CHEM 2101

Chapter - 2

Lecture Notes

Atoms, Molecules and Ions

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Chapter-2

Atoms, Molecules and Ions

2.5 The Modern View of Atomic Structure: An Introduction.
2.6 Molecules and Ions
2.7 An Introduction to Periodic Table
2.8 Naming Simple Compounds

CHRONOLOGY OF ATOMIC STRUCTURE

1.	Dalton (1808)	:	Discovery of atom
2.	Julius Plucker (1859)	:	First discovery of cathod rays.
3.	Goldstein (1886)	:	Discovered anode rays and protons.
4.	Sir. J. J. Thomson (1897)	:	Discovered electron and determined charge / mass (e/m) ratio for electron.
5.	Rutherford (1891)	:	Discovered nucleus and proposed atomic model.
6.	Max Planck (1901)	:	Proposed quantum theory of radiation.
7.	Robert Millikan (1909)	:	Determined charge of an electron.
8.	H.G.J. Mosely (1913)	:	Discovered atomic number.
9.	Niels Bohr (1913)	:	Proposed new model of atom.
10.	Clark Maxwell (1921)	:	Electromagnetic wave theory.
11.	De-Broglie (1923)	:	Established wave nature of particle.
12.	Pauli (1927)	:	Discovery of neutrino.
13.	Werner Heisenberg (1927)	:	Uncertainty Principle.
14.	James Chadwick (1932)	:	Discovery of neutron.
15.	Anderson (1932)	:	Discovery of positron.
16.	Fermi (1934)	:	Discovered antineutrino.
17.	Hideki Yukawa (1935)	:	Discovered mesons.
18.	Segre (1955)	:	Discovered antiproton.
19.	Cork and Association (1956)	:	Discovered antineutron.

PROGRESS OF ATOMIC MODELS

- **In 1803, John Dalton, proposed his atomic theory. He suggested that atoms were indivisible solid spheres.**

- **J.J. Thomson proposed that an atom was a solid sphere of positively charged material and negatively charged particles, electrons were embedded in it like the seeds in a guava fruit. But later this concept was proved wrong.**

- **Rutherford suggested the planetary model, but this model was rejected.**

- **In 1913, Neils Bohr proposed that electrons revolve around the nucleus in a definite orbit with a particular energy. Based on the facts obtained from spectra of hydrogen atom, he introduced the concept energy levels of atom.**

- **In 1916 Sommerfield Bohr's model by introducing elliptical orbits for electron path. He defined sub energy levels for every major energy level predicted by Bohr.**

- **The concept of quantum numbers was introduced to distinguish the orbital on the basis of their size, shape and orientation in space by using principal, azimuthal, magnetic and spin quantum numbers.**

- **From the study of quantum numbers, various rules are put forward for filling of electrons in various orbitals by following**

    - **Aufbau principle.**
    - **Pauli exclusion principle and**
    - **Hunds rule of maximum multiplicity.**

- **In 1921 Burry and Bohr gave a scheme for the arrangement of electrons in an atom. Further the nature of electron(s) is studies.**

IMPORTANT TERMS

Atom

The smallest chemically indivisible particle of a matter. It is regarded to be structureless hard, impenetrable particles.

Subatomic particles

Particles which are smaller than the atom. Example: Protons, Neutrons and Electrons.

Proton

A positively charged particle present inside the nucleus.

Electron

A negatively charged particle present around the nucleus of atom.

Neutron

A neutral particle, which is present in the nucleus, having the mass equal to that of a proton.

Mass Number (A)

The total number of nucleons is termed (protons and neutrons) as mass number (A) of the atom.

Atomic Number (Z)

The number of protons in an atom is equal to the number of electrons in an atom and it is known as atomic number (Z).

Orbit

Orbit is a circular path around the nucleus in which electron revolves.

Orbital

Orbital is the region where there is maximum probability of finding electron.

Periodic table

Periodic table is defined as the arrangement of various elements according to their properties in a tabular form.

Triad

A group of three elements having similar properties.

Dobereiners law of triads

When elements are arranged in the order of increasing atomic mass in a triad, the atomic mass of the middle element was found to be approximately equal to the arithmetic mean of the other two elements.

Newlands law of octaves

If the elements were arranged in order of their increasing atomic weights, the eighth element starting from a given one, possessed properties similar to the first like the eighth note in an octave of music.

Mendeleev's periodic law

The properties of the elements are the periodic function of their atomic weights.

Modern periodic law

The physical and chemical properties of the elements are periodic function of their atomic numbers.

Period

The horizontal row in the periodic table is known as period.

Groups

The vertical column in the periodic table is known as groups.

Periodicity

Repetition of properties of elements at regular intervals is called periodicity in properties.

Atomic radius

Atomic radius is defined as the distance between the center of the nucleus and the outermost shell of electrons in an atom.

Ionic radius

Ionic radius is defined as the distance between the center of the nucleus and the outermost shell of electrons in an ion.

Isoelectronic

Atoms or ions which contain the same number of electrons.

Ionization energy

Ionization energy is defined as the energy required to remove an electron from an atom (or electron gain enthalpy).

Electron affinity

Electron affinity is defined as the amount of energy released when an isolated gaseous atom accepts an electron to form a monovalent gaseous anion.

Electronegativity

Electronegativity is defined as the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.

Chemical bond

Chemical bond is the existence of a strong force of binding between two or many atoms to form a stable compound.

Different types of bond

(i) Ionic (or) electrovalent bond
(ii) Covalent bond
(iii) Co-ordinate covalent (or) dative bond

Ion

An ion is an atom or molecule, which has a positive or negative charge.

Ionic bond

The binding forces existing as a result of electrostatic attraction between the positive and negative ions.

Cation

A positive ion is called cation. Example: Na^{+}, Mg^{2+} and NH^{4+}.

Anion

An ion with a negative charge is called anion. Example: Cl^{-}, NO_3^{-}

Ionic compound

Ionic compounds are compounds formed from cations (positively charged ions) and anions (negatively charged ions).

Ionic solid

A solid consisting of oppositely charged ions is called ionic solid or a salt.

Covalent bond

Covalent bond is formed between two atoms by mutual sharing of pair of electrons to attain stable outer octet or electrons for each atom.

Co-ordinate covalent bond

Co-ordinate co-valent bond is a bond which is formed as a result of electron pair sharing with the pair of electrons being donated by only one atom of the bond.

Molecules

Molecules are a group of atoms of the same/different elements joined by covalent bonds.

Homonuclear diatomic molecule

Molecules having two identical atoms.

(e.g) H_2, O_2, Cl_2, N_2

Heteronuclear diatomic molecule

Molecules containing two different atoms.
(e.g) CO, HCl, NO, HBr

Homonuclear polyatomic molecule

Molecules containing many identical atoms.
(e.g) P_4, S_8

Heteronuclear polyatomic molecule

Molecules containing more than two atoms of different kinds.
(e.g)CH_4, NH_3, CH_3COOH, SO_2.

Molecular formula

It gives the exact number of atoms of all the elements present in a compound.

Empirical formula

It is a simplest formula of a compound which gives the ratio of all atoms present in it.

Structural formula

The structural formula of a compound is a graphical representation of the molecular structure showing how the atoms are arranged.

Isotopes

Atoms of the same element having same atomic number but different mass number are called isotopes.

Example: Hydrogen is having three kinds of isotopes – Protium, Deuterium and Tritium.

$${}_1H^1, {}_1H^2, {}_1H^3$$

Allotropes

Allotropes are distinct form of an element.
Example: Diamond and graphite are allotropes of carbon.

Periodic Table

Alkali metals, Alkaline earth metals, Transition metals, Halogens, Noble gases

1 (1A)	2 (2A)	3	4	5	6	7	8	9	10	11	12	13 (3A)	14 (4A)	15 (5A)	16 (6A)	17 (7A)	18 (8A)
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac†	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Uun	111 Uuu	112 Uub						

*Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
†Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

2.5 The modern view of Atomic Structure

Atomic theory has four assumptions:

1. Atoms make up all matter.

2. The atoms of one element are different from the atoms of another element.

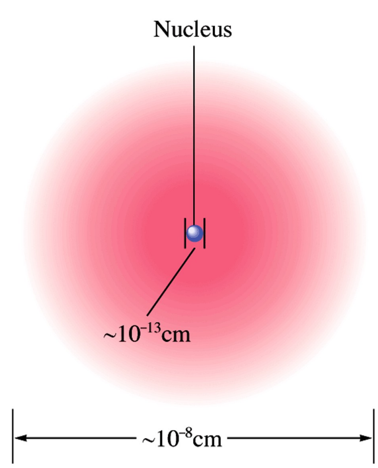
3. Atoms combine in definite ratios to make compounds.

