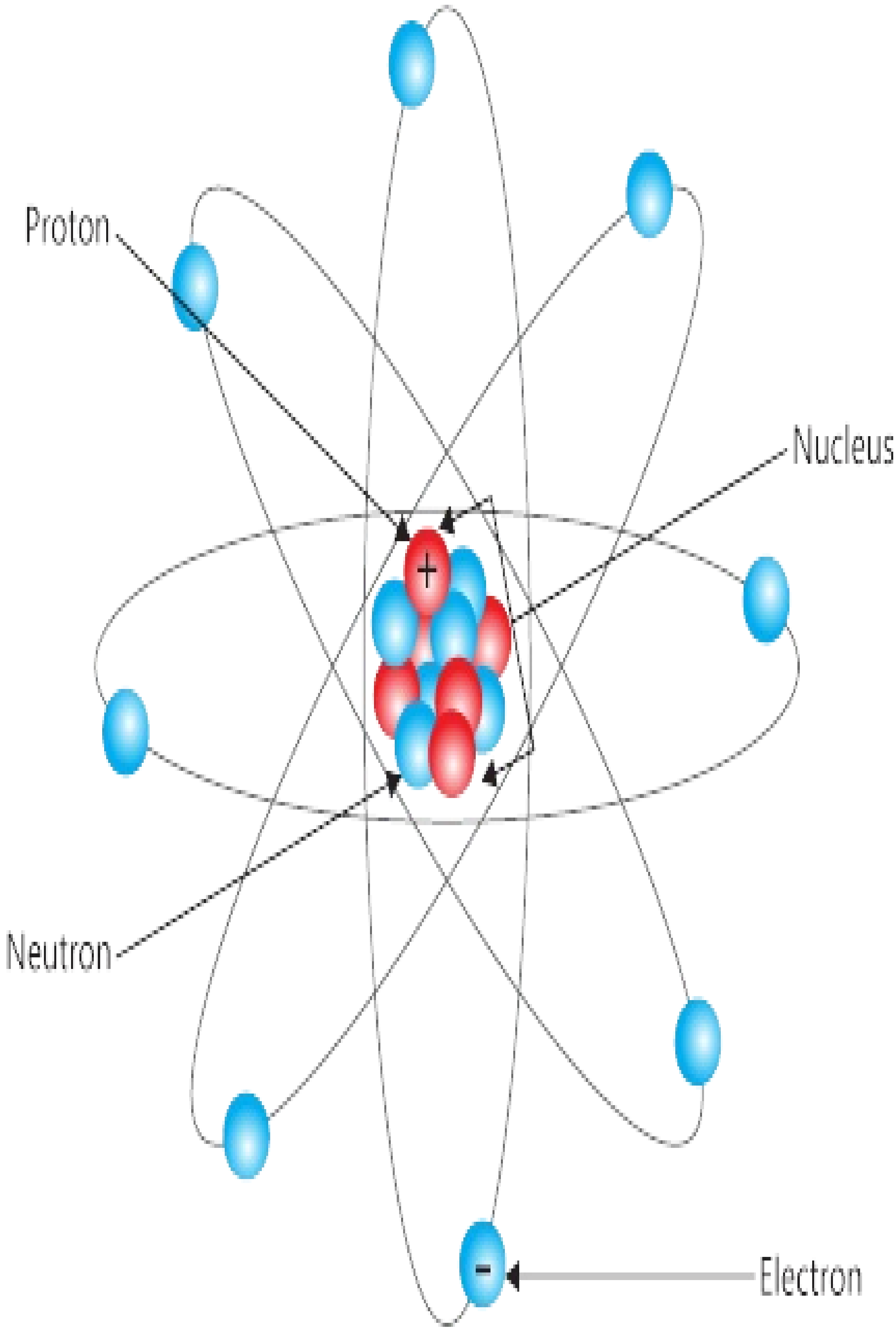

Proton
Nucleus
+
Neutron
-
Electron

Basics of Chemistry

Thank you for choosing our Books

Contents

Chapter 10: Acid & Bases

Chapter 11: Intermolecular Forces

KNOWLEDGE FLOW BOOKS

India

Introduction

Introduction

C

hemistry deals with the composition, formation and properties of matter. This slant can be best described and understood in terms of basic unit of matter: atoms and molecules. Any material substance with mass & volume is called matter. There are three phases of matter.

Solid - Solid has definite shape and definite volume.

Liquid - Liquid has indefinite shape, takes the shape of the container and definite volume.

Gas - Gas has indefinites shape, takes the shape of the container and indefinite volume that can expand and be compressed.

“Chemistry is the science of molecules and their transformations. It is the science not so much of the one hundred elements but of the infinite variety of molecules that may be built from them”

-Roald Hoffmann

Importance of Chemistry

Science can be viewed as a continuing human toil to systematize mastery for describing and understanding nature. For the cause of comfort science is sub-divided into various disciplines: chemistry, physics, biology, geology etc. Chemistry is the limb of science that studies the composition, properties and interaction of matter. Chemists are enquiring in knowing how chemical transformations occur.

Chemistry plays a mean role in science and is often entwine with other branches of science like physics, biology, geology etc. Chemistry also plays an important character in daily life. Chemical principles are important in diverse areas, such as: weather composition, functioning of brain and working of a computer. Chemical industries produce fertilizers, alkalis, acids, alloys, dyes, polymers, drugs, metals, detergents, soaps, salts and other inorganic and organic chemicals, including new materials, accord in a big way to the national economy.

Chemistry plays an chief role in meeting human necessitate for food, health care products and other materials aimed at improving the quality of life. This is typified by the large scale production of a variety of fertilizers, improved difference of pesticides and insecticides. Similarly many lives save medicine such as cisplatin and taxol, are effective in cancer therapy and AZT (Azidothymidine) used for helping AIDS casualty, have been isolated from plant and animal sources or prepared by synthetic methods.

Basics of Chemistry

Matter is the most important part of chemistry all matter in universe is composed of atoms. Any material substance with mass & volume is called matter. These elements are composed of only one type of atom. Atoms are mostly empty space.

Atoms have electrons which are very microscopic and are negatively charged and have a negligible mass (mass = 0). Electrons move in orbits around the center of the atom in relatively distinct region called energy levels - Aka, Orbits or Shells.

The farther from the center an electron is the more energy it has.

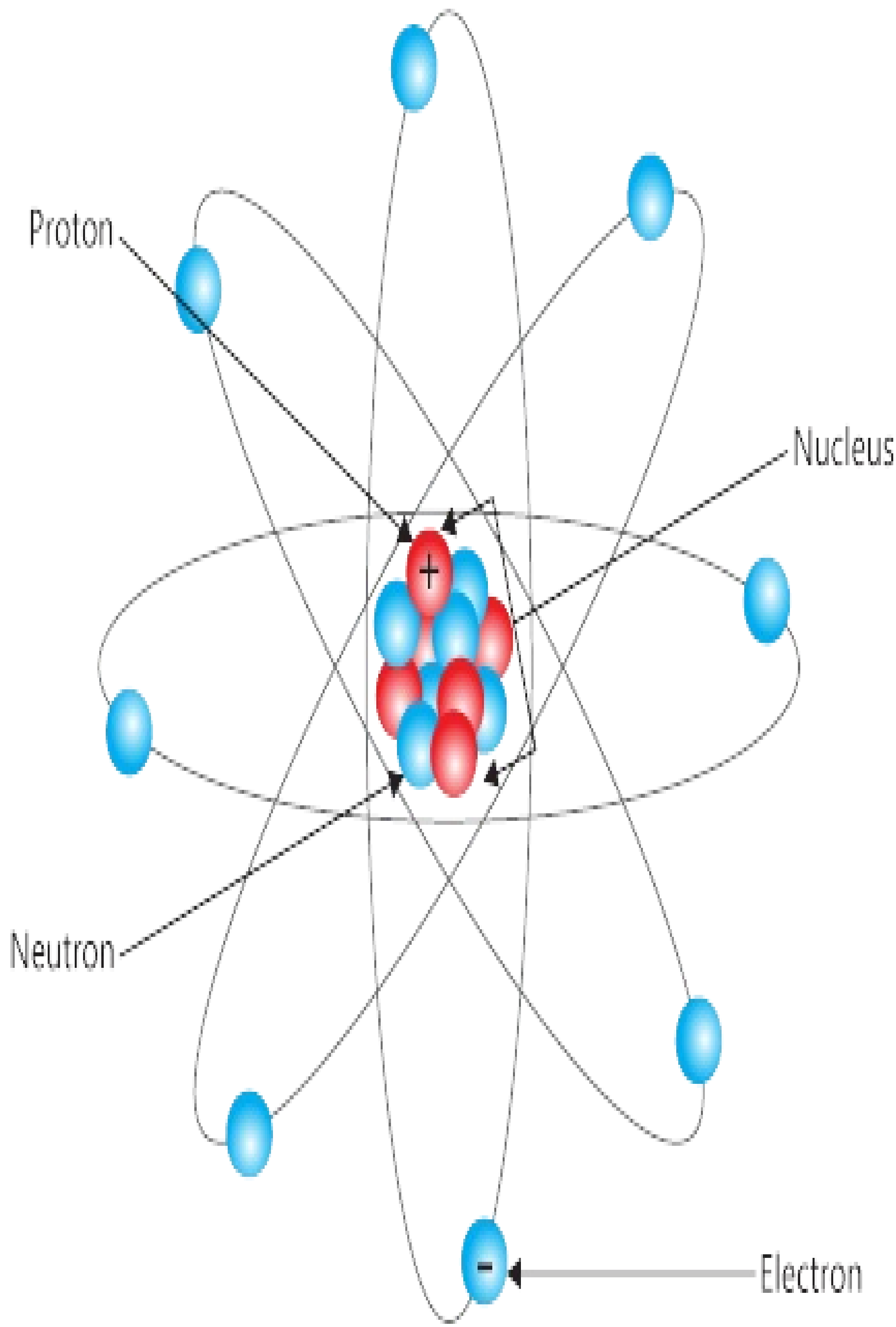
Electrons (& therefore atoms, can gain and lose energy) and do this by moving between energy levels.

Atoms Have a Nucleus Which Contains Protons and Neutrons

Protons are positively charged and have a mass =1

The number of protons in an atom's nucleus determines what element it is.

Neutrons have no charge and are therefore called neutral and have a mass = 1.


Proton
Nucleus
Neutron
Electron

Atomic structure

2

Organic Chemistry

Introduction

The basic information that environmental engineers and scientists need about organic chemistry is differs from that the organic chemist requisite. This difference is due to the fact that chemists are concerned chiefly with the mash-up of compounds, whereas environmental engineers and scientists are concerned, in the main, with how the organic mixture in liquid, solid, and gaseous wastes can be destroyed and how they react in the environment.

Organic chemistry is the branch of chemistry of carbon compounds; inorganic chemistry is also a branch of the chemistry of compounds of all elements other than carbon.

Chemistry of the compounds present in living organisms.

They all contain carbon. This organic chemistry is the chemistry of carbon.

An organic compound is one that has carbon as the uppermost element.

An inorganic element is any compound that is not an organic compound.

Carbon has unique property in chemistry.

It has 6 electrons in its outer shell arranges $1s^2 2s^2 sp^2$.

These organic compounds have discrete geometry throughout carbon to carbon bond.

If there are 4 atoms or groups throughout a carbon atom, it has a tetrahedral geometry.

General properties of Carbon Compounds Are:

Organic compounds are usually inflammable.

Organic compounds, commonly, have lower melting and boiling points.

Organic compounds are normally less soluble in water.

Reactions of organic compounds are generally molecular rather than ionic. As a result, they are often entirely slow.

The molecular weights of organic compounds may be very high-rise, often well over 1000.

Most organic compounds can serve as an inventor of food for bacteria.

Most carbon compounds are non-electrolytes.

The reaction rates of carbon compounds are usually slow.

Many carbon compounds oxidize slowly in air but quickly rapidly if heated.

Most carbon compounds are unsteady at high temperatures.

Natural Sources of Organic Compounds

All organic compounds are derived from three main sources.

Nature - Alkaloids, vegetable oils, animal oils and fats, Fibers, cellulose, starch, sugars, and so on.

Synthesis - Manufacturing process produced compounds and materials.

The Fermentation - All this glycerol, Alcohols, antibiotics, acetone, acids, and the like are obtain by the action of microorganisms upon organic matter.

The wastes create in the processing of natural organic materials and from the synthetic organic and fermentation industries compose a major part of the industrial and hazardous waste obstacle that environmental engineers and scientists are called upon to solve.

Living things → Carbohydrates / Proteins / Fats / Vitamins / Antibiotics

Development of Organic Chemistry as a Science

Organic chemistry is the branch of chemistry of carbon compounds (except CO, CO2, carbonates, hydrogen carbonates, carbides and cyanides) acquire from natural sources or synthesized in the laboratories.

Elements

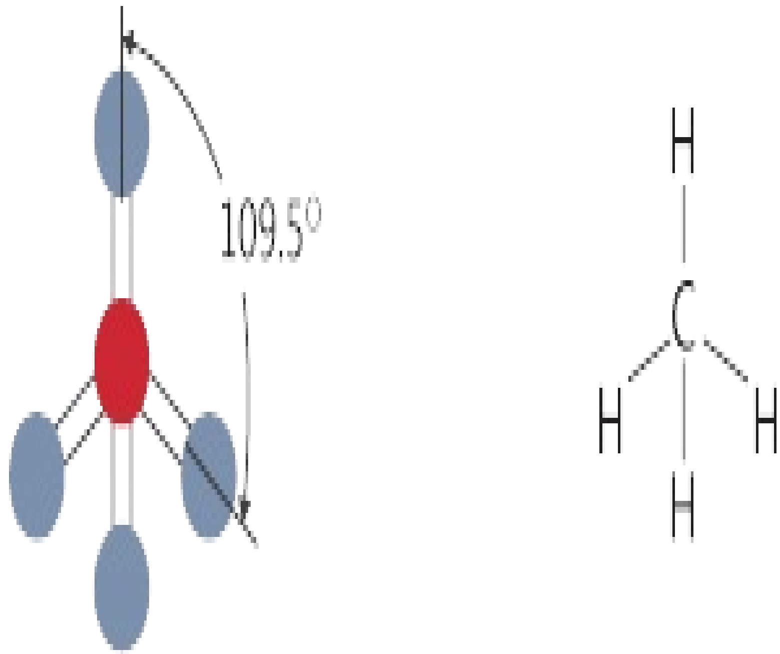
All organic compounds contain carbon in composition with one or more elements. The hydrocarbons contain only carbon and hydrogen. Many compounds contain carbon, hydrogen, and oxygen, and they are considered to be the prime elements. Minor elements in naturally arise compounds are nitrogen, phosphorus, and sulphur, and sometimes halogens and metals.

Organic Compounds

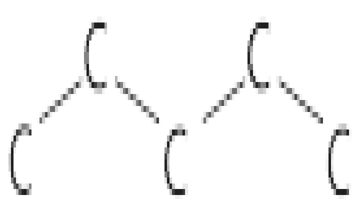
The carbon atom design bonds in a tetrahedral structure with a bond angle of 109.5O.

In carbon-to-carbon bond angles are 109.5O, so a chain of carbon atoms form a zigzag pattern.

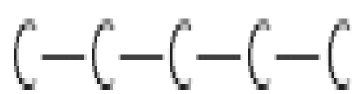
The unbranched chain of carbon atoms is normally clarify in a way that looks like a straight chain, but it is actually a zigzag, as shown in fig.



(A) Three-dimensional model



(B) An unbranched chain



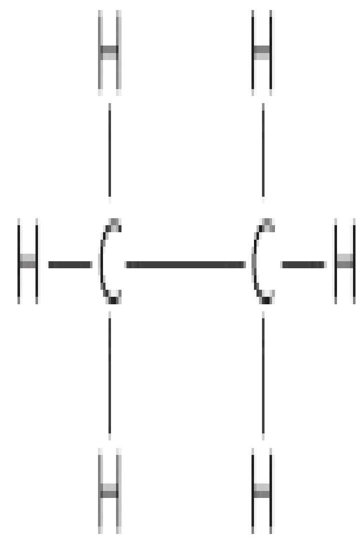
C) Simplified unbranched chain

Organic Compounds

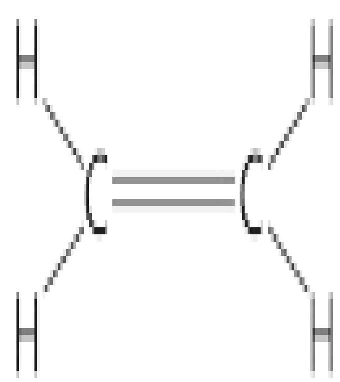
Hydrocarbons

A hydrocarbon is a compound contain of only hydrogen and carbon. In hydrocarbon carbon to carbon can be single, double, or triple bonds. These bonds are always nonpolar. Alkanes are hydrocarbons with only single bonds, alkanes is called a homologous series. Each successive compound differs from the one before it only by a CH_2.

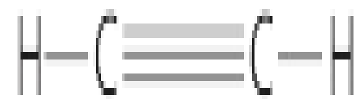
In this reaction Carbon-to-carbon bonds can be single (A), double (B), or triple (C). Note that in each example, each carbon atom has 4 dashes, which shows four bonding pairs of electrons, satisfying the octet rule.



(A) Ethane



(B) Ethene



C) Ethyne

Hydrocarbons

Compounds which has the same molecular formula, but unlike structures are called isomers.

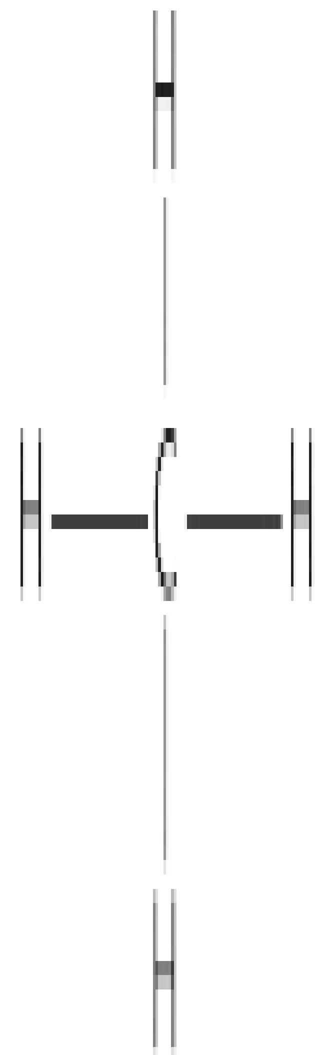
Naming Alkenes

Discover the longest continuous chain. The locations or other class of atoms attached to the longest chain are recognize and numbered by counting from the end of the molecule which retain the numbering system as low as possible. All hydrocarbon groups are attached to the longest constant chain and named using the parent name and changing the – ane suffix to – yl.

A molecular formula (A) represents the numbers of different types of atoms in a molecule, and a structural formula (B) shows a two-dimensional model of how the atoms are bonded to each other. Each dash represents a bonding pair of electrons.

CH_4

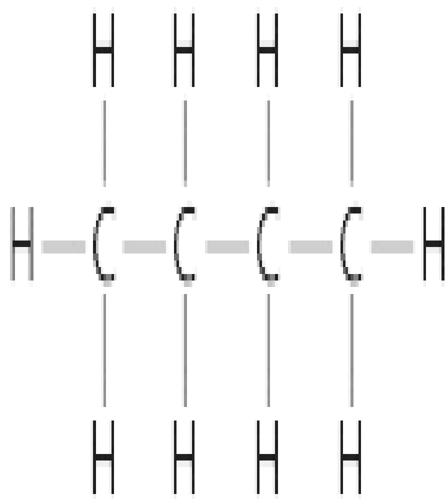
(A)



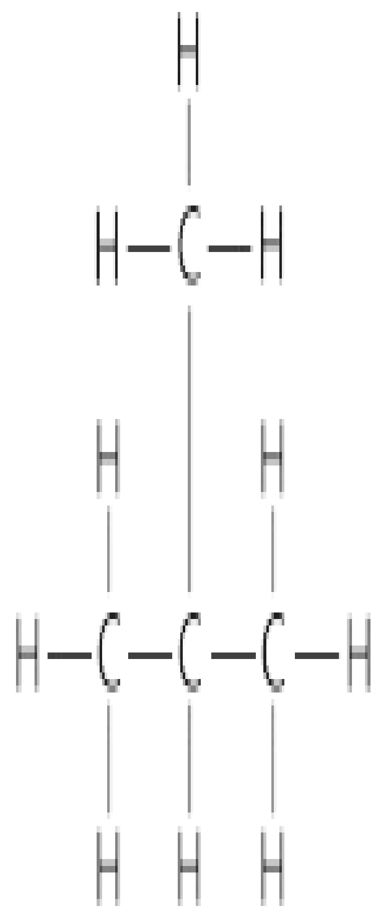
(B)

(A) Molecular formula and (B) Structural formula

(A) straight-chain alkane is recognized by the prefix n- for "normal" in the common naming system. (B) a branched-chain alkane isomer is recognized by the prefix iso- for "isomer" in the common naming system. In the IUPAC name, isobutane is 2-methylpropane, (Carbon bonds are generally the same length).



(A) *n*-butane, C_4H_{10}



(B) Isobutane (2-methylpropane), C_4H_{10}

Alkenes and Alkynes

Alkenes are hydrocarbons with minimum one paired carbon to carbon bond. To represent the presence of the paired bond, the –ane suffix from the alkane name is transmute to – ene. These alkenes are unsaturated with respect to hydrogen. In this process it does not have the maximum number of hydrogen atoms as it would if it were an alkane (a saturated hydrocarbon). Naming is homogeneous to naming alkenes except.

