Elements of METALLURGY and ENGINEERING ALLOYS

Edited by
F.C. Campbell

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Preface

This book is intended for those wishing to learn more about metallurgy and the metals and engineering alloys used throughout industry. It would be useful to almost anyone who deals with metals, including designers, structural engineers, materials and process engineers, manufacturing engineers, and production personnel. It was written so that both the fundamentals of metallurgy and the specifics of the important engineering alloys could be combined in a single book. Throughout, the emphasis of this book is on practical principles of metallurgy and the engineering alloys.

It should be noted that it is not the objective of this book to replace handbooks on engineering alloys. Rather, the goal is to gather information on engineering alloys in one book and edit it into a uniform format so that the reader is able to compare the metallurgy, properties, and applications of the most important engineering alloy systems. Many of the chapters are adapted from the *ASM Handbook* series. When this was not possible, I wrote chapters to meet this objective.

The book is separated into two parts. Part I covers the fundamentals of physical and mechanical metallurgy, while Part II goes into greater detail for specific metals and their alloys.

The first eleven chapters cover most of the basics of physical metallurgy. Physical metallurgy may be described as the study of how metallic structure and processing affect the final properties of metals and alloys. Chapters on physical metallurgy include metallic structure; crystalline imperfections and plastic deformation; solid solutions; diffusion; phase transformations; phase diagrams; solidification and casting; recovery, recrystallization, and grain growth; precipitation hardening; the iron-carbon system; and the heat treatment of steel. This is followed by five chapters dealing with mechanical metallurgy, which is the study of how loads, either during fabrication or in service, affect metals and their properties. Chapters on mechanical metallurgy include mechanical behavior, fracture, fatigue, creep, and deformation processing. The last two chapters in Part I cover physical properties and corrosion.

Part II deals with specific engineering alloys. The first six chapters cover the important ferrous metal alloys: plain carbon steels, alloy steels, surface hardening of steels, tool materials, stainless steels, and cast irons. The next seven chapters are on the nonferrous metals and their alloys: copper, aluminum, magnesium and zinc, titanium, nickel and cobalt, superalloys, and refractory metals. This is followed by a chapter on several miscellaneous nonferrous metals and alloys. The last chapter deals with metal-matrix composites.

A first course in materials science would be helpful in understanding the material in this book; however, most of the material is self-explanatory and builds on itself as one progresses through the book.

I would like to acknowledge the help and guidance of the ASM International staff. Steve Lampman and Charles Moosbrugger championed the book through the approval process at ASM. Eileen De Guire, with help from Charles Moosbrugger, did the majority of the editorial work. Ann Britton provided help and guidance with the permissions. The efforts of Madrid Tramble, Kathryn Muldoon, and Liz Marquard in converting the manuscript into a book are appreciated. I would also like to thank the ASM peer reviewers for their valuable critiques. Finally, I would like to thank my wife Betty for her support and encouragement during this project.

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CHAPTER 1

Metallic Structure

THE WORD *METAL*, derived from the Greek *metallon*, is believed to have originated as a verb meaning to seek, search after, or inquire about. Today, a metal is defined as any element that tends to lose electrons from the outer shells of its atoms. The resulting positive ions are held together in crystalline structure by the cloud of these free electrons in what is known as the metallic bond. The metallic bond yields three physical characteristics typical of solid metals: (1) metals are good conductors of electricity, (2) metals are good conductors of heat, and (3) metals have a lustrous appearance. In addition, most metals are malleable, ductile, and generally denser than other elemental substances. Those elements that do not display the characteristics of the metallic elements are called nonmetals. However, there are some elements that behave as metals under some circumstances and as nonmetals under different circumstances. These are now called semimetals but have also been called metalloids, meaning like metals. The boundaries separating the regions in the periodic table covered by the different classes of elements are not distinct, except that nonmetals never form positive ions. A simplified periodic table is shown in Fig. 1.1, highlighting the elements that are currently considered to be metals.