4. Combinations of atoms in compounds can change only when a chemical reaction happens. This means reactions alter atom combinations, but the identity of the atoms themselves remain the same.

Atom

Atom is the smallest particle of a chemical element that can take part in a chemical reaction. Atoms are made up of protons, neutrons and electrons.

The simplest view of the atom is that it consists of a tiny nucleus (with a diameter of about 10^{-13} cm) and electrons that move about the nucleus at an average distance of about (radius) 10^{-8} cm from it.



Subatomic particles

- **The atom is made up of three main subatomic particles, the electrons, protons and neutrons.**
- **The protons and neutrons form an extremely small nucleus, which is surrounded by a cloud of electrons.**
- **The Protons and neutrons are of similar mass and are approximately 1840 times the mass of electrons.**

The three simple subatomic particles are the electron, proton and neutron. The properties of these three are summarized here.

Properties of Subatomic particles

The Mass and Charge of the Electron, Proton and Neutron

Particle	**Mass**	**Charge**
Electron	9.11 x 10^{-31} kg	-1
Proton	1.67 x10^{-27} kg	+1
Neutron	1.67 x 10^{-27} kg	None

***The magnitude of the charge of the electron and the proton is 1.60 x 10^{-19}C.**

Exercises:

Which of the following particles has the smallest mass?

a. **Proton**
b. **Neutron**
c. **Electron**

Answere:

Sorry if you picked "a" . The proton has a mass of 1.0073 amu. This is roughly the same mass as a neutron and almost two thousand times the mass of the electron. The neutron has a mass of 1.0087 amu and the electron 5.486 x 10^{-4} amu. The proton has a mass of 1.6726 x 10^{-24} g, the neutron has a mass of of 1.6749 x 10^{-24} and the electron a mass of 0.00091094 x 10^{-24} g.

Sorry if you picked "b". The neutron has a mass close to that for a proton. A neutron is almost 2000 times more massive than an electron.

Good if you picked "c". You are right. The electron is the least massive of the three particles. The electron mass is 0.00091094 x 10^{-24} g

Which of the following has a negative charge?

- **a. Proton**
- **b. Neutron**
- **c. Electron**
- **d. Nucleus**

Sorry if you picked "a" . The proton has a relative plus one charge. The proton has a charge on an absolute scale of 1.6021773 x 10^{-19} Coulombs. This is exactly the same magnitude but opposite sign as the charge on an electron.

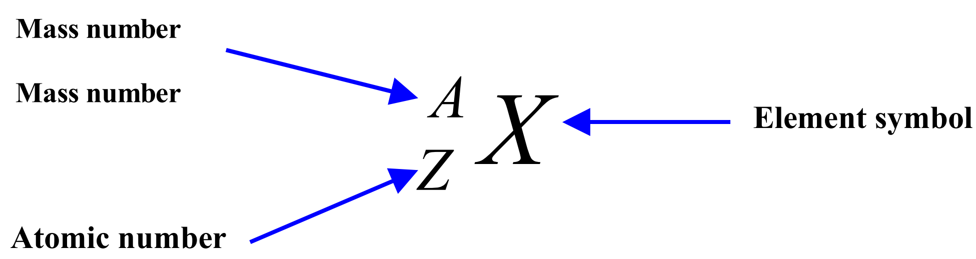
Sorry if you picked "b". The neutron has a zero charge.

Good if you picked "c". You are right. The electron has a relative minus one charge. The electron has an absolute charge of -1.6021773 x 10^{-19} Coulombs. This is exactly the same magnitude but opposite sign as the charge on a proton.

Sorry if you picked "d". All nuclei have a plus charge. The magnitude will be a multiple of the charge on a proton. The atomic number determines the multiple. Hydrogen has +1 charge on the nucleus while sodium has a +11 nucleus.

Atomic number, Z

The number of protons in an atom is equal to the number of electrons in an atom and it is known as **atomic number (Z). The atomic number Z is written as subscript.**



- **The identity of an element is controlled by the number of protons in the nucleus.**
- **The number of protons equals the atomic number.**
- **Every element has its own unique atomic number.**
- **The atomic number represented by the symbol, Z.**
- **The periodic table is arranged in sequence of increasing atomic number or proton count.**
- **Hydrogen has one proton and has atomic number one, Z=1.**
- **The atomic number increases by one unit for every additional proton.**
- **Helium has two protons and is therefore atomic number Z= 2.**

Example:

What is the atomic number for nitrogen, N?

Answer:

Nitrogen is in the seventh position in the periodic table. This means nitrogen atoms have 7 protons in the nucleus, and they have an atomic number of 7.

Mass number, A:

Mass Number, A (Nucleon Number), is the total number of protons and neutrons in the nucleus of an atom of the element.

The mass number A (the total number of protons and neutrons) is written as a superscript.

Mass number

Mass number

$${}^{A}_{Z}X$$

Element symbol

Atomic number

- **The count of neutrons and protons equals the mass number for an atom.**
- **This has its origin in the fact that the "massive" particles in the atom are protons and neutrons.**
- **The symbol for mass number is A.**
- **The entries in the periodic table do not show the mass numbers for atoms.**
- **The periodic table is useless when we want to figure mass numbers.**
- **The mass number for an atom is commonly indicated with a number after the element name.**

The mass number and atomic number are related as shown below.

Number of neutrons = A - Z

Number of neutrons = mass number - atomic number

Example:

How many neutrons and protons are in an atom of sodium-23?

Answer:

Sodium, Na, has atomic number 11. This means sodium atoms have 11 protons in the nucleus. The atom has a mass number equal to the number of protons and neutrons, so there must be 12 neutrons in the nucleus.

Number of neutrons = A - Z

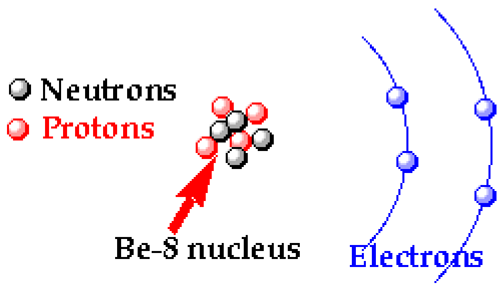
Number of neutrons = 23 - 11 = 12

Atomic number and electron count:

Atoms are neutral particles. The numbers of positively charged protons are canceled out by an equal number of negatively charged electrons.

Hydrogen has one proton and one electron.

The illustration shows a Beryllium-8 atom, Be-8, with four electrons and four protons.



Example:

1. How many electrons and protons are in a neutral sodium atom?

Answer:

Sodium, Na, has atomic number 11. This means sodium atoms have 11 protons in the nucleus. The neutral atom has to have equal numbers of protons and electrons, so there must be 11 electrons.

2. How many neutrons are in an atom of sulfur, S, with mass number 33?

Answer:

The atomic number for sulfur is 16.

The number of neutrons = A - Z
= 33 - 16 = 17

3. An atom contains 24 neutrons and 25 protons, what is the mass number of the atom?

Answer:

Mass number = A = number protons + number of neutrons = 24 + 25 = 49

4. An atom with a mass number of 39 contains 20 neutrons. What is the atomic number and identity of the element?

<u>Answer:</u>

The atomic number is Z = 39- 20 =19 .

The identity is potassium because K is element 19.

Isotopes

Atoms of the same element having same atomic number but different mass number are called isotopes.