The longest continuous chain must carry the double bond.

The base name now ends when find out –ene.

The carbons are numbered so as to keep the number for the double bond as short as possible.

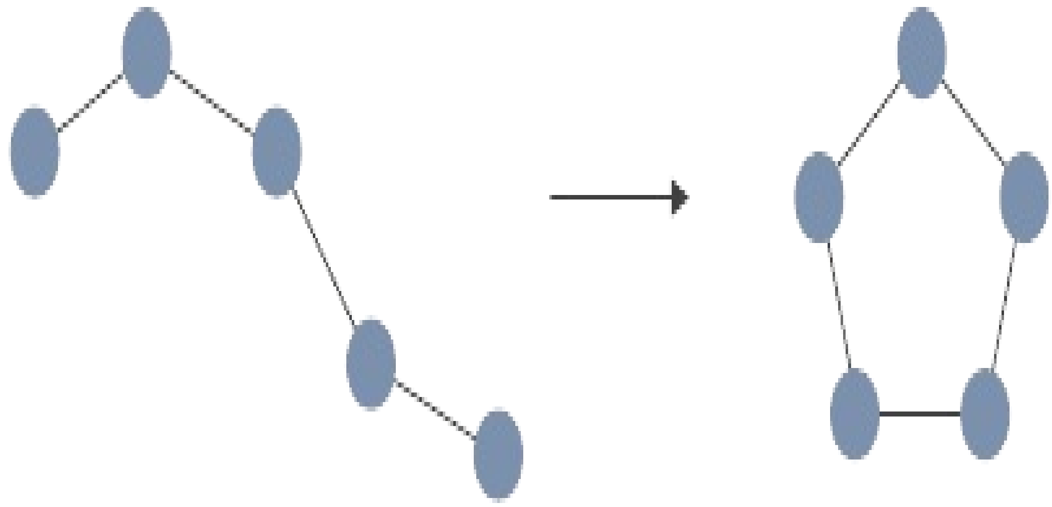
The base name is given a number which discover the location of the paired bond.

An alkyne is a hydrocarbon with minimum one carbon to carbon triple bond. Naming an alkyne is similar to the alkenes, except the base name ends in – yne.

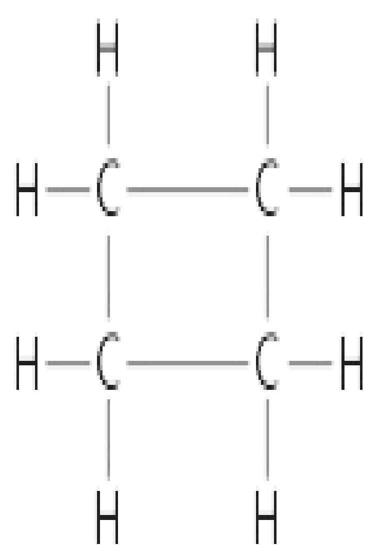
Cycloalkanes and Aromatic Hydrocarbons

Cycloalkanes are alkanes (create the single bonds between carbon to carbon) which create a ring structure. An aromatic (fragrant) compound is one that is based on the benzene ring. A benzene ring that is attached to another compound is given the name phenyl.

In this reaction (A) shows the "straight" chain has carbon atoms that are able to rotate freely through their single bonds, sometimes connect in a closed ring. (B) Ring compounds of the first four cycloalkanes.



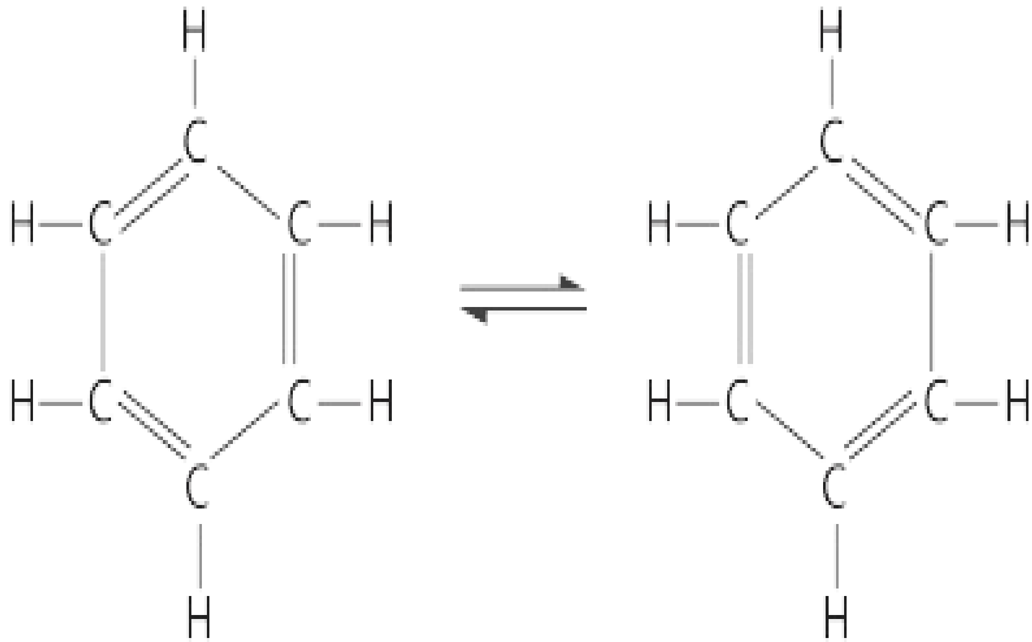
(A)



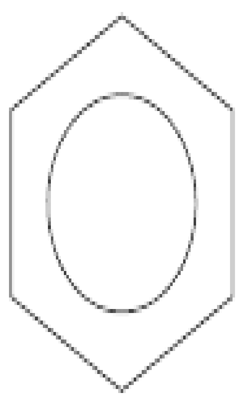
(B) Cyclobutane, C_4H_8

(A)The straight chain and (B) Ring compounds

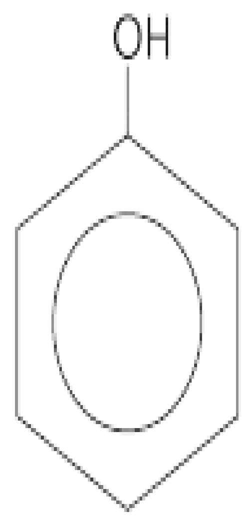
In this reaction (A) the bonds in C_6H_6 are something between single and double, which gives it different chemical properties than paired-bond hydrocarbons. (B) The six-sided symbol with a circle represents the benzene ring.



(A)



Benzene



Phenol

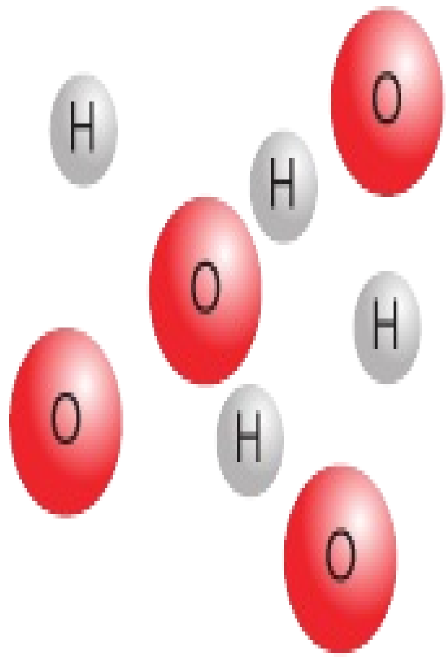
(B)

3

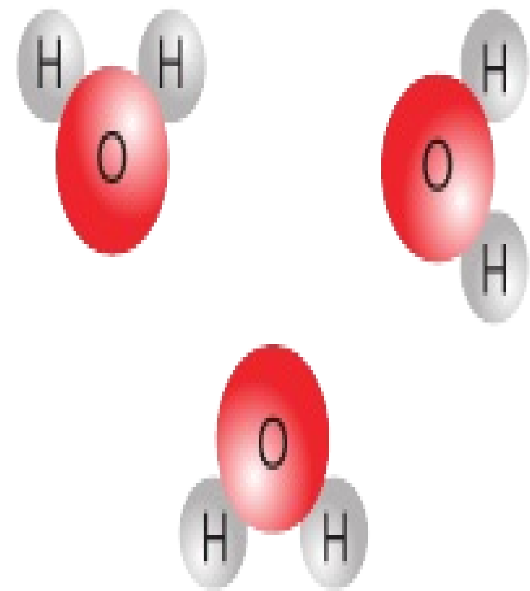
Atomic Theory

Dalton Atomic Model

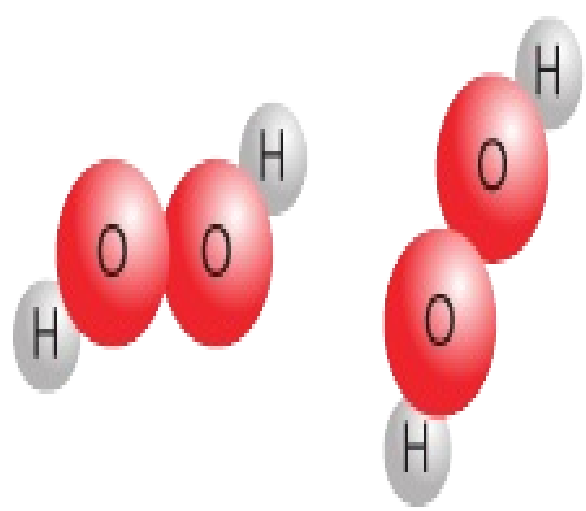
Elements are composed of minute, which indivisible particles called atoms. Atoms of the equal element are alike in masses and sizes. Atoms of distinct elements have different masses and sizes. Compounds are formed by the coupling of two or more atoms of different elements. Atoms merge to form compounds in simple numerical ratios. Atoms of two elements may merge in different ratios to form more than one compound.



(A)



(B)



(C)

Dalton atomic model

Structure of Atoms

In bonding mechanisms between atoms are jointly associated to the structure of the atoms themselves.

Atoms = nucleus (protons & neutrons) + electrons

Charges

Electrons and protons have charges that are negative and positive. These charges have the equal magnitude, 1.6 × 10-19 Coulombs. Neutrons (N) are electrically neutral.

Mass

Protons and Neutrons have the equal mass, 1.67 × 10-27 kg. Electron mass is much smaller, 9.11 × 10-31 kg and can be neglected in calculation of atomic mass.

The atomic mass (A) = mass of protons (P) + mass of neutrons (N)

Protons

Protons (P) give chemical identification of the element. Protons (P) are equal to atomic number (Z) and neutron defines isotope number.

Atomic Mass Units and Atomic Weight

Atomic Mass

The atomic mass unit (amu) is often used to explicit atomic weight. 1 amu is defined as 1/12 of the atomic mass of the most frequent isotope of carbon atom that has 6 protons (Z=6) and six neutrons (N=6). Mproton ≈ Mneutron = 1.66 x 10-24 g = 1 amu unit. The atomic mass of the 12C atom is 12 amu units.

Atomic Weight

Atomic weight of an element = weighted mean of the atomic masses of the atoms usually occurring isotopes. Atomic weight of carbon is 12.011 amu units. The atomic weight is often specified in mass per mole.

Mole

A mole is the amount of element that has a mass in grams uniform to atomic mass in amu unit of the atoms (A mole of carbon has a mass of 12 grams). Mole’s atoms number is called the Avogadro number, Nav = 6.023 × 1023.

1 amu unit/atom = 1 gram/mol.

Properties of Electric Charge

Charge may be positive or negative.

Unlike charges attract, like charges repel.

Charge may be transferred from one object to another, by contact.

The less the distance between two charges, the higher force of attraction (or repulsion) the force can be expressed.

$F = kq1q2/r2$

Where q1, q2: charge, r: distance, k: constant.

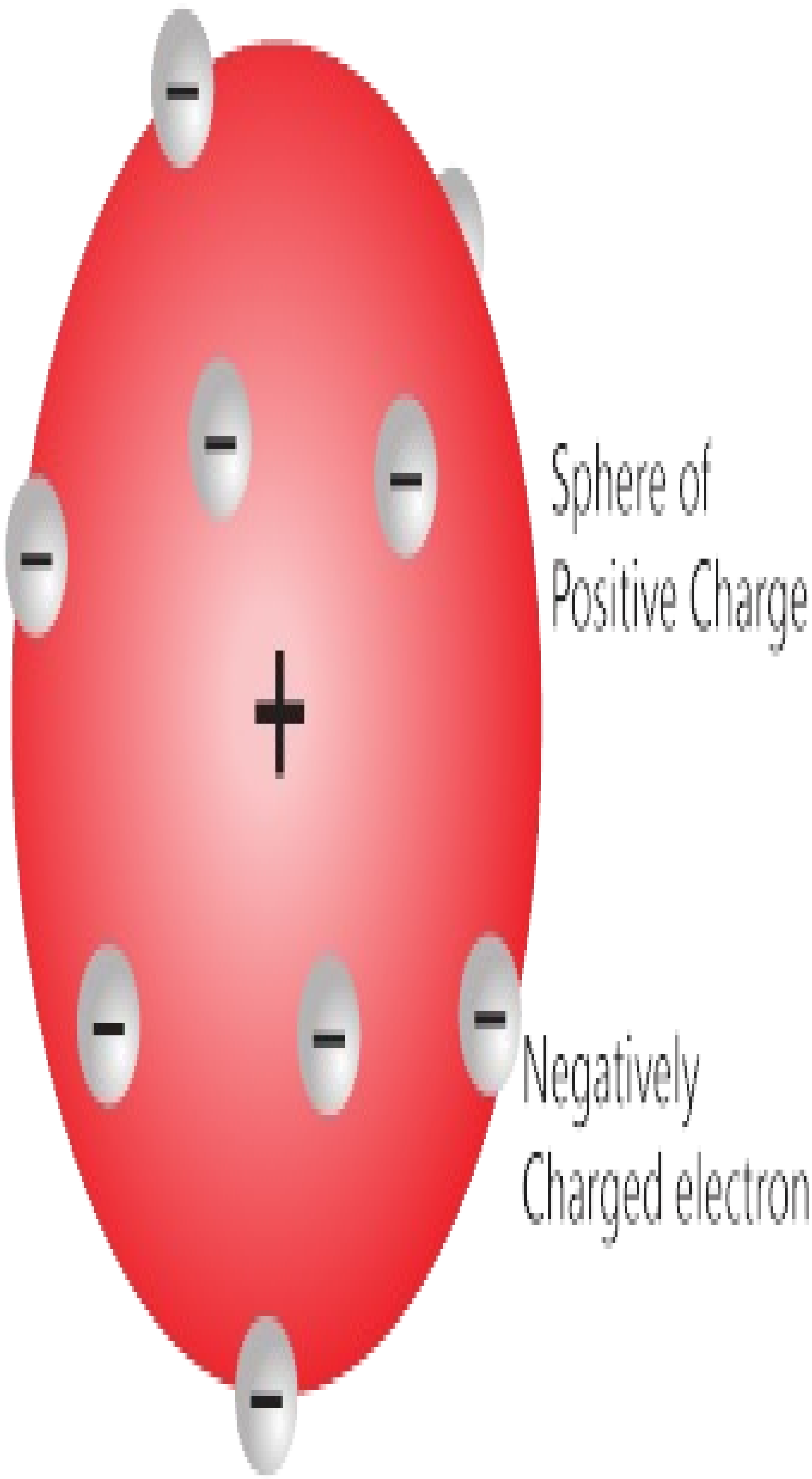
Discovery of Ions

According to the Faraday's rule certain substances conduct an electric current when dissolved in water. Ions - Arrhenius ion is an atom convey a positive or negative charge. Na^+ ions move toward the cathode, Cl^- ions move toward the anode. Positive (+) ion called cation. Negative (-) ion called anion.

NaCl Na+ + Cl-

Thomson Discovery of Electron - Thomson Model of Atom

The electrons and the protons are negatively charged particles embedded in the atomic sphere a neutral atom could become an ion by gaining or losing electrons.

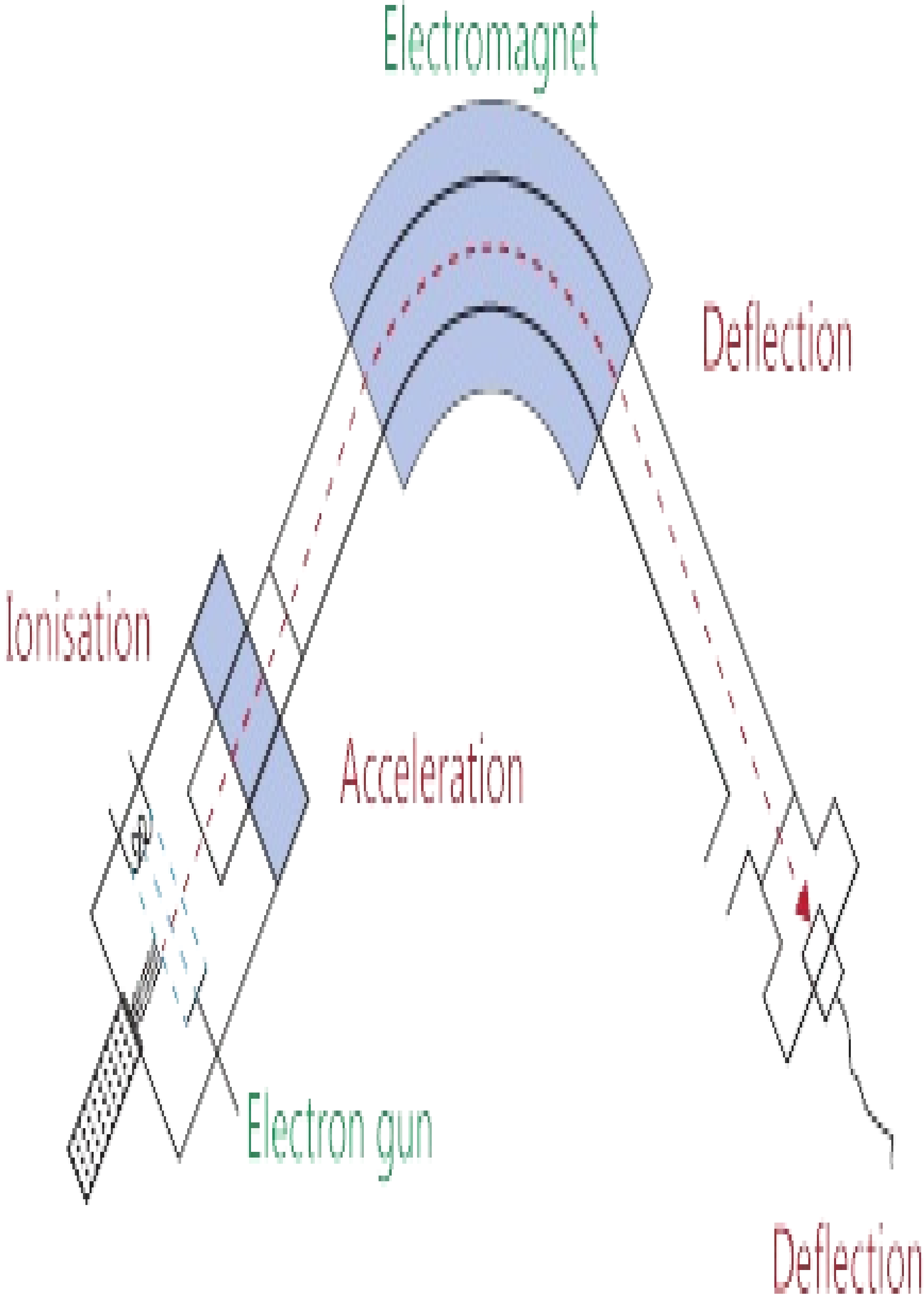

Sphere of
Positive Charge
Negatively
Charged electron

Thomson model of atom

Mass Spectrometer

This instrument is determining the masses of individual atoms. 1 atomic mass unit (amu) – defined as equal to exactly 1/12 of the mass of a 12C atom.

1 amu = 1.66 × 10-24 g, Number of neutrons = Mass number – Atomic number


Electromagnet
Deflection
Ionisation
Acceleration
Electron gun
Deflection

Mass spectrometer

4

Periodic Table

Introduction

T

he periodic table arranges the elements in a particular way. A great deal of information about an element can be assembling from its position in the period table. For example, you can predict with reasonably good accuracy the physical and chemical attribute of the element. Elements are organized on the periodic table according to their atomic number, generally found near the top of the square.