1.1 Periodic Table

In the periodic table, it is the number of electrons in the outer shell that affects the properties of the elements the most. Those elements that have the same number of electrons

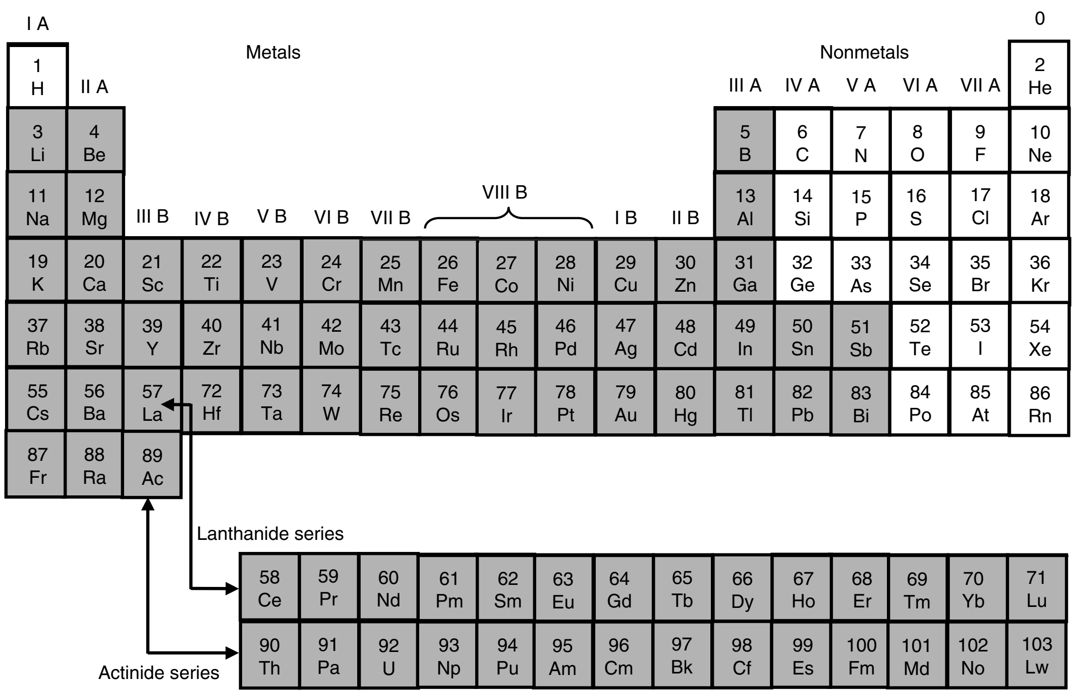

Fig. 1.1 Periodic table of the elements

in their outermost electron shells, and therefore have similar chemical behavior, are placed in columns. For example, lithium, sodium, and potassium each have a single electron in their outer shells and are chemically very similar. They all oxidize very rapidly and react vigorously with water, liberating hydrogen and forming soluble hydroxides. They are physically very similar, being soft, light metals with a somewhat silver color. At the other end of the periodic table, the gases fluorine and chlorine, with seven electrons in their outer shells, also have similar chemical properties. Both are gases with strong nonmetallic properties. At the far right side of the periodic table, the noble gases helium, neon, and argon contain eight electrons in their outer shells. Since this fills the shell, these gases are nonreactive, or inert, under normal circumstances. Therefore, the chemical interaction between elements is governed by the number of electrons present in the outer shell. When the outer shell is filled, the atom has no further tendency to combine or react with other atoms.