- **Isotopes of an element are atoms that have the same number of protons, but different numbers of neutrons.**

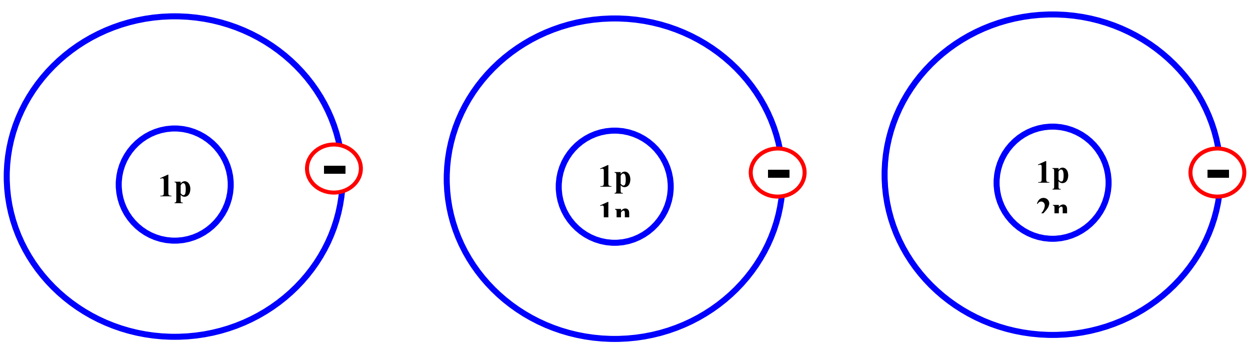
- **Isotopes of the same element have similar chemical properties.**

Greek "iso" means same and "tops" means place. This fits the idea that isotopes are in the same place in the periodic table, but have different masses. Periodic table entries provide the information shown here. The periodic table does not indicate isotope information.

Example-1:

There are three isotopes for hydrogen with mass number 1, 2, and 3, each possessing an atomic number of one.

1. **Protium or Hydrogen**
2. **Deuterium or heavy hydrogen**
3. **Tritium**

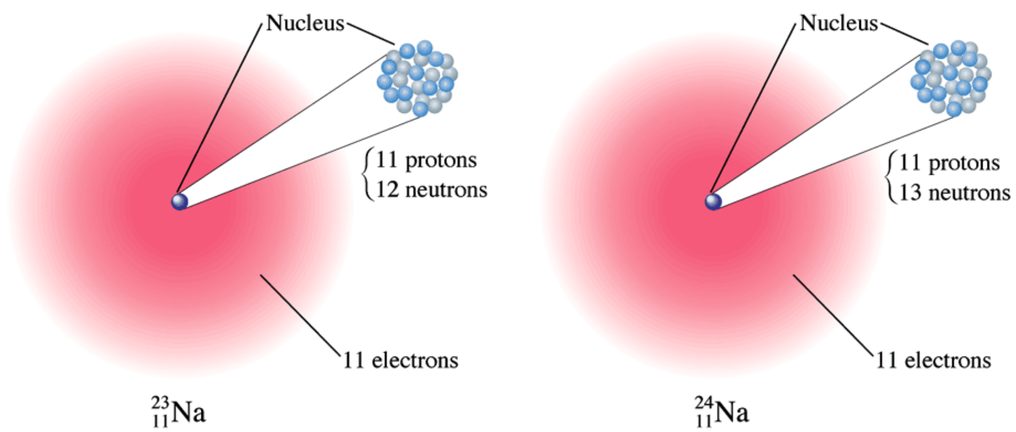


$_1H^1$ $_1H^2$ $_1H^3$

Protium **Deuterium** **Tritium**

Example-2:

Isotopes of sodium



All the atoms of an element have the same atomic number, but they can have different numbers of neutrons and different mass numbers.

Example-3:

Isotopes of Carbon.

There are three isotopes for carbon with mass number 12, 13, and 14, each possessing an atomic number of six.

Carbon-14; 8 neutrons

Carbon-13; 7 neutrons

Carbon-12; 6 neutrons

$$_6C^{12}\ _6C^{13}\ _6C^{14}$$

Allotropy

The elements can exist in two or more different forms are known as allotropes of that element.

Allotropes are different structural modifications of an element.

The phenomenon of allotropy is also called allotropism.

Example -1:

The element carbon has two common allotropes: Diamond and Graphite.

In diamond the carbon atoms are bonded together in a tetrahedral lattice arrangement.
In graphite the carbon atoms are bonded together in sheets of a hexagonal lattice arrangement.

Example-2:

The two allotropes of oxygen – dioxygen, O_2 and Ozone O_3.

2.6 Molecules and Ions

Molecules

In chemistry, a molecule is defined as a sufficiently stable electrically neutral group of at least two atoms in a definite arrangement held together by very strong chemical bonds.

It can also be defined as a unit of two or more atoms held together by covalent bonds.

Molecules containing same kind of atoms are called molecular elements and those containing more than one kind of atoms are called Molecular compound.

Example for Molecular Elements:

H_2 - Hydrogen
O_2 - Oxygen
O_3 - Ozone
Cl_2 - Chlorine
N_2 - Nitrogen

Example for Molecular Compounds:
H_2O - water
NH_3 - Ammonia
CH_4 - Methane
CO – Carbon monoxide

Types of Molecules

Homonuclear diatomic molecule

Molecules having two identical atoms.

Example: H_2, O_2, Cl_2, N_2.

Heteronuclear diatomic molecule

Molecules containing two different atoms.

Example: CO, HCl, NO, HBr.

Homonuclear polyatomic molecule

Molecules containing many identical atoms.

Example: P_4, P_8.

Heteronuclear polyatomic molecule

Molecules containing more than two atoms of different kinds.

Example: CH_4, NH_3, CH_3COOH, SO_2.

Ions

An ion is an atom or molecule, which has lost or gained one or more valence electrons, giving it a positive or negative electrical charge.

Example: Na^+, Mg^{2+}, Cl^-, Br^-

Cation:

A positive ion is called cation.

Example: Na^+, Mg^{2+}, NH_4^+

Anion:

Ions with negative charge are called anion.

Example: Cl^-, SO_4^{2-}

Chemical Bond

Chemical Bond is the existence of a strong force of binding between two or many atoms to form a stable compound.

Types of Chemical Bonds

- **Covalent bond,**
- **Ionic bond**
- **Co-ordinate-covalent bond**
- **Metallic bond.**

1. Covalent Bond

Covalent bond is formed between two atoms by mutual sharing of pairs of electron is termed as covalent bond.

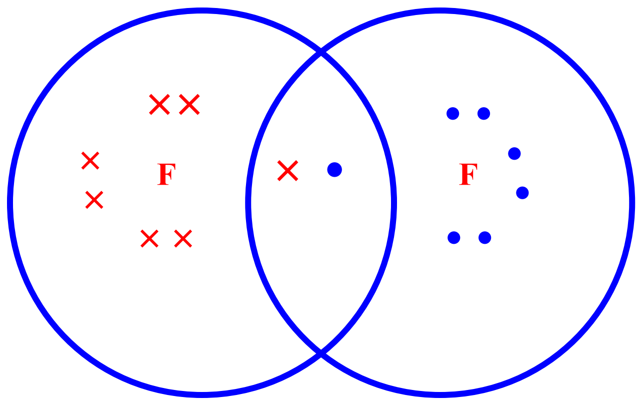
- **Each bond is the result of sharing of an electron pair between the atom.**
- **Each combining atom contributes one electron to the shared pair.**
- **The combining atom attain the outer filled shells of the noble gas configuration.**

Example

Examples of molecules that contain covalent bonds are

1. **Hydrogen (H_2)**
2. **Water (H_2O)**
3. **Oxygen (O_2)**
4. **Ammonia (NH_3)**
5. **Methane (CH_4)**

Example-1: **Single Bond**



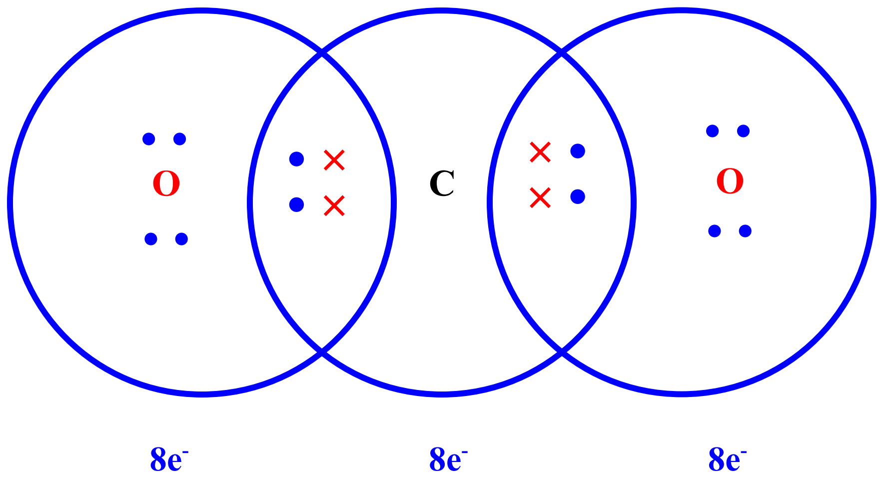
$8e^-$ $8e^-$ F_2 Molecule

If the two atoms share a pair of electrons, a single bond is formed.