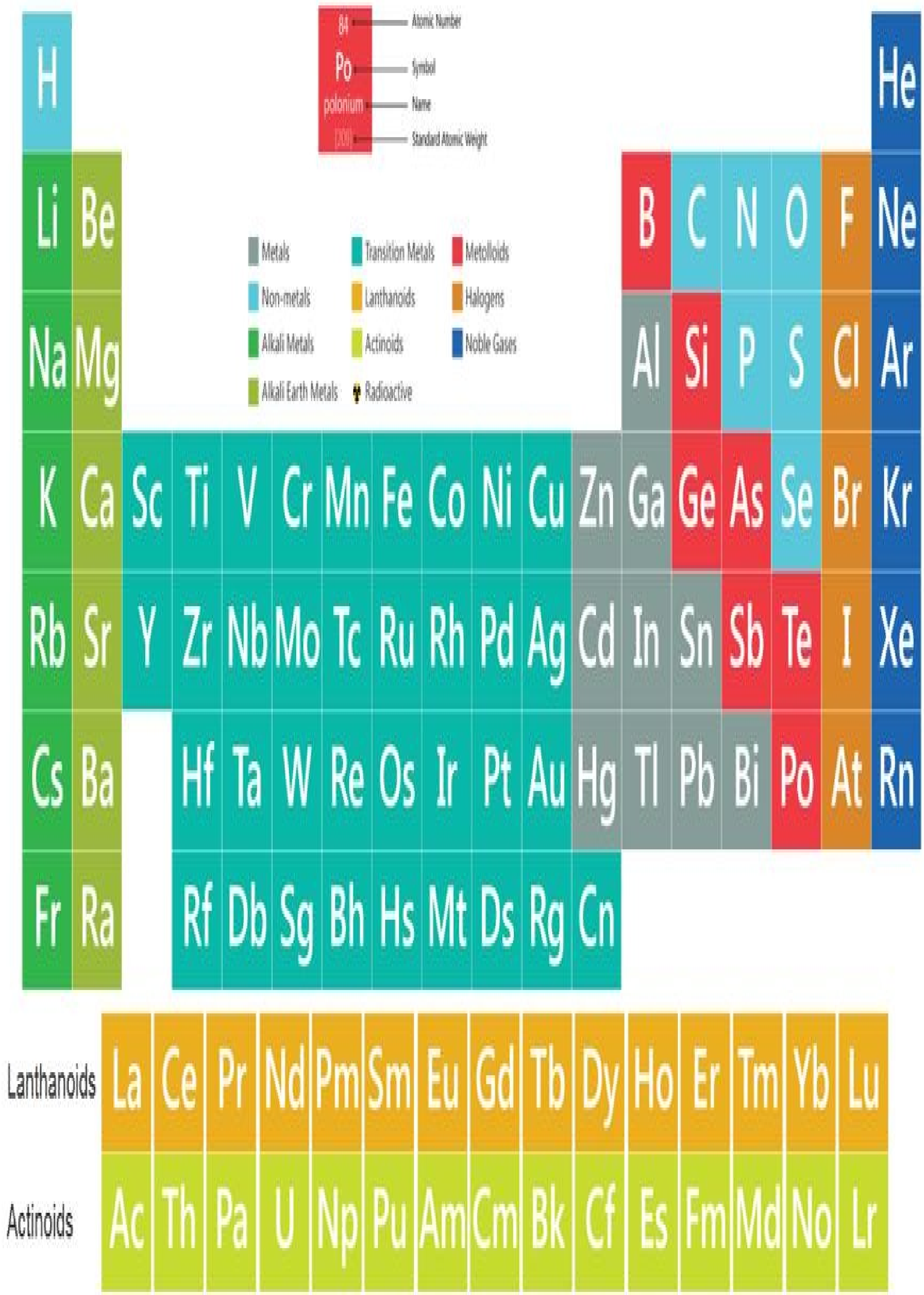
The atomic number refers to how many protons an atom of that element has.

For instance, hydrogen has 1 proton, so its atomic number is 1.

The atomic number is unique to that element. No two elements have the similar atomic number.

Scientists contribute a lot of time organizing information into useful patterns. Before they can organize information, however, they must acquire it, and it must be correct. Botanists had sufficient information about plants to organize their sector in the eighteenth century. Because of uncertainties in atomic masses and because many elements remained unexplored, chemists were not able to organize the elements until a century later.

We can determine one element from all others by its particular set of observable physical properties. For example, sodium has a low thickness of and a low melting point of 97.81 °C. No other element has this same combination of thickness and melting point. Potassium, though, also has a low thickness and low melting point (63.65 °C), much like sodium.



84
Po
polonium
[209]
Atomic Number
Symbol
Name
Standard Atomic Weight
Metals
Transition Metals
Metolloids
Non-metals
Lanthanoids
Halogens
Alkali Metals
Actinoids
Noble Gases
Alkali Earth Metals
Radioactive
H He
Li Be B C N O F Ne
Na Mg Al Si P S Cl Ar
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe
Cs Ba Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn
Fr Ra Rf Db Sg Bh Hs Mt Ds Rg Cn
Lanthanoids La Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu
Actinoids Ac Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr

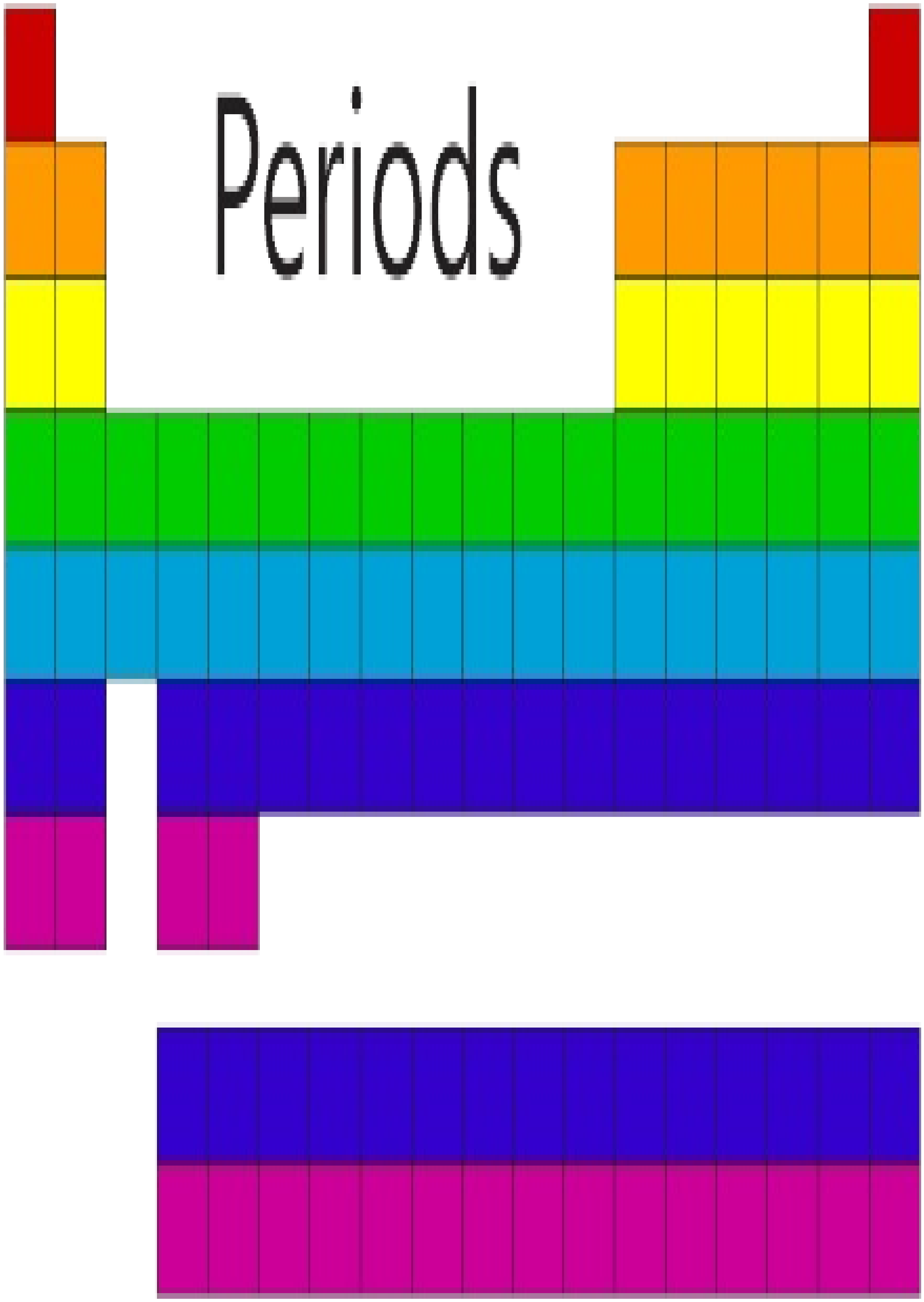
Periodic Table

Created by Mendeleev

The search for a systematic arrangement of the elements started with the discovery of individual elements. By 1860 about 60 elements were known and a procedure was needed for association. In 1869, Russian chemist Dimitri Mendeleev propounds arranging elements by atomic weights and properties. The table contained gaps but Mendeleev predicted the finding of new elements.

Arrangement

Periodic table contains rows (left to right) and columns (up and down). In periodic table, elements have something in common if they are in the same row (left to right). In this period all elements have the same number of atomic orbital. In top row (the first period) every element has one orbital for its electrons. All elements of the second row (the second period) have two orbital for their electrons. It goes down the periodic table like that. Periodic table has a special name for its columns, when a column goes from top to bottom, it's called a group.


Periods

Period (Period=Row)

In this Periodic Table Groups = Columns, The elements in a group have the same number of electrons in their outer orbital. In periodic table every element in the first column (group one) has one electron in its outer shell. Same as Every element on the second column (group two) of periodic table has two electrons in the outer shell. As you keep counting the columns of periodic table, you'll know how many electrons are in the outer shell. There are some exceptions to the order when you focus at the transition elements, but you get the general idea.

Structural Features

In this table there are 18 vertical columns called groups. They are numbered from 1 to 18. Every group has a unique configuration. There are seven horizontal rows. These rows are called periods thus the periodic table has seven periods, numbered from 1 to 7. There are a total of 114 elements known to us till today. All the known elements 90 are naturally arising and others are made through nuclear transformations or are synthesised artificially. Either way they are synthetic elements, but you will find the term specifically applied to transuranic elements (elements listed after uranium) only.

First period consists of only two elements (very short period). A second and third period consists of only 8 elements each (short periods). Fourth and fifth periods include 18 elements each (long periods). Sixth period consists of 32 elements (long period). Seventh period is yet incomplete and more and more elements are adjoining to be added as the scientific research advances. There are also snick names given to the groups or a cluster of groups on the basis of the similarity of their properties, as given below:

Group 1 element except hydrogen, are called alkali metals.

Group 2 elements are called alkaline earth metals.

Elements of group 3 to 12 are called transition metals.

Elements of group 16 are called halogens.

Elements of group 17 elements are called halogens.

Elements of group 18 elements are called noble gases.

Apart from this above elements with atomic numbers 58 to 71 are called lanthanoids – or Inner transition elements (First series). Elements of periodic table from atomic numbers 90 to 103 are called actinoids – Inner transition elements (Second series). This all elements except transition and inner transition elements are also collectively called Main group elements.

Alkali Metals

Alkali metals are very reactive metals that do not arise freely in nature malleable, ductile, good conductors of heat and electricity. Alkali metals can explode if they are exposed to water.

Li
Na
K
Rb
Cs
Fr

Alkali Metals

Alkaline Earth Metals

Second column of periodic table is Group 2. In second column reactive metals that are always combined with nonmetals in nature. In this group several elements are important mineral nutrients (such as Mg and Ca).

Transition Metals

In transition metals family elements are in groups 3-12. Transition metals less reactive harder metals. Transition metals include metals used in jewelry and construction.

Boron Family

In Boron family's elements are in group 13. In this family aluminum metal was once rare and expensive, not a "disposable metal."

Carbon Family

In carbon family's elements are in group 14. Carbon family contains elements important to life and computers. Carbon is the basis for an entire branch of chemistry. In carbon family silicon and germanium are important semiconductors.

Nitrogen Family

In nitrogen family's elements are in group 15. Nitrogen makes up over ¾ of the atmosphere. Nitrogen and phosphorus is both major key in living things. Most of the world's nitrogen is not available to living things. The red material on the tip of matches is phosphorus.

Oxygen Family or Chalcogens

In oxygen family elements are in group 16. Group 16's elements oxygen is necessary for respiration. Many things that stink carry sulfur (rotten eggs, garlic, skunks, etc.)

Halogens

These elements are in group 17. Halogens are very reactive, volatile, diatomic, nonmetals. Halogens are always found combined with other element in nature. Halogens used as disinfectants and to strengthen teeth.

Noble Gases

These elements are in group 18. Noble gases are very uncreative, monatomic gases. Noble gases used lighted “neon” signs. Noble gases used blimps to fix the hindenburg problem. Noble gases have a full valence shell.

Bonds

Introduction

A

chemical bonding is defined as the force of attraction which holds or binds the constituent atoms present in a molecule together. When two atoms present in a molecule tend to remain together strongly, it is said that they are in chemical bonding with each other. It is the tendency of atoms to acquire or achieve a condition of stable configuration which results in a chemical combination between them.

Types of Chemical Bonding

Ionic or Electrovalent bond

Covalent bond

Co-ordinate covalent bond

Metallic bond

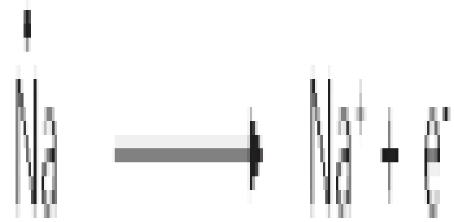
Ionic or Electrovalent Bond

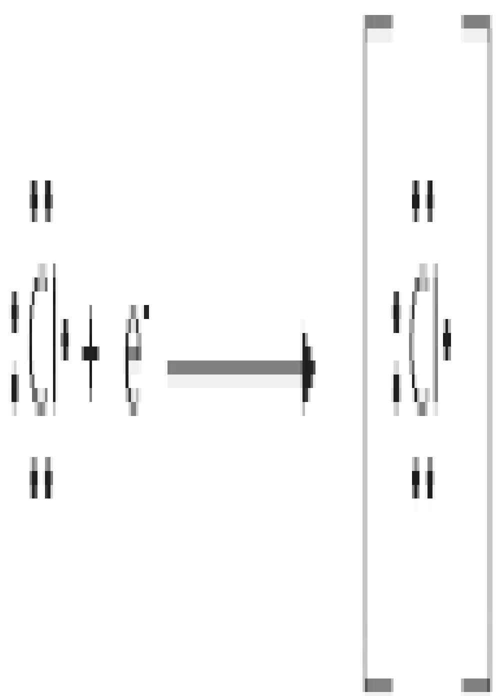
Ionic bond is developed when one or more electrons get completely transferred from one atom to other. This type of bond is generally found in inorganic compounds.

Example: Formation of Sodium Chloride (NaCl).

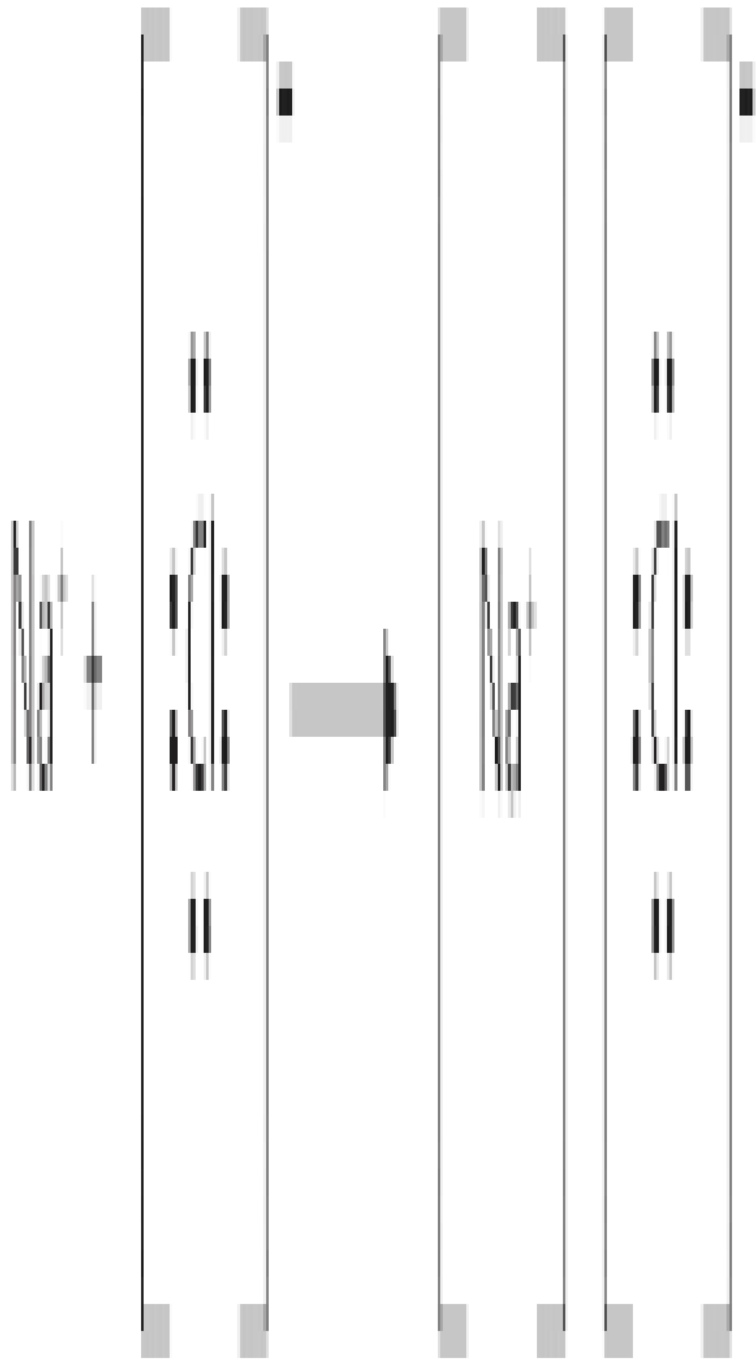
The atomic number of sodium is 11 and its electronic configuration is $1s^2$, $2s^2$, $2p^6$, $3s^1$ (2, 8, 1) and the atomic number of chlorine is 17 and its electronic configuration is $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^2$ (2, 8, 7).

The sodium has one electron in excess while chlorine is one electron short of acquiring the stable configuration.





Sodium by transferring one electron to chlorine gets the positive charge and chlorine by getting one electrons gets negative charge.



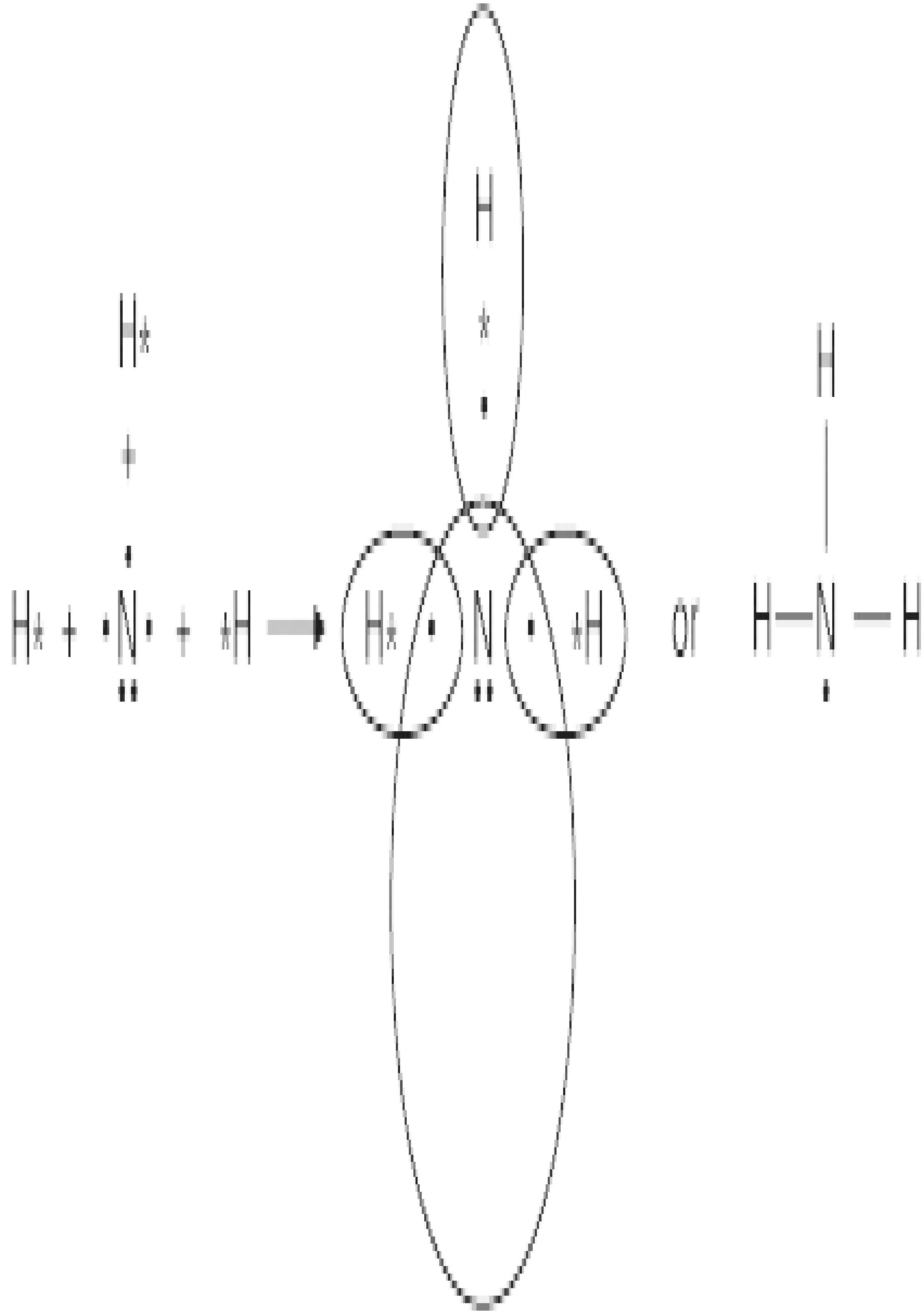
The bond formed in this way is called ionic or electrovalent bond.