Metallic properties depend on both the nature of their constituent atoms and the way in which they are assembled. Assemblies of atoms can be gases, liquids, or solids. When they are in the solid state, metals are normally arranged in a crystalline structure. The crystalline nature of metals is responsible for their ultimate engineering usefulness, and the crystalline arrangement strongly influences their processing. Although metals can exist as single crystals, they are more commonly polycrystalline solids with crystalline grains of repeating atomic packing sequences. Periodic crystalline order is the equilibrium structure of all solid metals. Crystalline structures are a dominant factor in determining mechanical properties, and crystal structures also play an important role in the magnetic, electrical, and thermal properties. The greatest bonding energy occurs when the atoms are closely packed, and the atoms in a crystalline structure tend to pack as densely as possible. In addition, total metallic bonding energy is increased when each atom has the greatest possible number of nearest neighbor atoms. However, due to a shared bonding arrangement in some metals that is partially metallic and partially covalent, some metals do not crystallize into these close-packed structures. Covalent tendencies appear as one moves closer to the nonmetals on the periodic table. As one moves rightward across the periodic table, progressively greater numbers of metals have looser-packed structures. Most metals bordering the nonmetals possess more complex structures with lower packing densities, because covalent bonding plays a large role in determining their crystal structures. The directionality of covalent bonding dictates fewer nearest neighbors than exist in densely packed metallic crystals. For metals near the nonmetals on the right side of the periodic table, where electronegativities are high, covalency becomes a major part of the bonding.

Properties important to the engineer are strongly influenced by crystal structure. One of the most important properties related to crystal structure is ductility. Densely packed structures usually allow motion on one or more slip planes, permitting the metal to deform plastically without fracturing. Ductility is vital for easy formability and for fracture toughness, two properties that give metals a great advantage over ceramic materials for many engineering uses.

1.2 Bonding in Solids

Bonding in solids may be classified as either primary or secondary bonding. Methods of primary bonding include the metallic, ionic, and covalent bonds. Secondary bonds are much weaker bonding mechanisms that are only predominant when one of the primary bonding mechanisms is absent. When two atoms are brought close to each other, there will be a repulsion between the negatively charged electrons of each atom. The repulsion force increases rapidly as the distance of separation decreases. However, when the separation is large, there is attraction between the positive nucleus charge and the negative charge of the electrons. At some equilibrium distance, the attractive and repulsive forces balance each other, and the net force is zero. At this equilibrium distance, the potential energy is at a minimum, as shown in Fig. 1.2. The magnitude of this energy is known as the bond energy, usually expressed in kJ/mol. Primary bond energies range from 100 to 1000 kJ/mol, while the much weaker secondary bonds are on the order of only 1 to 60 kJ/mol. The equilibrium distance, a_0, is the bond length. Strong primary bonds have large forces of attraction, with bond lengths of 1 to 2 Å, while the weaker secondary bonds have larger bond lengths of 2 to 5 Å. While it is convenient to discuss the four major types of bonding separately, it should be recognized that although metallic bonding may be predominant, other types of bonding, in particular covalent bonding, may also be present. A comparison of the some of the properties of the different bond types is given in Table 1.1.

1.2.1 Metallic Bonding

Metallic bonding occurs when each of the atoms of the metal contributes its valence electrons to the formation of an electron cloud that surrounds the positively charged metal ions, as illustrated in Fig. 1.3. Hence, the valence electrons are shared by all of the atoms. In this bond, the positively charged ions repel each other uniformly, so they arrange themselves into a regular pattern that is held together by the negatively charged electron cloud. Since the negative electron cloud surrounds each of the positive ions that make up the orderly three-dimensional crystal structure, strong electronic attraction holds the metal together. A characteristic of metallic bonding is the fact that every positive ion is equivalent. Ideally, a symmetrical ion is produced when a valence electron is removed from the metal atom. As a result of this ion symmetry, metals tend to form highly symmetrical, close-packed crystal structures. They also have a large number of nearest neighboring atoms (usually 8 to 12), which helps to explain their high densities and high elastic stiffness.

