



ManaBarnootaaHaayiliigi
ሃይሊግትምህርት ቤት
Heilig School

☎0221124055

✉1045 Adama, Ethiopia E-mail: heiligschool@gmail.com

Grade 9 Subject: chemistry

Date:- 24/05/2013 -06/06/2013 E.C.

1st Sem

Note: 4

1.3.1 Atomic Number and Mass Number

Atoms of different elements have different number of protons. Atoms of the same element have the same number of protons. The Atomic Number (Z) of an element is equal to the number of protons in the nucleus of an atom. It is also equal to the number of electrons in the neutral atom.

$$Z = p^+ \text{ or } Z = e^- \text{ (in a natural atom)}$$

For different elements, the nuclei of their atoms differ in the number of protons they contain. Therefore, they differ in the amount of positive charge that their atoms possess. Thus, the number of protons in the nucleus of an atom determines the identity of an atom. Each atom of an element is identified by its atomic number. In the periodic table, the atomic number of an element is indicated below its symbol. Notice that the elements are placed in order of increasing atomic number. Hydrogen, H, has atomic number 1, hence, atoms of the element hydrogen have one proton in the nucleus. Next is helium, He, which has two protons in its nucleus. Lithium, Li, has three protons. Beryllium, Be, has four protons, and so on. Mass Number (A) is the sum of the number of protons and neutrons in the nucleus of an atom. Collectively, the protons and neutrons of an atom are called its nucleons.

Mass Number (A) = Number of protons + Number of neutrons

$$A = p^+ + n^0$$

The name of an element, its atomic number and mass number can be represented by a shorthand symbol. For example, to represent neutral atoms of magnesium with 12 protons and 12 neutrons, the symbol is $^{24}_{12}\text{Mg}$. The atomic number is written as a subscript to the left of the symbol of magnesium. The mass number is written as a superscript to the left of the symbol. This method can be represented in general by the following illustration, in which X stands for any chemical symbol.

MassNumber A

AtomicNumber Z $^{24}_{12}\text{X}$

Mg can be read as “Magnesium-24”, and “Magnesium-25” represents Mg.

The number of neutrons in an atom is found by subtracting the atomic number from the mass number. For example, the number of neutrons for a silver atom ($^{108}_{47}\text{Ag}$) is:

Mass number – Atomic number = Number of neutrons

$$108 - 47 = 61$$

∴ The number of neutrons is 61.

Example 1

How many protons, electrons, and neutrons are found in an atom of bromine-80 if its atomic number is 35?

Solution:

Atomic number of bromine is 35.

Mass number of bromine is already given as 80. Now we can write bromine, Br in the form of:



35

Mass number – Atomic number = Number of neutrons

$$80 - 35 = 45 \text{ neutrons}$$

For a neutral atom: number of protons = number of electrons = atomic number. \ Number of protons = Number of electrons = 35

1.3.2 Isotopes and Atomic Mass

The simplest atoms are those of hydrogen. Each hydrogen atom contains one proton only. However, hydrogen atoms contain different numbers of neutrons. Three types of hydrogen atoms are known. The most common type of hydrogen is called **protium**. It accounts for 99.985% of the hydrogen atoms found on Earth. The nucleus of a protium atom consists of one proton only, and it has one electron moving around it. The second form of hydrogen atom is called deuterium, which accounts for

0.015% of the Earth's hydrogen atoms. Each **deuterium** atom has a nucleus containing one proton and one neutron. The third form of hydrogen atom is known as **tritium**, and it is radioactive. It exists in very small amounts in nature, but it can be prepared artificially. Each tritium atom contains one proton and two neutrons. Protium, deuterium and tritium are the three isotopes of hydrogen.

- ❖ Isotopes are the atoms of the **same element** that have the **same atomic number** but **different masses number**. The isotopes of a particular element have the same number of protons and electrons but different number of neutrons. Isotopes of an element may occur naturally, or they may be made in the laboratory (artificial isotopes). Since isotopes have the same number of protons and electrons, they have the same chemical properties.