Covalent Bond

Covalent bond is formed when a pair of electrons is mutually shared between two atoms. Both the atoms contribute to the shared electrons. This bond is indicated by a line called single bond (−).

Example: Formation of Ammonia (NH3).

The nitrogen is having atomic number as 7 and its electronic configuration is $1s^2$, $2s^2$, $2p^3$ and that of hydrogen is 1 and its electronic configuration is $1s^1$.


or

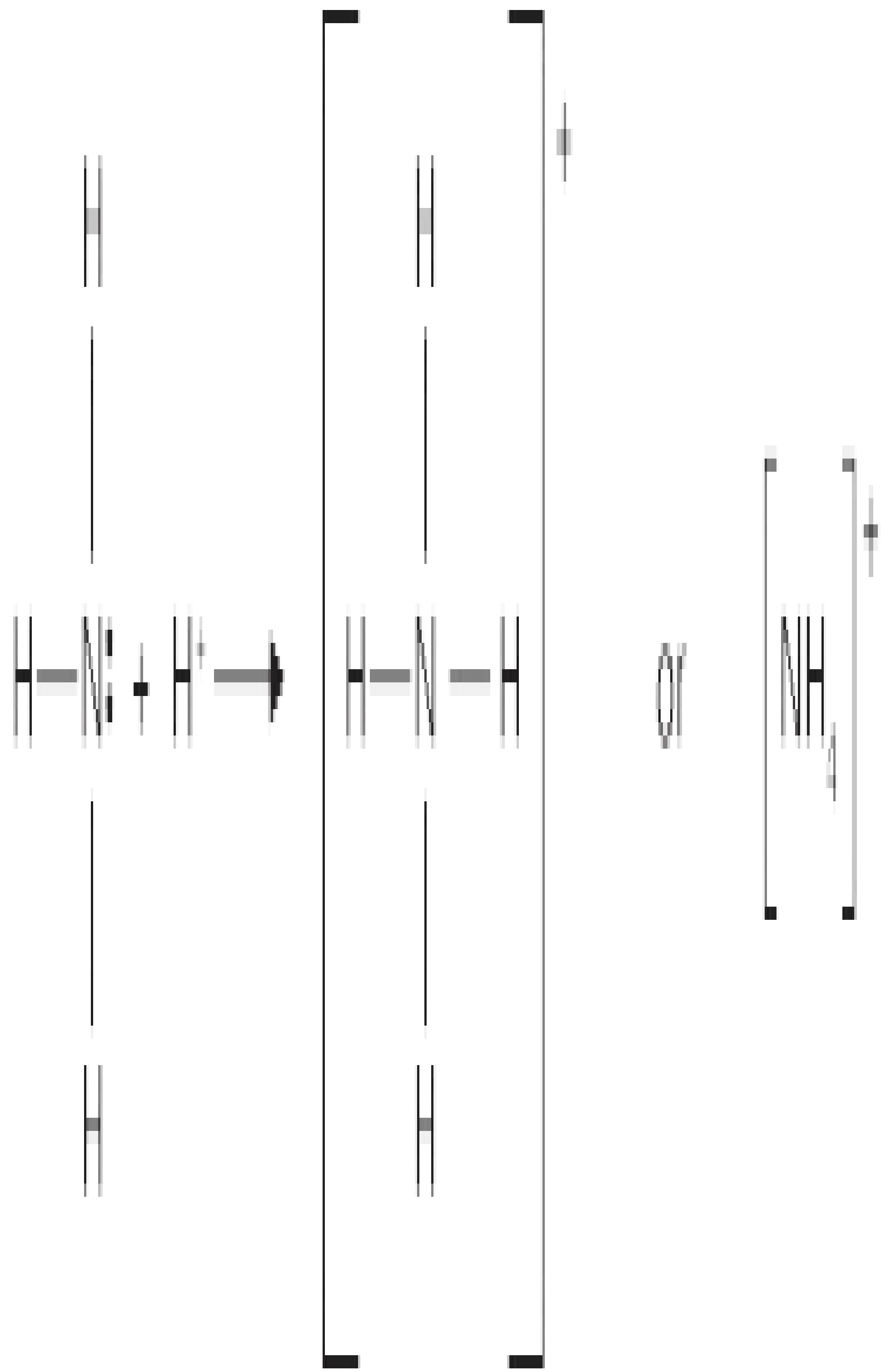
In order to acquire stable electronic configuration, nitrogen shares its 3 electrons with electrons of 3 atoms of hydrogen and thus ammonia is formed with three covalent bonds.

Co-ordinate Covalent Bond

In formation of co-ordinate covalent bond both the transfer as well as sharing of electrons takes place and the shared pair of electrons is supplied by one atom only while the other atom simply takes part in it. This bond is formed when a pair electron is contributed by one atom only and the sharing is done by both the combining atoms. The atom which supplies the shared pair of electrons is called the donor and the atom which accepts this pair is called the acceptor. The bond is indicated by an arrow mark (→). The arrow direction points to the acceptor atom.

Example: Formation of Ammonium ion $[NH_4]^+$.

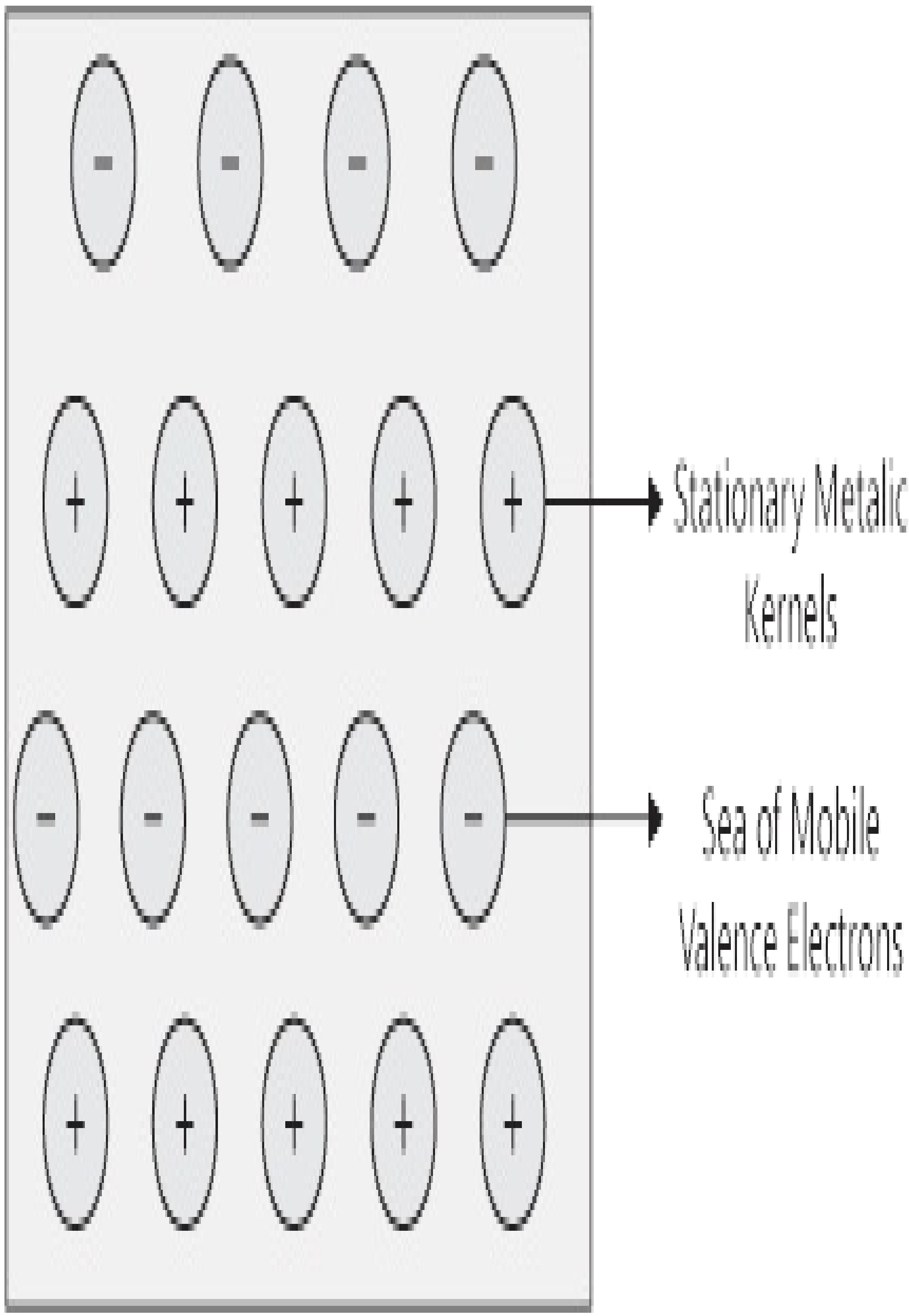
When Ammonia (NH_3) reacts with Hydrogen ion (H^+), nitrogen donates its unshared pair of electrons to hydrogen ion which is in need of two electrons to acquire stable configuration. Thus, the sharing of pair of electrons takes place between nitrogen atom and hydrogen ion. The donor atom which contributes its pair of electrons becomes positive and the acceptor atom becomes negative.



Metallic Bond

Metallic bond is defined as the force that binds the atoms without the valence electrons called the kernels to number of electrons within its reference sphere. In every metal, the valence electrons can move freely in the vacant valence orbits present around the nucleus. In simple words, the valence electrons move from one kernel to another kernel in the metallic crystals which results in number of stationary positively charged kernels in the metallic crystal.

An electrostatic force of attraction is developed between positively charged kernels and the negatively charged mobile valence electrons. This force of attraction that exists between the metal atoms is known as metallic bond.


Stationary Metalic Kernels
Sea of Mobile Valence Electrons

Metallic Bond

Lewis Theory of Bonding

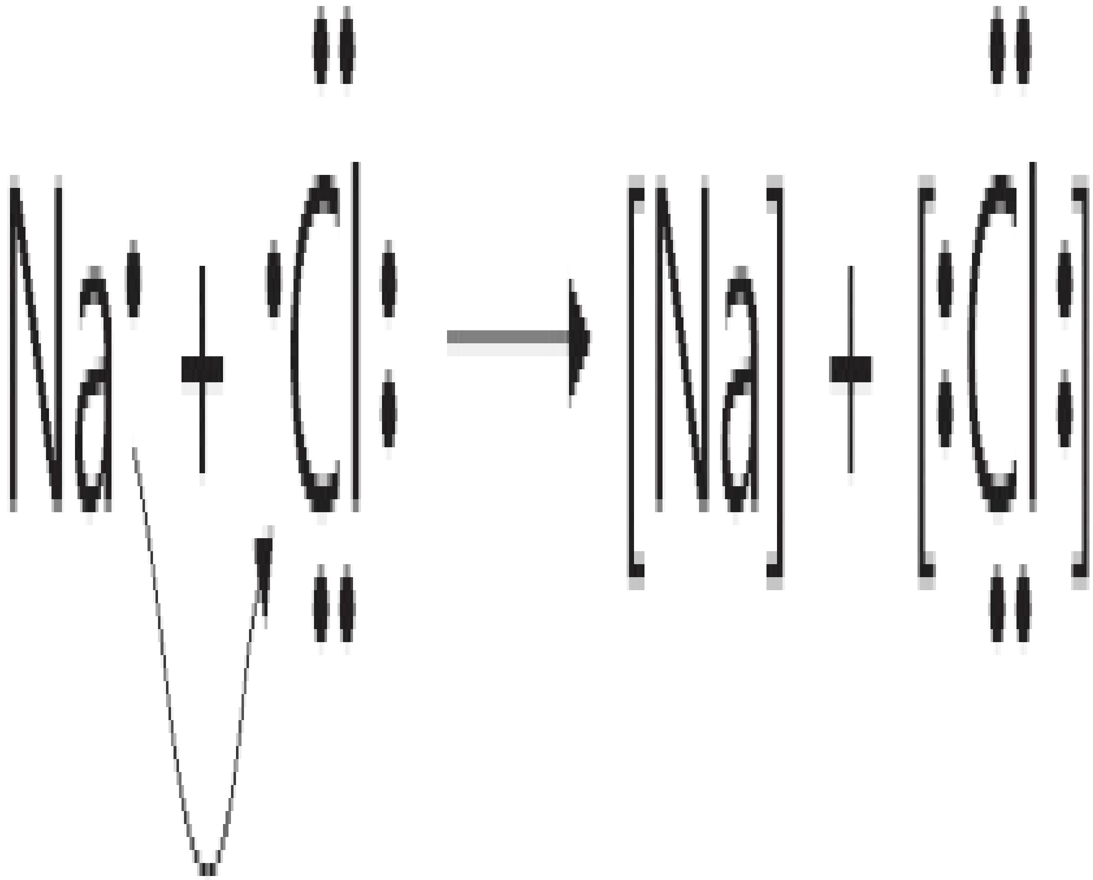
According to this theory electrons are most stable when they are paired. Atoms form chemical bonds to achieve a full valence integument of electrons. This may be achieved by two ways:

According to this theory an exchange of electrons between metal and non-metal atoms.

Split electrons between non-metal atoms.

Lewis Biagrams

The chemical symbol shows the nucleus and core electrons. Dots around the symbol show the valence electrons.

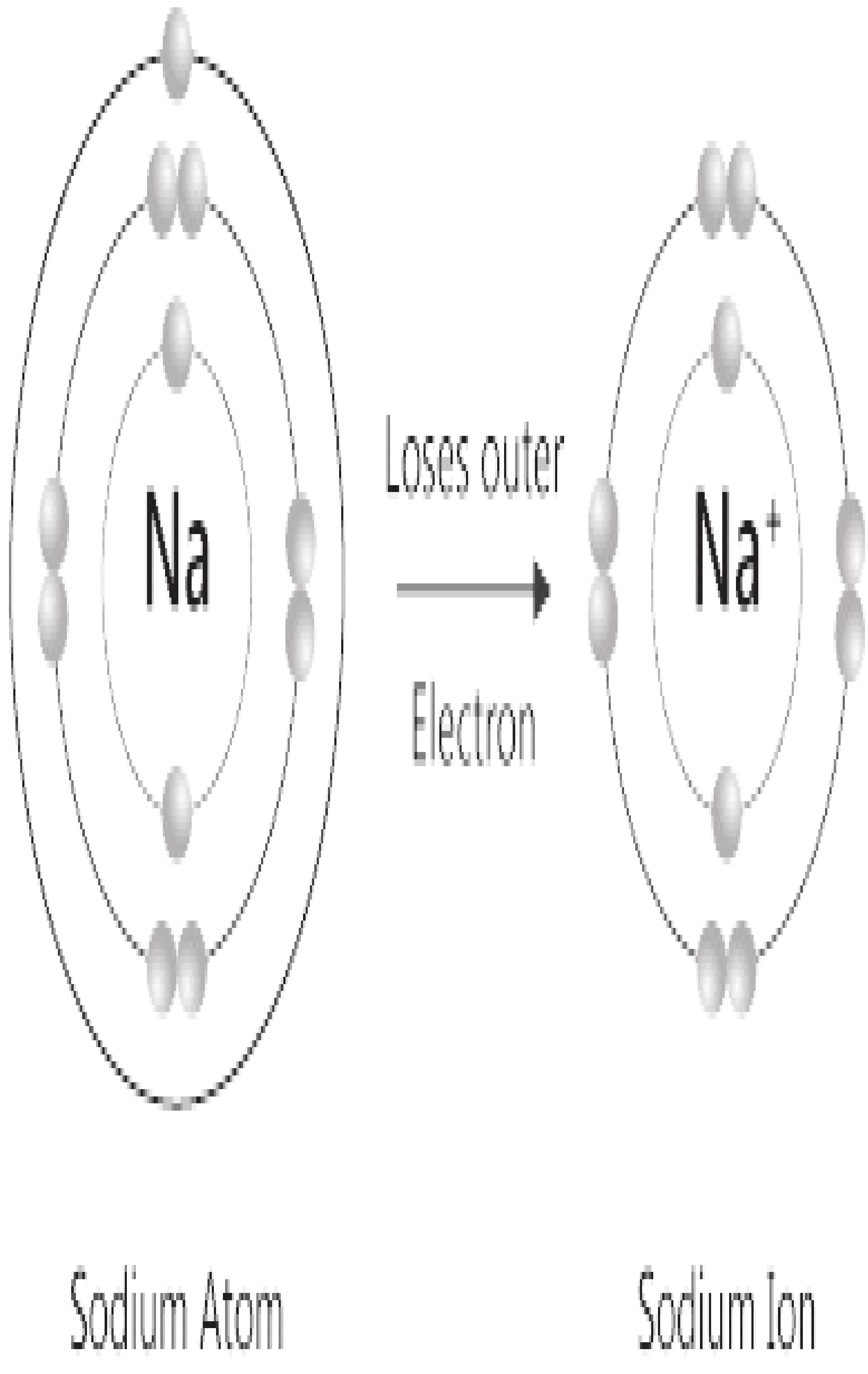


Electron transfer from
sodium to chlorine

Lewis diagram for ions

The Octet Rule

Atoms tend to gain, lose, or allocation electrons until they are isoelectronic with a noble gas (have the selfsame number of electrons as a noble gas). Octet rule is examination that many atoms move form the most stable substances when they are adjoining by eight electrons in their valence shell.


Na
Loses outer
Electron
Na+
Sodium Atom
Sodium Ion

Octet rule

Bond Lengths

Covalent bond lengths depend on two items.

1. Atomic radius of the atoms convoluted in the bond.

Example:

d(C – Cl) = 177 pm r(Cl) = 100 pm

r(C) = 77 pm

d(C – H) = 109 pm r(H) = 37 pm

d(C – F) = 133 pm r(F) = 72 pm

d(C – Br) = 194 pm r(Br) = 114 pm

$d(C – H) = 109$ pm $r(H) = 37$ pm

$d(C – I) = 213$ pm $r(I) = 133$ pm

2. Number of bonds between atoms.

Example:

$d(N = N) = 122$ pm $d(C = C) = 134$ pm

$d(N – N) = 146$ pm $d(C – C) = 154$ pm

$d(N \equiv N) = 110$ pm $d(C \equiv C) = 121$ pm

Normally, the shorter bond length, higher bond dissociation energy.

Example:

$d(C – C) = 153$ pm $E_{dis} = 348$ kJ/mol

$d(C = C) = 134$ pm $E_{dis} = 613$ kJ/mol

$d(C \equiv C) = 120\ pm\ Edis = 839\ kJ/mol$

6

Quantum Numbers

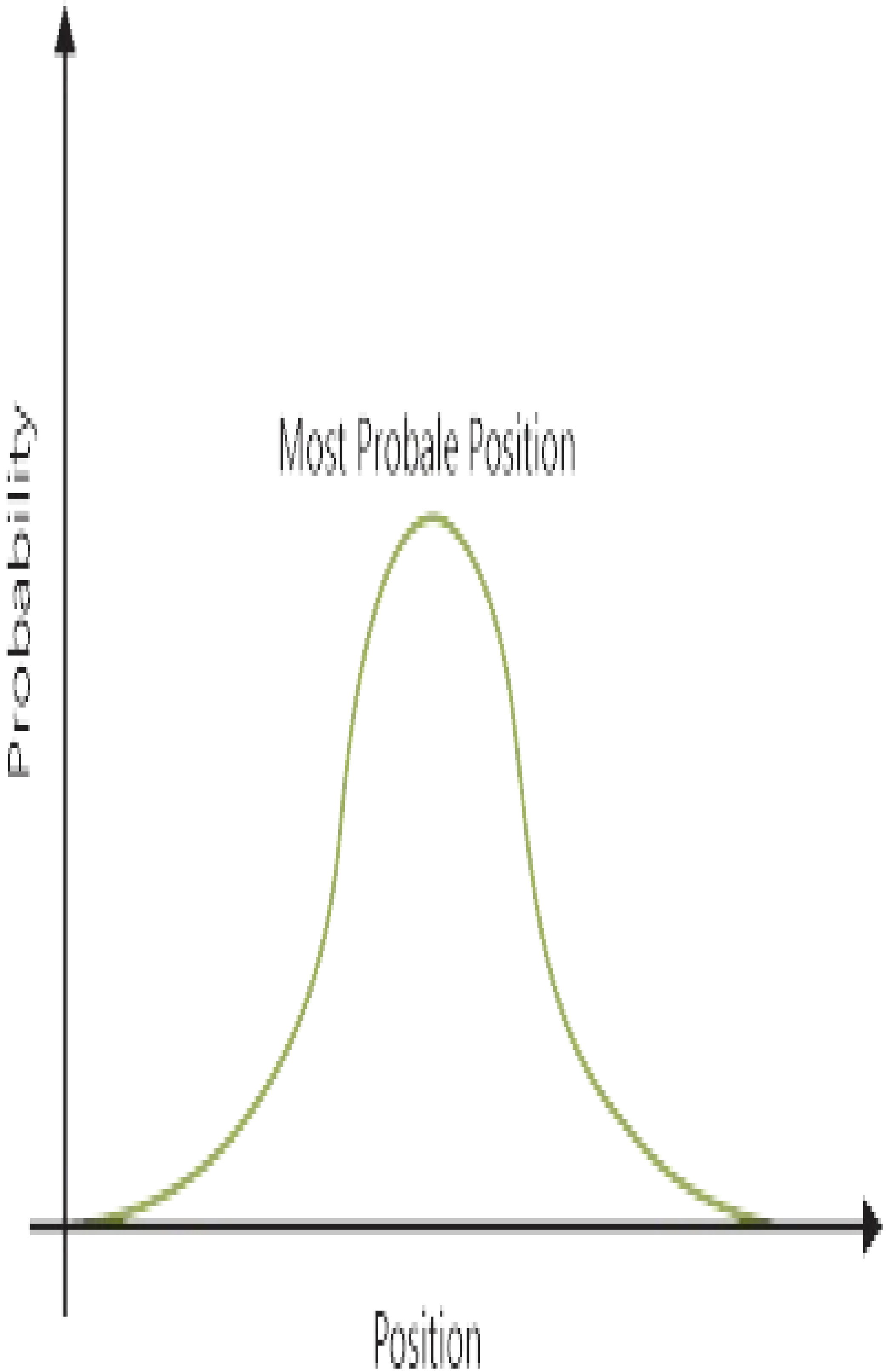
Quantum

A

quantum of energy is the amount of energy need to move an electron from one energy level to another. The energy levels are like the rungs of a ladder but are not same spaced.