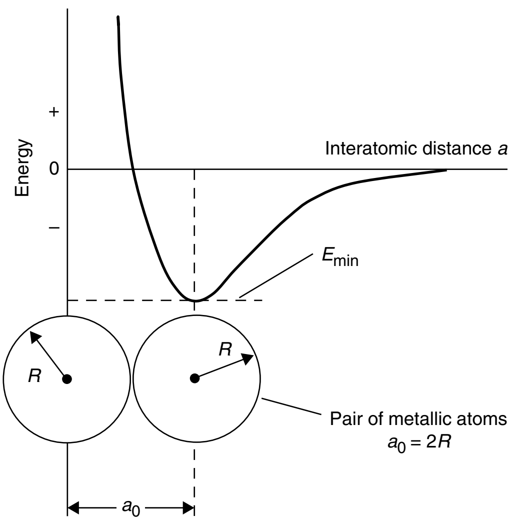

Fig. 1.2 Bond energy in metallic bond

Since the valence electrons are no longer attached to specific positive ions and are free to travel among the positive ions, metals exhibit high electrical and thermal conductivity. The opaque luster of metals is due to the reflection of light by the free electrons. A light wave striking the surface causes the free electrons to vibrate and absorb all the energy of the wave and prevent transmission. The vibrating electrons then reemit, or reflect, the wave from the surface. The ability of metals to undergo significant amounts of plastic deformation is also due to the metallic bond. Under the action of an applied shearing force, layers of the positive ion cores can slide over each other and reestablish their bonds without drastically altering their relationship with the electron cloud. The ability to alloy, or mix several metals together in the liquid state, is one of the keys to the flexibility of metals. In the liquid state, solubility is often complete, while in the solid state, solubility is generally much more restricted. This change in solubility with temperature forms the basis for heat treatments that can vary the strength and ductility over quite wide ranges.

In general, the fewer the valence electrons and more loosely they are held, the more metallic is

Table 1.1 General characteristics of bond types

Property	Metallic bond	Covalent bond	Ionic bond	Secondary bond
Example	Cu, Ni, Fe	Diamond, silicon carbide	NaCl, $CaCl_2$	Wax, Ar
Mechanical	Weaker than ionic or covalent bond Tough and ductile Nondirectional	Very hard and brittle Fails by cleavage Strongly directional	Hardness increases with ionic charge Fails by cleavage Nondirectional	Weak and soft Can be plastically deformed
Thermal	Moderately high melting points Good conductors of heat	Very high melting points Thermal insulators	Fairly high melting points Thermal insulators	Low melting points
Electrical	Conductors	Insulators	Insulators	Insulators
Optical	Opaque and reflecting	Transparent or opaque High refractive index	Transparent Colored by ions	Transparent

Source: Ref 1

the bonding. Metals such as copper and silver, which have few valence electrons, are very good conductors of electricity and heat, because their few valence electrons are highly mobile. As the number of valence electrons increases and the tightness with which they are held to the nucleus increases, the valence electrons become localized and the bond becomes more covalent. The transition metals, such as iron and nickel, have incomplete *d*-shells and exhibit some covalent bonding, which helps explain their relatively high melting points. Tin is interesting in that it has two crystalline forms, one which is mostly metallic and ductile and another which is mostly covalent and very brittle. Intermetallic compounds can also be formed between two metals in which the bonding is partly metallic and partly ionic. As the electronegativity difference between the two metals increases, the bonding becomes more ionic in nature. For example, both aluminum and vanadium have an electronegativity of 1.5 and the difference is 0, so the compound Al_3V is primarily metallic. On the other hand, aluminum and lithium (electronegativity of 1.0) have an electronegativity difference of 0.5; thus, when they