Relative Atomic Masses

Actual masses of atoms, measured in grams, are very small. An atom of oxygen-16, for example, has a mass of $2.657 \times 10^{-23}\text{g}$. For most chemical calculations, it is more convenient to use relative atomic masses. In order to set up a relative scale of atomic mass, one atom has been arbitrarily chosen as the standard and assigned a relative mass value. The masses of all other atoms are expressed in terms of this defined standard.

The standard used by scientists to determine units of atomic mass is the carbon-12 atom. It has been arbitrarily assigned a mass of exactly 12 atomic mass units, or 12a.m.u. One atomic mass unit, or 1 a.m.u, is exactly 1/12 the mass of a carbon-12 atom, or $1.660 \times 10^{-24}\text{g}$. The atomic mass of any atom is determined by comparing it with the mass of the carbon-12 atom. The hydrogen-1 atom has an atomic mass of about 1/12 of that of the carbon-12 atom, or about 1 a.m.u. The precise value of the atomic mass of a hydrogen-1 atom is 1.007 a.m.u. An oxygen- 16 atom has about 16/12 (or 4/3 the mass of a carbon-12 atom). Careful measurements show the atomic mass of oxygen-16 to be 15.994 a.m.u. The mass of a magnesium-24 atom is found to be slightly less than twice that of a carbon-12

atom. Its atomic mass is 23.985 a.m.u

Average Atomic Masses of Elements

Most elements occur naturally as mixtures of isotopes, as shown in the examples in The percentage of each isotope in the naturally occurring elements is nearly always the same, no matter where the element is found. The percentage at which each isotope of an element occurs in nature is taken into account when calculating the average atomic mass of the element. Average

atomic mass is the weighted average of the atomic masses of the naturally occurring isotopes of an element.

included in the calculation of atomic mass of hydrogen.

The average atomic mass of an element is calculated by summing up the products of relative atomic mass and percentage abundance of each isotope and dividing by 100.

Example

Chlorine has two isotopic compositions: 75.77 % of $^{35}_{17}\text{Cl}$ and 24.23 % of $^{37}_{17}\text{Cl}$. Determine the average atomic mass of chlorine.

Solution:

75.77 % of natural chlorine exists as chlorine 35, and the rest 24.23 % as chlorine 37.

The average atomic mass of chlorine is calculated as follows:

$$= \frac{(75.77 \times 35) + (24.23 \times 37)}{100}$$

$$= 35.45$$

Thus, the average atomic mass of chlorine is 35.45 a.m.u

Example

Calculate the average atomic mass of boron, using the following data:

Isotope	Relative masses	Percent abundance
---------	-----------------	-------------------

$^{10}_{5}\text{B}$

10.0134

19.70

$^{11}_{5}\text{B}$

11.0094

80.30

Solution:

The average atomic mass of boron is obtained by taking the sum of the products of the relative atomic mass and the percentage abundance of the first and the second isotopes of boron and dividing by 100.

$$\text{Average atomic mass of B} = \frac{10.0134 \times 19.70 + 11.0094 \times 80.30}{100}$$

$$= 10.813 \text{ a.m.u}$$

Exercise 1.6

Give appropriate answers for the following questions.

1. What are isotopes? Explain their symbolic designation using the isotopes of hydrogen.

Activity 1.16

2. The three isotopes of uranium are: $^{234}_{92}\text{U}$, $^{235}_{92}\text{U}$ and $^{238}_{92}\text{U}$. How many protons, neutrons and electrons are present in each isotope?

$^{234}_{92}\text{U}$, $^{235}_{92}\text{U}$ and $^{238}_{92}\text{U}$

$^{234}_{92}\text{U}$, $^{235}_{92}\text{U}$ and $^{238}_{92}\text{U}$

$^{234}_{92}\text{U}$. How many protons, neutrons and electrons are present in each isotope?

