Modern Construction Material
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Module - 02
Lecture - 02
Review of Atomic Bonding- I

This is the second lecture on Modern Construction Materials. In the first one, we gave an introduction to the course. We looked at the important consideration, in the choice of a civil engineering material. Today, we will going to the microstructure, and we look at the chemical bonds that are there between different elements and atoms. And we will see what is the link between, what happened in the chemical bonding and the material behavior that we see. Most of these you should have studied in high school and the first two years of college, but today we will look at it contacts of material behavior.

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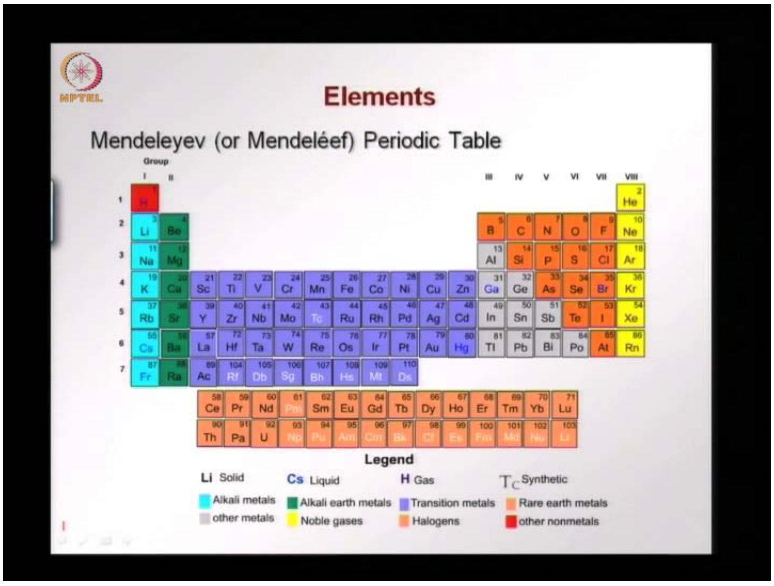

Elements

Mendeleyev (or Mendeléef) Periodic Table

Group	I	II											III	IV	V	VI	VII	VIII
1	H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	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
5	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
6	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
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds								

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
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

Legend

Li Solid
Cs Liquid
H Gas
Tc Synthetic

Alkali metals
Alkali earth metals
Transition metals
Rare earth metals
other metals
Noble gases
Halogens
other nonmetals

So we will start of with Mendeleyev periodic table, which you have all seen before and here we seen listing of elements in the order of that atomic number and they are put into periods and groups.

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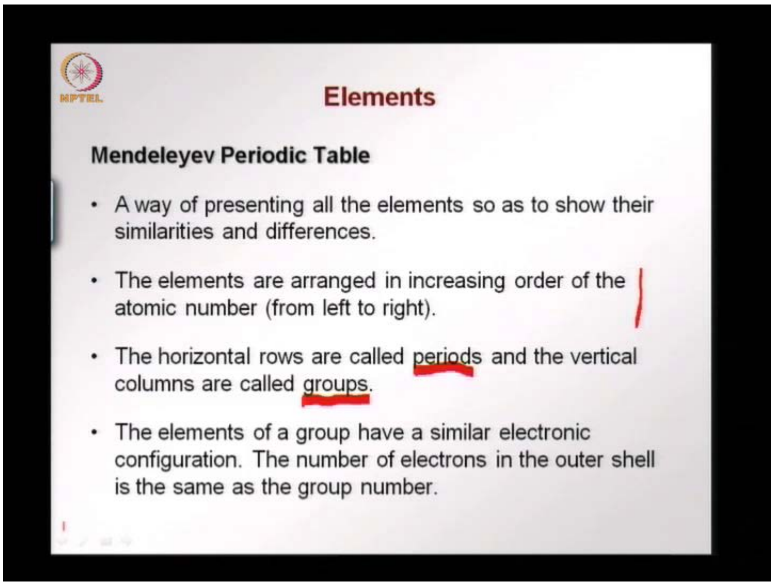

Elements

Mendeleyev Periodic Table

- A way of presenting all the elements so as to show their similarities and differences.
- The elements are arranged in increasing order of the atomic number (from left to right).
- The horizontal rows are called periods and the vertical columns are called groups.
- The elements of a group have a similar electronic configuration. The number of electrons in the outer shell is the same as the group number.

We see that the periodic table as I said presence all the elements are that we can identify similarities and differences. The elements are arranged in increasing order of that atomic number from left to right. And we have the horizontal rows which are called periods, vertical columns are called groups. Now what is of interest is that the elements of a group have the same electronic configuration that means the number of electrons in the outer shell of a group are the same and are as the same as the group number.

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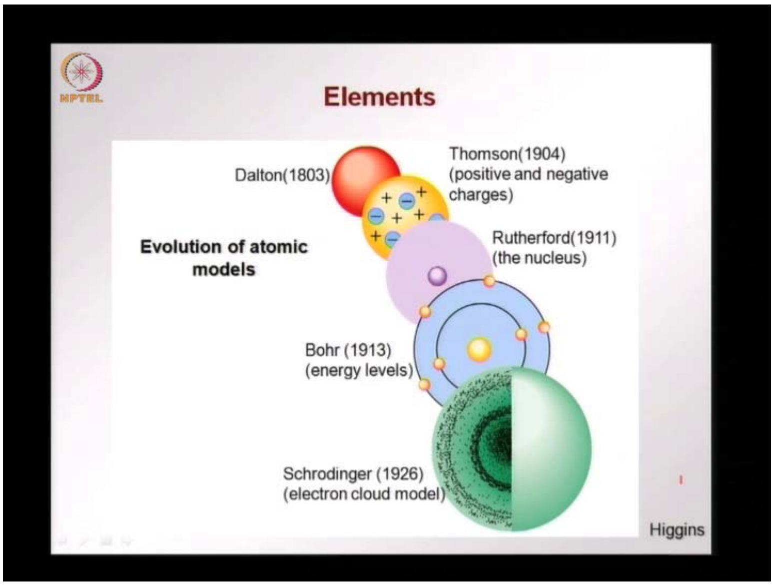

Elements

Evolution of atomic models

Higgins

Now the atomic model has evolved over a long period of time. Dalton in 1803 looked at the atom as an indivisible particle. This has changed a lot Thomson in 1904 found that they the atom had positive and negative charges; Rutherford found that the atom had nucleus and then Bohr found that they where energy levels with electrons moving around the nucleus in different shells. In 1926, Schrodinger develop this model further and came of with an electron cloud model saying that the electron does not have a very specific orbit but it moves in a certain area and this gave rise to the electron cloud model.

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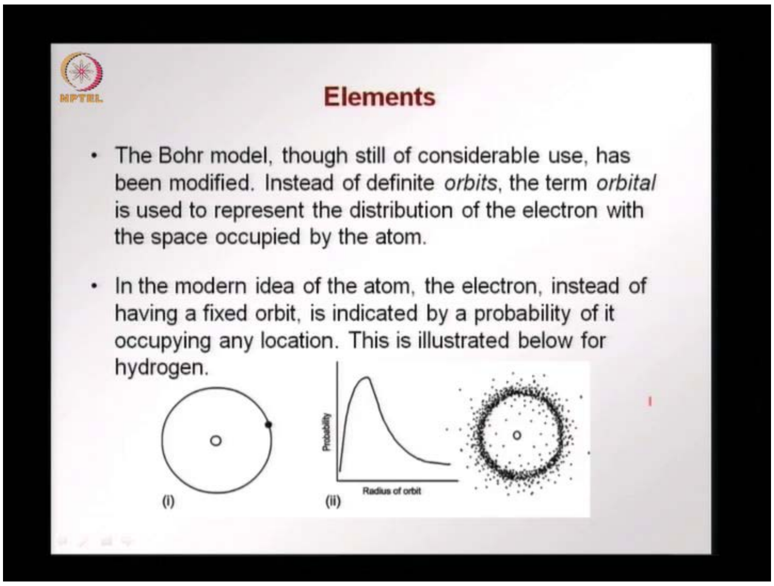

Elements

- The Bohr model, though still of considerable use, has been modified. Instead of definite *orbits*, the term *orbital* is used to represent the distribution of the electron with the space occupied by the atom.
- In the modern idea of the atom, the electron, instead of having a fixed orbit, is indicated by a probability of it occupying any location. This is illustrated below for hydrogen.