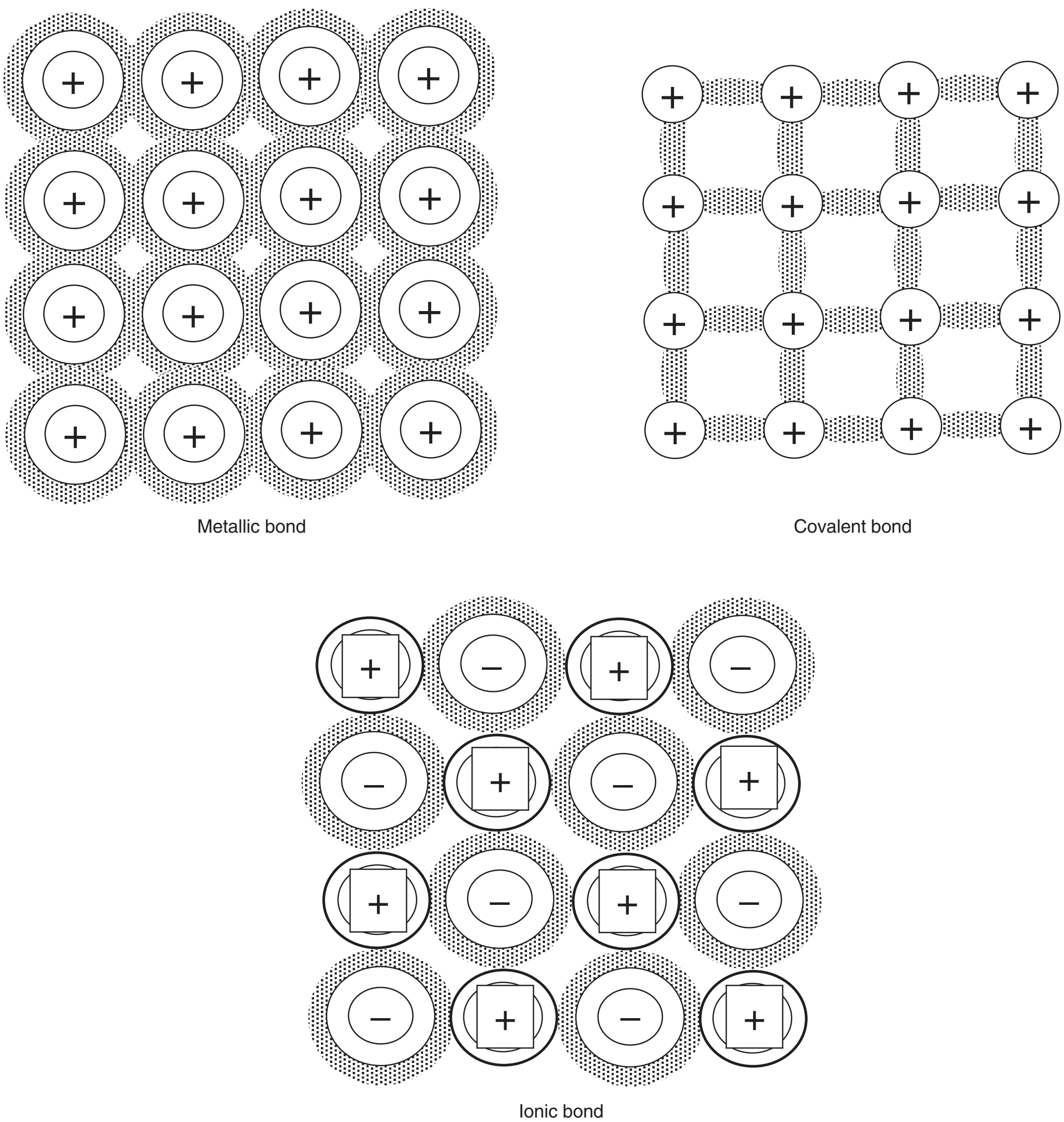

Fig. 1.3 Primary bonding mechanisms

form the compound AlLi, the bond is a combination of metallic and ionic.

1.2.2 Ionic Bonding

Ionic bonding, also shown in Fig. 1.3, is a result of electrical attraction between alternately placed positive and negative ions. In the ionic bond, the electrons are shared by an electropositive ion (cation) and an electronegative ion (anion). The electropositive ion gives up its valence electrons, while the electronegative ion captures them to produce ions having full electron orbitals or suborbitals. As a consequence, there are no free electrons available to conduct electricity. In ionically bonded solids such as salts, there are very few slip systems along which atoms can move. This is a consequence of the electrically charged nature of the ions. For slip in some directions, ions of like charge must be brought into close proximity to each other, and because of electrostatic repulsion, this mode of slip is very restricted. This is not a problem in metals, since all atoms are electronically neutral. No electrical conduction of the kind found in metals is possible in ionic crystals, but weak ionic conduction occurs as a result of the motion of the individual ions. When subjected to stresses, ionic crystals tend to cleave, or break, along certain planes of atoms rather than deform in a ductile fashion as metals do.