Answer

Isotopes	Protons	Electrons	Neutrons
----------	---------	-----------	----------

$^{234}_{92}\text{U}$

92

92

142

$^{235}_{92}\text{U}$

92

92

143

$^{238}_{92}\text{U}$

92

92

146

3. Indicate the number of fundamental particles present in an atom of $^{206}_{82}\text{Pb}$.

4. The metal thallium occurs naturally as 30% thallium-203 and 70% thallium-205. Calculate the atomic mass of thallium.

Critical Thinking

5. Why is the gravitational force in the nucleus so small?

6. Could a nucleus of more than one proton but no neutron exist? Explain.

1.4 ATOMIC MODELS

Atoms are too small to be observed directly. Hence, it is better to develop a tentative mental picture (model) of the atomic concept. The ideas about atoms have changed many times. But, at present we can reduce these ideas into **five** models: Dalton's Model, Thomson's Model, Rutherford's Model, Bohr's Model, and the Quantum Mechanical Model.

i) Dalton's Atomic Model

This model of atoms was accepted for 100 years without serious challenge. He thought of atoms as solid indestructible spheres. He called it the billiard ball model.

ii) Thomson Model

After learning that atoms contain electrons, Thomson proposed a new model of the atom, which is shown in Figure 1.8. It is sometimes called the plum-pudding model. In his model of an atom, Thomson proposed that electrons were embedded inside a positively charged sphere, just like plums in a pudding. Today we might call Thomson's model the chocolate chip ice-cream model.

iii) Rutherford Model

Why was Thomson's model of atom discarded and replaced by Rutherford's model?

By 1911, evidence was collected that allowed scientists to modify Thomson's model of an atom. Ernest Rutherford (1871–1937), a former student of Thomson, performed one of the classic experiments of scientific history. As indicated earlier in this unit, he bombarded different types of matter with high-energy, positively charged alpha particles. Based on the Thomson's model of an atom, Rutherford hypothesized that the alpha particles should go through very thin metal foils undeflected. However, after performing an experiment on gold leaf, he found that a small but significant number of the alpha particles were deflected through large angles by the gold atoms. These

results showed Rutherford that the Thomson model of the atom was not valid. A uniform positive charge with embedded negatively charged electrons would not interact with the alpha particles in such a manner. The Rutherford model of the atom has a small, dense, positively charged nucleus around which electrons whirl at high speeds and at relatively long distances from it.

He compared the structure of an atom with the solar system, saying that the nucleus corresponds to the sun, while the electrons correspond to the planets. This picture of the atom is also called the planetary atom. In other words, Rutherford gave us the nuclear model of the atom.

Even though Rutherford showed that the atom has a nucleus, he did not know how the electrons were arranged outside the nucleus. This was the major drawback of Rutherford's model of an atom.

A Nobel Prize winner, Niels Bohr, was known not only for his own theoretical work, but also as a **mentor to younger physicists** who themselves made important contributions to physical theory. As the director of the Institute for Theoretical Physics at the University of Copenhagen, Bohr gathered together some of the finest minds in the physics community, including Werner Heisenberg and George Gadow. During the 1920's, the Institute was the source of many important works in quantum mechanics and in theoretical physics in general.

iv) Bohr Model

Niels Bohr proposed an atomic model in which the electrons moved around the nucleus in circular paths called orbits. He assumed the electrons to be moving around the nucleus in a circular orbit as the planets move around the sun. Based on the Rutherford's atomic model, Bohr made the following modifications.

- The electrons in an atom can exist only in a restricted number of stable orbits with energy levels in which they neither absorb nor emit energy. These orbits designated by a number called the principal quantum number, n . The principal quantum number has the values of 1, 2, 3, . . . The energy levels are also called shells. They are represented by K, L, M, . . . etc. The K-shell is the first shell, the L-shell is the second shell, the M-shell is the third shell, etc.
- When an electron moves between orbits it absorbs or emits energy. When an electron jumps from lower to higher states it absorbs a fixed amount of energy. When an electron falls from a higher (excited) state to a lower (ground) state it emits a fixed amount of energy.
- The electrons move around the nucleus in energy levels.
- According to Bohr, for each element, the number of energy levels or shells is fixed. Also, each different energy level or shell of an atom can only accommodate a certain number of electrons. The maximum number of electrons that the main energy level can have is given by the general formula $2n^2$, where n is equal to the main energy level.