However, we still use the modified Bohr model instead of definite orbits for the electrons, we use the term orbital to represent the distribution of the electron within a certain space around the nucleus in the atom. So in the modern idea of the atom, the electron instead of having the fixed orbit, is indicated by a probability of the electron occupying any location. And this is given for hydrogen then we have the electron in different possible locations these got you see here around when you place all the different possible places that the hydrogen electron could begin.

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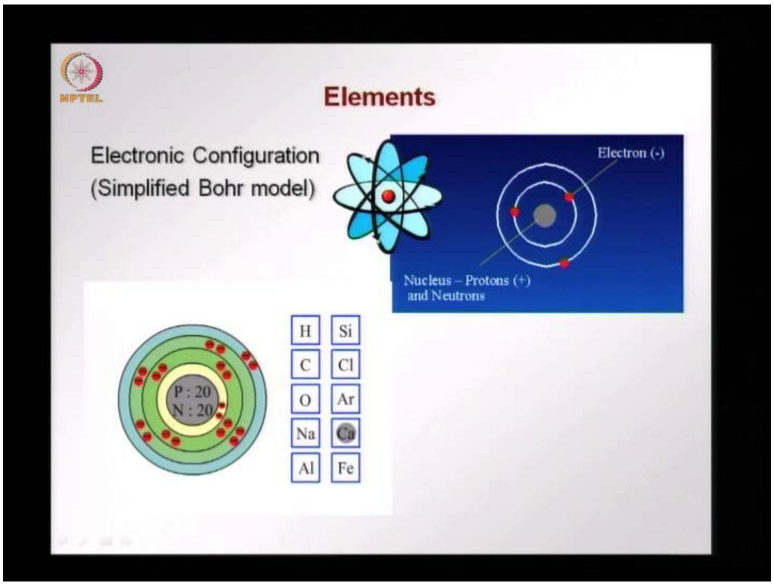

Elements

Electronic Configuration
(Simplified Bohr model)

Now when we look at different element we have different electronic configurations here I will put the figures for some of them. We saw hydrogen having a single electron, carbon has four electrons in the outer shell. These are the protons and the neutrons, oxygen, sodium, aluminum, iron, silicon, chlorine, calcium and argon is the reference for an atom which has eight electrons in the outer shell and that is called the octal configuration. And when an atom reaches or has eight electrons in the outer shell of the octal configuration, we considered is to be stable that it does not need to bond with any other atom to become stable.

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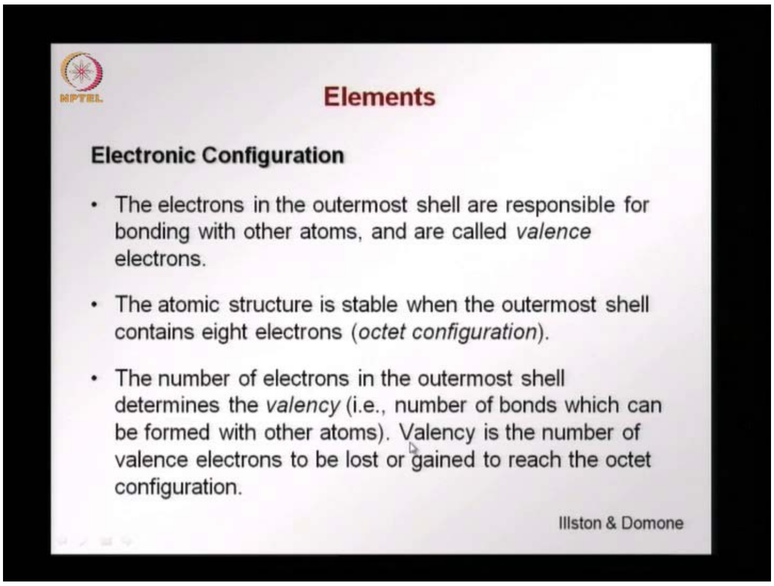

Elements

Electronic Configuration

- The electrons in the outermost shell are responsible for bonding with other atoms, and are called *valence* electrons.
- The atomic structure is stable when the outermost shell contains eight electrons (*octet configuration*).
- The number of electrons in the outermost shell determines the *valency* (i.e., number of bonds which can be formed with other atoms). Valency is the number of valence electrons to be lost or gained to reach the octet configuration.

Illston & Domone

The electrons in the outermost shell are called the valence electrons. And as I said when we have eight electrons in the outer shell that is called the octet configuration and that is a stable configuration. The number of electrons in the outermost shell that is the valance electrons, determines the valence of an element that is the number of bonds which can be formed with other atoms, this depends on the valency, and the valence depends on the number of electrons in the outer the shell. So, the valence is the number of valence electrons that can be lost or gained to reach the octet configuration. So, valance electrons can either be lost or gained to reach the octet configuration.

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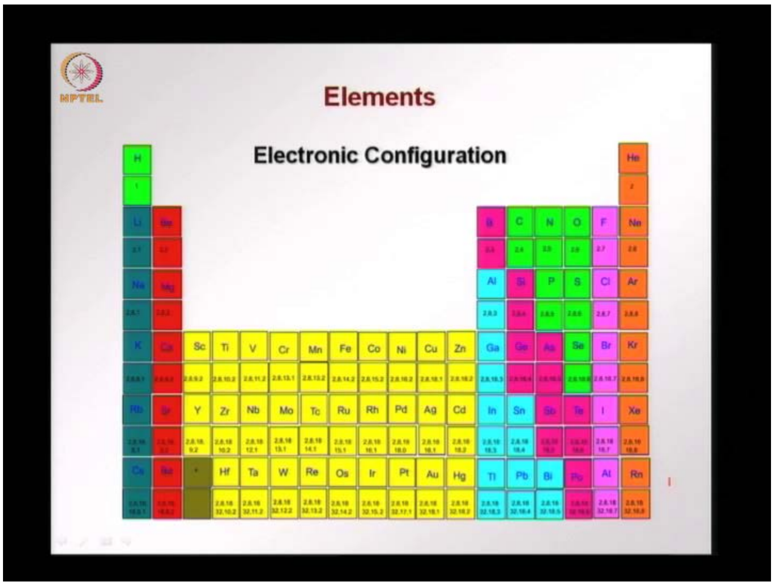

Elements

Electronic Configuration

And from the electronic configuration table, we can find from the last number below each of these elements symbols, you find the valance electrons and therefore, you can find the valency of each of the atoms. So, this gives an idea of the atomic configuration, the electronic configuration and how possibly the bonds can develop between different atoms.

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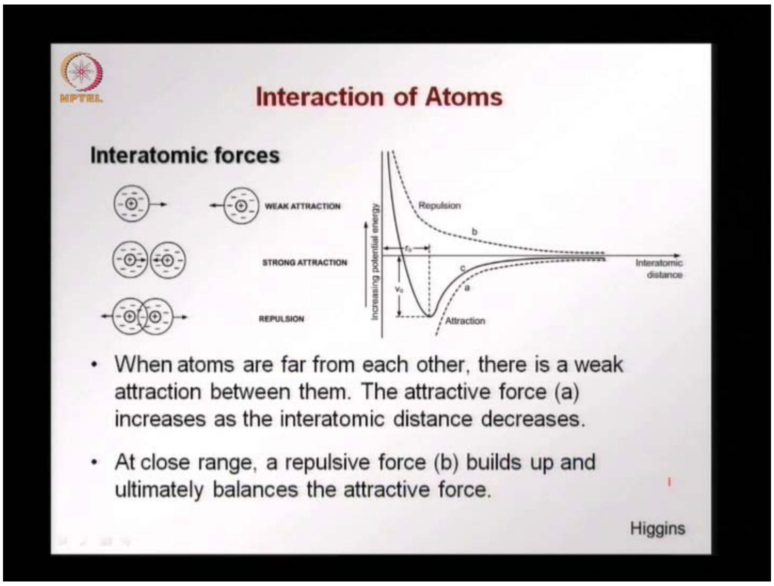

Interaction of Atoms

Interatomic forces

- When atoms are far from each other, there is a weak attraction between them. The attractive force (a) increases as the interatomic distance decreases.
- At close range, a repulsive force (b) builds up and ultimately balances the attractive force.

Higgins

Now, what happens, when atoms come together; initially when the atoms are far away you have week attractive force between the atoms. When they come closer together, the attractive force is becoming larger, there is a stronger attraction. But when they are very close to each other, when we are they are push together, even further, then there is a repulsive force. And this is given in this diagram; showing on the y-axis, the potential energy increasing potential energy. In the x-axis gives the distance between the atoms the inter-atomic distance. The bottom dash curve is the attractive forces gives the attractive forces, how they change the potential energy. The top dash line gives the effect of the repulsive forces that is curve b.

