CHAPTER1

Elementary notions of the structure of matter

I. ATOM

All the materials available either in solid, liquid or gaseous form is made up of atoms

Atom is the small est indivisible particle of the matter. Atom is made of electron, proton and neutrons. Atoms are the basic building blocksof matter that make up everyday objects. Main parts of an atom: Nucleus-99.9% of the atom's mass Electron cloud or energy rings

-Atoms are made of subatomic particles:protons, neutrons, & electrons

a- Electron: is a particle charged with negative electricity qe= -1.6 10-19 C of very low mass m= 9.1085 10-31kg

b- Proton: is a particle charged with positive electricity qp= +1.6 10-19C of very large charged with negative electricity qe= -1.6 10-19 C of very low mass m= 9.1085 10-31kg.

b- Proton: is a particle charged with positive electricity qp= +1.6 10-19C of very large mass compared to that of the electron mp =1800me

c- neutron: is an electrically neutral particle, its mass is close to that of the proton (mn = mp).

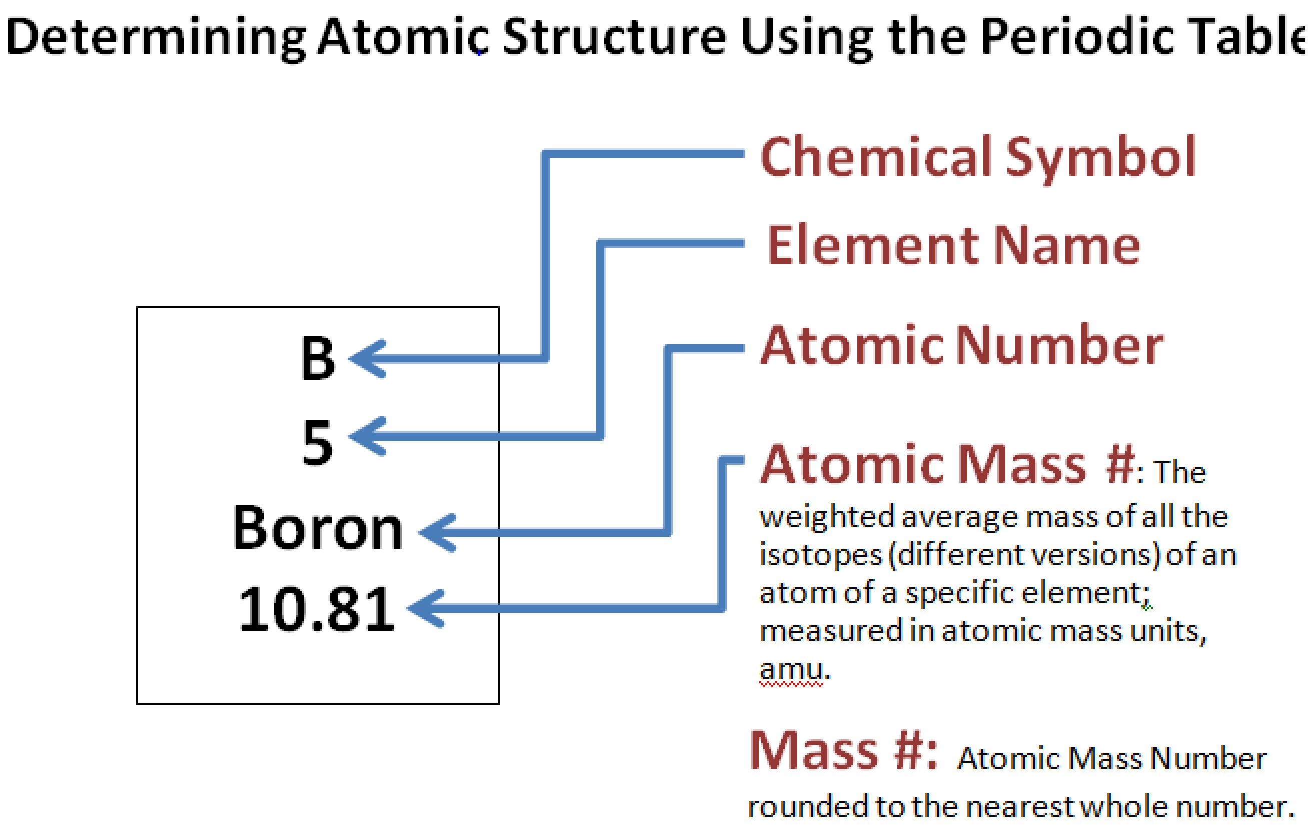
I.1 ATOMIC NUMBER AND MASSNUMBER

I.1.1 ATOMIC NUMBER

Protons are present in the nucleus of an atom. It is the number of protons of an atom, which determines its atomic number. 'Z' denotes it. All atoms of an element have the same atomic number, Z. In fact, the number of protons they possess defines elements. For hydrogen, Z = 1, because in hydrogen atom, only one proton is present in the nucleus. Similarly, for carbon, Z = 6. Therefore, the atomic numbeis defined as the total number of protons present in the nucleus of an atom.