F_2 Molecule.

F – F

Example-2: Double Bond



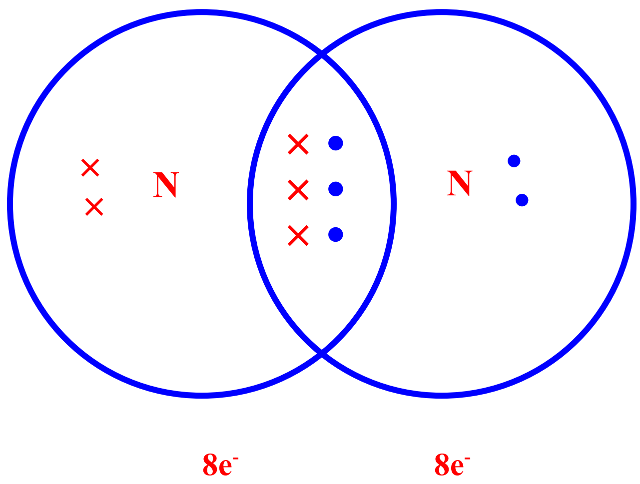
If the two atoms share two pairs of electrons, a double bond is formed.

Example:

CO_2 Molecule

O = C = O

Example-3 - Triple Bond



If the two atoms share three pair of electrons, a triple bond is said to be formed.

Example:

N_2 Molecule

$N \equiv N$

2. Electrovalent or Ionic Bond:

The electrostatic attraction force existing between the cation and the anion produced by the electron transfer from one atom to the other is known as the ionic or electrovalent bond.

The compounds containing such a bond are referred as ionic (or) electrovalent compounds.

- **Electropositive elements like (Na, K), (Mg, Ba) etc. belongs to group 1 and 2 posses the tendency to loose their valence electrons to attain the stable octet of noble gas electronic configuration and easily transformed as cations such as (Na^+, K^+), Mg^{2+} etc.**

- **Electro-negative elements like (F, Cl), (O, S) etc., belongs to 17, 16 (mostly p-block) possess the tendency to gain electrons into their valence shell to attain the stable octet of noble gas electronic configuration and easily transformed as anions such as F^-, Cl^-, O^{2-}, S^{2-} etc.**

- **The force of attraction that binds the cations and the anions together as a molecule and forms an ionic bond.**

Example -1:

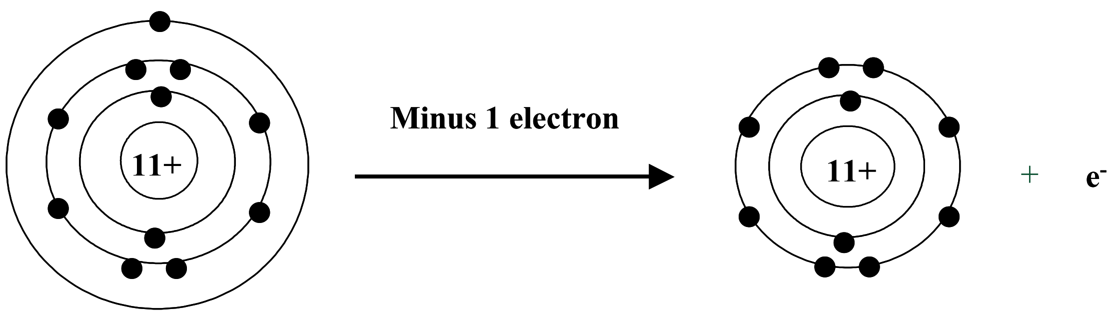
NaCl is formed by the electron ionization of sodium atom to Na^+ ion due to its low ionization potential value and chlorine atom to chloride ion by capturing the odd electron due to high electron affinity.

Thus NaCl (ionic compound) is formed. In NaCl, both the atoms possess unit charges.

Step 1: Ionization (Formation of Cation)

Na (g)	$\rightarrow$	Na^+ (g)	+	e^-
$2s^2 2p^6 3s^1$		$2s^2 2p^6$		
		Sodium cation		

Neutral Sodium atom (Na)　　　　Sodium ion (Na^+)



Na　　　　Na^+

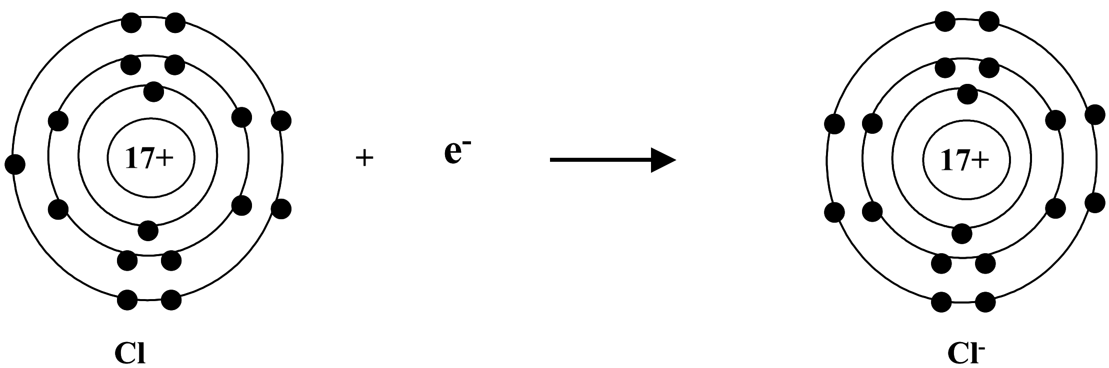
Na
11 electrons

Na^+
10 electrons

Step 2: Affinity (Formation of anion)

Cl (g) +	e^-	$\rightarrow$	Cl^-.
$3s^2\ 3p^5$			$3s^2\ 3p^6$
			Chloride anion

Neutral Chlorine atom (Cl) **Chloride ion (Cl^-)**



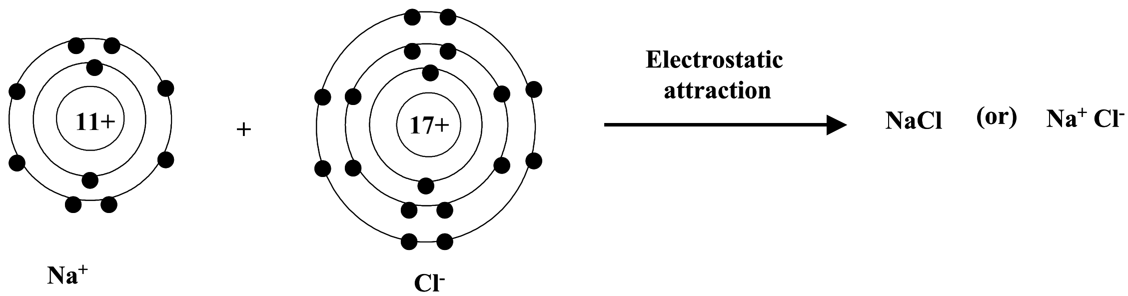
Cl
17 electrons

Cl^-
18 electrons

Step 3: (Electrostatic attraction)

$$Na^+ + Cl^- \xrightarrow{\text{Electrostatic attraction}} NaCl \text{ (or) } Na^+ Cl^-$$

Sodium ion **Chloride ion** **Ionic / crystalline compound is formed**



Na^+ Sodium ion **Cl^- Chloride ion** **Ionic or crystalline NaCl**

Example 2:

In CaO, which is an ionic compound, the formation of the ionic bond involves two electron transfers from Ca to O atoms. Thus, doubly charged positive and negative ions are formed.