The resultant of curves a and b is curve c. This is the resultant of the effects of the attractive, and the repulsive forces. Note that both the attractive and repulsive forces act together, and you get this curve as a resultant of that. What we see is that r as the atomic distance decreases the attractive potential energy is more and then as the atoms are push together or come very close to each other, the potential energy increases and the repulsive part of the potential energy also increases. r zero is the point for the inter atomic separation when the potential energy is minimum or the attractive part of the potential energy is maximum.

So, when the atoms are far from each other, there is a weak attraction between them. The attractive forces given by curve a increases as the inter atomic distance decreases and when the atoms are very close to each other are repulsive force forms and this ultimately balances the attractive force and when you have balance as the attractive force and when you have very close inter atomic distances the repulsive force is very high.

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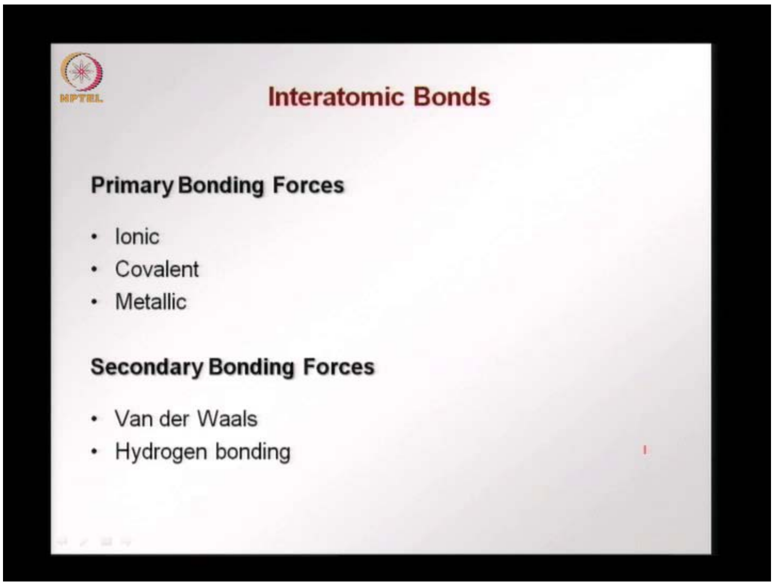

Interatomic Bonds

Primary Bonding Forces

- Ionic
- Covalent
- Metallic

Secondary Bonding Forces

- Van der Waals
- Hydrogen bonding

Now there are five types of bonding forces; three of them are called primary bonding forces, and two secondary bonding forces. The primary bonding forces we will see how much stronger then the secondary bonding forces. So, we look in little bit of detail in each of these cases and see what effects they could have on the material behavior.

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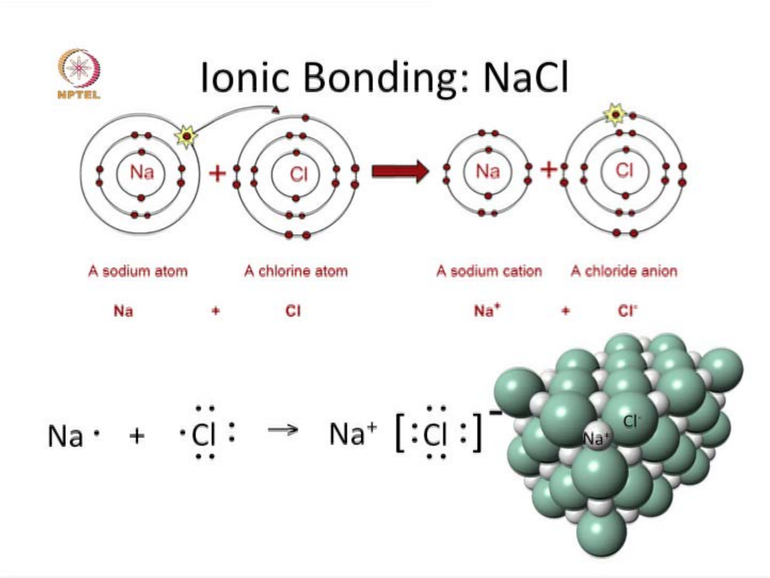

Ionic Bonding: NaCl

A sodium atom | A chlorine atom | A sodium cation | A chloride anion

$Na + Cl \rightarrow Na^{+} + Cl^{-}$

$Na\cdot + \cdot\ddot{\underset{\cdot\cdot}{Cl}}: \rightarrow Na^{+}\,[:\ddot{\underset{\cdot\cdot}{Cl}}:]^{-}$

We start with the ionic bonding with the example of sodium chloride. So, here we have an atom of sodium with one electron in the outer shell; chlorine has seven electrons in the outer shell, and to be stable each of these atoms have to reach the octet configuration. So, in the case of ionic bonding, what happens is the single electron of the sodium atom in the outer shell is lost. What we now have is a positively charged sodium ion – cation, and this electron is now transfer to chlorine atom and it becomes a chloride anion with that is negatively charged. So, we have chloride anion hence. So, that is given figuratively at the bottom. We have sodium with a single electron with this outer shell combining with chlorine which has seven electron in its outer shell to give a sodium chloride. Where sodium now exist as a sodium cation positively charge ion and then you have chlorine which is existing, as a chloride anion for a negatively charged ion.

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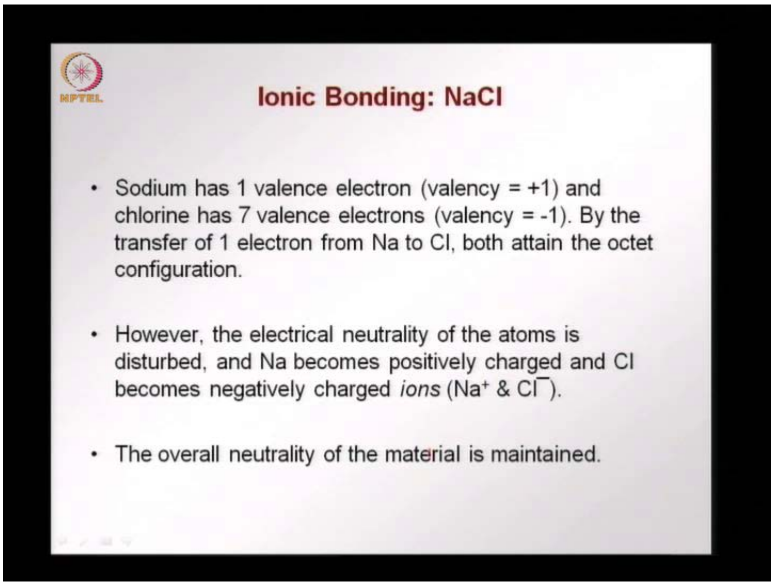

Ionic Bonding: NaCl

- Sodium has 1 valence electron (valency = +1) and chlorine has 7 valence electrons (valency = -1). By the transfer of 1 electron from Na to Cl, both attain the octet configuration.
- However, the electrical neutrality of the atoms is disturbed, and Na becomes positively charged and Cl becomes negatively charged *ions* (Na^+ & Cl^-).
- The overall neutrality of the material is maintained.

So, to summarize sodium as one valance electron; that means, it has valency of plus one, chlorine had seven valance electrons a valency of minus one that is it needs one electron to reach the octet configuration. So, by transferring one electron from sodium to chlorine, both get the octet configuration. This transfer of electrons changes the neutrality of the atom is such, the sodium becomes positively charged; that means, it becomes a cation; and chlorine becomes negatively charged, it becomes a chloride an ion. However, the overall neutrality of the material is maintained; sodium chloride has such is neutral, because it has the cations and an ions together. So, the overall neutrality of the material is maintained.