I.1.2 MASS NUMBER

After studying the properties of the sub- atomic particles of an atom, we can conclude that mass of an atom is practically due to protons and neutrons alone. These are present in the nucleus of an atom. Hence protons and neutrons are also called nucleons. Therefore, the mass of an atom resides in its nucleus. For example, mass of



I.2 ISOTOPES

In nature, a number of atoms of some elements have been identified, which have the same atomic number but different mass numbers. For example, take the case of hydrogen atom, it has three atomic species, namely protium (^{1}H), deuterium (^{2}H or D) and tritium (^{3}H or T). The atomic number of each one is 1, but the mass number is 1, 2 and 3, respectively. Other such examples are i carbon, ^{12}C and ^{14}C, (ii) chlorine, ^{35}Cl and ^{37}Cl, etc.

I.3 ISOBARS

Let us consider two elements — calcium, atomic number 20, and argon, atomic number 18. The number of electrons in these atoms is different, but the mass number of both these elements is 40. That is, the total number of nucleons is the same in the atoms of this pair of elements. Atoms of different elements with different atomic numbers, which have the same mass number, are known as isobars.

II. Elements are the fundamental substances of chemistry and are composed of atoms.

~115 different elements have been identified, eg. hydrogen, phosphorus, oxygen, nitrogen, sulfur, helium, carbon, calcium, iron, sodium, chlorine.

CHAPITRE I
Elementary notions of the structure of matter

Elements cannot be decomposed or converted to simpler substances or other elements by any common form of energy, eg. heat, light, electricity, sound, magnetism.
Only neutron bombardment can induce fission of some nuclei (ie ^{236}U) causing decay to other elements and release of energy.
Atoms, once thought to be the ultimate indivisible particles that make up all matter, are among the fundamental particles of the science of chemistry
John Dalton (1776-1844) (UK) proposed that all matter was composed of atoms-he was correct!

II.1 Molecules are groups of two or more atoms held together by the forces of chemical bonds. H_2 and O_2 are molecules but not compounds.

An **ion** is an atom or group of atoms that carries an electrical charge.

An **anion** is a negatively charged ion.
A **cation** is a positively charged ion.

A **compound** is a pure substance formed by the chemical combination of two or more different elements in a specific ratio.

- Dalton used the percentages of elements in compounds and the chemical formulas to deduce the **relative** masses of atoms
- Unit is the **amu**.
  - **atomic mass unit**
  - **1 amu = 1.66 x 10^{-24}g**
- We define the masses of atoms in terms of atomic mass units
  - **1 Carbon atom = 12.01 amu,**
  - **1 Oxygen atom = 16.00 amu**
  - **1 O_2 molecule = 2(16.00 amu) = 32.00 amu**
- Atomic masses allow us to convert weights into numbers of atoms

If our sample of carbon weighs 3.00 x 10^{20} amu we will have 2.50 x 10^{19} atoms of carbon

$$3.00 \times 10^{20} \text{ amu} \times \frac{1 \text{ C atom}}{12.01 \text{ amu}} = 2.50 \times 10^{19} \text{ C atoms}$$

Since our equation tells us that 1 C atom reacts with 1 O_2 molecule, if I have 2.50 x 10^{19} C atoms, I will need 2.50 x 10^{19} molecules of O_2
Example #1
Calculate the Mass (in amu) of 75 atoms of Al

- Determine the mass of 1 Al atom

1 atom of Al = 26.98 amu

- Use the relationship as a conversion factor

$$75 \text{ Al atoms x } \frac{26.98 \text{ amu}}{1 \text{ Al atom}} = 2024 \text{ amu}$$

Chemical Packages – Moles

- We use a package for atoms and molecules called a **mole**
- A mole is the number of particles equal to the number of Carbon atoms in 12 g of C-12
- One mole = 6.022×10^{23} units
- The number of particles in 1 mole is called **Avogadro's Number**
- 1 mole of C atoms weighs 12.01 g and has 6.02×10^{23} atoms

Lead bar
207.2 g

Silver bars
107.9 g

Pile of copper
63.55 g

Figure 8.1: All these samples of pure elements contain the same number (a mole) of atoms: 6.022×10^{23} atoms