Ionic bonds form between electropositive metals and electronegative nonmetals. The further apart the two are on the periodic table, the more likely they are to form ionic bonds. For example, sodium (Na) is on the far left side of the periodic table in group I, while chlorine (Cl) is on the far right side in group VII. Sodium and chlorine combine to form common table salt (NaCl). As shown in Fig. 1.4, the sodium atom gives up its outer valence electron, which is transferred to the outer electron shell of the chlorine atom. Since the outer shell of chlorine now contains eight electrons, similar to the noble gases, it is an extremely stable configuration. In terms of symbols, the sodium ion is written as Na^+, and the chlorine ion is written as Cl^-. When the two combine to form an ionic bond, the compound (NaCl) is neutral since the charges balance. Since the positively charged cation can attract multiple negatively charged anions, the ionic bond is nondirectional.

1.2.3 Covalent Bonding

Many elements that have three or more valence electrons are bound into crystal structures by forces arising from the sharing of electrons. The nature of this covalent bonding is shown schematically in Fig. 1.3. To complete the octet of electrons needed for atomic stability, electrons must be shared with 8-*N* neighboring atoms, where *N* is the number of valence electrons in the given element. High hardness and low electrical conductivity are general characteristics of solids of this type. In covalently bonded ceramics, the bonding between atoms is specific and directional, involving the exchange of electron charge between pairs of atoms. Thus, when covalent crystals are stressed to a sufficient extent, they exhibit brittle fracture due to a separation of electron pair bonds, without subsequent reformation. It should also be noted that ceramics are rarely either all ionically or covalently bonded; they usually consist of a mix of the two types of bonds. For example, silicon nitride (Si_3N_4) consists of approximately 70% covalent bonds and 30% ionic bonds.

Covalent bonds also form between electropositive elements and electronegative elements. However, the separation on the periodic table is not great enough to result in electron

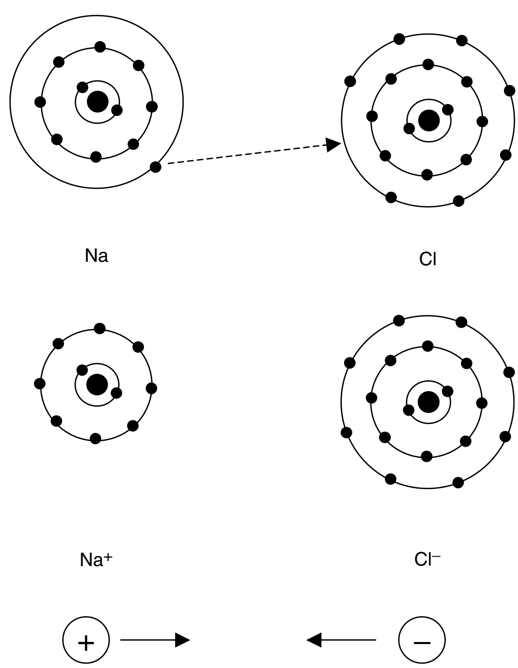

Fig. 1.4 Ionic bonding in NaCl

transfer, as in the ionic bond. Instead, the valence electrons are shared between the two elements. For example, a molecule of methane (CH_4), shown in Fig. 1.5, is held together by covalent bonds. Note that hydrogen, in group I on the periodic table, and carbon, in group IV, are much closer together than sodium and chlorine, which form ionic bonds. In a molecule of methane gas, four hydrogen atoms are combined with one carbon atom. The carbon atom has four electrons in its outer shell, and these are combined with four more electrons, one from each of the four hydrogen atoms, to give a completed stable outer shell of eight electrons held together by covalent bonds. Each shared electron passes from an orbital controlled by one nucleus into an orbital shared by two nuclei. Covalent bonds, since they do not ionize, will not conduct electricity and are nonconductors. Covalent bonds form the basis for many organic compounds, including long-chain polymer molecules. As the molecule size increases, the bond strength of the material also increases. Likewise, the strength of long-chain molecules also increases with increases in chain length.