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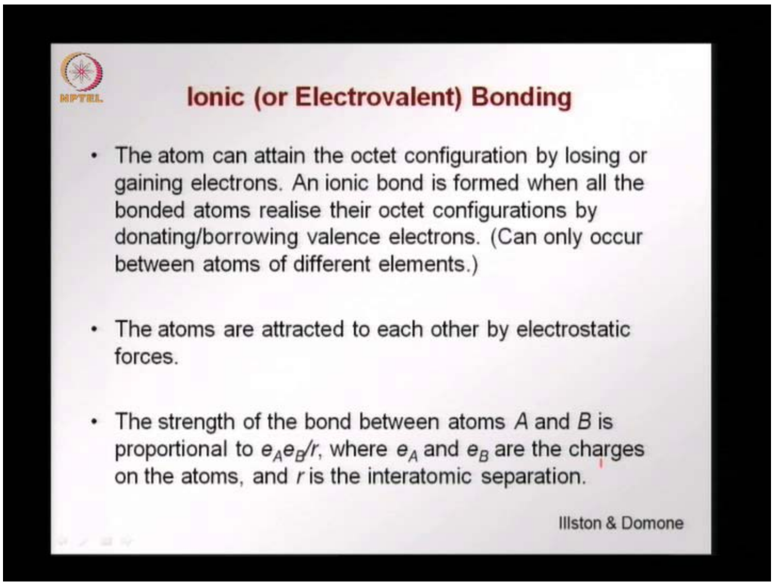

Ionic (or Electrovalent) Bonding

- The atom can attain the octet configuration by losing or gaining electrons. An ionic bond is formed when all the bonded atoms realise their octet configurations by donating/borrowing valence electrons. (Can only occur between atoms of different elements.)
- The atoms are attracted to each other by electrostatic forces.
- The strength of the bond between atoms *A* and *B* is proportional to $e_A e_B/r$, where e_A and e_B are the charges on the atoms, and r is the interatomic separation.

Illston & Domone

So, in one ionic or electrovalent bonding what we see is the atom can get the octet configuration either by losing or gaining electrons. And ionic bond is formed when all the bonded atoms reach their octet configurations by either donating or borrowing valance electrons. And this; obviously, can only occurred between atoms of different valance, because this atoms of the same element will have the same valency and an ionic bond cannot form between them. The atoms in the ionic bond are attracted to each other by electrostatic forces. And the strength of the forces and the strength of the bond which develops is given by e of a times e of b divided by r. Where e sub A and e sub B are the charges of the two atoms that participate in the bond, and r is the inter atomics separation. And what we see here is when the inter atomic separation is larger the attractive forces the electrostatic forces that I keeping them together decrease.

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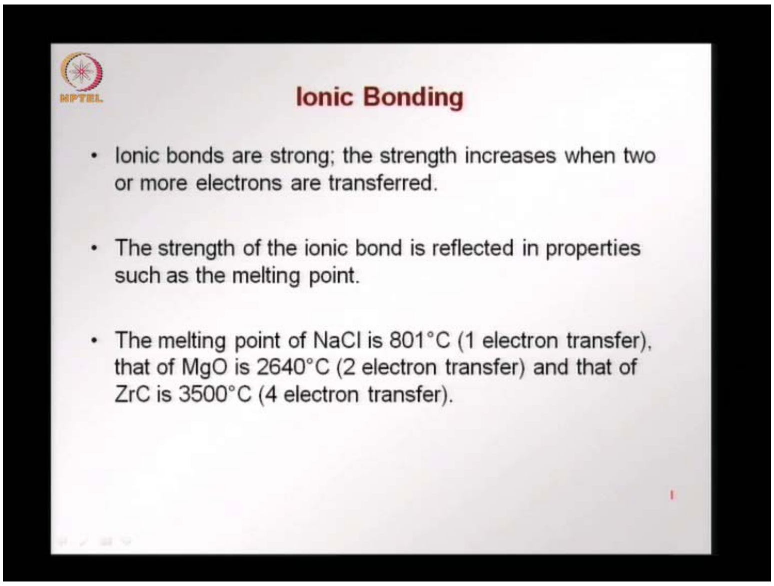

Ionic Bonding

- Ionic bonds are strong; the strength increases when two or more electrons are transferred.
- The strength of the ionic bond is reflected in properties such as the melting point.
- The melting point of NaCl is 801°C (1 electron transfer), that of MgO is 2640°C (2 electron transfer) and that of ZrC is 3500°C (4 electron transfer).

Now as we said initially, the ionic bond is a strong bond, and the strength of the bond increases when more electrons has transferred from one atom to the other. The strength of the ionic bond increases when two or more electrons are transferred. And this strength of the ionic bond is reflected in the properties such as the melting point. The melting point of sodium chloride is about 800 degree centigrade. It as we saw already involves the transfer of one electron; that of magnesium oxide is 2640 degrees Celsius two electrons are involved, two electrons are transferred in the ionic bond between magnesium oxygen. And that of zirconium carbide is 3500 degrees Celsius, four electrons has transferred in the ionic bond between zirconium and carbide. So, what we see is more electrons transferred stronger is the bond, and when the bond is strong, the melting point is higher, because more energy has to be put in to break the bond and separate the elements that from the bond.

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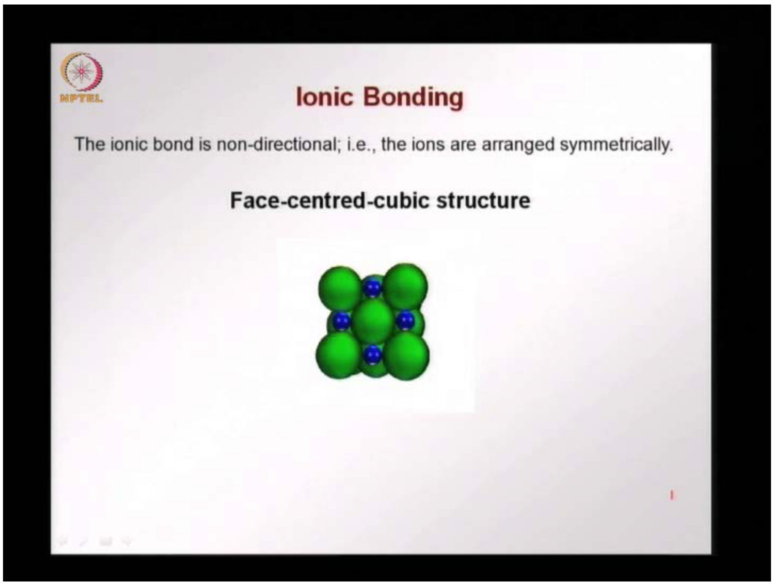

Ionic Bonding

The ionic bond is non-directional; i.e., the ions are arranged symmetrically.

Face-centred-cubic structure

Another feature of the ionic bond is that it is non-directional. The ions in the ionic bond can be are arranged symmetrically that is there is no preferential direction of growth of materials that are bonded with ionic bond. You see the same structure in all directions for example, in this face centered cubic structure.

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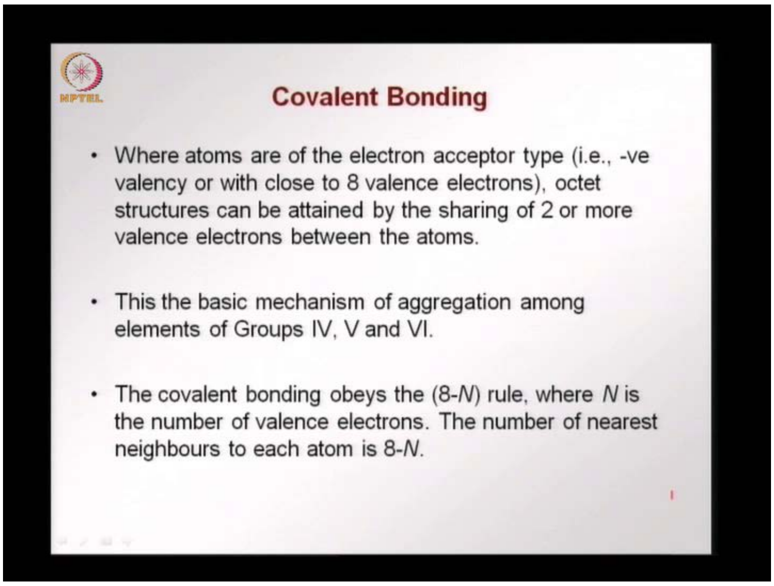

Covalent Bonding

- Where atoms are of the electron acceptor type (i.e., -ve valency or with close to 8 valence electrons), octet structures can be attained by the sharing of 2 or more valence electrons between the atoms.
- This the basic mechanism of aggregation among elements of Groups IV, V and VI.
- The covalent bonding obeys the (8-*N*) rule, where *N* is the number of valence electrons. The number of nearest neighbours to each atom is 8-*N*.

Another type of strong bond is the covalent bond, where atoms or all of the electron acceptor type that is they have a negative valence and close to eight valence electrons. The octet structure is attained by sharing of two or more valence electrons between the atoms. So, here instead of the atom transferring the electron like we saw in the ionic bond, the octet structures achieved by sharing of the valence electrons between atoms, and this basically happens in bonds involving elements of groups four, five and six. And if you remember, we said that the number indicates the number of electrons in the outer shell or the valence electrons. Covalent bonding obits what is call the eight minus N rule, where N is the number of electrons and what we see is the number of nearest neighbors, when atoms want together the covalent bond is eight minus N.