Therefore, the maximum number of electrons in each shell in the Bohr model is:

first energy level ($n = 1$) is $2(1)^2 = 2$ electrons

second energy level ($n = 2$) is $2(2)^2 = 8$ electrons

third energy level ($n = 3$) is $2(3)^2 = 18$ electrons

fourth energy level ($n = 4$) is $2(4)^2 = 32$ electrons

.Principal quantum number (n)----- 1 2 3 4

.Shell ----- K L M N

.Maximum possible number of electrons --2 8 18 32

- When electrons fill the various energy levels in an atom, they first occupy the shell with the lowest energy level. When the lowest energy level is filled, according to the $2n^2$ rule, then the electrons enter the next higher energy level.

The outermost shell of an atom cannot accommodate more than eight electrons, even if it has the capacity to hold more electrons. This is because having more than eight electrons in the outermost shell makes the atom unstable.

For example, let us write the electron configuration of magnesium that has 12 electrons. The first 2 electrons occupy the K-shell (the first energy level). The L-shell (the second energy level) is occupied with 8 electrons. The remaining 2 electrons enter the M-shell (the third energy level).

Exercise 1.7

Draw Bohr's diagrammatic representations for the following atoms:

- Nitrogen (atomic number = 7)
- Sulphur (atomic number = 16)
- Potassium (atomic number = 19)
- Calcium (atomic number = 20)

v) The Quantum Mechanical Model

During the 1920's, the discovery of the wave-like properties of electrons led to the wave-mechanical model or quantum-mechanical model of atoms. In this model, the electrons are

associated with definite energy levels, but their locations cannot be pinpointed. Instead, they are described in terms of the probability of being found in certain regions of space about the nucleus. These regions of space are called orbitals.

An orbital is a particularly shaped volume of space where the probability of finding an electron is at a maximum.

According to this model, the electrons in an atom are in a series of energy levels or shells. Each energy level or shell is described by a number called the principal quantum number, n , which is related to the size of the energy level. The larger the energy value of n , the farther the electrons are found from the nucleus.

The quantum mechanical model introduces the concept of sublevel for each main energy level. For atoms of the known elements, there are four types of sublevels designated by the letters s, p, d, and f. The number of sublevels within each energy level or shell is equal to the numerical value of n .

The principal quantum number specified with shell, sublevels and maximum number of electrons.

.Principal quantum number (n)	.Shell	.Number and type of sublevels	.Maximum number of electrons in each shell ($2n^2$)
1	K	1 (s)	2
2	L	2 (s, p)	8
3	M	3 (s, p, d)	18
4	N	4 (s, p, d, f)	32
5	O	5 (s, p, d, f, g)	50
6	P	6 (s, p, d, f, g, h)	72

Any orbital can accommodate a maximum of two electrons. Accordingly,

s-sublevel has only 1 orbital and can hold a maximum of 2 electrons,

p-sublevel has 3 orbitals and can hold a maximum of 6 electrons,

d-sublevel has 5 orbitals and can hold a maximum of 10 electrons, and

f-sublevel has 7 orbitals and can hold a maximum of 14 electrons.

Electron Configurations

The arrangement of electrons in an atom is known as the electron configuration of the atom.

Because atoms of different elements have different numbers of electrons, a distinct electronic configuration exists for the atoms of each element. Like all systems in nature, electrons in atoms tend to assume arrangements that have the lowest possible energies. The lowest energy arrangement of the electrons in an atom is called the ground state electron configuration. A few simple rules, combined with the quantum number relationships discussed below, allow us to determine these groundstate electron configurations.

The quantum mechanical model is designated by the following notation: a coefficient which shows the main energy level, a letter that denotes the sublevel that an electron occupies, and a superscript that shows the number of electrons in that particular sublevel.

Designation

$1s^2$

Number of electrons =2

Sublevel =s

Main energy level =1

For example, the electron configuration of lithium (${}^3\text{Li}$) is: $1s^2 2s^1$.

This indicates that there are 2 electrons in the first s-sublevel and 1 electron in the 2nd s-sublevel.

The configuration for sodium (${}_{11}^{23}\text{Na}$) atom is: $1s^2 2s^2 2p^6 3s^1$. This indicates that there are 2 electrons in the first s-sublevel, 2 electrons in the second s-sublevel, 6 electrons in the second p-sublevel, and 1 electron in the third s-sublevel.

Form a group and perform the following task in groups and present your findings to the class.

