**Metallic Bonding**

The metallic bond is a type of chemical bond that occurs between atoms of metallic elements. It gives metals their unique properties that we don't see in non-metal substances, as you'll learn in this lesson.

**Definition**

If you have already learned about covalent and ionic bonding, you know that these bonds occur between two atoms. When two atoms share electrons, they form a covalent bond. When one atom takes an electron away from another and the resulting positive and negative ions are attracted to each other, those atoms have formed an ionic bond.

A metallic bond is pretty different from covalent and ionic bonds, but the goal is the same: to achieve a lower energy state. Instead of a bond between just two atoms, a metallic bond is a sharing of electrons between many atoms of a metal element.

Take a look at your desk and see if you can find a small piece of metal like a paper clip or a staple. All of the atoms in that small piece of metal are sharing a big pool of valence electrons known as a sea of electrons or delocalized electrons. The big pool is like a free-for-all in that any valence electron can move to any atom within the material.

The metallic bond is not the easiest type of bond to understand, so an analogy might help. Imagine filling your bathtub with golf balls. Fill it right up to the top. The golf balls will arrange themselves in an orderly fashion as they fill the space in the tub. Do you see any spaces between the balls? If you turn on the faucet and plug the drain, the water will fill up those spaces. What you now have is something like metallic bonding. The golf balls are the metal atoms, and the water represents the valence electrons shared by all of the atoms.

Once the valence electrons detach from their original atomic owners and float around in the sea, the metal atoms become positive ions. The result is an orderly structure of positive metal atoms surrounded by a sea of negative electrons that hold the ions together like glue.

**The nature of metallic bonding**

The combination of two phenomena gives rise to metallic bonding: [delocalization of electrons](https://en.wikipedia.org/wiki/Delocalized_electron) and the availability of a far larger number of delocalized energy states than of delocalized electrons.[[clarification needed](https://en.wikipedia.org/wiki/Wikipedia%3APlease_clarify)] The latter could be called [electron deficiency](https://en.wikipedia.org/wiki/Electron_deficiency).

**In 2D**

[Graphene](https://en.wikipedia.org/wiki/Graphene) is an example of two-dimensional metallic bonding. Its metallic bonds are similar to [aromatic bonding](https://en.wikipedia.org/wiki/Aromaticity) in [benzene](https://en.wikipedia.org/wiki/Benzene), [naphthalene](https://en.wikipedia.org/wiki/Naphthalene), [anthracene](https://en.wikipedia.org/wiki/Anthracene%22%20%5Co%20%22Anthracene), [ovalene](https://en.wikipedia.org/wiki/Ovalene%22%20%5Co%20%22Ovalene), and so on.

**In 3D**

[Metal aromaticity](https://en.wikipedia.org/wiki/Metal_aromaticity) in [metal clusters](https://en.wikipedia.org/wiki/Metal_cluster) is another example of delocalization, this time often in three-dimensional entities. Metals take the delocalization principle to its extreme and one could say that a crystal of a metal represents a single molecule over which all conduction electrons are delocalized in all three dimensions. This means that inside the metal one can generally not distinguish molecules, so that the metallic bonding is neither intra- nor intermolecular. 'Nonmolecular' would perhaps be a better term. Metallic bonding is mostly non-polar, because even in [alloys](https://en.wikipedia.org/wiki/Alloys) there is little difference among the [electronegativities](https://en.wikipedia.org/wiki/Electronegativity%22%20%5Co%20%22Electronegativity) of the [atoms](https://en.wikipedia.org/wiki/Atom) participating in the bonding interaction (and, in pure elemental metals, none at all). Thus, metallic bonding is an extremely delocalized communal form of covalent bonding. In a sense, metallic bonding is not a 'new' type of bonding at all, therefore, and it describes the bonding only as present in a chunk of condensed matter, be it crystalline solid, liquid, or even glass. Metallic vapors by contrast are often atomic ([Hg](https://en.wikipedia.org/wiki/Mercury_%28element%29)) or at times contain molecules like [Na2](https://en.wikipedia.org/wiki/Sodium) held together by a more conventional covalent bond. This is why it is not correct to speak of a single 'metallic bond'.[[clarification needed](https://en.wikipedia.org/wiki/Wikipedia%3APlease_clarify%22%20%5Co%20%22Wikipedia%3APlease%20clarify)]

The delocalization is most pronounced for [s- and p-electrons](https://en.wikipedia.org/wiki/Azimuthal_quantum_number). For [caesium](https://en.wikipedia.org/wiki/Caesium) it is so strong that the electrons are virtually free from the caesium atoms to form a gas constrained only by the surface of the metal. For caesium, therefore, the picture of Cs+ ions held together by a negatively charged [electron gas](https://en.wikipedia.org/wiki/Nearly-free_electron_model) is not too inaccurate. For other elements the electrons are less free, in that they still experience the potential of the metal atoms, sometimes quite strongly. They require a more intricate quantum mechanical treatment (e.g., [tight binding](https://en.wikipedia.org/wiki/Tight_binding)) in which the atoms are viewed as neutral, much like the carbon atoms in benzene. For d- and especially f-electrons the delocalization is not strong at all and this explains why these electrons are able to continue behaving as [unpaired electrons](https://en.wikipedia.org/wiki/Unpaired_electron) that retain their spin, adding interesting [magnetic properties](https://en.wikipedia.org/wiki/Magnetism) to these metals.