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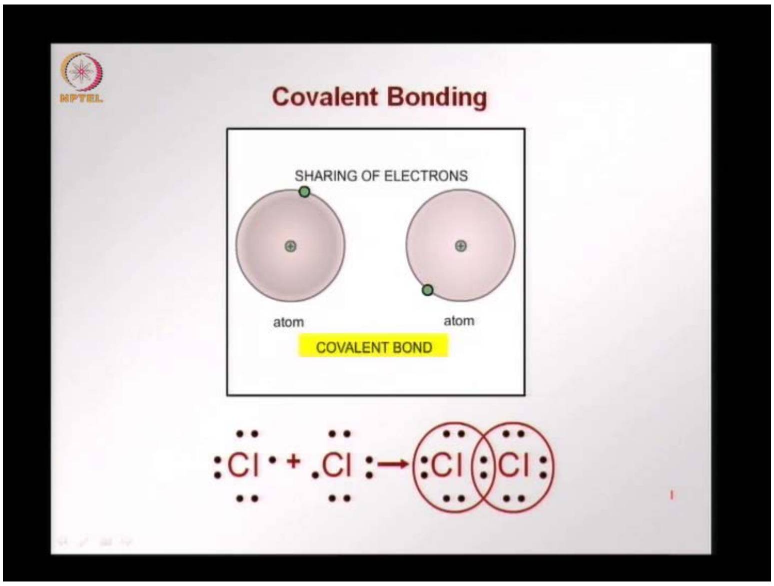

Covalent Bonding

Let us look at a case of chlorine. chlorine has seven electrons in the outer shell, and we find that when two atoms of chlorine get-together then a covalent bond is formed two electrons are shared between the two atoms and you have the octet configuration and you have a chlorine molecule.

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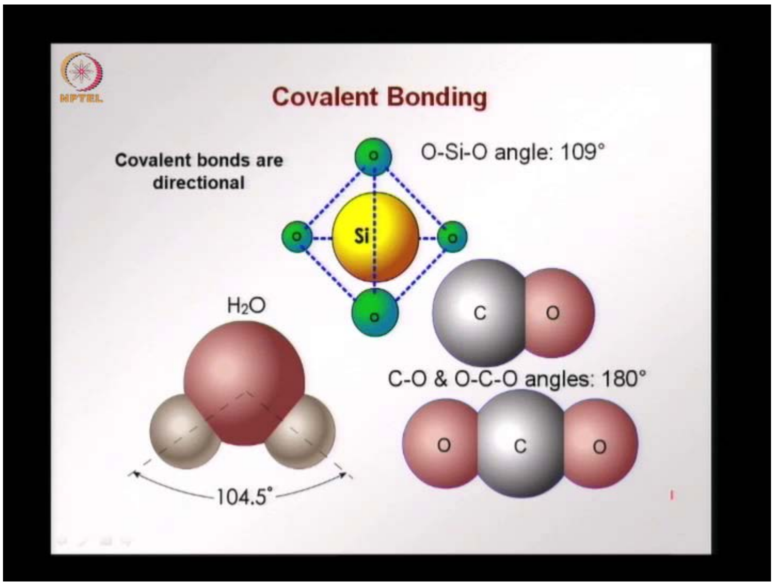

Covalent Bonding

Covalent bonds are directional

Covalent bonds are direction is already that ionic bonds are not direction they can form a symmetrically in all directions covalent bonds are directional that mean that the bond that forms between the atoms and should have a particular configuration for example, we see here in the water molecule we see that there is an angle of about hundred and five degrees between the hydrogen oxygen hydrogen configuration within the water molecule similarly when we have a bond between carbon and oxygen the angular one eighty degrees; that means, we have lengthwise configuration of the molecules in the case of carbon monoxide and carbon dioxide in the tetrahedral. We find that the angle between the oxygen silicon oxygen configuration is hundred and nine degrees; that means, that whichever bond that happens with these elements will always have this type of configuration and this gives rise to a specific shape of the molecule that forms due to covalent bonding and this is very important and will say this later because this specific type of directionality give rise to chains and states covalently bonded materials.

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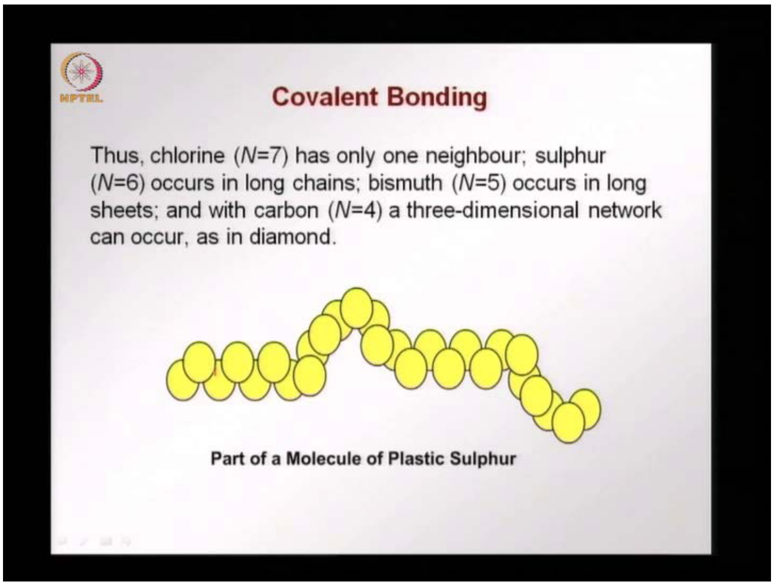

Covalent Bonding

Thus, chlorine (N=7) has only one neighbour; sulphur (N=6) occurs in long chains; bismuth (N=5) occurs in long sheets; and with carbon (N=4) a three-dimensional network can occur, as in diamond.

Part of a Molecule of Plastic Sulphur

So, in chlorine the valence electrons in the outer shell as seven. So, the valence is minus one it has only one neighbor following the eight minus n role eight minus seven giving one so; that means, chlorine needs only one neighbor to be stable sulphur occurs in long chains sulphur has six electrons in its outer shell. So, it valence e is minus two the eight minus n role gives us a value of two so; that means, the see that for each sulphur atom we have two neighboring atoms to make it stable this much as n equal to five. So, it occurs in long sheets carbon can form a three-dimensional network the valence is four. So, it as it needs four neighboring atoms to form a stable system. So, you have a three-dimensional network with carbon as in diamond.

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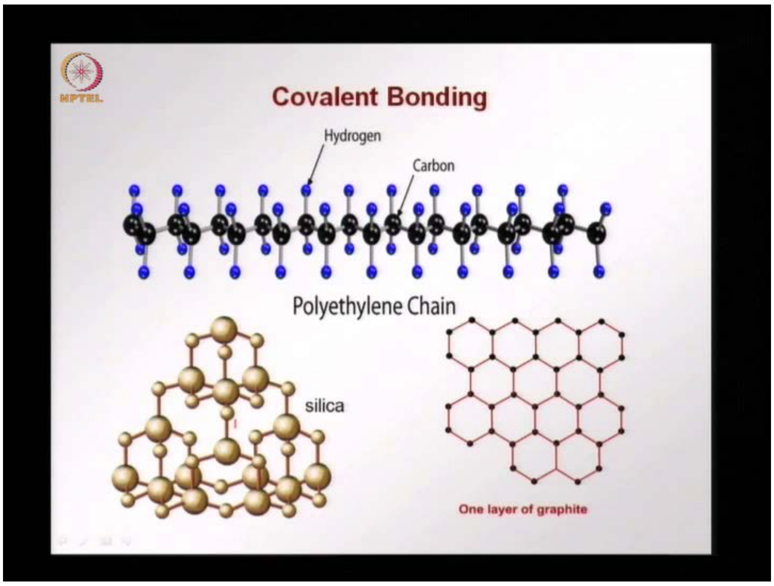

So, as we said we can have long chains occurring like in polymers, where we have here in polyethylene a chain of carbon and hydrogen the block spheres represent carbon and the blue spheres represent hydrogen and you find that the angle between the bonds have to be specific, because they are directional and due to this we find that the polyethylene chain can deform, can change is configuration without changing the angle between the individual atoms. So, this can give lies rise to long chains which can bend and have a certain amount of flexibility at the bottom right we have a layer of graphite with carbon atoms and we find that graphite forms in sheets within a sheet.

We have covalent bonds now this sheet is stable enough such that it does not have to react with anything else out of the plane of the sheet silica on the other hand as a more of a three-dimensional network where we have silicone and actions and atoms in a three-dimensional network again the angle between the oxygen silicone oxygen configuration is predetermined or its fixed.