Quantum Mechanics

Erwin Schrödinger obtains a complex mathematical formula to integrate the wave and particle properties of electrons. In quantum mechanics wave behavior is described with the wave function ψ. The prospect of finding an electron in a certain region of space is proportional to ψ^2 and is called electron density.


Most Probale Position
Probability
Position

Quantum wave functions

Schrödinger equation prescribes possible energy states an electron can occupy in a hydrogen atom. The energy plight and wave functions are described by a set of quantum numbers. Instead of referring to orbits as in the Bohr model, quantum numbers and wave functions illustrate atomic orbital.

Quantum Numbers

Bohr model of quantum number was a one-dimensional model that used one quantum number to shows the distribution of electrons in the atom. Only data that was important was the size of the orbit, which was represent by the n quantum number. Schrodinger's model of quantum number allowed the electron to occupy three-dimensional space. It therefore necessary three coordinates, or three quantum numbers, to describe the orbital in which electrons can be found. There are three types of quantum numbers:

Principal (n)

Angular (l) and

Magnetic (m) quantum numbers

Quantum numbers represent the size, shape, and orientation in space of the orbital on an atom.

Principal Quantum Number (n)

Principal quantum number shows by (n). Principal quantum number describes the size of the orbital. Principal quantum number indicates main energy levels n = 1, 2, 3, 4. In this quantum number each main energy level has sub-levels. The principal quantum number therefore indirectly represents the energy of an orbital. In Principal quantum number the maximum number of electrons in a principal energy level is given by-

Max # electrons = 2n2

Where n= the principal quantum number.

Angular Quantum Number (l)

This angular quantum number (l) represents the shape of the orbital. Orbital have structure that are best represents as convex (l = 0), polar (l = 1), or cloverleaf (l = 2). They can even grip on more complicated structure as the value of the angular quantum number turn larger.

There is only one technique in which a convex (l = 0) can be oriented in expanse. Orbital's that have polar (l= 1) or cloverleaf (l = 2) structure, although, can point in different directions. Therefore we need a third quantum number, known as the magnetic quantum number (m).

Magnetic Quantum Number (ml)

Magnetic quantum number (m) to represents the orientation in area of a particular orbital. It is called the magnetic quantum number because the effect of different alignment of orbital was first observed in the presence of a magnetic field. The values of ml are integers that depend on the value of the angular moment quantum number: l,…0,…+l

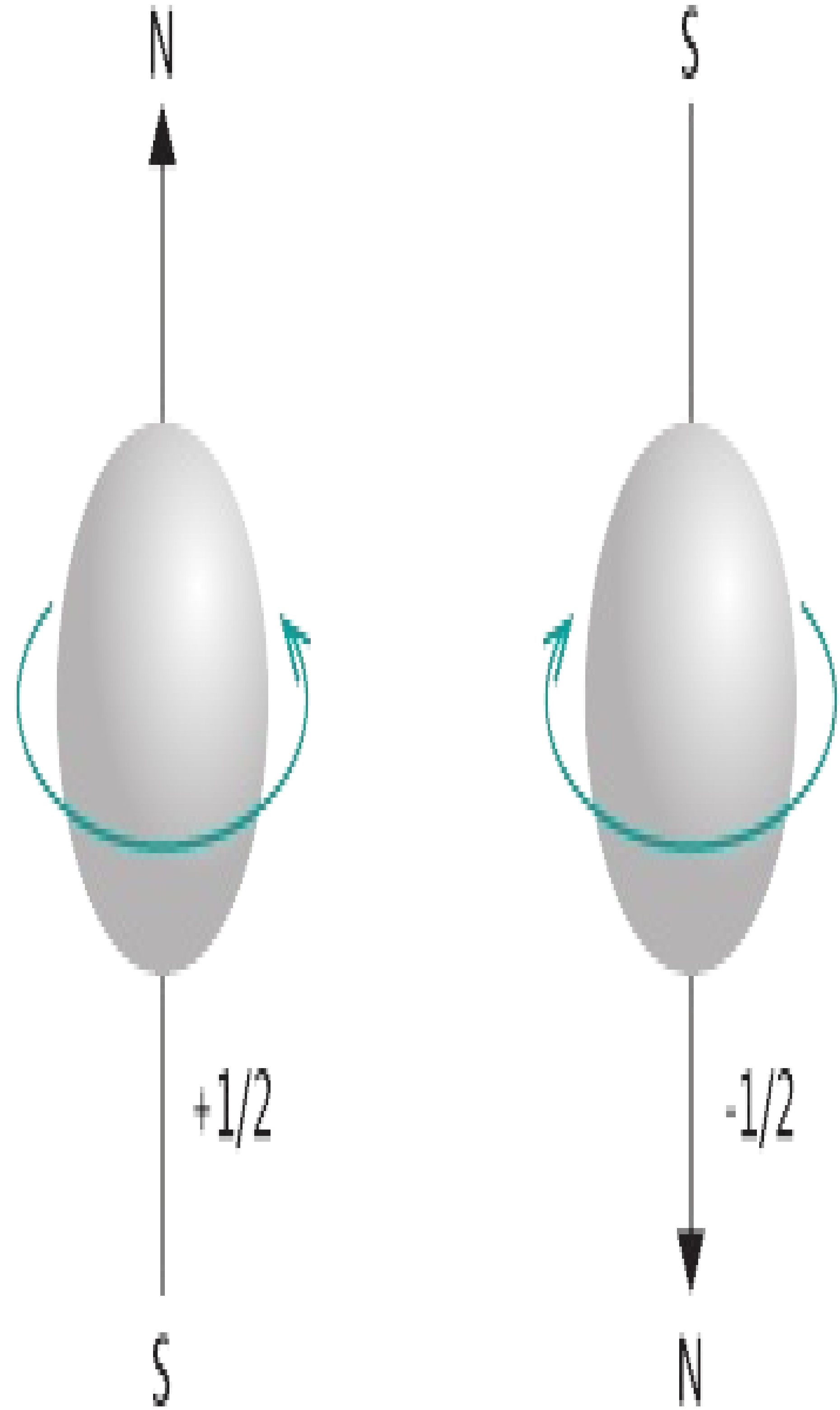
When n is	l can be	When l is	ml can be
1	Only 0	0	Only 0
2	0 or 1	0 1	Only 0 -1,0, or +1
3	0, 1 or 2	0 1 2	Only 0 -1,0, or +1 -2, -1,0, +1, or +2
4	0,1,2 or 3	0 1 2 3	Only 0 -1,0, or +1 -2, -1,0, +1, or +2 -3, -2

▪

Allowed values of the quantum numbers n, l and m

Electron Spins Quantum Number (ms)

The electron revolve quantum number (ms) is used to specify an electron’s revolve. There are two possible directions of revolve. In this quantum number (ms) allowed values of ms are +½ and −½.


N
S
+1/2
-1/2
S
N

Electrons spin quantum number

A beam of atoms is break by a magnetic field. Statistically, half of the electrons revolve clockwise, the other half revolve counterclockwise.

Rules Governing Allowed Combinations of Quantum Numbers

The three quantum numbers (n, l, and m) that represents an orbital are integers: 0, 1, 2, 3, and So on. The principal quantum number (n) cannot be nil. Permit values of n are therefore 1, 2, 3, 4, and so on. In this quantum number (I) is between any integer 0 and n - 1. If n = 3, for example, l can be either 0, 1, or 2. In this quantum number (m) between any integer -l and +l. If l = 2, m can be 2, -1, 0, +1, or +2.

Shells and Sub-shells of Orbitals

Quantum numbers designate shells, sub shells, and orbitals. Orbital those have the same face value of the principal quantum number form a shell. Orbital within a shell are split into sub shells that have the same value of the angular quantum number. Chemists represent the shell and sub shell in that case an orbital associate with a two-character code such as 2p or 4f. The first character shows the shell (n = 2 or n = 4).

And the second character recognizes the sub shell. By convention, the following lowercase letters are used to indicate different sub shells.

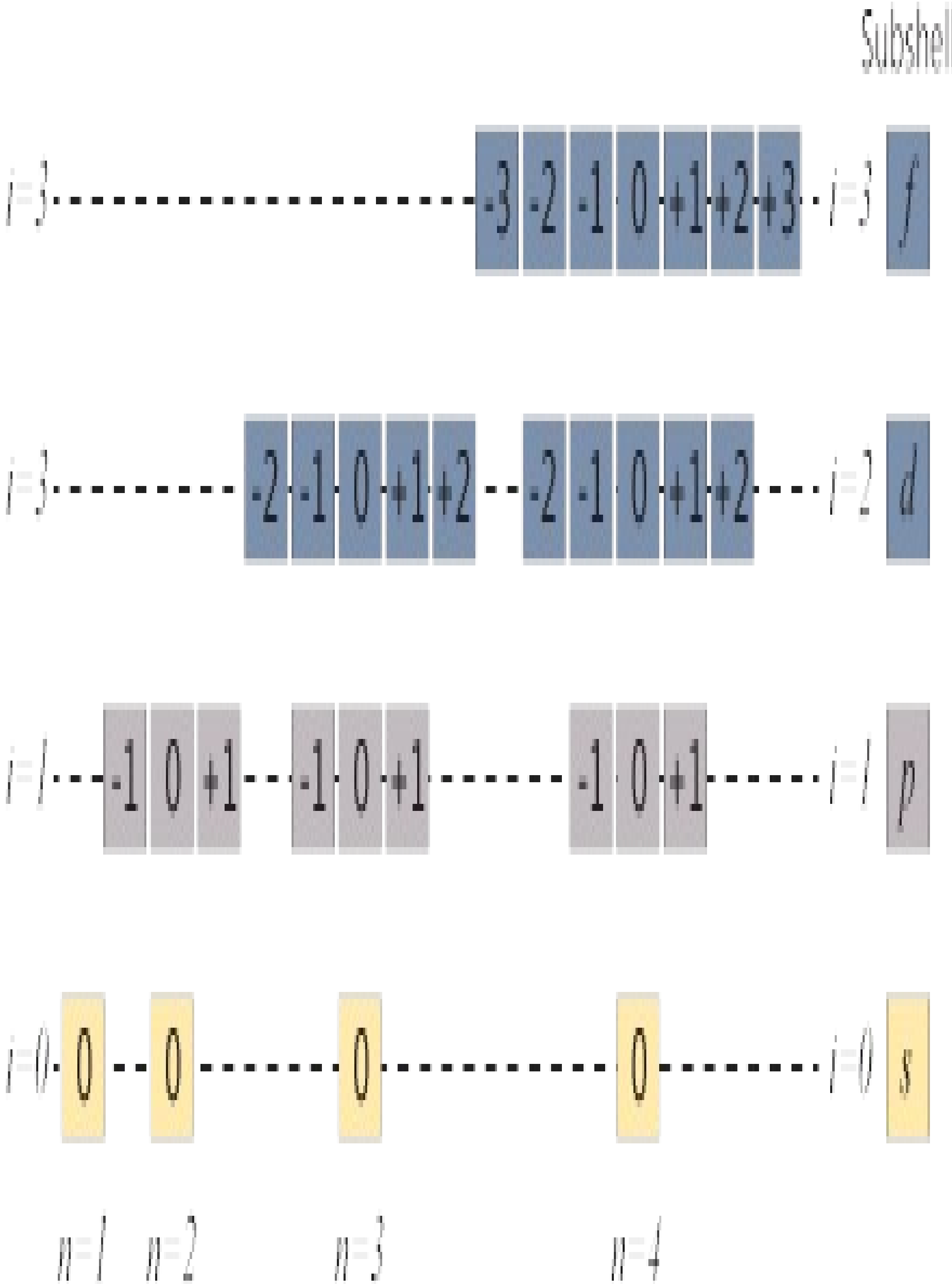
s: $l = 0$

p: $l = 1$

d: $l = 2$

f: $l = 3$

There is no blueprint in the first four letters (s, p, d, f). The letters progression alphabetically from that point (g, h, and so on). Some of the allowed composition of the n and l quantum numbers is shown in the figure below.



Subshell
-3 -2 -1 0 +1 +2 +3
f
-2 -1 0 +1 +2
-2 -1 0 +1 +2
d
-1 0 +1
-1 0 +1
-1 0 +1
p
0
0
0
0
s
n=1
n=2
n=3
n=4

Quantum Numbers Designate Shells, Sub-shells and Orbitals

The third rule of this limiting allowed combinations of the n, l, and m quantum numbers has a prime consequence. It intensity the number of sub shells in a shell to be equal to the principal quantum number for the shell. The n = 3 shell, for example, hold three sub shells: the 3s, 3p, and 3d orbital.

Equations & Reactions

Introduction

A

chemical reaction shows a chemical change. For example, one chemical characteristics of hydrogen is that it will act with oxygen to make water. We can represent this chemical change more succinctly as

$$\textit{Hydrogen} + \textit{oxygen} \rightarrow \textit{water}$$

The + sign in above reactions means that the two medium interact chemically with each other and the → symbol implicit that a chemical reaction takes place. But medium can also be shows by chemical formulas. Remembering that hydrogen and oxygen both prevail as diatomic molecules, we can rewrite our chemical change as: H2 + O2→H2O.

This is the example of a chemical equation-1, which is an incisive way of representing a chemical reaction. The initial medium is called reactants2, and the final medium is called products3.

A chemical change or chemical reaction is an activity in which one or more pure substances are converted into one or more dissimilar pure substances.

Chemical changes lead to the evolution of substances that help expand our food, make our existence more productive, rehabilitate our heartburn, and much, much more.

For example, nitric acid, HNO_3, which is used to form fertilizers and explosives? Is formed in the chemical reaction of the ammonia (NH_3), and oxygen (O_2). Silicon dioxide (SiO_2), gasses acts with carbon (C), at high temperature to yield silicon (Si)—which can be used for to make computers—and carbon monoxide (CO). An antacid tablet might carry calcium carbonate, $CaCO_3$, which merge with the hydrochloric acid in your stomach to yield calcium chloride, $CaCl_2$, water, and carbon dioxide (CO_2).

Represent Chemical Equations

$$NH_3 (g) + 2O_2 (g) \rightarrow HNO_3 (aq) + H_2O (l)$$

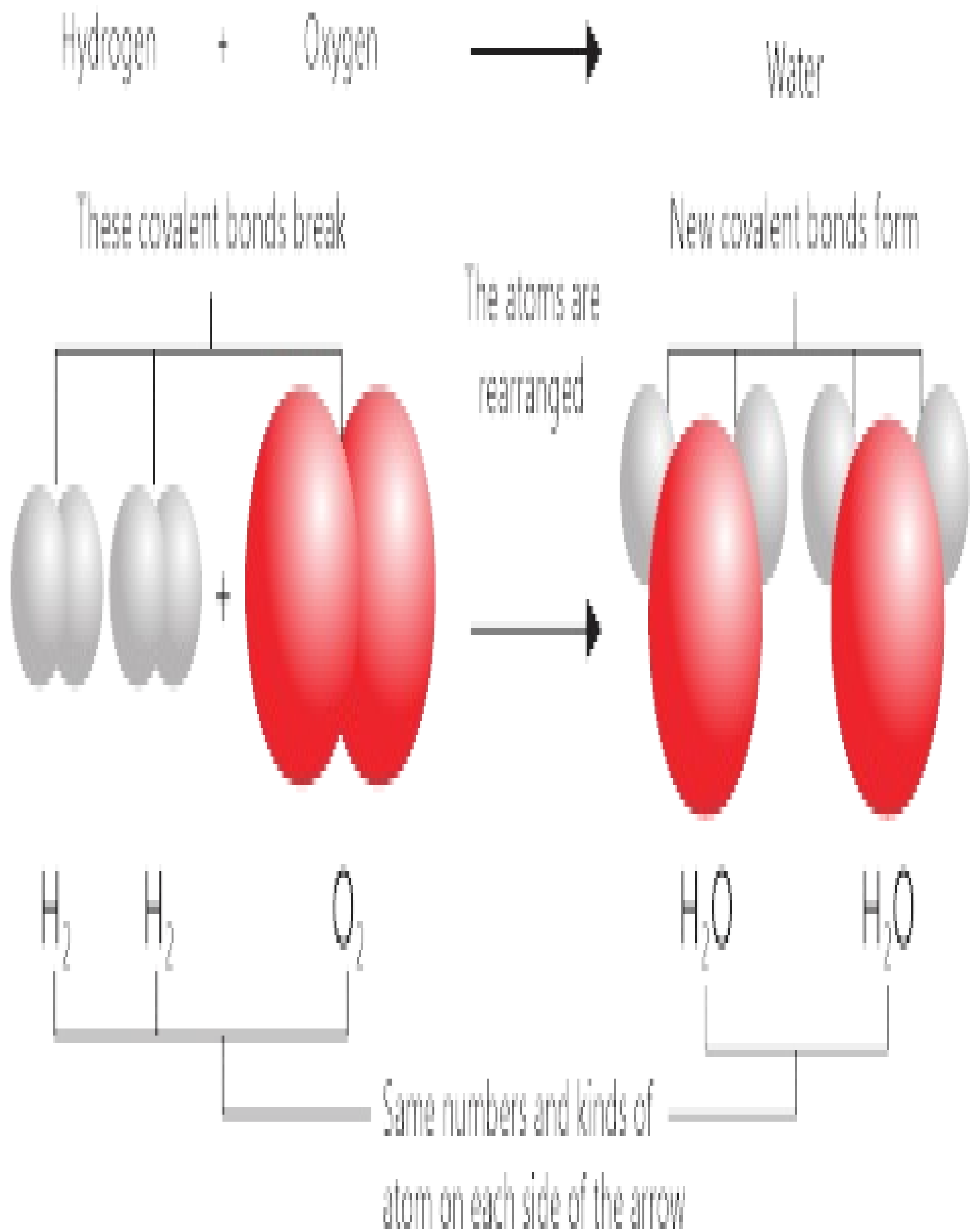
$$SiO_2(s) + 2C(s)\ Si (l) + 2CO (g)$$

$$CaCO_3(s) + 2HCl (aq) \rightarrow CaCl_2 (aq) + H_2O (l) + CO_2 (g)$$

Interpreting Chemical Equation

In these chemical reactions, atoms are reposition and regrouped through the breaking and making of chemical bonds.

For example, when hydrogen gas, H_2 (g), is burned in the appearance of gaseous oxygen, O_2 (g), a new substance, liquid water, H_2O (l), forms. The covalent bonds inside the H_2 molecules and O_2 molecules split, and new covalent bonds form between oxygen atoms and hydrogen atoms.


Hydrogen
+
Oxygen
Water
These covalent bonds break
New covalent bonds form
The atoms are rearranged
+
H_2
H_2
O_2
H_2O
H_2O
Same numbers and kinds of atom on each side of the arrow

The formation of water from hydrogen and oxygen

This chemical equation is a shorthand representation of a chemical reaction. The following equation shows the burning of hydrogen gas to form liquid water.

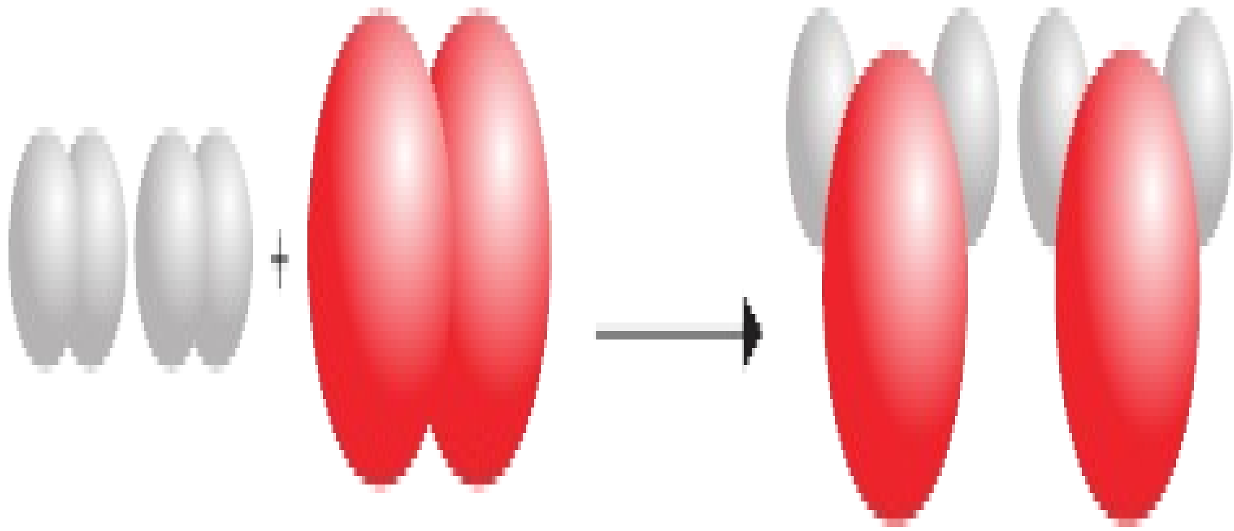
$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$$

Chemical Equations Give the Following Information about Chemical Reactions

Chemical equations describe the formulas for the substances that take part in the reaction. The formulas on the left side of the arrow show the reactants, the substances that change in the reaction. The formulas on the right side of the arrow show the products, the substances that are formed in the reaction. If there is more than one reactant they are separated by plus signs.