**Electron deficiency and mobility**

Metal [atoms](https://en.wikipedia.org/wiki/Atoms) contain few [electrons](https://en.wikipedia.org/wiki/Electron) in their [valence shells](https://en.wikipedia.org/wiki/Electron_shell#Valence_shells) relative to their periods or [energy levels](https://en.wikipedia.org/wiki/Energy_level). They are [electron deficient](https://en.wikipedia.org/wiki/Electron_deficiency) elements and the communal sharing does not change that. There remain far more available energy states than there are shared electrons. Both requirements for conductivity are therefore fulfilled: strong delocalization and partly filled energy bands. Such electrons can therefore easily change from one energy state into a slightly different one. Thus, not only do they become delocalized, forming a sea of electrons permeating the structure, but they are also able to migrate through the structure when an external electrical field is imposed, leading to electrical conductivity. Without the field, there are electrons moving equally in all directions. Under the field, some will adjust their state slightly, adopting a different [wave vector](https://en.wikipedia.org/wiki/Wave_vector). As a consequence, there will be more moving one way than the other and a net current will result.

The freedom of conduction electrons to migrate also give metal atoms, or layers of them, the capacity to slide past each other. Locally, bonds can easily be broken and replaced by new ones after the deformation. This process does not affect the communal metallic bonding very much. This gives rise to metals' typical characteristic phenomena of [malleability](https://en.wikipedia.org/wiki/Malleability) and [ductility](https://en.wikipedia.org/wiki/Ductility). This is particularly true for pure elements. In the presence of dissolved impurities, the defects in the structure that function as cleavage points may get blocked and the material becomes harder. Gold, for example, is very soft in pure form (24-[karat](https://en.wikipedia.org/wiki/Carat_%28purity%29)), which is why alloys of 18-karat or lower are preferred in jewelry.

Metals are typically also good conductors of heat, but the conduction electrons only contribute partly to this phenomenon. Collective (i.e., delocalized) vibrations of the atoms known as [phonons](https://en.wikipedia.org/wiki/Phonon) that travel through the solid as a wave, contribute strongly.

However, the latter also holds for a substance like [diamond](https://en.wikipedia.org/wiki/Diamond). It conducts heat quite well but not electricity. The latter is not a consequence of the fact that delocalization is absent in diamond, but simply that carbon is not electron deficient. The electron deficiency is an important point in distinguishing metallic from more conventional covalent bonding. Thus, we should amend the expression given above into: Metallic bonding is an extremely delocalized communal form of electron deficient covalent bonding.

**Metallic radius**

Metallic radius is defined as one-half of the distance between the two adjacent metal ions in the metallic structure. This radius depends on the nature of the atom as well as its environment—specifically, on the [coordination number](https://en.wikipedia.org/wiki/Coordination_number) (CN), which in turn depends on the temperature and applied pressure.

When comparing periodic trends in the size of atoms it is often desirable to apply so-called Goldschmidt correction, which converts the radii to the values the atoms would have if they were 12-coordinated. Since metallic radii are always biggest for the highest coordination number, correction for less dense coordinations involves multiplying by x, where 0 < x < 1. Specifically, for CN = 4, x = 0.88; for CN = 6, x = 0.96, and for CN = 8, x = 0.97. The correction is named after [Victor Goldschmidt](https://en.wikipedia.org/wiki/Victor_Goldschmidt) who obtained the numerical values quoted above. The radii follow general [periodic trends](https://en.wikipedia.org/wiki/Periodic_trends): they decrease across the period due to increase in the [effective nuclear charge](https://en.wikipedia.org/wiki/Effective_nuclear_charge), which is not offset by the increased number of [valence electrons](https://en.wikipedia.org/wiki/Valence_electrons). The radii also increase down the group due to increase in [principal quantum number](https://en.wikipedia.org/wiki/Principal_quantum_number). Between rows 3 and 4, the [lanthanide contraction](https://en.wikipedia.org/wiki/Lanthanide_contraction) is observed – there is very little increase of the radius down the group due to the presence of poorly [shielding](https://en.wikipedia.org/wiki/Electron_shielding) [f orbitals](https://en.wikipedia.org/wiki/Atomic_orbitals).

**Strength of the bond**

The atoms in metals have a strong attractive force between them. Much energy is required to overcome it. Therefore, metals often have high boiling points, with [tungsten](https://en.wikipedia.org/wiki/Tungsten) (5828 K) being extremely high. A remarkable exception is the elements of the [zinc group](https://en.wikipedia.org/wiki/Group_12_element): Zn, Cd, and Hg. Their electron configuration ends in ...ns2 and this comes to resemble a noble gas configuration like that of [helium](https://en.wikipedia.org/wiki/Helium) more and more when going down in the periodic table because the energy distance to the empty np orbitals becomes larger. These metals are therefore relatively volatile, and are avoided in [ultra-high vacuum](https://en.wikipedia.org/wiki/Ultra-high_vacuum) systems.