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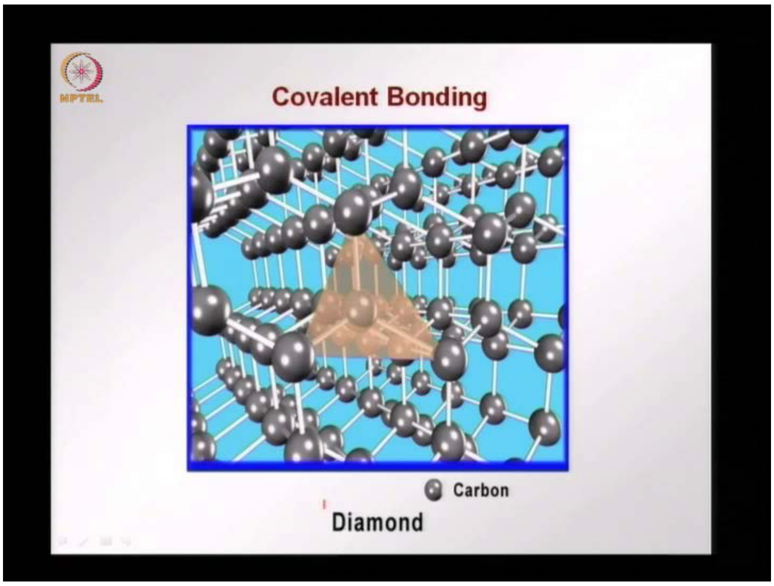

Covalent Bonding

Diamond

We saw that in the case of carbon, because of the valence see being four, we have a three-dimensional network which is a rare case for covalently bonded materials diamond is an exemption, where we have a three-dimensional network forming between the carbon atoms.

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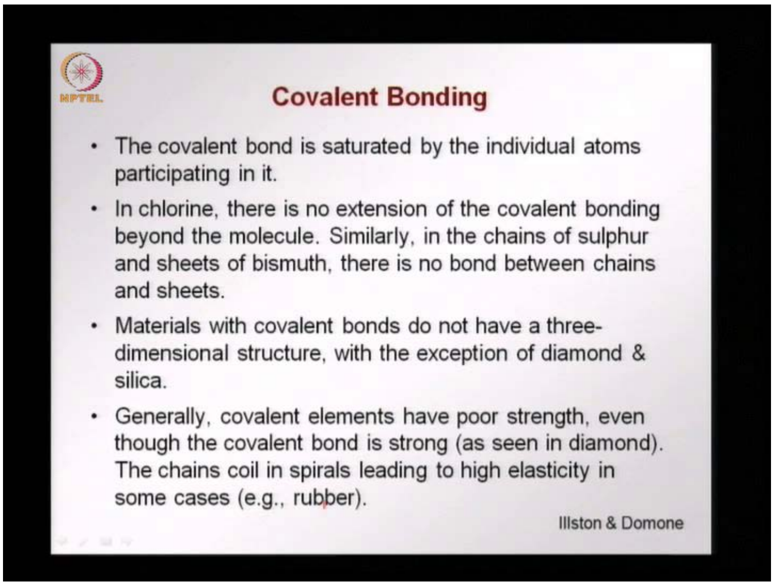

Covalent Bonding

- The covalent bond is saturated by the individual atoms participating in it.
- In chlorine, there is no extension of the covalent bonding beyond the molecule. Similarly, in the chains of sulphur and sheets of bismuth, there is no bond between chains and sheets.
- Materials with covalent bonds do not have a three-dimensional structure, with the exception of diamond & silica.
- Generally, covalent elements have poor strength, even though the covalent bond is strong (as seen in diamond). The chains coil in spirals leading to high elasticity in some cases (e.g., rubber).

Illston & Domone

So, to summarize the covalent bond we find that the covalent bond is saturated by individual atoms meaning that once you have a covalently bonded material the molecules do not have to bond with anything else it is saturated by the individual bonds for example, in chlorine. When you have a chlorine molecule, there is no extension, there are no other bonds form beyond what is needed to form the molecule similarly in chains of sulphur sheets of bismuth there is no bond between different chains or different sheets. So, this chains and sheets remain independently within the material as I said before covalent bonds do not lead to a lot of three-dimensional structures the exceptions being diamond and silica both of which are very strong materials. In the other cases is there is not much bonding not a strong bond between the different chains and sheets generally we find that covalent elements have poor strength.

Even though the covalent bond itself is strong, since we do not have a strong bond between the sheets and between the chains. We see that covalent elements have poor strength, the exception being diamond and silica, and because of these long chains and sheets forming and the possibility of them bending while maintaining the directionality, we find that covalently bonded elements leads to chains that can coil in spirals leading to high elasticity and flexibility like in the case of rubber.

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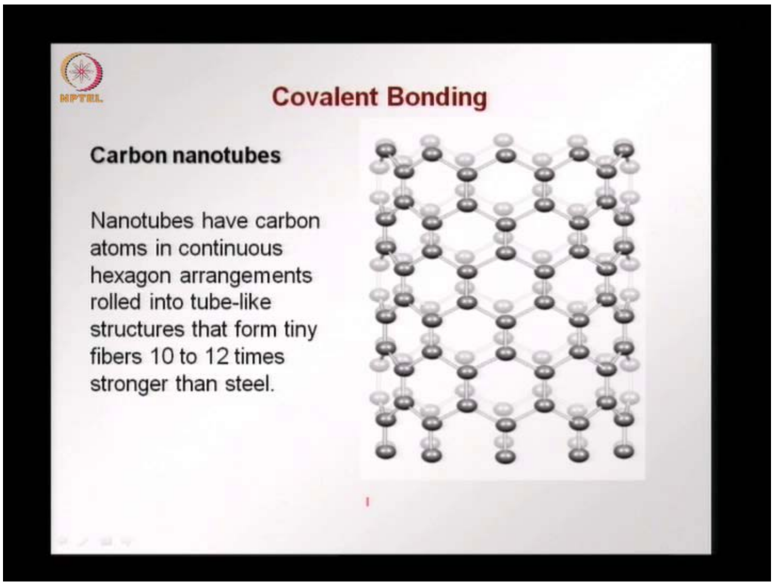

Covalent Bonding

Carbon nanotubes

Nanotubes have carbon atoms in continuous hexagon arrangements rolled into tube-like structures that form tiny fibers 10 to 12 times stronger than steel.

One interesting case of covalent bonding is in carbon nano tubes where we have a continuous hexagon arrangements of carbon atoms leading to a tube like structure. So, this is a this could be the top of the tube and these tubes or very strong they are tiny, but have strength of ten to twelve times that of steel.

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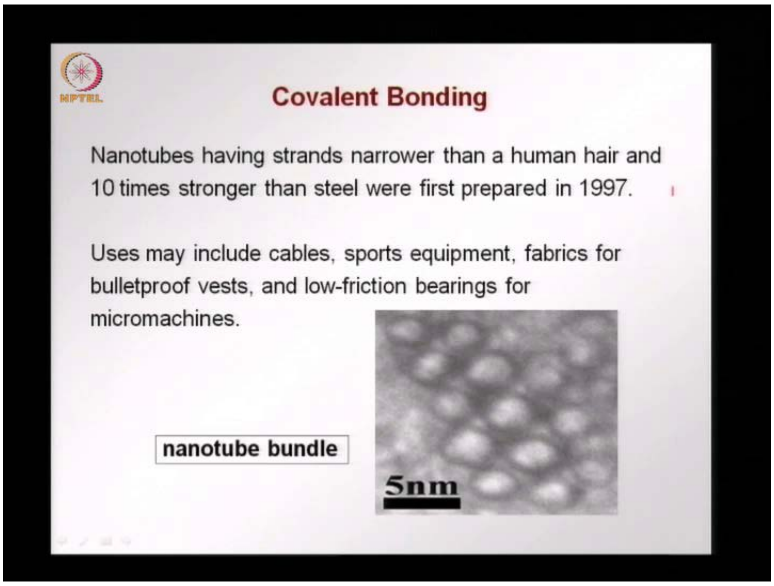

Covalent Bonding

Nanotubes having strands narrower than a human hair and 10 times stronger than steel were first prepared in 1997.