Example #2
Compute the number of moles
and number of atoms in 10.0 g of Al

¬ Use the Periodic Table to determine the mass of 1 mole of Al
1 mole Al = 26.98 g
Use this as a conversion factor for grams-to-moles

$$10.0 \text{ g Al x } \frac{1 \text{ mol Al}}{26.98 \text{ g}} = 0.371 \text{ mol Al}$$

Example #2
Compute the number of moles

and number of atoms in 10.0 g of Al

® Use Avogadro's Number to determine the number of atoms in 1 mole

1 mole Al = 6.02 x 10^{23} atoms

¯ Use this as a conversion factor for moles-to-atoms

$$0.371\,\text{mol Al x}\,\frac{6.02\,\text{x}\,10^{23}\ \text{atoms}}{1\,\text{mol Al}} = 2.23\,\text{x}\,10^{23}\ \text{Al atoms}$$

Example #3

¬ Use Avogadro's Number to determine the number of atoms in 1 mole

1 mole Al = 6.02 x 10^{23} atoms

Use this as a conversion factor for atoms-to-moles

Compute the number of moles

and mass of 2.23 x 10^{23} atoms of Al

$$2.23\,\text{x}\,10^{23}\ \text{Al atoms}\ \ \text{x}\,\frac{1\,\text{mol Al}}{6.02\,\text{x}\,10^{23}\ \text{atoms}} = 0.370\,\text{mol Al}$$

Example #3

Compute the number of moles

and mass of 2.23 x 10^{23} atoms of Al

® Use the Periodic Table to determine the mass of 1 mole of Al

1 mole Al = 26.98 g

¯ Use this as a conversion factor for moles-to-grams

$$0.370\,\text{mol Al x}\,\frac{26.98\,\text{g}}{1\,\text{mol Al}} = 9.99\,\text{g Al}$$

II.2 Molar Mass

- The **molar mass** is the mass in grams of one mole of a compound
- The relative weights of molecules can be calculated from atomic masses

water = H_2O = 2(1.008 amu) + 16.00 amu

= 18.02 amu

- 1 mole of H_2O will weigh 18.02 g, therefore the molar mass of H_2O is 18.02 g
- 1 mole of H_2O will contain 16.00 g of oxygen and 2.02 g of hydrogen

Percent Composition

- Percentage of each element in a compound

– By mass
- Can be determined from
Ê the formula of the compound or
Ë the experimental mass analysis of the compound
- The percentages may not always total to 100% due to rounding

$$\text{Percentage} = \frac{\text{part}}{\text{whole}} \times 100\%$$

Example #4
Determine the Percent Composition
from the Formula C_2H_5OH
¬ Determine the mass of each element in 1 mole of the compound
2 moles C = 2(12.01 g) = 24.02 g
6 moles H = 6(1.008 g) = 6.048 g
1 mol O = 1(16.00 g) = 16.00 g
Determine the molar mass of the compound by adding the masses of the elements
1 mole C_2H_5OH = 46.07 g
Example #4
Determine the Percent Composition
from the Formula C_2H_5OH

® Divide the mass of each element by the molar mass of the compound and multiply by 100%

$$\frac{24.02\text{g}}{46.07\text{g}} \times 100\% = 52.14\%\text{C}$$

$$\frac{6.048\text{g}}{46.07\text{g}} \times 100\% = 13.13\%\text{H}$$

$$\frac{16.00\text{g}}{46.07\text{g}} \times 100\% = 34.73\%\text{O}$$

II.3 Empirical Formulas

- The simplest, whole-number ratio of atoms in a molecule is called the **Empirical Formula**
  - can be determined from percent composition or combining masses
- The Molecular Formula is a multiple of the Empirical Formula

$$\%\,A \xrightarrow{100g} \text{mass A (g)} \xrightarrow{MM_A} \text{moles A}$$
$$\%\,B \xrightarrow{100g} \text{mass B (g)} \xrightarrow{MM_B} \text{moles B}$$
$$\longrightarrow \frac{\text{moles A}}{\text{moles B}}$$

Example #5
Determine the Empirical Formula of
Benzopyrene, $C_{20}H_{12}$

- Find the greatest common factor (GCF) of the subscripts

factors of 20 = (10 x 2), (5 x 4)
factors of 12 = (6 x 2), (4 x 3)
GCF = 4
Divide each subscript by the GCF to get the empirical formula
$C_{20}H_{12} = (C_5H_3)_4$
Empirical Formula = C_5H_3
Example #6
Determine the Empirical Formula of
Acetic Anhydride if its Percent Composition is
47% Carbon, 47% Oxygen and 6.0% Hydrogen

- Convert the percentages to grams by assuming you have 100 g of the compound
  - Step can be skipped if given masses

$$100g \times \frac{47gC}{100g} = 47gC$$

$$100g \times \frac{47gO}{100g} = 47gO \qquad 100g \times \frac{6.0gH}{100g} = 6.0gH$$