Step 1: Ionisation

$Ca \rightarrow Ca^{2+} + e^-$

$3p^6 4s^2$

Calcium cation

Step 2: Affinity

$O + 2e^- \rightarrow O^{2-}$.

$2s^2\ 2p^4$ $\qquad$ $2s^2\ 2p^6$

Oxide anion

Step 3: Electrostatic attraction

$$Ca^{2+} + O^{2-} \xrightarrow{\text{Electrostatic attraction}} CaO$$

Ionic compound is formed

Example 3

Similarly formation of MgO may be shown to occur by the transfer of two electrons as:

Loss of e^-:

Mg	$\rightarrow$	Mg^{2+}	+	$2e^-$
[Ne] $3s^2$		**[Ne]**		

where [Ne] = electronic configuration of Neon
= $2s^2\ 2p^6$

Gain of e-:

O	+	$2e^-$	$\rightarrow$	O^{2-}.
[He] $2s^2\ 2p^4$				**[He] 2s2 2p6 (or) [Ne]**

[He] = electronic configuration of Helium

$$Mg^{2+} + O^{2-} \xrightarrow{\text{Electrostatic attraction}} MgO \text{ (or) } Mg^{2+}\, O^{2-}$$

Ionic compounds

Ionic compounds are compounds formed from cations (positively charged ions) and anions (negatively charged ions).

E.g. NaCl is composed of Na^+ and Cl^- ions.

When sodium metal reacts with chlorine gas (nonmetal) to form sodium chloride, electrons are transferred from sodium atoms to chlorine atoms to form sodium cations and chloride anions.

$$Na \rightarrow Na^+ + e^-$$
$$Cl + e^- \rightarrow Cl^-$$
$$Na^+ + Cl^- \rightarrow NaCl$$

- Ionic compounds do not have discrete molecular unit but have the cations and anions arranged in a three dimensional networks.
- The formulas of ionic compounds are always their empirical formulas.
- Ionic compounds are electrically neutral.
- If the charges of the cations and the anions are numerically different, the subscript of the cation is numerically equal to the charge of the anion and the subscript of the anion is numerically equal to the charge of the cation.

Al^{3+}, O^{2-} $\quad$ Al_2O_3

Mg^{2+}, Br^- $\quad$ $MgBr_2$.

- If the charges of the cations and the anions are numerically the same, then subscript are not necessary.

Na^+	Cl^-	NaCl
Ca^{2+}	O^{2-}	CaO

Ionic Solid

A solid consisting of oppositely charged ions is called an ionic solid, or a salt.

Example:

Sodium chloride – Simple ionic solid.

Ammonium nitrate (NH_4NO_3) – Polyatomic ions.

Chemical Formula

Chemical Formula is a representation of the composition of molecules and ionic compounds in terms of symbol.

- **The symbol of the elements represents the atoms present.**
- **The subscripts indicate the relative numbers of atoms.**

Example:

NH_3 contains N and H in the ratio of 1:3

Molecular Formula

It gives the exact number of atoms of all the elements present in a compound.

Example:

M.F of Benzene is C_6H_6.

Empirical formula

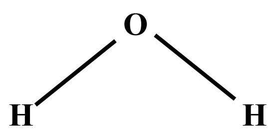
It is simplest formula of a compound, which gives the ratio of all atoms present in it.

Substance	Molecular formula	Empirical Formula
Water	H_2O	H_2O
Methane	CH_4	CH_4
Benzene	C_6H_6	CH
Sulfur	S_8	S

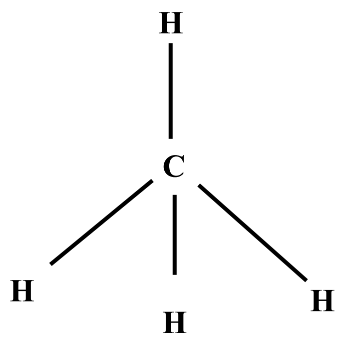
Structural Formula

The structural formula of a chemical compound is a graphical representation of the molecular structure showing how the atoms are arranged.

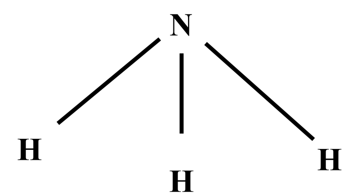
Example for Structural Formula



H_2O - Water



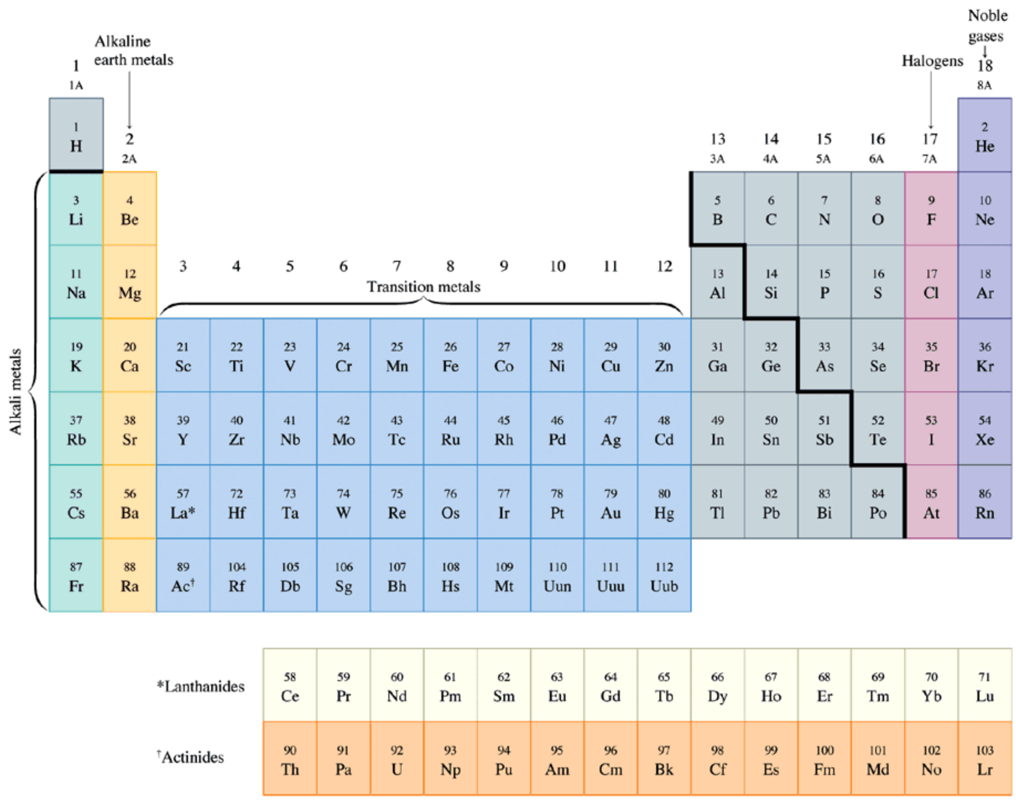
CH_4 - Methane



NH_3 - Ammonia

2.7 An Introduction to Periodic Table

Periodic Table



Structure of the Table

- **The horizontal rows of the periodic chart are called periods or rows.**
- **There are seven periods, each of which begins with an atom having only one valence electron and ends with a complete outer shell structure of an inert gas.**
- **The first three periods are short periods consisting of 2, 8, and 8 elements respectively.**
- **Periods 4 and 5 are long periods of 18 each, while period 6 has 32 elements, and period 7 is incomplete with only 17 elements most of which are radioactive and do not occur in nature.**

Periodic Table

Periodic table is defined as the arrangement of various elements according to their properties in a tabular form.

Mendeleev's periodic law

The physical and chemical properties of the elements are a periodic function of their atomic weights.

Modern periodic law

The physical and chemical properties of the elements are periodic functions of their atomic number.

Period

The horizontal row in the periodic table is known as period.

Groups

The vertical column in the periodic table is known as groups.

Periodicity in properties or Periodic properties

Repetition of properties of elements at regular intervals is called periodicity of properties.

The periodicity is due to similar electronic configuration of outermost shell.

Example:

1. **Atomic radii**
2. **Ionisation potential**
3. **Electron affinity**
4. **Electronegativity**

Triad

A group of three elements having similar properties.

