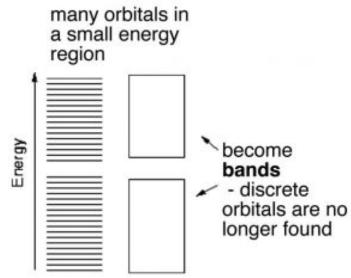
3.4: Metals

Metals have quite a wide range of properties at normal temperatures, from liquid (like mercury) to extremely hard (like tungsten). Most are shiny but not all are colorless. For example gold and copper have distinct colors. All metals conduct electricity but not all equally. How can we explain all these properties? Let us use aluminum (Al) as an example because most of us have something made of aluminum such as a pan or aluminum foil. With modern instrumentation it is quite easy to visualize atoms and a variety of techniques have been used to image where the aluminum atoms are in the solid structure. What emerges is a picture of aluminum nuclei and their core electrons, packed like spheres where one layer of spheres rests in the interstices of the underlying and overlying layers; where the positions of the electrons are within this structure not well defined.



In $H-H$ or diamond the electrons involved in bonding are located (most probably) between the two nuclei. In contrast in aluminum and other metals the valence electrons are not closely associated with each nucleus. Instead they are dispersed over the whole macroscopic piece of metal. Imagine that instead of two or three or four atomic orbitals combining to form MOs, a mole (6×10^{23}) of atomic orbitals were combined to produce a mole of MOs. As more and more MOs are formed the energies between them gets smaller and smaller. For a macroscopic piece of metal (one you can see) the energy gap between the individual bonding MOs will be negligible for all intents and purposes. These orbitals produce what is essentially a continuous band of (low-energy) bonding MOs and a continuous band of (higher-energy) anti-bonding MOs. The energy gap between the bonding and anti-bonding orbitals is called the band-gap and in a metal this band-gap is quite small (recall that the gap between the bonding and anti-bonding MOs in diamond is very large). Moreover in metals the bonding MOs (known as the valence band) are able to accommodate more electrons. This is because in metals there are typically fewer electrons than there are atomic orbitals. Consider aluminum: it has three valence electrons and in the ground (lowest energy) state has an electron configuration of $3s^2 3p^1$. This suggests that it has two unoccupied $3p$ orbitals. We can consider the bonding MOs in aluminum to be formed from all the available atomic orbitals, which means that there are many bonding MOs that are not occupied by electrons. The physical consequences of this are that the valence electrons can move relatively easily from one MO to another because their energies are very close together. Whereas nuclei and core electrons remain more or less locked in position the valence electrons can spread out to form a kind of electron sea within the metal. When an electrical potential is applied across the metal, electrons from an external source can easily enter the valence band and electrons can just as easily leave the metal. Electrical conductivity is essentially a measure of how easily electrons can flow through a substance. Metals typically have high conductivity due to the ease with which electrons can move from one MO to another and the fact that each MO extends throughout the whole piece of metal. Because the numbers of electrons entering and leaving are the same, the piece of metal remains uncharged.

In this model the atomic cores are packed together and surrounded by a cloud of electrons that serve as the “glue” that binds them together. There are no discrete bonds in this type of structure. When a piece of metal is put under physical stress (for example it is stretched or deformed) the atoms can move relative to one another but the electrons remain spread throughout the structure. Metals can often be slowly deformed into different shapes without losing their structural integrity or electrical conductivity—they are malleable! They can be melted (increased atomic movement), become liquid, and then allowed to cool until they solidify; throughout this process they retain their integrity and their metallic properties and so continue to conduct electricity.[17] This is quite different from how other substances (such as diamond or water) behave. The hardness of a solid metal depends on how well its atoms packed together and how many electrons are contributed to the valence band of orbitals.