Example #6
Determine the Empirical Formula of
Acetic Anhydride if its Percent Composition is
47% Carbon, 47% Oxygen and 6.0% Hydrogen
Convert the grams to moles

$$47g\,C \times \frac{1\,mol\,C}{12.01g} = 3.9\,mol\,C \qquad 6.0\,g\,H \times \frac{1\,mol\,H}{1.008g} = 6.0\,mol\,H$$

$$47\,g\,O \times \frac{1\,mol\,O}{16.00g} = 2.9\,mol\,O$$

Example #6

Determine the Empirical Formula of
Acetic Anhydride if its Percent Composition is
47% Carbon, 47% Oxygen and 6.0% Hydrogen

® Divide each by the smallest number of moles

$3.9\,\text{mol C} \div 2.9 = 1.3$ $6.0\,\text{mol H} \div 2.9 = 2$

$2.9\,\text{mol O} \div 2.9 = 1$

Molecular Formulas

- The molecular formula is a multiple of the empirical formula
- To determine the molecular formula you need to know the empirical formula and the molar mass of the compound

Example #7
Determine the Molecular Formula of Benzopyrene
if it has a molar mass of 252 g and an
empirical formula of C_5H_3

¬ Determine the empirical formula
 - May need to calculate it as previous

C_5H_3

Determine the molar mass of the empirical formula

5 C = 60.05 g, 3 H = 3.024 g
C_5H_3 = 63.07 g

Example #7
Determine the Molecular Formula of Benzopyrene
if it has a molar mass of 252 g and an
empirical formula of C_5H_3

® Divide the given molar mass of the compound by the molar mass of the empirical formula
 – Round to the nearest whole number

$$\frac{252g}{63.07g} = 4$$

Example #7
Determine the Molecular Formula of Benzopyrene
if it has a molar mass of 252 g and an
empirical formula of C_5H_3

¯ Multiply the empirical formula by the calculated factor to give the molecular formula

$(C_5H_3)_4 = C_{20}H_{12}$

III.Binding Energy

The equivalence of mass and energy becomes apparent when we study the binding energy of systems like atoms and nuclei that are formed from individual particles.The potential energy associated with the force keeping the system together is called **the** binding energy E_B

The binding energy is *the difference between the rest energy of the individual particles and the rest energy of the combined bound system*.The binding energy is the work required to pull the particles out of the bound system intoseparate, free particles at rest

Conservation of energy

$E = mc^2$

In the special theory of relativity, Einstein showed that energyand mass are equivalent

If a body has a mass *m*, then it contains an amount of energy:

$E = mc^2$

probably the most famous equation in physics. This means that *if* the mass of a nucleus is *less* than the mass ofits constituents, then those constituents are in a *lower energy state* when they are bound together inside the nucleus.This difference in mass, expressed as energy (normally MeV), isthe *binding energy* of those constituents inside the nucleus. For example consider the carbon nucleus, ^{12}C shown in the previous table. Let's calculate its binding energy.

It contains 6 protons and 6 neutrons so $A = 12$.

From the table the mass of a carbon *atom* is 12u (by definitionof the *atomic mass unit*, u).This is the mass of the carbon atom, i.e. the mass of the nucleus *plus the six electrons in the neutral carbon atom.* From the table, the mass of the ^{12}C *nucleus* is therefore:

$$12 - 6 \times 5.486 \times 10^{-4} = 11.9967 \text{ u}$$

$$6 \times 1.007276 + 6 \times 1.008665 - 11.9967 = 0.0989\text{u} \times 931.494\text{MeV}/c^2$$

Mass of aproton Mass of aneutron Mass of ^{12}C nucleus

This is 92.2 MeV in total Normally we express this as the binding energy per nucleon.In this case it is 7.68 MeV per nucleon.

It is found that the mass of a nucleus is always *less than* the sum of the masses of its constituent neutrons and protons (nucleons).

What is the reason for this? Well Einstein showed that mass and energy are equivalent. The lower mass shows that the nucleons in the nucleus are in a *lower energy state* than if they were all separate, isolated particles. It means that there must be some force between the nucleons that binds them together in the nucleus. This is the *strong force*.

This decrease in mass (known as the mass decrement) gives the *binding energy* of the particular nucleus in terms of the equivalent *mass*. Working in terms of the actual binding energy, we calculate as follows.Say for example if we have a nucleus with *Z* protons and *N* neutrons and mass M_A, where $A = Z + N$ then its binding energy in MeV is given by:

$$E_b(MeV) = (Zm_p + Nm_n - M_A) \times 931.494 \text{ MeV/u}$$