1.2.4 Secondary Bonding

Secondary, or van der Waals, bonding is weak in comparison to the primary metallic, ionic, and covalent bonds. Bond energies are typically on the order of only 10 kJ/mol (0.1 eV/atom). Although secondary bonding exists between virtually all atoms or molecules, its presence is usually obscured if any of the three primary bonding types is present. While van der Waals forces only play a minor role in metals, they are an important source of bonding for the inert gases that have stable electron structures, some molecular compounds such as water, and thermoplastic polymers where the main chains are covalently bonded but are held to other main chains by secondary bonding. Van der Waals bonding between two dipoles is illustrated in Fig. 1.6.

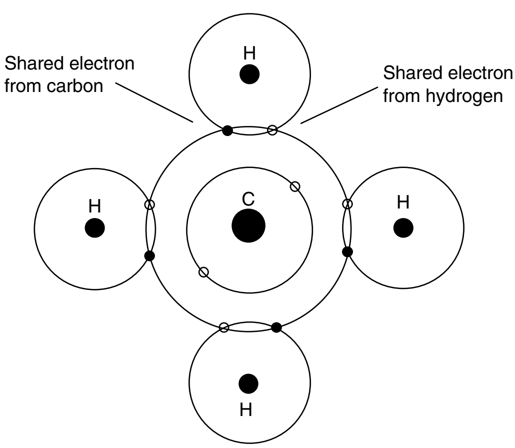

Fig. 1.5 Covalent bonding in methane

1.3 Crystalline Structure

When a substance freezes on cooling from the liquid state, it forms a solid that is either an amorphous or a crystalline structure. An amorphous structure is essentially a random structure. Although there may be what is known as short-range order, in which small groups of atoms are arranged in an orderly manner, it does not contain long-range order, in which all of the atoms are arranged in an orderly manner. Typical amorphous materials include glasses and almost all organic compounds. However, metals, under normal freezing conditions, normally form long-range, orderly crystalline structures. Except for glasses, almost all ceramic materials also form crystalline structures. Therefore, metals and ceramics are, in general, crystalline, while glasses and polymers are mostly amorphous.

1.3.1 Space Lattices and Crystal Systems

A crystalline structure consists of atoms, or molecules, arranged in a pattern that is repetitive in three dimensions. The arrangement of the atoms or molecules in the interior of a crystal is called its crystalline structure. A distribution of points (or atoms) in three dimensions is said to form a space lattice if every point has identical surroundings, as shown in Fig. 1.7. The intersections of the lines, called lattice points, represent locations in space with the same kind of atom or group of atoms of identical composition, arrangement, and orientation. The geometry of a space lattice is completely specified by the lattice constants *a*, *b*, and *c* and the interaxial angles α, β, and γ. The unit cell of a

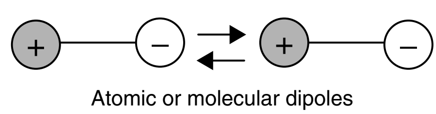

Fig. 1.6 Van der Waals bonding between two dipoles

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crystal is the smallest pattern of arrangement that can be contained in a parallelepiped, the edges of which form the *a*, *b*, and *c* axes of the crystal. Appendix C, "Crystallographic Planes and Directions," in this book describes how the Miller indices, a system for specifying crystallographic planes within the unit cell, are determined.

When discussing crystal structure, it is usually assumed that the space lattice continues to infinity in all directions. In terms of a typical crystal (or grain) of, for example, iron that is 0.2 cm^3 (0.01 $in.^3$) in size, this may appear to be a preposterous assumption, but when it is realized that there are 10^{18} iron atoms in such a grain, the approximation to infinity seems much more plausible.