Law of triads:

According to Dobereiner's law of triads, "when elements are arranged in the order of increasing atomic mass in a triad, the atomic mass of the middle element was found to be approximately equal to the arithmetic mean of the other two elements.

TABLE 2.2 **The Symbols for the Elements That Are Based on the Original Names**

Current Name	Original Name	Symbol
Antimony	Stibium	Sb
Copper	Cuprum	Cu
Iron	Ferrum	Fe
Lead	Plumbum	Pb
Mercury	Hydrargyrum	Hg
Potassium	Kalium	K
Silver	Argentum	Ag
Sodium	Natrium	Na
Tin	Stannum	Sn
Tungsten	Wolfram	W

Classification of Elements

Elements can be divided into three categories

- **Metals**
- **Non-Metals**
- **Metalloids**

Metals

- **Metals are good conductor of heat and electricity.**
- **Metals tend to lose electrons to form positive ions (Cations).**
- **Metals are solids at room temperature. (Exception Hg-liquid).**
- **They have melting and boiling points.**
- **They are malleable and ductile.**
- **They form ionic compounds.**
- **Metals are found on the left side of the periodic table with the most active metal in the lower left corner.**
- **Since the most active metals react with water to form bases, the Group I metals are called alkali metals. As you proceed to the right, the base-forming property decreases and the acid forming properties increase. The metals in the first two groups are the light metals, and those toward the center are called heavy metals.**

Non-metals

- **Non-metals are poor conductor of heat and electricity.**
- **Non-metals tend to gain electrons to form negative ions (anions).**
- **They are solids or gases at room temperature. (Exception Br_2-liquid).**
- **They have low melting and boiling points.**
- **They are neither malleable nor ductile. They are brittle.**
- **They form covalent compounds.**

- **Nonmetals are found on the right side with the most active nonmetal in the upper right hand corner of the chart.**

Metalloids

- **The elements, which show properties of both metals and non-metals, are known as metalloids (or) semi metals.**
- **The elements found along the dark line in the periodic table are called metalloids.**
- **Some examples of metalloids are boron, silicon, arsenic, and tellurium.**

- **In moving from left to right across the periodic table the physical and chemical properties change gradually from metallic to nonmetallic.**

FAMILIES OF ELEMENTS

Alkali Metals or Group IA or Group 1s Block Elements.

- **Alkali Metals are the elements in Group IA of the periodic table.**
- **The members of the family are lithium, sodium, potassium, rubidium, cesium, and francium.**
- **All six elements have the properties of metals except they are softer and less dense.**
- **They can be cut with a knife.**
- **They are the most reactive metals.**
- **They are so reactive that they are never found in nature.**
- **They are always combined with other elements.**
- **The alkali metals have only one electron in their outermost shell, so alkali metals form positive ions.**

General Characteristics of Alkali Metals

- **Electronic configuration: The elements Li, Na, K, Rb, Cs and Fr having the outer electronic configuration as ns^1 are termed as s-block elements.**
- **They readily form M^+ ions because of their low ionization enthalpy.**
- **The ionization enthalpy decreases down the group. Because of this, the electropositive character increases down the group.**
- **They exhibit +1 oxidation state in all their compounds.**
- **Because of their low ionization enthalpy the elements are good reducing agents.**
- **The metallic character increases down the group.**
- **Chemical reactivity: Alkali metals are highly reactive metals and the reactivity increases down the group.**

Alkaline Earth Metals (Group-2s Block Elements or Group 2A)

- **Alkaline Earth Metals** are **beryllium, magnesium, calcium, strontium, barium, and radium.**

- **These elements, which are harder and more dense than the alkali metals, also have higher melting points and boiling points.**

- **They are highly reactive, but not as active as the alkali metals.**

- **Like the alkali metals, the alkaline metals are never found free in nature.**

- **The alkaline earth metals have two electrons in their outermost energy level, so they also form positive ions.**

- **The word earth was applied in old days to a metallic oxide and because the oxides of calcium, strontium and barium produced alkaline solutions in water, these metals are called the alkaline earth metals.**

General Characteristics of Alkaline Earth Metals

.

Metallic Properties:

The alkaline earth metals are harder than alkali metals – possess high electrical and thermal conductivity.

Atomic and ionic radii

The atomic and ionic radii increases down the group.

Oxidation state

The alkaline earth metals exhibit +2 oxidation state. They form +2 ions (cations) when they react with nonmetals.

Group VII A or Halogens

- **Halogens** are the elements in family VIIA.
- They are strongly nonmetallic.
- The halogens include **fluorine, chlorine, bromine, iodine, and astatine**.
- They are the most active nonmetals.
- The chemical reactivity of the halogens is due to the number of electrons in the outermost energy levels of their atoms.
- Fluorine is the most active halogen.
- They have low melting points and boiling points.
- In the gas phase they exist as diatomic elements.
- Halongens combine readily with metals to form a class of compounds known as salts.
- The term halogen is derived from the Greek halos (=salt) and genes (=born) meaning salt producers because most of these elements exist in sea water, notably in the form of their sodium compounds.

Elements of Zero Group or Rare Gases or Group 8A

- **Zero group of the periodic table consists of six elements namely**

Helium,	**He (Z = 2)**
Neon,	**Ne (Z=10)**
Argon,	**Ar (Z = 18)**
Krypton,	**Kr (Z = 36)**
Xenon,	**Xe (Z = 54)**
Radon,	**Rn (Z = 86)**

- **These elements are also called by other names like inactive gases, inert gases, rare gases and noble gases.**
- **They have ns^2 np^6 electronic configuration.**
- **They have completely filled octet configuration.**
- **Due to that they are chemically inert.**
- **Noble Gasses** are colorless gasses that are extremely unreactive.
- Because do not readily combine with other elements to form compounds, the noble gasses are called inert.
- All the noble gasses are found in small amounts in the earth's atomsphere.
- One important property of the noble gasses is their inactivity.
- They are inactive because their outermost energy level is full.

Reasons for placing in zero group

- **These elements are chemically inert.**
- **The existence of such group is expected from the fact that there must be a separating group between strongly electronegative halogens of VIIA group and Strongly electropositive alkali metals of 1 A group.**
- **Therefore, these elements should be placed in between VIIA group and IA group, i.e., in the zero group.**

2.8 Naming Simple Compounds

When Chemistry was an infant science, there was no system for naming compounds.

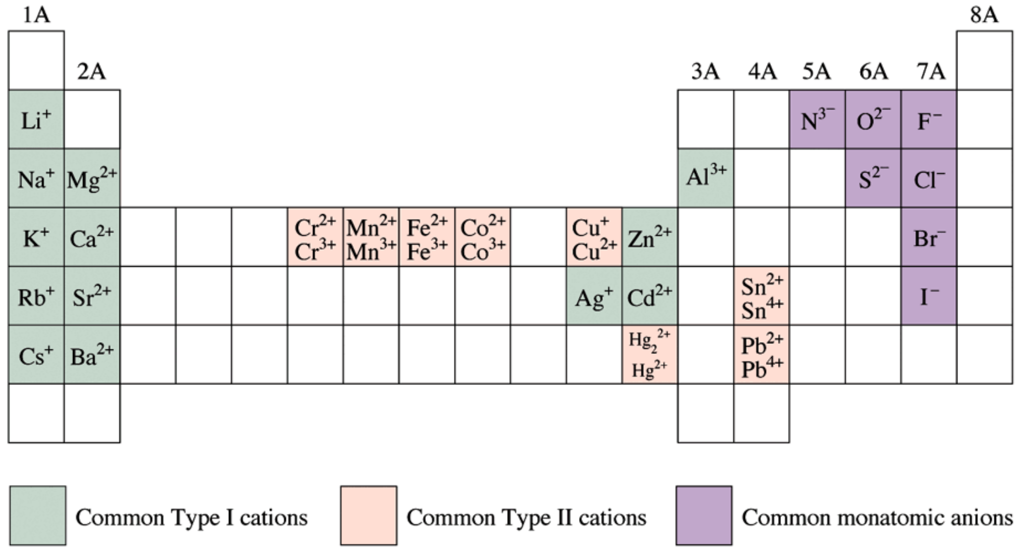
As chemistry grew, it became clear that using common names for compounds would lead to unacceptable one. Nearly 5 million chemical compounds are currently known. Memorizing common names for these compounds would be an impossible task.