Uses may include cables, sports equipment, fabrics for bulletproof vests, and low-friction bearings for micromachines.

nanotube bundle

These were first prepared in the late nineteen nineties, and they were found to be narrower than a human hair, but very very strong. We have add some research going on where the people have try to use nano tubes carbon nano tubes in cementation is materials in other applications civil engineering, but it is still in an arson stage what you see here at the bottom is a photographs taken of the web where you see a nano tube bundle this scale shown is five nanometers. So, you see that each of these nano tubes is even smaller than five nanometers right now the uses that are being done and in progress look at cables a equipment which require very high strength and flexibility fabrics for bullet-proof vests and some may be low friction bearings for micro machines and. So, on in terms of mechanical applications.

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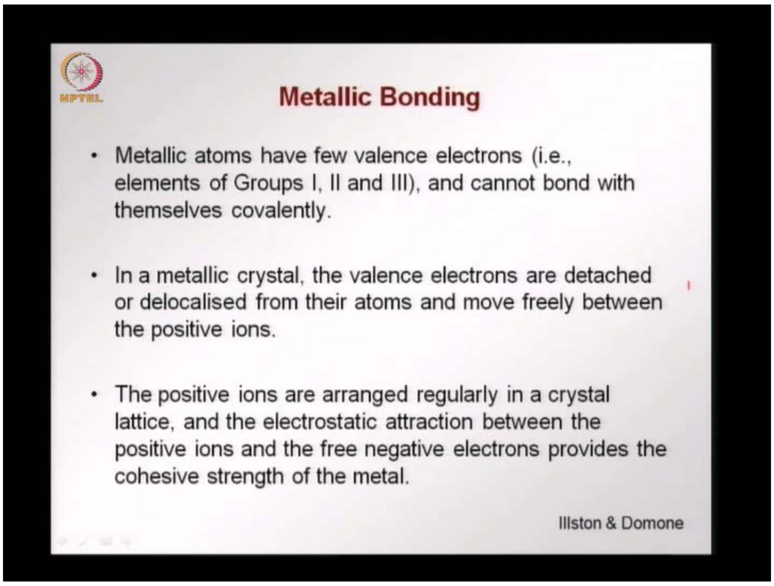

Metallic Bonding

- Metallic atoms have few valence electrons (i.e., elements of Groups I, II and III), and cannot bond with themselves covalently.
- In a metallic crystal, the valence electrons are detached or delocalised from their atoms and move freely between the positive ions.
- The positive ions are arranged regularly in a crystal lattice, and the electrostatic attraction between the positive ions and the free negative electrons provides the cohesive strength of the metal.

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The third strong bond is the metallic bond and this is of consequence to civil engineering because many of the materials that are very popularly used in civil engineering and in construction. A general or metals and these derive their properties from metallic bonds metallic atoms have few valence electrons that is they are elegant of groups I, II and III and cannot bond with themselves covalently. In a metallic structure, we have valence electrons that detached or delocalized from their atoms and move freely between the positive ions. So, instead of having atoms being shared as we saw in the covalent bond what we have here in a metallic bond is that the valence electrons detached separate from the atoms and move freely.

Within the metallic crystal between the positive ions, the positive ions or the cations arrange regularly or arrange regularly in the lattice in the crystal lattice. Lattice is the formation of the crystal, the electrostatic attraction between the positive ions and the freely floating negative ions provides the cohesive strength of the metal. So, this means that that even though there is no bond between two or more atoms specifically, the whole crystal structure is help together by the positive ions that are arranged in a crystal lattice and the negative ions that freely floating around in a cloud configuration.

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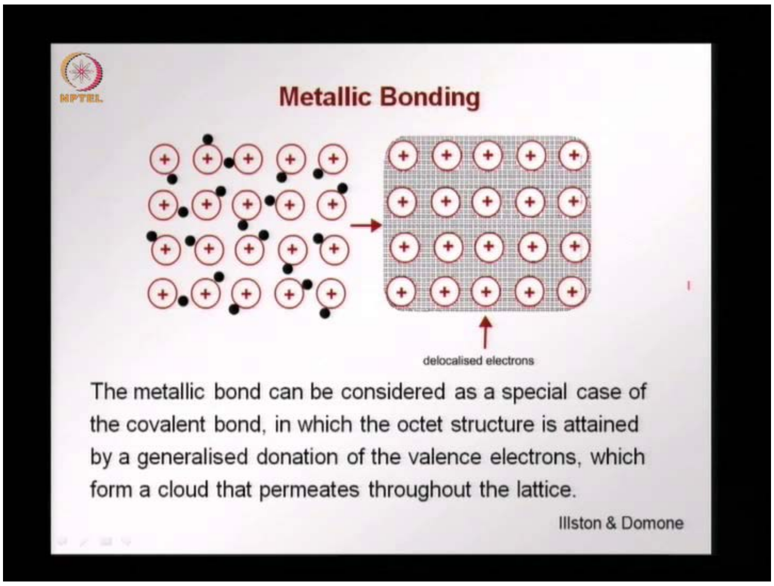

Metallic Bonding

The metallic bond can be considered as a special case of the covalent bond, in which the octet structure is attained by a generalised donation of the valence electrons, which form a cloud that permeates throughout the lattice.

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So, you see that here represented in this diagram. We have the positive ions the cat ions of the metallic atoms all arranged in a certain configuration which makes up. The crystal lattice around these cat ions we have the freely floating electron the electro-static forces between the positively charged ions and all the electrons keeps the whole crystal lattice in place and gives rise to strong bond this can be considered as a special case of the covalent bond where the opted structure is attained by a generalized donation of valence electrons forming a cloud that permeates through the lattice. So, we have a cloud of electrons throughout the crystal lattice that is keeping the bond together.

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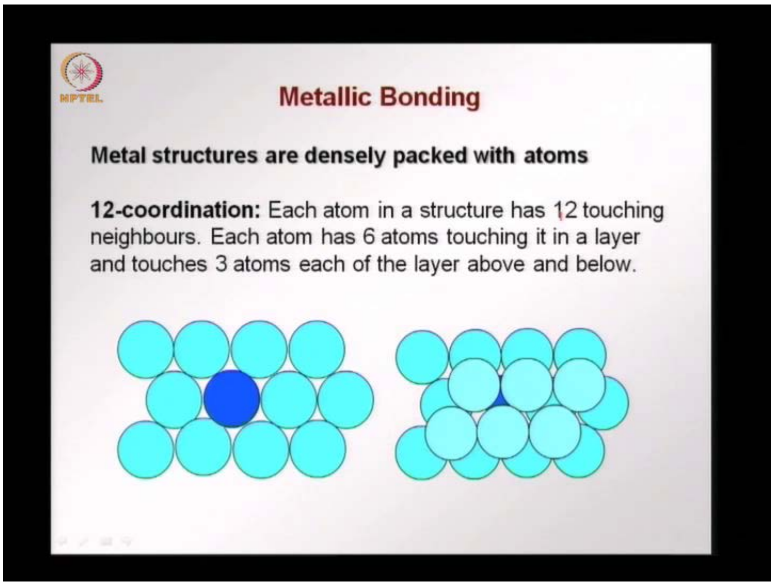

Metallic Bonding

Metal structures are densely packed with atoms

12-coordination: Each atom in a structure has 12 touching neighbours. Each atom has 6 atoms touching it in a layer and touches 3 atoms each of the layer above and below.

Now we can have different configurations, what we saw in the previous slide is that the positively charged ions have to be tact together to give the crystal configuration. So, this packing together can occur a different ways and one such is what is called the twelve coordination where we have each atom having twelve neighbors touching. So, in one plane we have this atom surrounded by six other atoms now when we think of the plane above and below it we will have three atoms touching this particular atom. So, we have twelve atoms touching the one in the middle giving rise to what we call the twelve coordination which is a densely packed structure. So, you can imagine that the atoms are all densely packed in this particular configuration.

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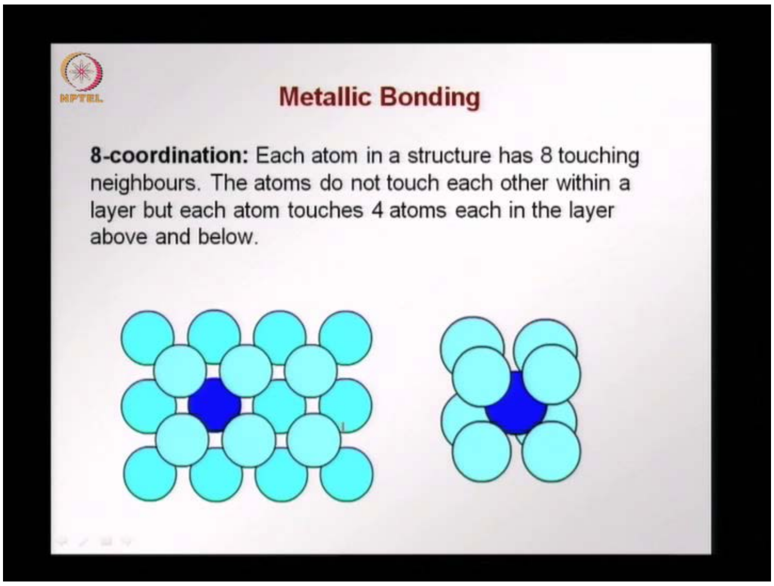

Metallic Bonding

8-coordination: Each atom in a structure has 8 touching neighbours. The atoms do not touch each other within a layer but each atom touches 4 atoms each in the layer above and below.

