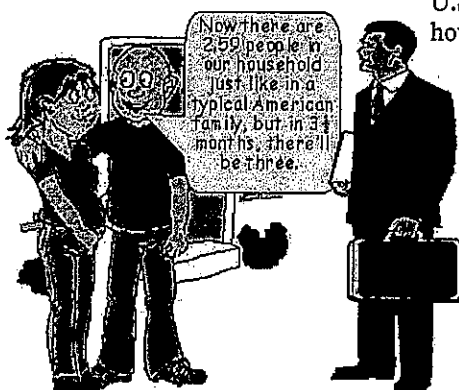
Each horizontal line represents an orbital in a sublevel. Each arrow represents an electron in an orbital. This is called orbital notation.

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            |                        |               |                  |

## 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.



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.

| KEY                      |         |    |
|--------------------------|---------|----|
| Atomic Mass →            | 12.0111 | -4 |
| Symbol →                 | C       | +2 |
| Atomic Number →          | 6       | +4 |
| Electron Configuration → | 2-4     |    |

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  $^{12}\text{C}$ , while only 1.108% are  $^{13}\text{C}$ . That is why the average mass is so close to 12.

### Average Atomic Mass

$$m_{\text{avg}} = p_1m_1 + p_2m_2 + \dots p_nm_n$$

$m_{\text{avg}}$  – 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?

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

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% SI-750 and 90.25% SI-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