The arrow sign separating the reactants from the products can be read as “goes to” or “yields” or “produces.” The physical position of the reactants and products are provided in the equation. Gases are represents by (G). Solids are represents by (s). Liquids are represents by (l). When a substance is dissolved in water, it is represents with (aq) for aqueous, which means “mixed with water.” The respective numbers of particles of each reactant and product are indicated by numbers placed in front of the formulas. These respective numbers are called coefficients. These equation hold correct coefficients are called a balanced equation.

For example, the 2’s in fore-end of H2 and H2O in the equation we saw the above coefficients. If this formula in a balanced equation has no declare coefficient, these coefficient is understood to be 1, as is the case for oxygen O2 in the equation above.



Coefficient   Coefficient

'Goes to'

$$2H_2(g) + O_2(g) \xrightarrow{\text{'Goes to'}} 2H_2O(l)$$

Physical states   Physical states

Chemical equation for the formation of water from hydrogen and oxygen

If special conditions are mandatory for a reaction to take place, they are often defining above the arrow. Some examples of unusual conditions are electric current, high temperature, high pressure, and light.

A special condition

$$2H_2O(l) \xrightarrow{\text{Electric current}} 2H_2(g) + O_2(g)$$

To indicate that a chemical reaction needs the continuous addition of heat in order to proceed, we place an upper case Greek delta, Δ, in the above equations. For example, the transformation of potassium chlorate (a fertilizer and food additive) to potassium chloride and oxygen needs the continuous addition of heat:

This indicates continous addition of heat

$$2KClO_3(s) \xrightarrow{\Delta} 2KCl(s) + 3O_2(g)$$

Balancing Chemical Equations

In chemical reactions, atoms are neither generate nor destroyed; they merely change partners. Thus the number of atoms of an insider in this reaction's products is equivalent to the number of atoms of that element in the main reactants. The coefficients we usually place in for-end of one or more of the formulas in a chemical equation return this fact. They are used whenever compulsory to balance the number of atoms of a particular element on either side of the arrow in the equations.

For example, let's return to the reaction of hydrogen gas and oxygen gas to make liquid water. In these equation for the reaction betwixt $H_2(g)$ and $O_2(g)$ to form $H_2O(l)$ represents there are two atoms of oxygen in the diatomic O_2 molecule to the left arrow, so there should also be 2 atoms of oxygen in the product to the right of the arrow. Because each water molecule, H_2O, carry only 1 oxygen atom, 2 water molecules must form for each oxygen molecule that acts.

In this equation Coefficient 2 in front of the H_2O (l) make this clear. But these 2 water molecules carry 4 hydrogen atoms, which mean that 2 hydrogen molecules must be present on the reactant side of the equation for the numbers of H atoms to balance it.

$$H_2\ (g) + O_2\ (g) \rightarrow H_2O_2\ (l)$$

Where water is H_2O, whereas H_2O_2 is hydrogen peroxide, it is very different substance from water.

Example for Balancing Equations

Balance the following equation so that it exactly shows the reaction for the formation of dinitrogen oxide, this is an anesthetic used in dentistry and surgery.

$$NH_3 (g) + O_2 (g) \rightarrow N_2O (g) + H_2O (l)$$

Solution

The following table represents that the atoms are not balanced yet.

Element	Left	Right
N	1	2
H	3	2
O	2	2

▪

Table

Nitrogen is the first member in the first formula. It is establish in 1 formula on each side of the arrow, so we can try to balance it now. Two nitrogen atoms on the right side of the equation and only one on the left; we convoy them into equilibrium by placing a 2 in front of NH_3.

$$2NH_3\ (g) + O_2\ (g) \rightarrow N_2O\ (g) + H_2O\ (l)$$

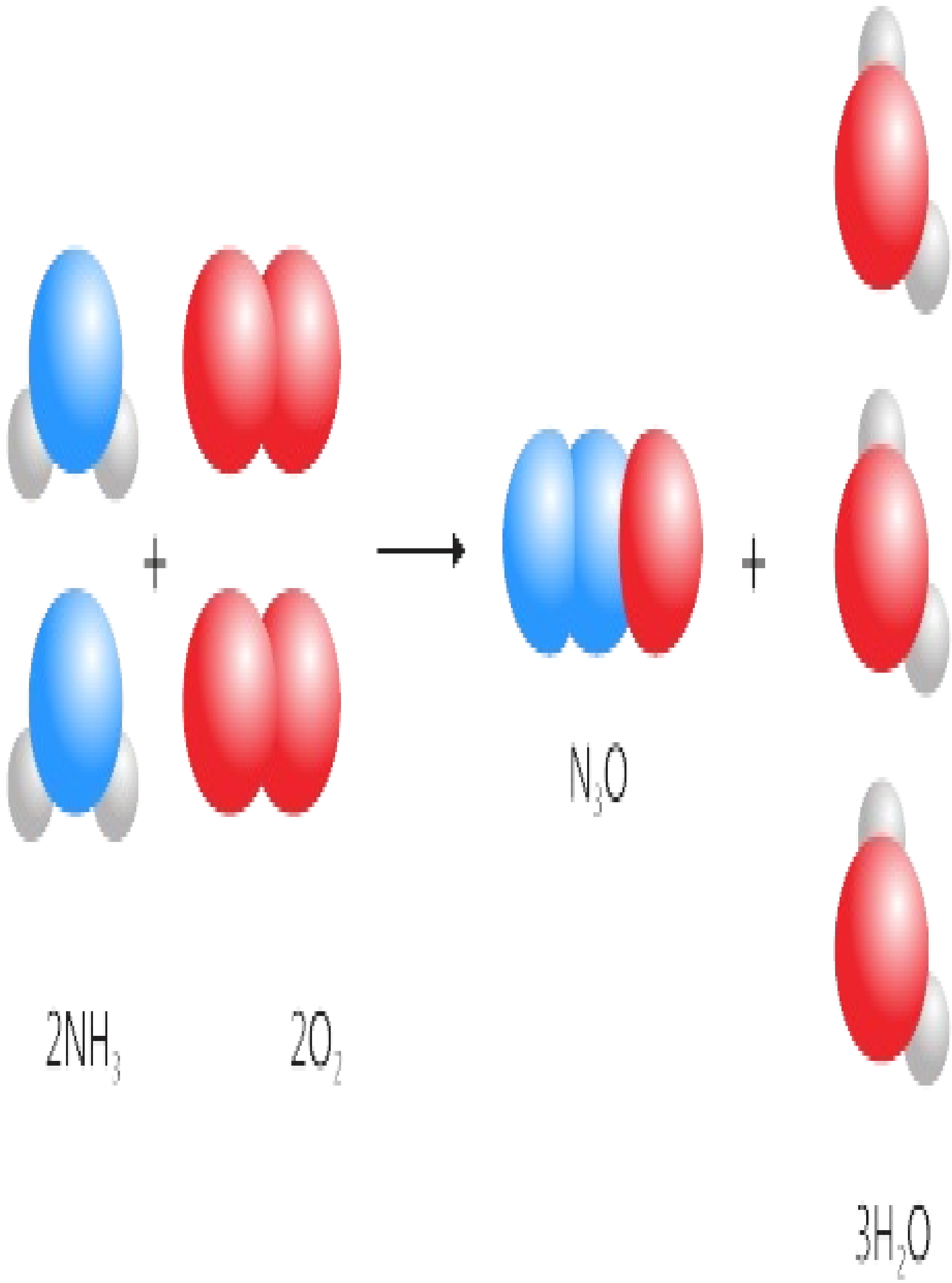
There are now 6 hydrogen atoms on the left side of the arrow and only 2 H's on the right, so we can balance hydrogen atoms by arranging a 3 in front of the H_2O. This gives 6 atoms of hydrogen on all side.

$$2NH_3\ (g) + O_2\ (g) \rightarrow N_2O\ (g) + 3H_2O\ (l)$$

There are now 2 oxygen atoms on the left and 4 on the right, so we balance the oxygen atoms by placing a 2 in front of the O_2.

$$2NH_3\ (g) + 2O_2\ (g) \rightarrow N_2O\ (g) + 3H_2O\ (l)$$

The following space filling models represent how you might envision the relative number of particles participating in that reaction. You can see that the atoms regroup but are neither created nor destroyed.



$2NH_3$
$2O_2$
N_2O
$3H_2O$

Space filling model

The following table represents that the atoms are now balanced.

-

Element	Left	Right
N	2	2
H	6	6
O	4	4

Table

Orbital Filling

Atomic Orbital

L

ike the Bohr model, the energy level in quantum mechanics represents locations where you are likely to find an electron. Orbitals structure seems to be graphically structure around the nucleus where electrons are found. Quantum mechanics determine the probabilities where you are “likely” to discover electrons.

So scientists admit to limit these calculations to locations where there was at least a 90% chance of discover an electron. Think of orbitals as sort of a "border” for spaces throughout the nucleus inside which electrons are permitted. No more than two electrons can ever be in one orbital. The orbital just defines a “region” where you can find an electron.

Energy Levels

In quantum mechanics has a principal quantum number. It is represented by a small “n”. It shows the “Energy Level” similar to the Bohr’s model. Where n=1 represents the first energy level and n=2 represents the second energy level etc. Every energy level represents a period or row on the periodic table.

Sub-Levels is Equal to Specific Atomic Orbital

There are mainly four types of atomic orbitals s, p, d and f. Each of these sub-levels shows the blocks on the periodic table. n = 3 mean three represents 3 sub-levels or s, p and d.

Every energy level has “sub-levels” which shows the specific “atomic orbitals” for that level.

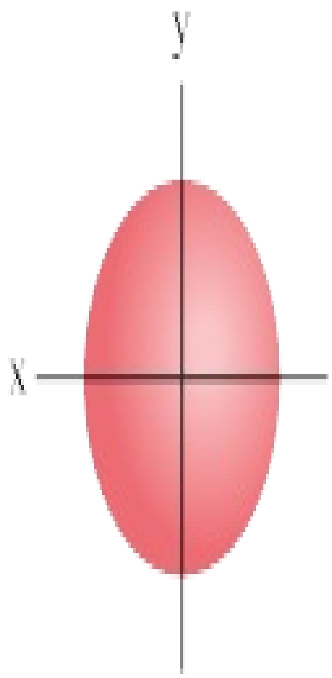
n = 1 represents 1 sub-level or s orbital.

n = 2 represents 2 sub-levels or s and p.

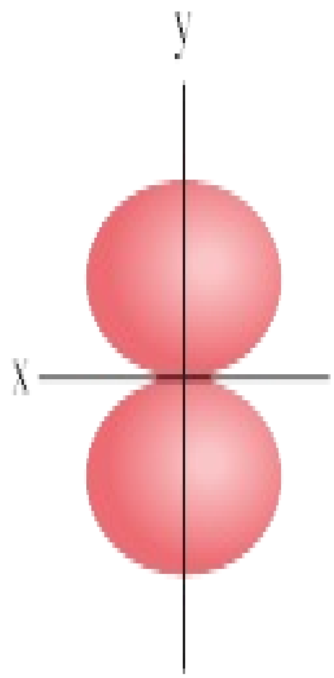
This n = 3 shows three sub-levels (Specific Atomic Orbitals) that is s, p and d.

n = 4, means four sub-levels (Specific Atomic Orbital) that is s and p, d and f.

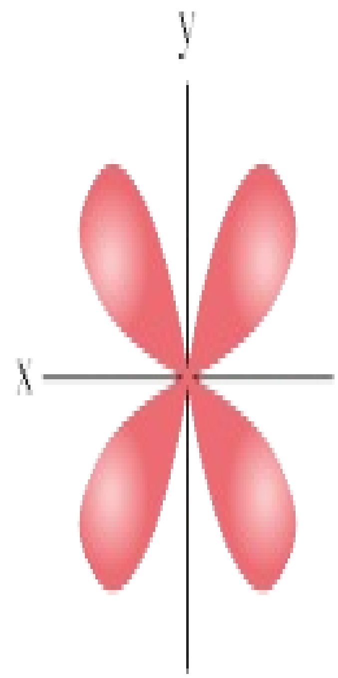
Orbital



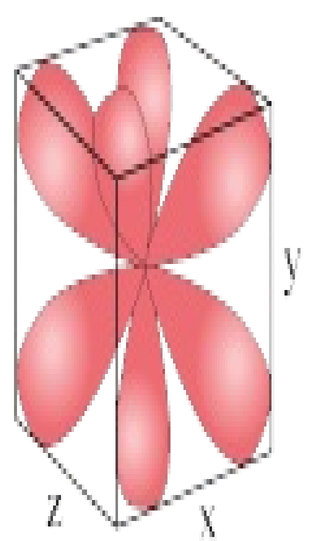
s orbital



p orbital



d orbital



f orbital

In the s block, electrons are proceeding into s orbitals. In the p block, the s orbitals are full new electrons are preceding into the p orbitals. In the d block, the s and p orbitals are full. New electrons are proceeding into the d orbital.

Energy level	Sub levels	Total orbital
n=1	s	1 (1s orbital)
n=2	S p	1 (2s orbital) 3 (2p orbitals)
n = 3	s p d	1 (3s orbital) 3 (3p orbitals) 5 (3d orbitals)
n = 4	s p d f	1 (4s orbital) 3 (4p orbitals) 5 (4d orbitals) 7 (4f orb

Orbital Table

Filling Order Rules of Orbital

There are three main rules.

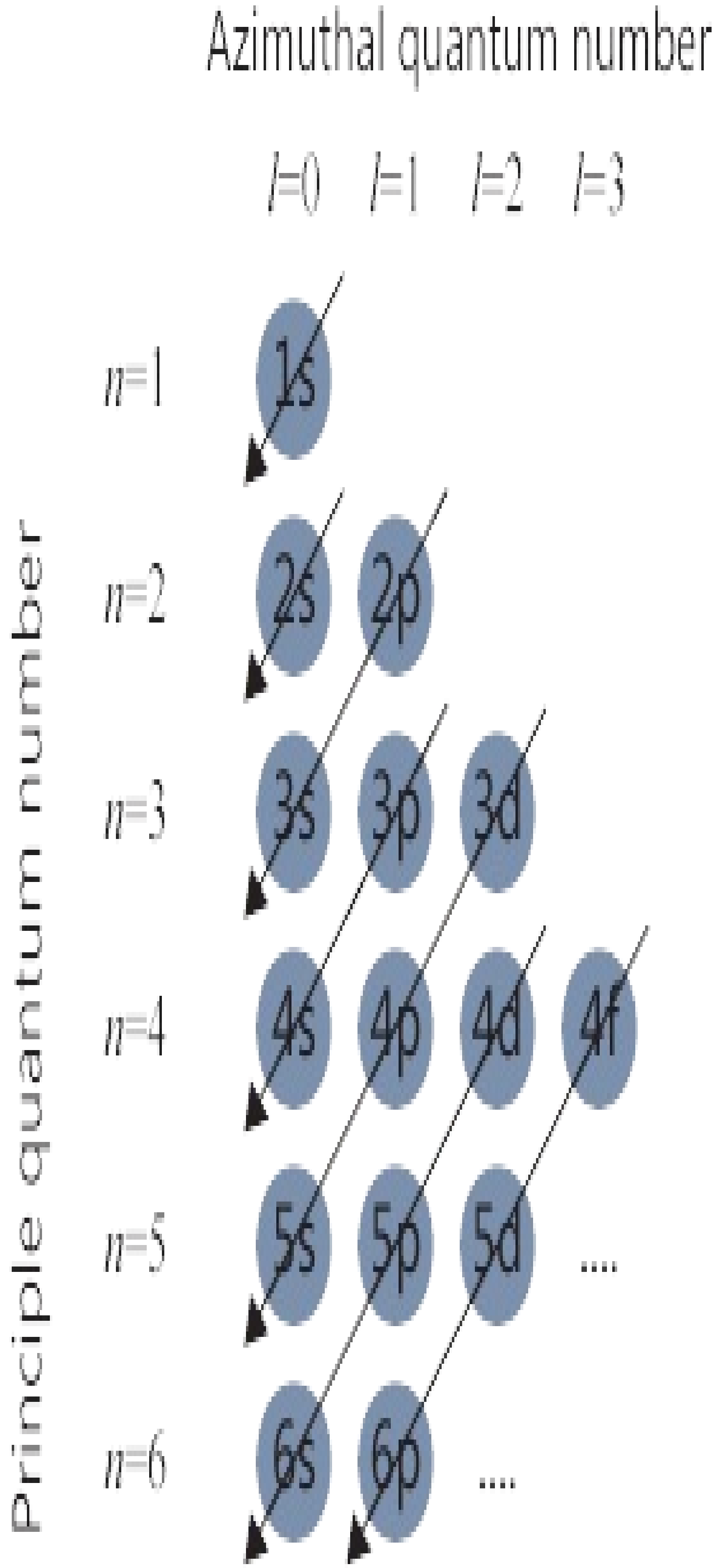
The Aufbau Principle.

The Pauli Principle.

Hund’s Rule.

Aufbau Principle

Aufbau principle represents the order in which energy sublevels fill with electrons. The sublevels in yellow are unwanted because there are no atoms with sufficient electrons to fill these levels. But, it doesn’t injure to draw them in if it’s easier to remember.


Azimuthal quantum number
l=0 l=1 l=2 l=3
Principle quantum number
n=1
n=2
n=3
n=4
n=5
n=6
1s
2s 2p
3s 3p 3d
4s 4p 4d 4f
5s 5p 5d ....
6s 6p ....

Aufbau principle

To use this diagram, follow the arrows, and recall the maximum number of electrons every type of sublevel can hold: s2, p6, d10, and f14.

Pauli Exclusion Principle

Pauli principle describe that a single orbital, regardless of expanse or shape, may contain no more than two electrons and those electrons must have opposite spin. It is view that opposing spins decrease the repulsion between the electrons, which are both negatively charged. Electrons are like spinning magnets, and must have opposite alignment. We “imagine” this as one “↑” and one “↓”. Every single orbital takes 2 electrons only.

There is one s orbital = 2 electrons.

There are three p orbitals = 6 electrons.

There are five d orbitals = 10 electrons.

There are seven f orbitals = 14 electrons.

Hund’s Rule

Hund’s Rule describes that when filling a pair of equal energy orbitals, every orbital gets one electron before any orbital gets 2.

Examples:

When multiple orbitals are available, each orbital gets filled separately at first, and pairing begins.

The single electrons all have the uniform alignment.

If there are multiple orbitals at the uniform energy, they fill separately first, before electrons pair.

Representing Electrons

There are three types of electron diagrams.

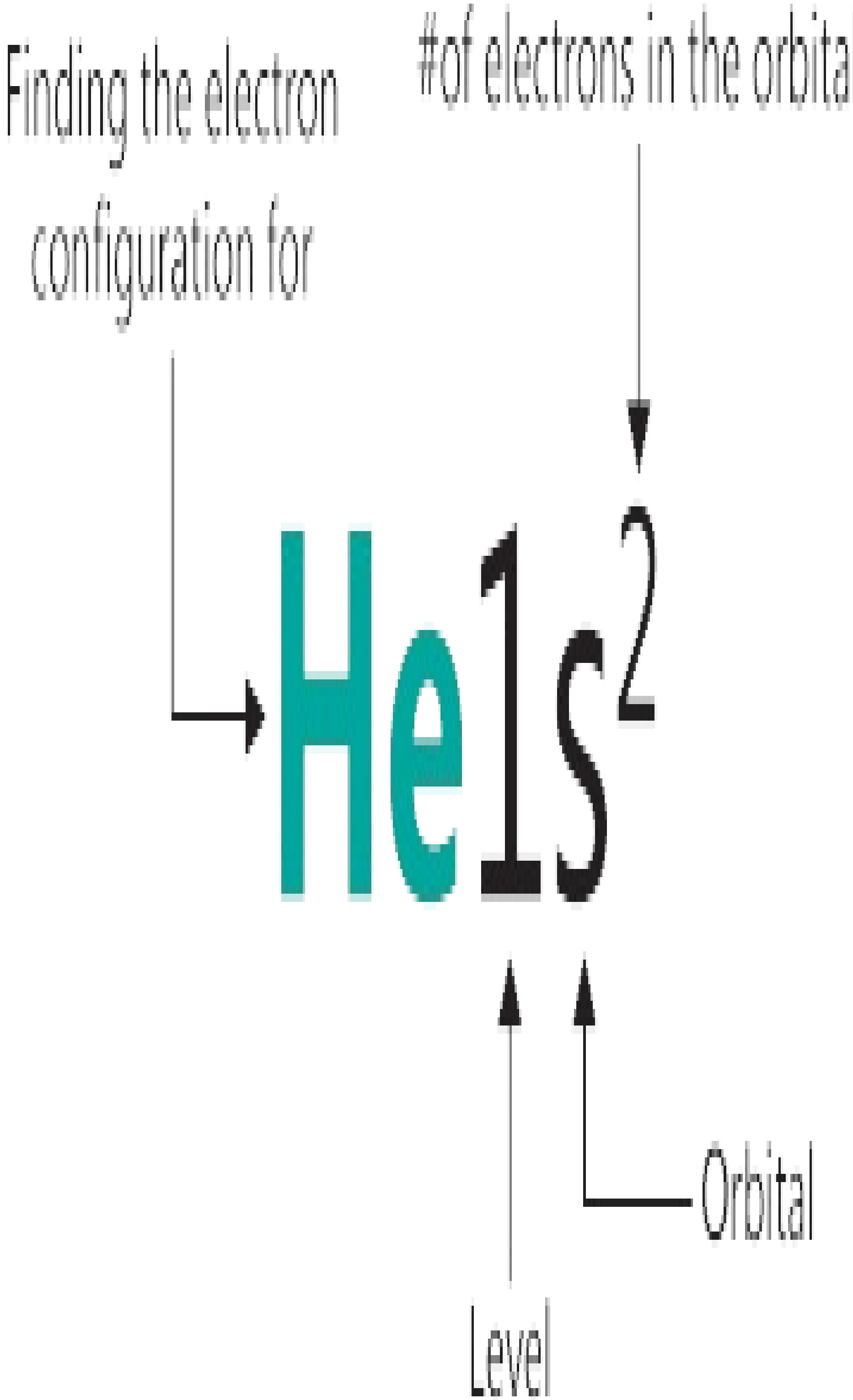
Electron configuration diagrams.