Another configuration which also gives rise to a densely packed structures is the eight coordination where each atom as a structure where eight neighboring atoms are touching. So, you have here in in one plane you have the atom and on the plane above and below it you have four atoms touching this atom denoted here and so, here you see how the dark blue atom as eight other atoms touching it this is not as closely packed as a twelve coordination, but also gives rise to a dense structure within the crystal lattice.

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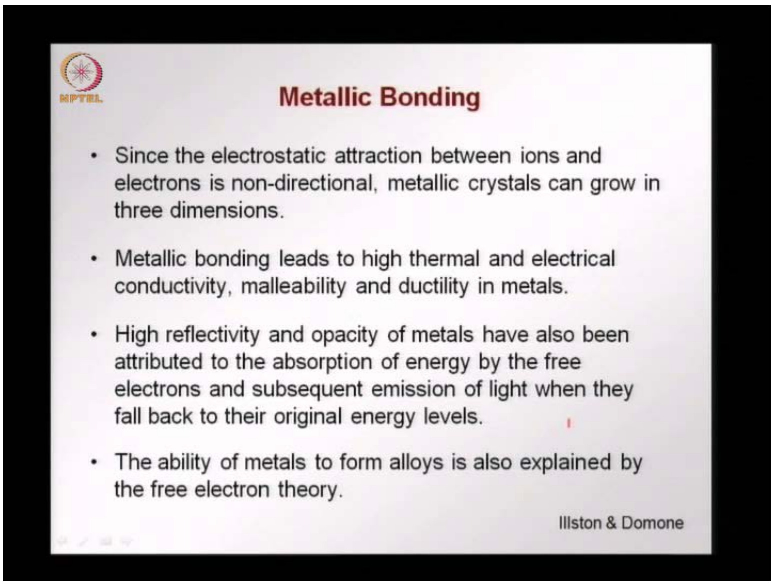

Metallic Bonding

- Since the electrostatic attraction between ions and electrons is non-directional, metallic crystals can grow in three dimensions.
- Metallic bonding leads to high thermal and electrical conductivity, malleability and ductility in metals.
- High reflectivity and opacity of metals have also been attributed to the absorption of energy by the free electrons and subsequent emission of light when they fall back to their original energy levels.
- The ability of metals to form alloys is also explained by the free electron theory.

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Now in terms of directionalities, the metallic bond is non-directional. The electrostatic attraction is such that between the ions electrons such that it gives rise to a non-directional bond and therefore, metallic crystals can grow in three-dimensions in all directions. So, this is very important because this gives rise to the fact that we have isotropic properties within the metal; that means, the properties are same in all directions and we can have different shapes and sizes being built up with a metal. The type of bond and its properties also gives rise to the heigh thermal electrical conductivity of the metallic bond, its malleability and ductility and the thermal conductivity and the electrical conductivity come about because we have electrons that are freely floating around in the crystal lattice.

And when a electricity has to be transferred this electrons have a good amount of mobility within the crystal. So, they can move and translate the heat from one and of the body to the other and the malleability and the ductility comes about, because we saw that the crystal arrangements was in sheets that are closely packed and these sheets later on we will seen more details can slip on one another. So, this gives rise to both malleability and ductility malleability is the ability to create any form needed ductility is the ability to strain a lot before failing that it is a lot of elongation. So, we find that these crystal planes can slip without breaking of many bonds and this gives rise to malleability and ductility also we find that the metallic bond due to its nature gives high reflectivity and opacity to metals.

We know that metals are generally shiny; that means, they have high reflectivity; that means, they reflect light and they are all. So, opaque they do not let like through and this is attributed to the energy absorption by the free electrons and the emission of light when the energy levels fall back that is the electrons can observe energy and give it back. So, this gives rise to the reflectivity. They also do not let the light through the body of the metal making it opaque, also we find that since there is this sort of a lose structure within the crystal in the sense that there are electrons floating around there is not a one to one bond between the different atoms the metals form alloys easily or relatively easily we can mix different metal atoms together to form an alloy. So, these are some of the very important features of the metallic bond that gives rise to the behavior that we see later in the metals.

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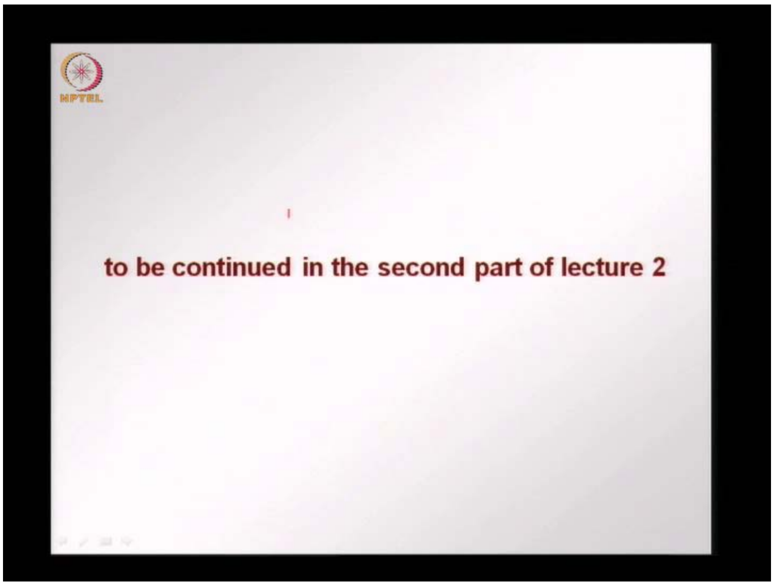

to be continued in the second part of lecture 2

So, to summarize we have looked at three types of strong bonds - the ionic bond, the covalent bond and the metallic bond. In the ionic bond, we had the bond forming by the donation or the capture or the acceptance of electrons. We looked at the examples of sodium chloride sodium atom loses one electron to chlorine. And we have an ionic bond form between sodium and chlorine, and we saw that the ionic bond is strong and become stronger with more electrons being transferred or exchange in the process of the ionic bond. The ionic bond also we saw has non directional bond and it gives rise to symmetric structure that can grow in all direction.

Next what we look at was the covalent bond where we had the sharing of electrons to form the stable octet structure, and we saw that this was a directional bond it gives rise to chains and sheets except in the case of silica and diamond where we had a network structure of the three-dimensional structure, and some of the important aspects of the covalent bond is that the covalently bonded materials are generally strong, because there is not a strong bond between the sheets or between the chains. We can have long chains large sheets, but they do not bond to each other like for example, graphite and that is the reason why we have the sheets which can slip against each other. And the long molecular chains which can bend bend we put some force on it and the directionalities maintained, when they are bending will come back and talk about this when we look at the polymer structure later on.

Then the third strong bond that we looked as the metallic bond which as a special type of the covalent bond very instead of having specific bonds between the atoms, what we have is a loosely bonded material in the sense that we have the electrons floating around with the cat ions in a certain configuration, but the electrons floating around with then the crystal lattice and this gives rise to very special properties that we use a lot in engineering. For example, the malleability and the ductility where we have the sheets the planes within the crystal that can slip giving it deformability the metal crystal deform a lot without breaking because these sheets slip against one another. We also have very good thermal and electrical conductivity, because we have these electrons floating around in the crystal structure and this can transfer heat from one part of the material to the other. We also have due to the metallic bond the fact that metals have very high reflectivity and opacity the electron observe the energy and give it back as the fallback to another energy level.

So, metals as we all know can have very shiny surfaces they can reflect light back and they also do not let like through because you have a compact structure you have the electrons absorbing the energy and not letting it through. So, these where the three strong bonds we also saw in the introduction that there are two weak bonds and we will talk about it in next part of this lecture will talk about the van der waals bonds and the hydrogen bonds both of these are weak bonds or what they have called as also as secondary bonds with a strong bonds being called the primary bonds. So, we will take that up in the second part of the lecture.

Thank you.