Solid State Physics
Lecture 20
Metallic bonds

Hello, we have so far discussed three different kind of interaction, in the that are responsible for binding of solids. So, we have discussed the van der Waals and Repulsive that is Lennard Jones kind of interaction. We have also discussed ionic interaction between two different ions unlike and also, similar ions the. So, the different ions, different charges that contribute to the attractive force and similar ions that contribute to the repulsive force. We have also discussed the covalent bonds. (Refer Slide Time: 01:07)

After having discussed all these things, now we will discuss about Metallic bonds. As the name suggests you can easily understand that metallic bonds are found in metals. So, what happens in metals? It is one kind of element say iron that forms a metal. There is no ion in there. So, there is no two dissimilar ions of different charges and that attract and similar charges that repel, nothing like that happens. And you can notice that iron is pretty strong, it is not as weak as the inert gas crystals. So, the interaction must be something very strong. What contributes to this interaction in case of metals? Let us try to understand that. So, metal is characterized by its high electrical conductivity and a large number of electrons in a metal are free to move around. So, of course, 1 or 2 electrons per atom is free to move around; but if you consider a metal block, 1 or 2 per atom would give you something of the order of Avogadro number per mole, which is pretty large number and those many electrons can move around almost freely without experiencing much of a potential. So, the strength of the metal must come from this something like free electron movement in case of the metal. These electrons that are free to move around are called the conduction electrons. Now, what happens? In some metals, the interaction of the ion cores with the conduction electrons that makes a large contribution in the binding energy; but the characteristic feature of metallic bonding is lowering the energy of the system, because of the valence electrons compared to the free atoms. That means, motion of these electrons that makes the total energy of the system lower than that of elemental atoms. The binding energy of let us say an alkali metal examples are potassium, sodium and so on. So, if you consider an alkali metal crystal, it is much the binding energy is much smaller than for example, an alkali halide like sodium chloride. So, sodium chloride is much more strongly bound than sodium. So, if you are also familiar with their reactivity, if you have sodium in open air; but and if it comes in contact with some moisture from the air, it starts an aggressive reaction. Sodium chloride does nothing of that sort; sodium chloride just dissolves into water. But sodium elemental sodium crystal that starts a fire, once you put it into water and that happens because its binding energy is much smaller than for example sodium chloride. The bond formed by the conduction electron that is very strong. The inter atomic distances that is relatively large in alkali metal because the kinetic energy of the conduction electrons that is lower at large inter atomic distances. Now, metals tend to crystallize in relatively closed packed structures. So, they have hcp formation or fcc formation, iron in normal temperature is found in bcc kind of a lattice; but nothing more sparse than this. As opposed to that if you consider a diamond structure which is found for carbon in diamond form, silicon germanium, it is rather sparse the packing fraction is much less for a diamond structure compared to hcp, fcc, bcc, this kind of structures. And in case of transition metals like iron, there is additional binding from the inner electron shell. For example, in case of iron you would see there are many d-electrons in the system and those electrons offer a stronger binding of the system. That is the reason iron is so strong, transition metals are so strong as opposed to alkali metals. The large d shell in the transition metals contribute to the large binding energy making iron very strong as we have already discussed. So, this is the metallic bond that we wanted to discuss. This is something very important and one other kind of bond is left in our discussion so far that is hydrogen bond. Hydrogen bond is certainly not the bond between two hydrogen atoms to form a hydrogen molecule. We have already discussed that the bond between two hydrogen atoms is the covalent kind of a bond. So, what is hydrogen bond? Because natural hydrogen gas has natural hydrogen not in gas; in elemental form, it has only 1 electron, it should form a covalent bond with only one other atom. However, there are situations, where a hydrogen atom is attracted by rather strong forces of two atoms. Now, hydrogen bond is largely ionic in character and that is formed because of in when hydrogen comes in contact with electro negative systems. So, if you consider fluorine, oxygen or nitrogen, these are very electronegative materials and when they form a covalent bond, they attract most of the electrons to their nuclei and in that kind of a situation, it is a pseudo ionic kind of arrangement that is there. It is never completely covalent kind of a bond. Well, in this context, I must say that pure covalent bonds are formed between like atoms. For example, two hydrogen atoms in between them, the pure covalent bond can be formed; two carbon atoms, it can be formed. But when it comes to oxygen and hydrogen, it is not pure covalent kind of bond; it is some it is partially ionic, partially covalent that kind of a bond is formed in there and oxygen attracts most of the electrons. So, what happens? In the extreme ionic form of the hydrogen bond, the hydrogen atom loses its electron to another atom in the molecule and then the bare proton is forms the hydrogen bond with multiple of other atoms like fluorine, oxygen, nitrogen whatever. The atom adjacent to the proton are so close that no more than two of them would get to each other and thus, hydrogen bond connects only two other atoms; for example, two oxygen atoms. So, in what circumstances, do we see hydrogen bond in water? In water, liquid from hydrogen bond is not there, when we condense water into ice, hydrogen bonds are formed and because hydrogen bonds are formed, the volume expands instead of contracting upon formation of solid. This is something pretty interesting in case of water. After discussing hydrogen bond, let us discuss atomic or ionic radii. So, what is atomic or ionic radii? As the name suggests it would be the radius of a sphere around in which the atom is confined or the ion is confined. So, the distance between atoms in crystal that can be measured very accurately by X-ray diffraction we have seen that. However, can we make any estimate of the spread of the electron cloud? In other words, can a definite meaning be assigned to the radius of an atom or an ion irrespective of the nature and the composition of the crystal? Strictly speaking, the answer is no, we cannot do that; but the charge, we cannot do that because the charge distribution around an atom is not limited to a rigid spherical boundary. But the concept of an atomic radius is still fruitful in predicting interatomic spacing if. So, it is rather a self-consistent scheme in which we form we describe atom, we associate an atom with an atomic radius under certain circumstances and that somewhat predicts what would be the lattice constant for a new compound formed by two of the atoms that we choose or more than two of the atoms or ions that we choose. To make prediction of lattice constant, it is convenient to assign sets of self consistent radii to various types of bonds; a set for ionic crystals with 6 coordination say, another set for crystals with 12 coordination, another set for covalent bonds like that and that helps us understand the packing of a crystal. So, here, we conclude our discussion about the binding of crystals.