Orbital filling diagram, and

A Lewis or electron dot diagram.

Electron configuration diagram

This electron configuration diagram represents PEL, sublevel, and how many electrons are in every sublevel.


Finding the electron configuration for
#of electrons in the orbital
He1s2
Orbital
Level

Electron configuration diagram

Orbital Filling Diagram

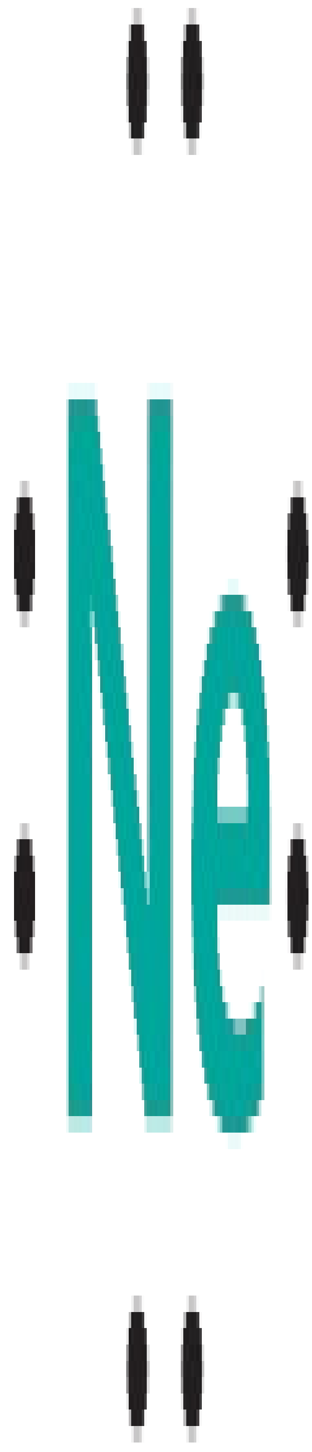
Orbital filling diagram uses a box or line to shows each atomic orbital. Underneath each box or line is the orbital designation. Arrows pointer is used to show 1 or 2 electrons and their spin.

2S

Orbital filling diagram

Lewis or Electron Dot Diagram

Lewis or electron dot diagram represents a series of 1-8 dots in pairs throughout the element symbol. Only the electrons in the highest numbered PEL are appearing. This is called valence or outer shell electrons.


Ne

Lewis or electron dot diagram

Filling in Electrons

n = 1 → 1s orbital → 2 electrons

n = 2 → 2s orbital → 2 electrons

→ 2p orbital→ 6 electrons

Low energy rows have few orbitals available; therefore there are few elements there.

n = 3 → 3s orbital → 2 electrons+

→ 3p orbital →6 electrons

→3d orbital →10 electrons

Beyond n = 4, all levels have s, p, d and f orbitals.

n = 4 → 4s orbital → 2 electrons

→ 4p orbital → 6 electrons

→4d orbital →10 electrons

→4f orbital → 14 electrons

Solution

Introduction

When two or more substances are mixed in different proportions to form a homogeneous mixture, the mixture is called a solution. There are two main components of a solution:

Solute – The substance that is dissolved in the solvent to get a homogeneous mixture is called a solute. It is present in solution in small quantity and forms a minor part of solution.

Solvent – The substance in which solute is dissolved to get a homogeneous mixture is called a solvent. It is present in solution in large quantity and forms a major part of solution.

Thus, a solution can be said as a homogeneous mixture of a solute and solvent.

Measuring the Concentration of Solution

Different units are used for measuring a solution's concentration:

Percentage by mass

Normality

Molarity

Molality

Mole-fraction

Percentage by Mass

The number of grams of a solute which are present in 100 gram of solution is called percentage by mass.

Normality

The number of gram equivalents of a solute which are present in 1 litre or 1000 ml of solution is called the normality. It is denoted by alphabet N.

$$\text{Normality (N)} = \frac{\text{Mass of solute x 1000}}{\text{Equivalent mass of solute x Volume of solution}}$$

Molarity

The number of moles of a solute which are present in 1 liter or 1000 ml of solution is called the molarity. It is denoted by alphabet M. A solution containing 1 mole of solute in 1 liter solution is called a molar solution.

$$\text{Molarity (M)} = \frac{\text{Mass of solute} \times 1000}{\text{Molecular mass of solute} \times \text{Volume of solution}}$$

Molality

The number of moles of a solute which are present in 1 kg or 1000 gm of solvent is called the molality. It is denoted by alphabet m. A solution containing 1 mole of solute in 1 kg or 1000 gm of solvent is called a molal solution.

$$\text{Molality (m)} = \frac{\text{Mass of solute} \times 1000}{\text{Molecular mass of solute} \times \text{Mass of solvent}}$$

Mole-fraction

Mole-fraction of solvent

It is the ratio of number of moles present in a solvent to total number of moles present in the solution (solute + solvent).

$$\text{Mole fraction of solvent} = \frac{\text{Number of moles in solvent}}{\text{Total moles present in the solution}}$$

Mole-fraction of solute

It is the ratio of number of moles present in a solute to total number of moles present in the solution (solute + solvent).

$$\text{Mole fraction of solute} = \frac{\text{Number of moles in solute}}{\text{Total moles present in the solution}}$$

10

Acid & Bases

Introduction

A

cids and bases are commonly found solutions that can exist everywhere. Almost every type of liquid that we see in our daily lives consists of properties which are acidic or basic except water. They possess completely different properties and can be neutralized to form H2O. Solutions are classified into acids and bases depending upon the type of ions produced in the solution. The acidic and basic characteristics or properties of matter are essential for the study of chemical reactions.

Acids

An acid is a substance that gives out hydrogen ions (H+) in an aqueous solution.

$$HCl \rightleftharpoons H^+ + Cl^-$$

$$HNO_3 \rightleftharpoons H^+ + NO_3^-$$

Bases

A base is a substance that gives out hydroxide ions (OH^-) in an aqueous solution.

$$NaOH \rightleftharpoons Na^+ + OH^-$$

$$KOH \rightleftharpoons K^+ + OH^-$$

Theories of Acids and Bases

There are certain theories based on which the acidic and basic nature of the matter or solution can be easily explained. They are: Arrhenius theory, Lowry-Bronsted theory and Lewis theory.

Arrhenius theory

This theory was proposed in 1827. According to this theory:

An acid is a substance that gives out hydrogen ions (H+) in an aqueous solution.

$$HCl \rightleftharpoons H^+ + Cl^-$$

$$H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-}$$

A base is a substance that gives out hydroxide ions (OH^-) in an aqueous solution.

$$NaOH \rightleftharpoons Na^{+} + OH^{-}$$

$$Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^{-}$$

Acidic and basic strengths of solutions depend upon the extent upto which the solution gives H^+ or OH^- during process of ionization.

Limitations

This theory is valid only for aqueous solutions. Basic properties of compounds which are non-hydroxyl like NH_3 and Na_2CO_3 is difficult to explain.

Lowry-Bronsted Theory

This theory is based on protonic concept. According to this theory:

An acid is one which donates proton (H^+) to other substances. Hence, it is considered as a proton donor. A base is one which accepts proton (H^+) from other substances. Hence, it is considered as a proton acceptor.

$$\text{Acid} \rightleftharpoons H^+ + B^-$$

Proton Donor

Proton Acceptor

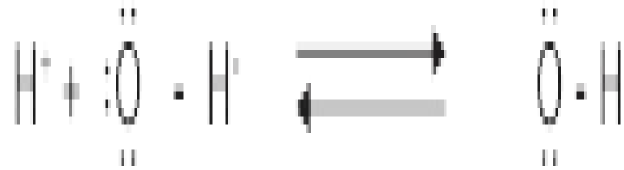
Limitations

It explains acidic and basic nature of a substance only in presence of a solvent like water. Neutralization of acidic oxides by basic oxides cannot be explained.

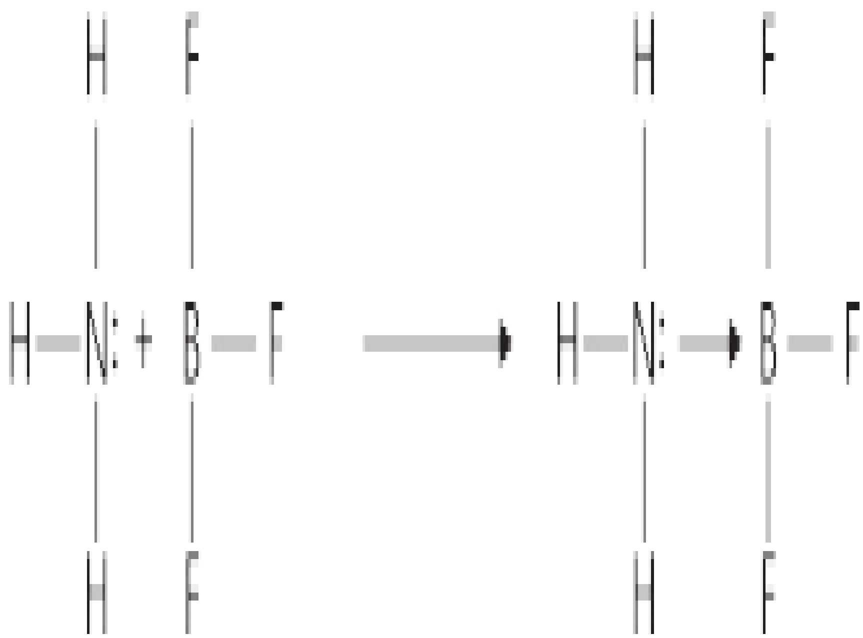
Lewis Theory

It is also known as electronic concept as it involves the transfer of electrons during process of base formation. As per this theory:

An acid is one which accepts the pair of electrons. Hence, it is considered as electron pair acceptor. A base is one which donates the pair of electrons. Hence, it is considered as electron pair donor.



Acid Base



Base Acid

In the above reaction, OH- and NH3 are bases as they donate a pair of electrons and H+ and BF3 are acids as they accept a pair of electrons.

Advantages

It explains the acidic character of substances which do not contain hydrogen. The acidic properties and nature of substances like BF, FeCl, AlCl etc can be explained by this theory.

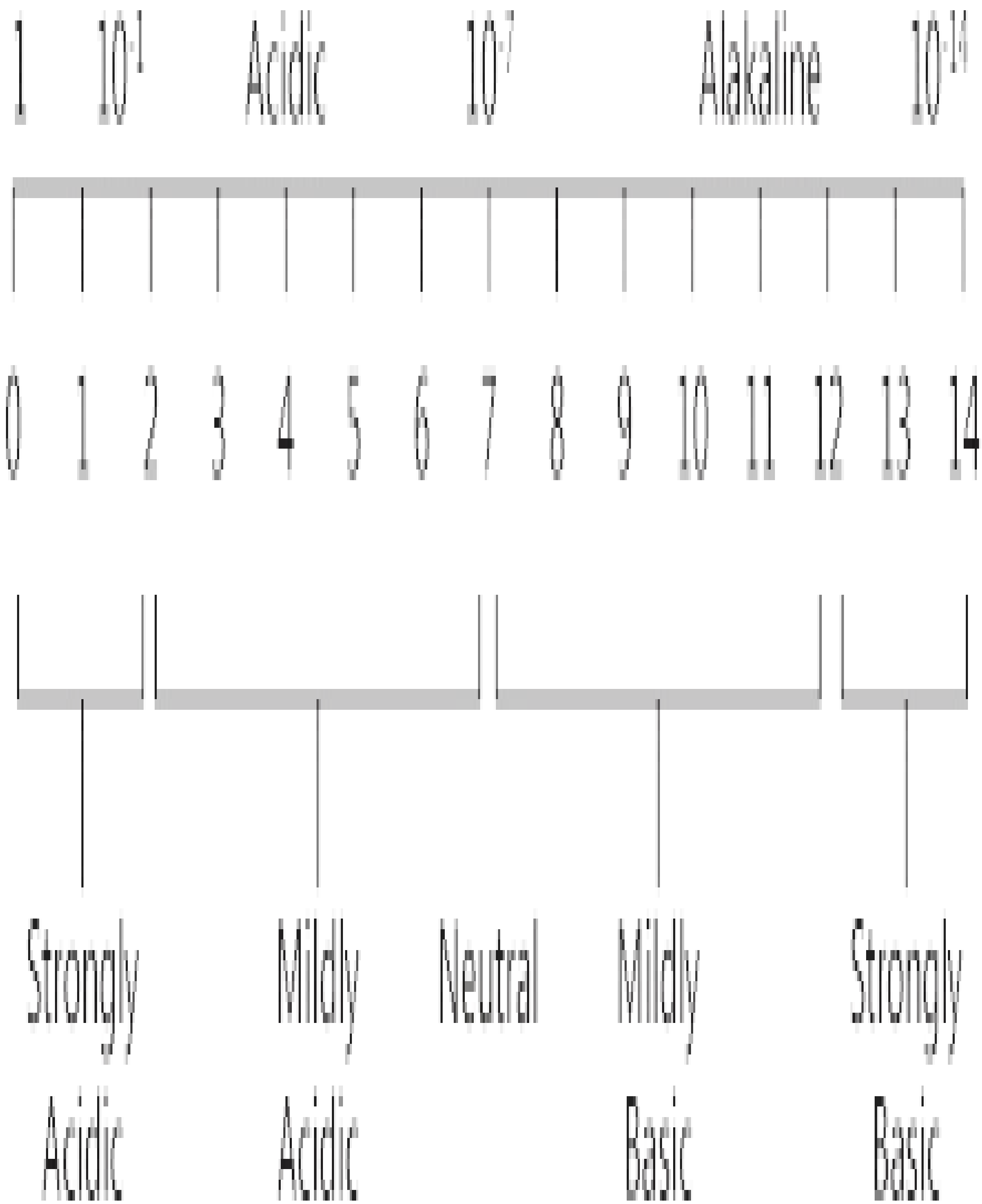
The neutralization of acidic and basic oxides can also be explained by this theory.

pH Value

pH value is used to express the acidity or concentration of hydrogen ion (H+) in a solution. pH is expressed as the negative logarithm to the base 10 of concentration of hydrogen ion (H+) in a solution.

$$pH = - \log 10 [H+]$$

For higher concentration of H+, the pH value will be lower.


1
10^{-1}
Acidic
10^{-7}
Alakaline
10^{-14}
0 1 2 3 4 5 6 7 8 9 10 11 12 13 14
Strongly Acidic
Mildly Acidic
Neutral
Mildly Basic
Strongly Basic

pH Scale

Buffer Solution

Buffer solution is a solution which is used to maintain a constant pH even after adding small amount of acid or alkali to the solution. Buffer solution is of two types:

Acidic Buffer – It is obtained by mixing a weak acid with its salt and maintains pH value between 0 to 7.

CH3COOH (Acetic acid) + CH3COONa (Sodium acetate)

Basic Buffer - It is obtained by mixing a weak base with its salt and maintains pH value between 7 to 14.

NH4OH (Ammonium hydroxide) + NH4Cl (Ammonium chloride)

Electronic Oxidation and Reduction

Electronic oxidation is a process which involves the removal of electrons from the elements.

$$\mathrm{HCL} \xrightarrow{-1e^-} \mathrm{Na^+}$$

$$\mathrm{K} \xrightarrow{-1e^-} \mathrm{K^+}$$

$$\mathrm{H} \xrightarrow{-1e^-} \mathrm{H^+}$$

Na, K and H are getting oxidized into Na^+, K^+, H^+ respectively.

Electronic reduction is a process which involves the addition of electrons to the elements.

$$Cl \xrightarrow{+1e^-} Cl^-$$

$$F \xrightarrow{+1e^-} F^-$$

$$Br \xrightarrow{+1e^-} Br^-$$

Cl, F and Br are getting reduced to Cl-, F-, Br- respectively.

11

Intermolecular Forces

Intermolecular Forces

A

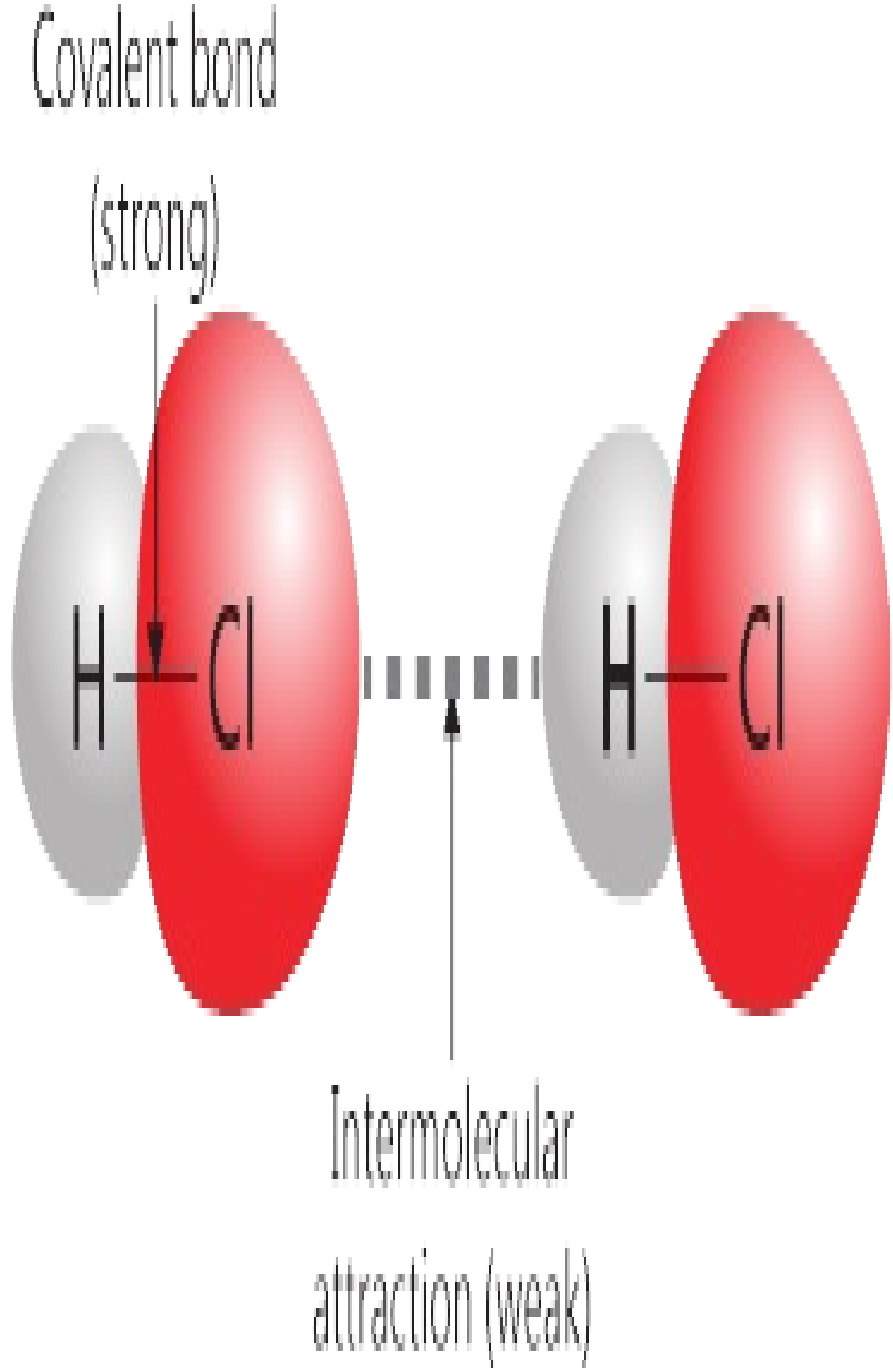
ttractiveness between molecules is not nearly as strong as the intra molecular attractiveness that holds compounds together. They are, however, strong sufficient to control physical characteristics such as boiling and melting points, viscosities, and vapour pressures. An attraction betwixt molecules, intermolecular forces are the forces within the molecule themselves is known as intermolecular forces.