Therefore it is to adopt a system for naming compounds in which the name tells something about the composition of the compound.

After learning the system, a chemist given a formula should be able to name the compound or, given a name, should be able to construct the compounds formula.

In this section we will specify the most important rules for naming compounds.

The common cations and anions



Naming inorganic Binary Compounds

Binary Compounds:

- **Compounds composed of two elements - Which we classify into various types for easier recognition.**
- **We will consider both ionic and covalent compounds.**

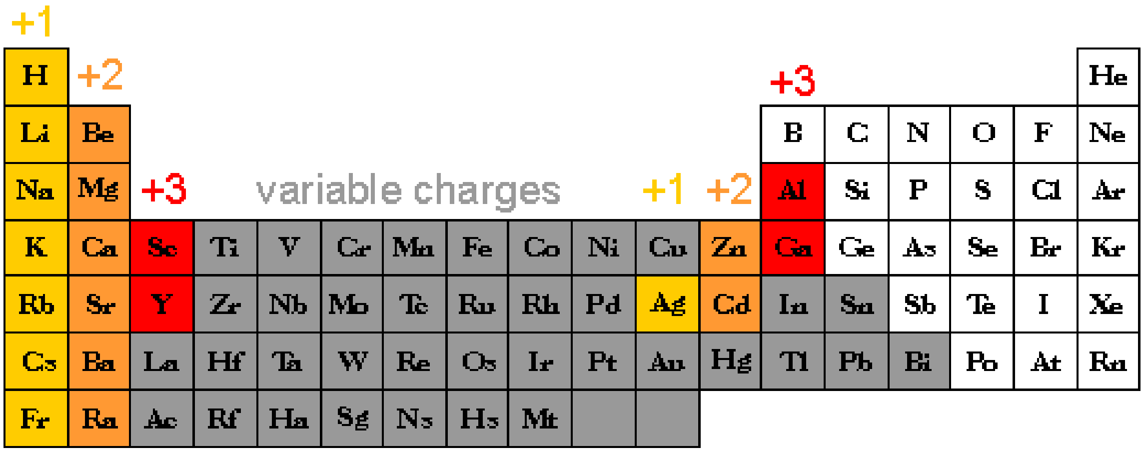
Binary Ionic Compounds (Type I)

Binary ionic compounds contain a positive ion (cation) always written first in the formula and a negative ion (anion). In naming these compounds, the following rules apply:

- **The cation is always named first and the anion second.**
- **A monatomic (meaning "one atom") cation takes its name from the name of the element.**

  Example:
  Na^+ is called sodium in the names of compounds containing this ions.
- **A monoatomic anion is named by taking the root of the element name and adding -ide. Thus Cl^- ion is called chloride.**

Elemental Cations

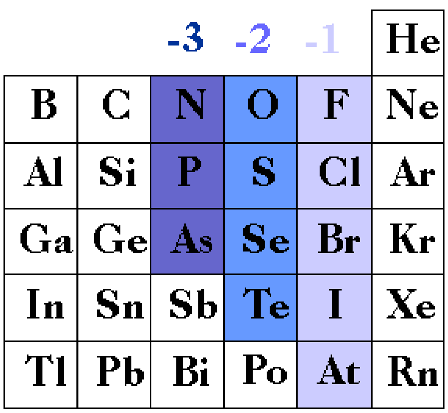


Fe^{2+} iron(II) ion
Fe^{3+} iron(III) ion
K^{+} potassium ion
Al^{3+} aluminum ion
Ba^{2+} barium(~~II~~) ion

Indicate the charge with a Roman numeral…

…unless the element forms only one cation.

Elemental Anions



H^{-} hydride ion
F^{-} fluoride ion
Cl^{-} chloride ion
Br^{-} bromide ion
O^{-2} oxide ion
S^{-2} sulfide ion
N^{-3} nitride ion

TABLE 2.3 Common Monatomic Cations and Anions

Cation	Name	Anion	Name
H^+	Hydrogen	H^-	Hydride
Li^+	Lithium	F^-	Fluoride
Na^+	Sodium	Cl^-	Chloride
K^+	Potassium	Br^-	Bromide
Cs^+	Cesium	I^-	Iodide
Be^{2+}	Beryllium	O^{2-}	Oxide
Mg^{2+}	Magnesium	S^{2-}	Sulfide
Ca^{2+}	Calcium	N^{3-}	Nitride
Ba^{2+}	Barium	P^{3-}	Phosphide
Al^{3+}	Aluminum		
Ag^+	Silver		

The rules for naming binary ionic compounds are illustrated by the following examples:

In formulas of ionic compounds, simple ions are represented by the element symbol: Cl means Cl^-, Na means Na^+, and so on.

Compound	Ions Present	Name
NaCl	Na^+, Cl^-	Sodium chloride
KI	K^+, I^-	Potassium iodide
CaS	Ca^{2+}, S^{2-}	Calcium sulfide
Li_3N	Li^+, N^{3-}	Lithium nitride
CsBr	Cs^+, Br^-	Cesium bromide
MgO	Mg^{2+}, O^{2-}	Magnesium oxide

Naming Type I Binary Compounds

1. CsF
2. $AlCl_3$
3. LiH

Solution

1. CsF - Cesium fluoride
2. $AlCl_3$ - Aluminium chloride
3. LiH - Lithium hydride

Note that, in each case, the cation is named first, and then the anion is named.

Binary Ionic Compounds (Type II)

- Type II binary ionic compounds contain a metal that can form more than one type of cation.
- In this case the charge on the metal ion must be specified.
- A compound must be electrically neutral.
- A compound containing a transition metal usually requires a Roman numeral in its name.

Example:

The compound $FeCl_2$ contains Fe^{2+} ions, and the compound $FeCl_3$ contains Fe^{3+} ions.

Therefore the systematic name of these two compounds are

$FeCl_2$ - Iron(II) chloride.
$FeCl_3$ – Iron(III) chloride.

The ion with the higher charge has a name ending in *–ic*, and the one with the lower charge has a name ending in *–ous*.

Example:

$FeCl_3$ - Ferric chloride.
$FeCl_2$ – Ferrous chloride.

List of systematic names for many common Type II cations.

TABLE 2.4 Common Type II Cations

Ion	Systematic Name
Fe^{3+}	Iron(III)
Fe^{2+}	Iron(II)
Cu^{2+}	Copper(II)
Cu^{+}	Copper(I)
Co^{3+}	Cobalt(III)
Co^{2+}	Cobalt(II)
Sn^{4+}	Tin(IV)
Sn^{2+}	Tin(II)
Pb^{4+}	Lead(IV)
Pb^{2+}	Lead(II)
Hg^{2+}	Mercury(II)
Hg_2^{2+}*	Mercury(I)
Ag^{+}	Silver†
Zn^{2+}	Zinc†
Cd^{2+}	Cadmium†

*Note that mercury(I) ions always occur bound together to form Hg_2^{2+} ions.

†Although these are transition metals, they form only one type of ion, and a Roman numeral is not used.

Naming Type II Binary Compounds

Give the systematic name of each of the following compounds.

1.
CuCl
2. **HgO**
3. **Fe_2O_3**
4. **MnO_2**
5. **$PbCl_2$**

All these compounds include a metal that can form more than one type of cation; thus we must first determine the charge on each cation.

This can be done by recognizing that a compound must be electrically neutral; that is, the positive and negative charges must exactly balance.

Solution

1. **CuCl - Copper(I) chloride.**
2. **HgO - Mercury(II) oxide.**
3. **Fe_2O_3 - Iron(III) oxide.**
4. **$PbCl_2$ - lead(II) chloride.**

Naming Binary Compounds

Give the systematic name of each of the following compounds:

$CoBr_2$
$CaCl_2$
Al_2O_3
$CrCl_3$

Solution

Compound	Name	Comment
$CoBr_2$	Cobalt(II) bromide	Cobalt is a transition metal; the compound name must have a Roman numeral. The two Br- ions must be balanced by a Co^{2+} ion.
$CaCl_2$	Calcium chloride	Calcium, an alkaline earth metal, forms only the Ca^{2+} ion. A Roman numeral is not necessary.
Al_2O_3	Aluminium oxide	Aluminium forms only Al^{3+}. A Roman numeral is not necessary.
$CrCl_3$	Chromium(III) chloride	Chromium is a transition metal. The compound name must have a Roman numeral. $CrCl_3$ contain Cr^{3+}.