Otherwise, metallic bonding can be very strong, even in molten metals, such as [Gallium](https://en.wikipedia.org/wiki/Gallium). Even though gallium will melt from the heat of one's hand just above room temperature, its boiling point is not far from that of copper. Molten gallium is, therefore, a very nonvolatile liquid thanks to its strong metallic bonding. The strong bonding of metals in the liquid form demonstrates that the energy of a metallic bond is not a strong function of the direction of the metallic bond; this lack of bond directionality is a direct consequence of electron delocalization, and is best understood in contrast to the directional bonding of covalent bonds. The energy of a metallic bond is thus mostly a function of the number of electrons which surround the metallic atom, as exemplified by the [Embedded atom model](https://en.wikipedia.org/wiki/Embedded_atom_model).[[8]](https://en.wikipedia.org/wiki/Metallic_bonding#cite_note-8) This typically results in metals assuming relatively simple, close-packed crystal structures, such as FCC, BCC, and HCP. Given high enough cooling rates and appropriate alloy composition, metallic bonding can occur even in [glasses](https://en.wikipedia.org/wiki/Metallic_glass) with an amorphous structure.

Much biochemistry is mediated by the weak interaction of metal ions and biomolecules. Such interactions and their associated [conformational change](https://en.wikipedia.org/wiki/Conformational_change) has been measured using [dual polarisation interferometry](https://en.wikipedia.org/wiki/Dual_polarisation_interferometry).

**Solubility and compound formation**

Metals are insoluble in water or organic solvents unless they undergo a reaction with them. Typically this is an oxidation reaction that robs the metal atoms of their itinerant electrons, destroying the metallic bonding. However metals are often readily soluble in each other while retaining the metallic character of their bonding. Gold, for example, dissolves easily in mercury, even at room temperature. Even in solid metals, the solubility can be extensive. If the structures of the two metals are the same, there can even be complete solid [solubility](https://en.wikipedia.org/wiki/Solubility), as in the case of [electrum](https://en.wikipedia.org/wiki/Electrum), the alloys of silver and gold. At times, however, two metals will form alloys with different structures than either of the two parents. One could call these materials [metal compounds](https://en.wikipedia.org/wiki/Intermetallics), but, because materials with metallic bonding are typically not molecular, Dalton's [law of integral proportions](https://en.wikipedia.org/wiki/Law_of_definite_proportions) is not valid and often a range of stoichiometric ratios can be achieved. It is better to abandon such concepts as 'pure substance' or 'solute' is such cases and speak of [phases](https://en.wikipedia.org/wiki/Phase_%28matter%29) instead. The study of such phases has traditionally been more the domain of [metallurgy](https://en.wikipedia.org/wiki/Metallurgy) than of [chemistry](https://en.wikipedia.org/wiki/Chemistry), although the two fields overlap considerably.

**Localization and clustering: from bonding to bonds**

The metallic bonding in complicated compounds does not necessarily involve all constituent elements equally. It is quite possible to have an element or more that do not partake at all. One could picture the conduction electrons flowing around them like a river around an island or a big rock. It is possible to observe which elements do partake, e.g., by looking at the core levels in an [X-ray photoelectron spectroscopy](https://en.wikipedia.org/wiki/X-ray_photoelectron_spectroscopy) (XPS) spectrum. If an element partakes, its peaks tend to be skewed.

Some intermetallic materials e.g. do exhibit [metal clusters](https://en.wikipedia.org/wiki/Metal_cluster), reminiscent of molecules and these compounds are more a topic of chemistry than of metallurgy. The formation of the clusters could be seen as a way to 'condense out' (localize) the electron deficient bonding into bonds of a more localized nature. [Hydrogen](https://en.wikipedia.org/wiki/Hydrogen) is an extreme example of this form of condensation. At high pressures [it is a metal](https://en.wikipedia.org/wiki/Metallic_hydrogen). The core of the planet [Jupiter](https://en.wikipedia.org/wiki/Jupiter) could be said to be held together by a combination of metallic bonding and high pressure induced by gravity. At lower pressures however the bonding becomes entirely localized into a regular covalent bond. The localization is so complete that the (more familiar) H2 gas results. A similar argument holds for an element like boron. Though it is electron deficient compared to carbon, it does not form a metal. Instead it has a number of complicated structures in which [icosahedral](https://en.wikipedia.org/wiki/Icosahedron) B12 clusters dominate. [Charge density waves](https://en.wikipedia.org/wiki/Charge_density_wave) are a related phenomenon.

As these phenomena involve the movement of the atoms towards or away from each other, they can be interpreted as the coupling between the electronic and the vibrational states (i.e. the phonons) of the material. A different such electron-phonon interaction is thought to cause a very different result at low temperatures, that of [superconductivity](https://en.wikipedia.org/wiki/Superconductivity). Rather than blocking the mobility of the charge carriers by forming [electron pairs](https://en.wikipedia.org/wiki/Electron_pair) in localized bonds, Cooper-pairs are formed that no longer experience any resistance to their mobility.