When an object evaporates, melts the molecules gain kinetic energy from an outside source, and this allows them to overcome the intermolecular forces holding them close together there are no intermolecular forces in ionic composite because ionic composite do not carry molecules. Ionic composite are solids at room temperature.

In an ionic crystal ionic bonds hold all the ions together; there is no difference between the bonds holding the composite together and those holding the crystal together. Ionic bonds are usually very strong and account for the high melting point ionic crystal for NaCl. Many molecular objects are gases at room temperature; others are liquids with knee-high boiling points or solids that melt easily.

Forces betwixt the molecules in molecular solids and liquids are relatively weak;

addition of thermal energy can smoothly overcome these intermolecular forces. The strength of intermolecular forces determines the physical characteristics of molecular compounds state, melting point, boiling point, surface tension, hardness, texture and solubility. As the intermolecular forces between molecules arise, the compounds m. p., b .p. and surface tension also arise.


Covalent bond
(strong)
H—Cl
H—Cl
Intermolecular
attraction (weak)

Intermolecular forces

Types of Intermolecular Forces

Dipole-dipole forces.

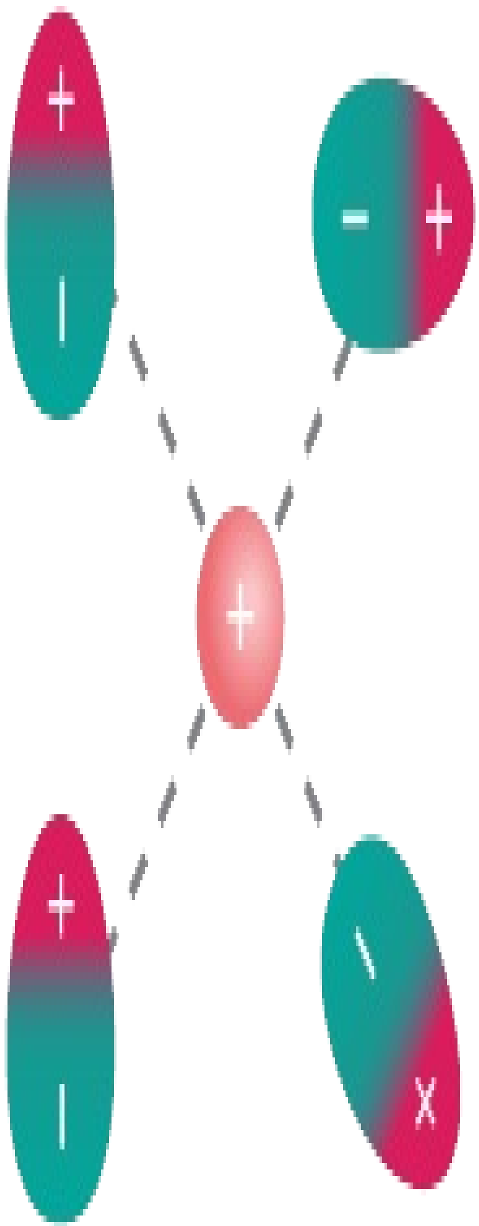
Hydrogen bonding.

London dispersion forces.

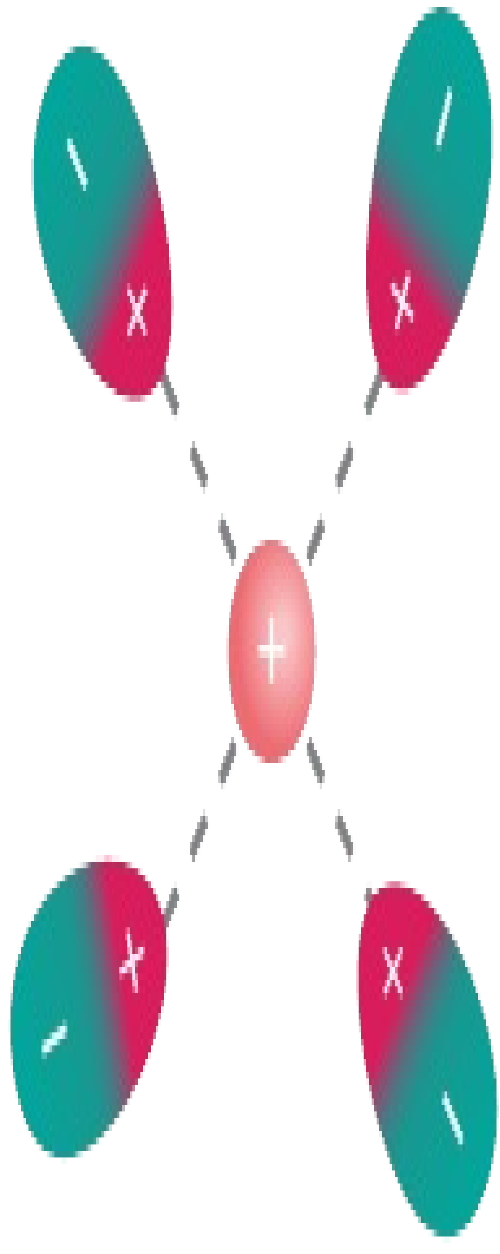
Dipole-Dipole Forces

Ion-dipole interactions

Ion-dipole interactions are a prime force in solutions of ions. Stability of these forces is what makes it possible for ionic object to dissolve in polar solvents.



Cation-dipole attractions

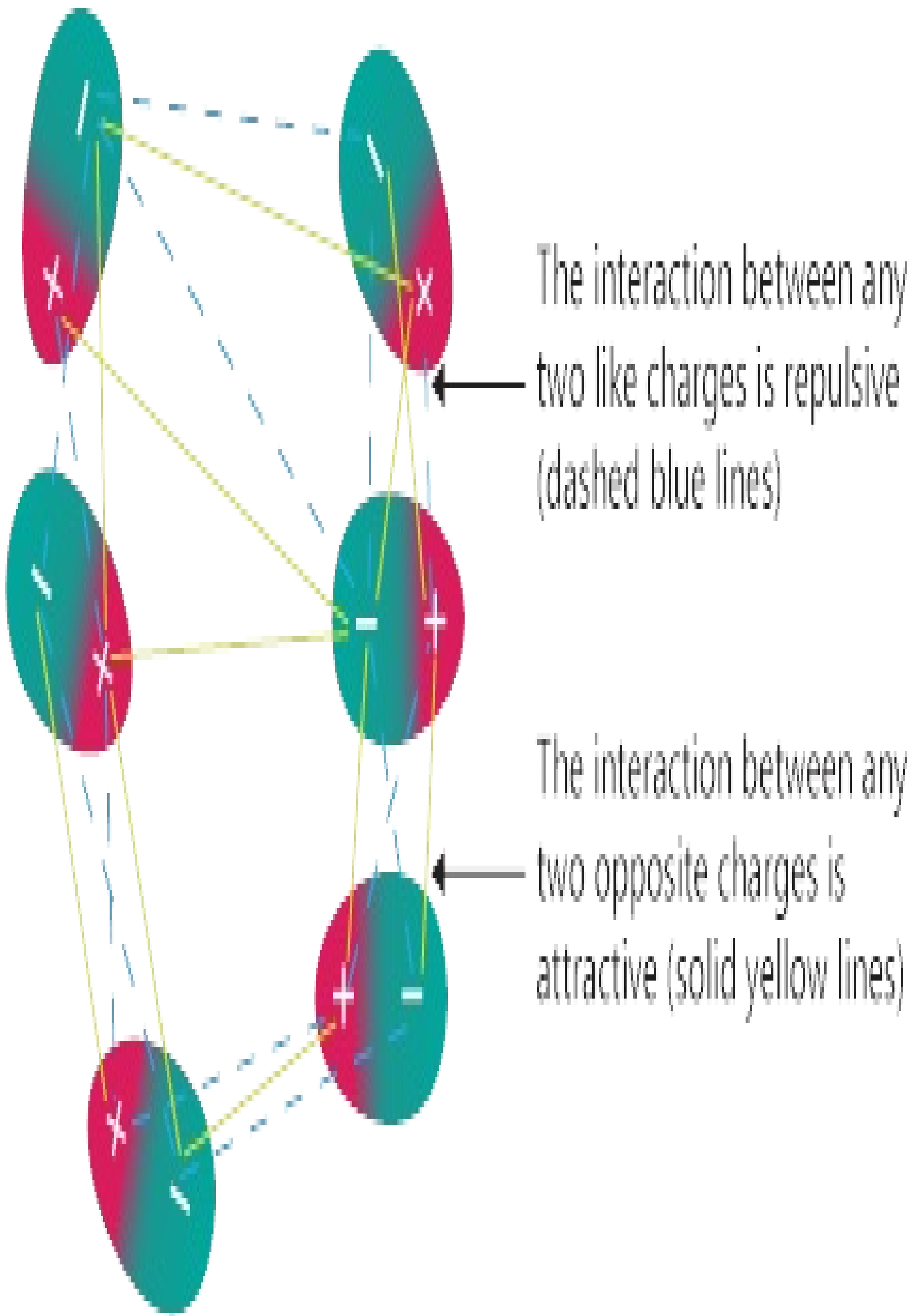


Anion-dipole attractions

Ion-dipole interactions

Dipole-Dipole Interactions

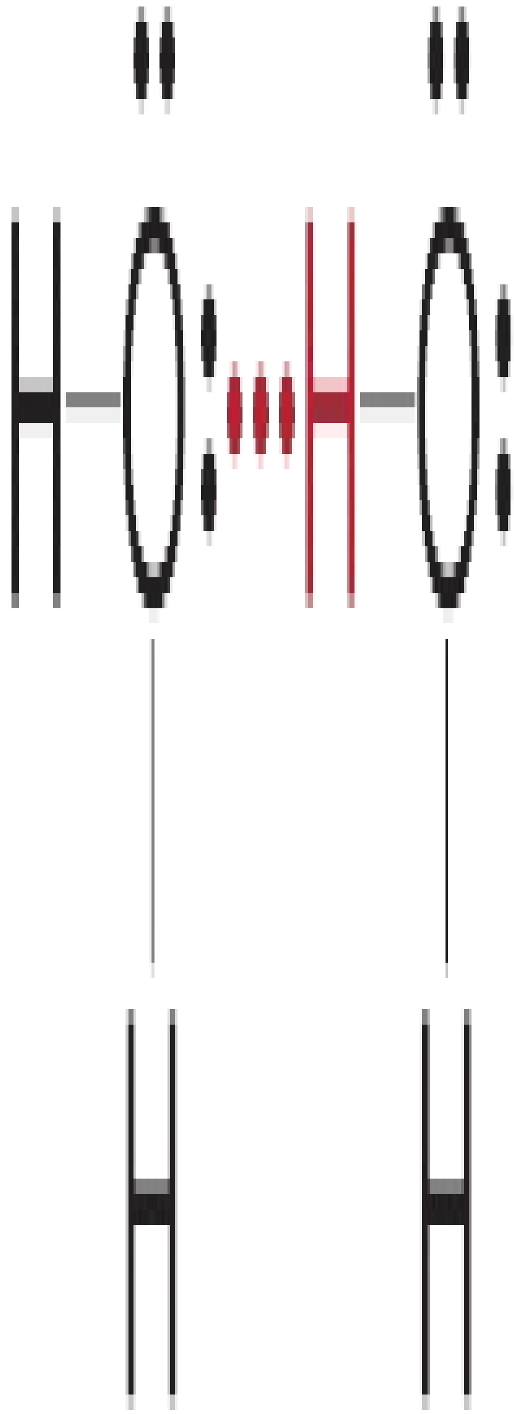
In dipole-dipole interactions molecules that have permanent dipoles are attracted to each other. Positive point of one is attracted to the negative point of the other and vice versa. These forces are only important when the molecules are close to each other.


The interaction between any two like charges is repulsive (dashed blue lines)
The interaction between any two opposite charges is attractive (solid yellow lines)

Dipole-dipole interactions

Hydrogen Bonding

The dipole-dipole interactions experienced when H is bonded to N, O, or F are unusually strong. This type of interactions is called hydrogen bonds. Hydrogen bonding is a significant type of molecular attraction betwixt the hydrogen atom in a polar bond and nonbonding electron couple on a nearby small electronegative ion and atom (e.g. F, O or N).



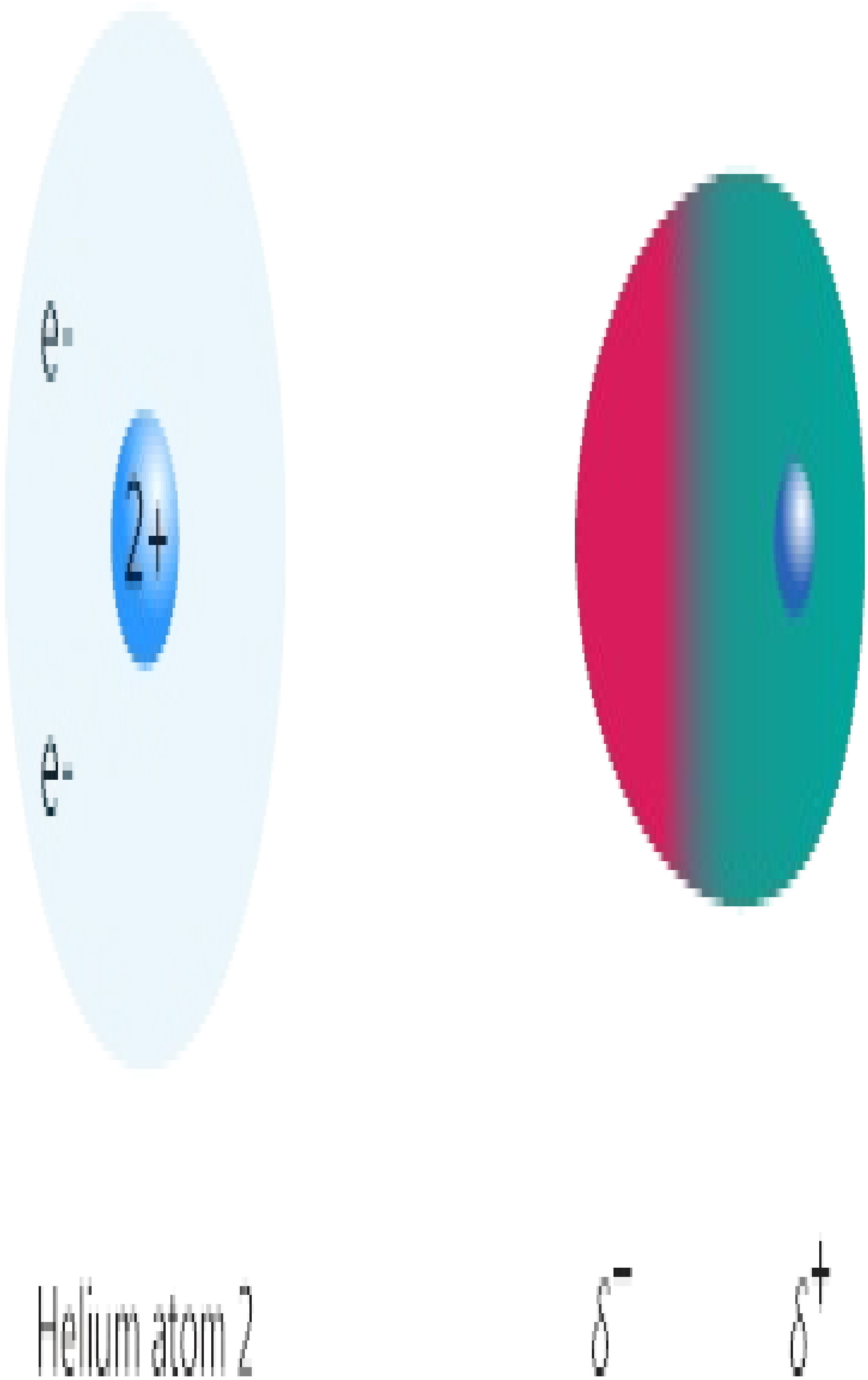
Hydrogen bonding arises in part from the high-rise electro negativity of nitrogen, oxygen, and fluorine. Also, when hydrogen is bonded to one of those very electronegative elements, the hydrogen nucleus is reveal.

London Dispersion Forces

In London dispersion forces all gases will condense to from liquids when cooled enough which signify that there is some type of attraction betwixt the entities when they are cooled. These attractions are caused by the temporary shifts in the electron gloom in an atom or molecule. This shift creates a temporary dipole for a fraction of a second that will also cause a temporary dipole in a neighbouring atom or molecule.

Very short lived as the attraction betwixt them continually forms and breaks, as a result these London forces are very frail. Larger the molecule, electrons and protons they are attracting one another, stronger the forces, high-rise the melting point. London dispersion forces exist betwixt all molecules even polar molecules. In polar composite the London forces are inconsequential compared to the much stronger dipole forces. Dipole-dipole forces and London forces are called “Van der Waals” forces.

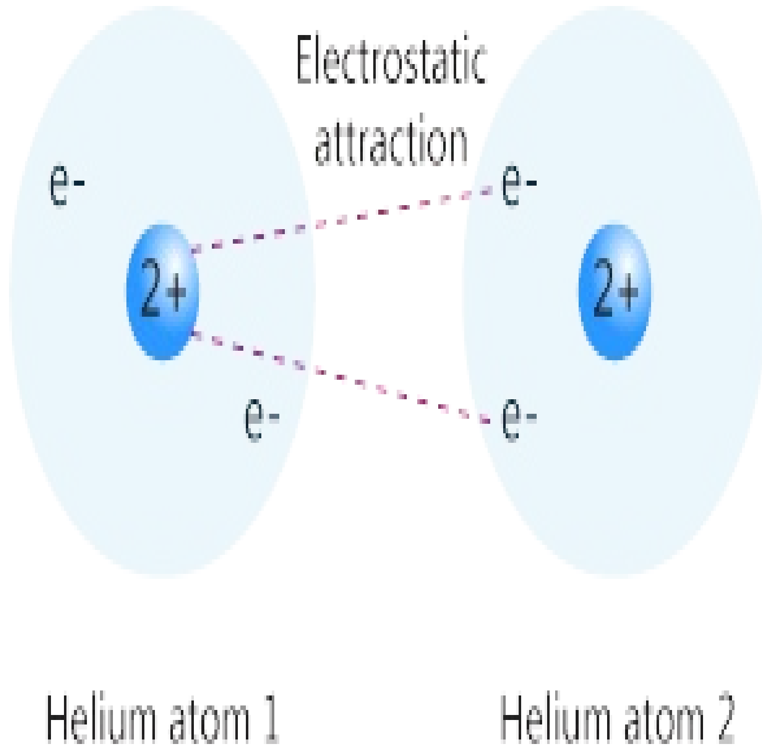
Example:

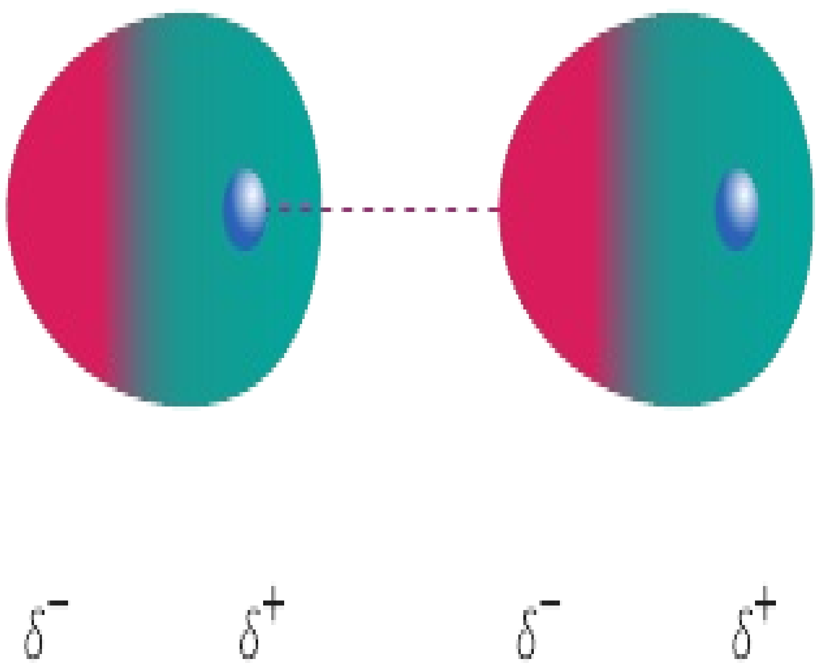


e^-
2+
e^-
Helium atom 2
δ^-
δ^+

London dispersion forces

At this instant, then, helium atom is polar, with a surplus of electrons on the left end and a shortage on the right end. Another helium nearby, then, would have a dipole convinced in it, as the electrons on the left end of helium atom 2 repulse the electrons in the cloud on helium atom 1.


Electrostatic attraction
e-
e-
2+
e-
e-
2+
Helium atom 1
Helium atom 2


δ^-
δ^+
δ^-
δ^+

Electrostatic attraction

London dispersion forces are attractions betwixt an instantaneous dipole and an induced dipole. London dispersion forces exist in every molecule, even if they are polar or nonpolar. Propensity of an electron cloud to deform in this manner is called polarizability.

Thank you for reading!