Ionic Compounds with Polyatomic Ions

Polyatomic ions are assigned special names that must be memorized to name the compounds containing them.

The most important Polyatomic Ions and their names are listed in Table 2.5

Common Polyatomic Ions

TABLE 2.5 Common Polyatomic Ions

Ion	Name	Ion	Name
Hg_2^{2+}	Mercury(I)	NCS^-	Thiocyanate
NH_4^+	Ammonium	CO_3^{2-}	Carbonate
NO_2^-	Nitrite	HCO_3^-	Hydrogen carbonate (bicarbonate is a widely used common name)
NO_3^-	Nitrate		
SO_3^{2-}	Sulfite		
SO_4^{2-}	Sulfate	ClO^-	Hypochlorite
HSO_4^-	Hydrogen sulfate (bisulfate is a widely used common name)	ClO_2^-	Chlorite
		ClO_3^-	Chlorate
		ClO_4^-	Perchlorate
OH^-	Hydroxide	$C_2H_3O_2^-$	Acetate
CN^-	Cyanide	MnO_4^-	Permanganate
PO_4^{3-}	Phosphate	$Cr_2O_7^{2-}$	Dichromate
HPO_4^{2-}	Hydrogen phosphate	CrO_4^{2-}	Chromate
$H_2PO_4^-$	Dihydrogen phosphate	O_2^{2-}	Peroxide
		$C_2O_4^{2-}$	Oxalate

Naming Compounds Containing Polyatomic Ions

Give the systematic name of each of the following compounds:

1. Na_2SO_4
2. KH_2PO_4
3. $Fe(NO_3)_3$
4. $Mn(OH)_2$
5. Na_2SO_3
6. Na_2CO_3
7. $NaHCO_3$
8. $CsClO_4$
9. $NaOCl$
10. Na_2SeO_4
11. $KBrO_3$

Compound	Name	Comment
Na_2SO_4	Sodium sulfate	
KH_2PO_4	Potassium dihydrogen phosphate	
$Fe(NO_3)_3$	Iron(III) nitrate	Transition metal – name must contain a Roman numeral. Fe^{3+} ion balances three NO_3^- ions.
$Mn(OH)_2$	Manganese(II) hydroxide	Transition metal – name must contain a Roman numeral. Mn^{2+} ion balances two OH^- ions.
Na_2SO_3	Sodium sulfite	
Na_2CO_3	Sodium carbonate	
$NaHCO_3$	Sodium hydrogen carbonate	Often called sodium bicarbonate
$CsClO_4$	Cesium perchlorate	
$NaOCl$	Sodium hypochlorite	
Na_2SeO_4	Sodium selenate	Atoms in the same group, like sulfur and selenium, often form similar ions that are named similarly. Thus SeO_4^- is selenate, like SO_4^{2-} (Sulfate).
$KBrO_3$	Potassium bromate	As above, BrO_3^- is bromate, like ClO_3^- (chlorate)

Binary Covalent Compounds (Type III)

In binary covalent compounds, the element names follows the same rules as for binary ionic compounds.

Binary covalent compounds are formed between two nonmetals. Although these compounds do not contain ions, they are named very similarly to binary ionic compounds.

Naming of binary covalent compounds, the following rules apply:

The first element in the formula is named first, using the full element name.

The second element is named as if it were an anion.

Prefixes are used to denote the numbers of atoms present. These prefixes are given in Table 2.6.

The prefix mono- is never used for naming the first element. For example, CO is called carbon monoxide, not monocarbon monoxide.

Prefixes Used to Indicate Number in Chemical Names

Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Names of the several covalent compounds formed by nitrogen and oxygen

Compound	Systematic Name	Common Name
N_2O	Dinitrogen monoxide	Nitrous oxide
NO	Nitrogen monoxide	Nitric oxide
NO_2	Nitrogen dioxide	
N_2O_3	Dinitrogen trioxide	
N_2O_4	Dinitrogen tetroxide	
N_2O_5	Dinitrogen pentaoxide	

In order to avoid awkward pronunciations, we often drop the final *o* or *a* of the prefix when the element begins with a vowel.

For example:

N_2O_4 is called dinitrogen tetroxide, not dinitrogen tetraoxide.

CO is called carbon monoxide, not carbon monooxide.

Some compounds are always referred to by their common names.

H_2O – Water

NH_3 - Ammonia

Naming Type III Binary Compounds

Sl. No.	Compound	Name
1	PCl_5	Phosphorus pentachloride
2	PCl_3	Phosphorus trichloride
3	SF_6	Sulphur hexafluoride
4	SO_3	Sulfur trioxide
5	SO_2	Sulfur dioxide
6	CO_2	Carbon dioxide

Naming Various Types of Compounds

Sl. No.	Compound	Name	Comment
1	P_4O_{10}	Tetraphosphorus decaoxide	Binary covalent compound (Type III) so prefixes are used. The *a* in *deca-* is sometimes dropped.
2	Nb_2O_5	Niobium(V) oxide	Type (II) binary compound containing Nb^{5+} and O^{2-} ions. Niobium is a transition metal and requires a Roman numeral.
3	Li_2O_2	Lithium peroxide	Type I binary compound containing the Li^+ and O_2^{2-} (peroxide) ions.
4	$Ti(NO_3)_4$	Titanium(IV) nitrate	Not abinary compound. Contains the Ti^{4+} and NO_3^- ions. Titanium is a transition metal and requires a Roman numeral.

Writing Compound Formulas from Names

Name	Chemical Formula	Comment
Vanadium(V) fluoride	VF_5	The compound contains V^{5+} ions and requires five F^- ions for charge balance
Dioxygen difluoride	O_2F_2	The prefix *di-* indicates two of each atom.
Rubidium peroxide	Rb_2O_2	Since rubidium is in Group 1A, it forms only 1^+ ions. Thus two Rb^+ ions are needed to balance the 2- charge on the peroxide ion (O_2^{2-})
Gallium oxide	Ga_2O_3	Since gallium is in Group 3A ions. Two Ga^{3+} ions are required to balance the charge on three O^{2-} ions.

ACIDS

- Acids can be recognized by the hydrogen that appears first in the formula.
- When dissolved in water, certain molecules produces free H^+ ions (protons) in the solution.
- An acid can be viewed as a molecule with one or more H^+ ions attached to an anion.

Rules for Naming Acids

If the anion does not contain oxygen, the acid named with *prefix hydro-* and the *suffix –ic*.

Example:

HCl - Hydrochloric acid
HCN - Hydrocyanic acid
H_2S - hydrosulfuric acid.

Names of Acids that do not contain oxygen

Acid	Name
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
HCN	Hydrocyanic acid
H_2S	Hydrosulfuric acid

When anion contains oxygen

When anion contains oxygen, the acidic name is formed from the root name of anion with a suffix of *-ic* or *–ous*, depending on the name of the anion.

1. If the anion name ends in *–ate*, the *suffix –ic* is added to the root name.

Example:

H_2SO_4 contains sulfate anion (SO_4^{2-}) and is called sulfuric acid.

H_3PO_4 contains the phosphate anion (PO_4^{3-}) and is called phosphoric acid.

CH_3COOH contains the acetate ion CH_3COO^- and is called acetic acid.

2. If the anion has an *–ite ending*, the *–ite* is replaced by *–ous*.

Example:

H_2SO_3 which contains sulfite (SO_3^{2-}), is named as sulfurous acid.

HNO_2, which contains nitrite (NO_2^-), is named nitrous acid.

Names of some oxygen containing acids

Acid	Name
HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid
H_2SO_3	Sulfurous acid
H_3PO_4	Phosphoric acid
CH_3COOH	Acetic acid

The application of these rules can be seen in the names of the acids of the oxyanions of chlorine.

Acid	Anion	Name
$HClO_4$	Perchlor*ate*	Perchlor*ic* acid
$HClO_3$	Chlor*ate*	Chlor*ic* acid
$HClO_2$	Chlor*ite*	Chlor*ous* acid
$HClO$	Hypochlor*ite*	Hypochlor*ous* acid