The following ground state electron configurations are incorrect. Identify in each case the mistakes that have been made and write the correct electron configurations.

i) $1s^2 2s^2 2p^4 3s^2$

ii) $1s^2 2s^1 2p^6$

iii) $1s^2 2s^2 2p^6 3p^3 3s^2$

iv) $1s^2 2s^2 2p^6 3s^2 3p^3 4s^2 3d^4$

The following table lists the first 14 elements in the periodic table and shows the electron configuration of each.

The ground state electron configuration of the first 10 elements.

Element	Electron Configuration
---------	------------------------

Hydrogen	$1s^1$
----------	--------

Helium	$1s^2$
--------	--------

Lithium	$1s^2 2s^1$
---------	-------------

Beryllium	$1s^2 2s^2$
-----------	-------------

Boron	$1s^2 2s^2 2p^1$
-------	------------------

Carbon	$1s^2 2s^2 2p^2$
--------	------------------

Nitrogen	$1s^2 2s^2 2p^3$
----------	------------------

Oxygen	$1s^2 2s^2 2p^4$
--------	------------------

Fluorine	$1s^2 2s^2 2p^5$
----------	------------------

Neon	$1s^2 2s^2 2p^6$
------	------------------

Exercise 1.8

Write the ground state electron configuration for the following elements.

a Phosphorus (atomic number = 15)

b Sulphur (atomic number = 16)

c Chlorine (atomic number = 17)

d Argon (atomic number = 18)

Rules Governing Electron Configuration

Aufbau principle{building up}—an electron occupies the lowest energy orbital available before entering a higher energy orbital.

The atomic orbitals are filled in order of increasing energy. The orbital with the lowest energy is the 1s orbital. The 2s orbital is the next higher in energy, then the 2p orbitals, and so on.

Beginning with the third main energy level, $n = 3$, the energies of the sublevels in the different main energy levels begin to overlap.

The diagonal rule is a guide to the order of filling energy sublevels. It is particularly helpful for atoms with atomic numbers higher than 18. This is because for atoms with higher atomic numbers their sublevels are not regularly filled.

If you list the orbitals following the direction of arrows, you can obtain the electron configuration of most atoms. Therefore, the order of filling the sublevels is given as:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6...$

The outermost occupied energy level of an atom is called the valence shell and the electrons that enter into these energy levels are referred to as **valence electrons**.

Example 1

What is the electron configuration of magnesium, $^{24}_{12}\text{Mg}$?

Solution:

24

12 Mg = 1s² 2s² 2p⁶ 3s² = sublevel

2 8 2 = main energy level

Noble gas notation = [Ne] 3s²

Example 2

sublevel for Cu-z=29 is 1s² 2s² 2p⁶ 3s² 3p⁶ 4s¹ 3d¹⁰

Noble gas notation = [Ar] 4s¹ 3d¹⁰

QUESTIONS

In a certain atom, the last electron has a 3d² electron configuration. On the basis of this:

a write the electron configuration of the element using a noble gas as a core.

b write the electron configuration using the main energy levels.

c what is the atomic number of the atom?

d what are the valence shell electrons of the atom?

e how many protons does the atom have?

9. What is the relative atomic mass of an element whose isotopic composition is 90% of ^{20}X and 10% of ^{22}X ?

1. $^{\text{A}}_{\text{Z}}\text{X}$ What is A? _____

2. Let x – Y be any element with atomic mass of 32 & its atomic number is 15 what is its.

a. p⁺ _____

b. e⁻ _____

c. n = _____

3. What is the measurement of very small particles? _____

4. Which one of the following atoms are isotopes?

$^{13}_5\text{X}$

$^{10}_6\text{X}$

$^{10}_5\text{X}$

$^{13}_9\text{X}$

5. An atom has a net charge of +3. It has 13 electrons & 14 neutrons find the mass number _____.

6. What are the two basic differences of Dalton & modern atomic theory?

_____, _____

7. An element X has isotopes $^{22}_{16}\text{X}$ & $^{26}_{16}\text{X}$ they exist in nature in a 1:1 ratio what is the atomic mass of X? follow the correct step

8. Calculate the relative atomic mass of $^{12}_6\text{X}$ which constitute 23.62 % & $^{14}_6\text{X}$ constitute 76.38% what is the relative atomic mass of X?