So why do some elements behave as metals and others do not? For example graphite conducts electricity but it is not malleable and can't be heated and molded into other shapes. The answer lies in the behavior of the MOs and the resulting bonds they can produce. Graphite has a rigid backbone of carbon–carbon bonds that makes it strong and stable but overlaying those bonds is the set of delocalized MOs that spread out over the whole sheet. As a result graphite has some properties that are similar to diamond (stability and strength), some that are similar to metals (electrical conductivity), and some that are a consequence of its unique sheet structure (slipperiness).

Why Are Metals Shiny?

We see things because photons hit the back of our retinas and are absorbed by specialized molecules (proteins and associated pigment molecules). This leads to changes in protein structure and initiates a cascade of neuron-based cellular events that alters brain activity. So where do these photons come from? First and foremost they can be emitted from a source (the Sun, a light bulb, etc.) that appears to shine and can be seen in the dark. Alternatively, photons can be reflected off a surface; in fact most of the things we see do not emit light, but rather reflect it. A red T-shirt appears red because it absorbs other colors and reflects red light. Photons can also be refracted when they pass through a substance. A cut diamond sparkles because light is refracted as it passes through the material and exits from the many facets. Refraction is caused when photons bump into electrons, are absorbed, and then (very shortly thereafter) are re-emitted as they travel through a material. These processes take time, so the apparent speed of light slows down. It can take a photon many thousands of years to move from the core to the surface of the Sun because of all the collisions that it makes during the journey.[18]

To explain why metals (and graphite) are shiny, we invoke a combination of reflection, refraction, and the energy levels of MOs. When a photon of light is absorbed and reemitted, the electron moves from one orbital to another. Let us consider a piece of metal at room temperature. When a photon arrives at the metal's surface it encounters the almost continuous band of MOs. Most photons, regardless of their wavelength, can be absorbed because there is an energy gap between orbitals corresponding to the energy of the photon. This process promotes electrons up to a higher energy level. As the electrons drop back down to a lower energy level, the photons are re-emitted, resulting in the characteristic metallic luster. Metals actually emit light, although this does not mean metals glow in the dark (like a light bulb or the Sun). Instead, metals absorb and re-emit photons, even at room temperature.

The color of a particular metal depends upon the range of wavelengths that are re-emitted. For most metals the photons re-emitted have a wide range of wavelengths which makes the metallic surface silvery. A few metals, such as copper and gold, absorb light in the blue region and re-emit light with wavelengths that are biased toward the red end region of the spectrum ($400-700\ \mathrm{nm}$) and therefore they appear yellowish. This is due to relativistic effects way beyond the scope of this book, but something to look forward to in your future physical chemistry studies!

Now we can also understand why metals emit light when they are heated. The kinetic energy of the atoms increases with temperature which promotes electrons from low to higher energy orbitals. When these electrons lose that energy by returning to the ground state, it is emitted as light. The higher the temperature the shorter the wavelength of the emitted light. As a filament heats up, it first glows red and then increasing whiter as photons of more and more wavelengths are emitted.

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Questions to Answer

- What properties indicate that a substance is metallic?
- Why are metals shiny?
- How can metallic properties be explained by the atomic-molecular structure of Al (for example)?
- Why can we see through diamond but not aluminum? How about graphite?
- Why does aluminum (and for that matter all metals) conduct electricity? What must be happening at the atomic-molecular scale for this to occur?
- What does the fact that diamond doesn't conduct electricity tell you about the bonding in diamond?
- How do the bonding models for diamond and graphite explain the differences in properties between diamond, graphite, and a metal like aluminum?
- Why is it OK to use different models to describe bonding in different species?

This chapter has brought us to a point where we should have a fairly good idea of the kinds of interactions that can occur among atoms of the same element. We have seen that the properties of different elements can be explained by considering the structure of their atoms and in particular the way their electrons behave as the atoms interact to form molecules or large assemblies of atoms (like diamond.) What we have not considered yet is how atoms of different elements interact to form compounds (substances that have more than one element). In Chapter 4 we will take up this subject and much more.