All crystal systems can be grouped into one of seven basic systems, as defined in Table 1.2, which can be arranged in 14 different ways,

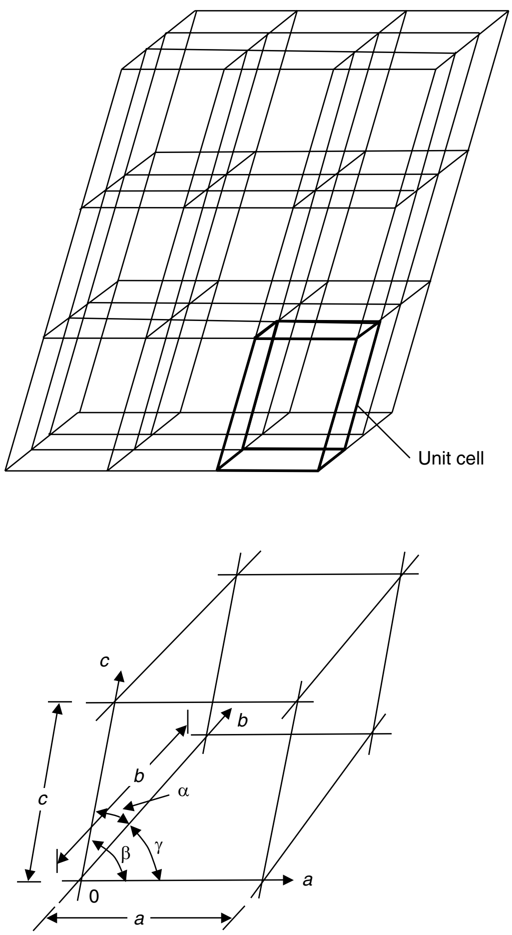


Fig. 1.7 Space lattice and unit cell

called Bravais lattices, as shown in Fig. 1.8. Almost all structural metals crystallize into one of three crystalline patterns: face-centered cubic, (fcc) hexagonal close-packed, (hcp) or body-centered cubic, (bcc) illustrated in Fig. 1.9 to 1.11. It should be noted that the unit cell edge lengths and axial angles are unique for each crystalline substance. The unique edge lengths are called lattice parameters. Axial angles other than 90° or 120° can also change slightly with changes in composition. When the edges of the unit cell are not equal in all three directions, all unequal lengths must be stated to completely define the crystal. The same is true if all axial angles are not equal.

1.3.2 Face-Centered Cubic System

The face-centered cubic (fcc) system is shown in Fig. 1.9. As the name implies, in addition to the corner atoms, there is an atom centrally located on each face. Since each of the atoms located on the faces belong to two unit cells and the eight corner atoms each belong to eight unit cells, the number of atoms belonging to a unit cell is four. The atomic packing factor (the volume of atoms belonging to the unit cell divided by the volume of the unit cell) is 0.74 for the fcc structure. This is the densest packing that can be obtained. The fcc structure is the most efficient, with 12 nearest atom neighbors (also referred to as the coordination number, or CN); that is, the fcc structure has a CN = 12. Methods of calculating atomic packing factors and coordination numbers are given in Appendix B, "Crystalline System Calculations," in this book. As shown in Fig. 1.12, the stacking sequence for the fcc structure is *ABCABC*. The fcc structure is found in many important metals such as aluminum, copper, and nickel.

1.3.3 Hexagonal Close-Packed System

The atoms in the hexagonal close-packed (hcp) structure (Fig. 1.10) are also packed along close-packed planes. It should also be noted

Table 1.2 Seven crystal systems

Crystal system	Edge length	Interaxial angle
Triclinic	$a \neq b \neq c$	$\alpha \neq \beta \neq \gamma \neq 90°$
Monoclinic	$a \neq b \neq c$	$\alpha = \gamma = 90° \neq \beta$
Orthorhombic	$a \neq b \neq c$	$\alpha = \beta = \gamma = 90°$
Tetragonal	$a = b \neq c$	$\alpha = \beta = \gamma = 90°$
Hexagonal	$a = b \neq c$	$\alpha = \beta = 90°, \gamma = 120°$
Rhombohedral	$a = b = c$	$\alpha = \beta = \gamma \neq 90°$
Cubic	$a = b = c$	$\alpha = \beta = \gamma = 90°$

that both the fcc and hcp structures are what is known as close-packed structures with crystallographic planes having the same arrangement of atoms; however, the order of stacking the planes is different. Atoms in the hcp planes (called the basal planes) have the same arrangement as those in the fcc close-packed planes. However, in the hcp structure, these

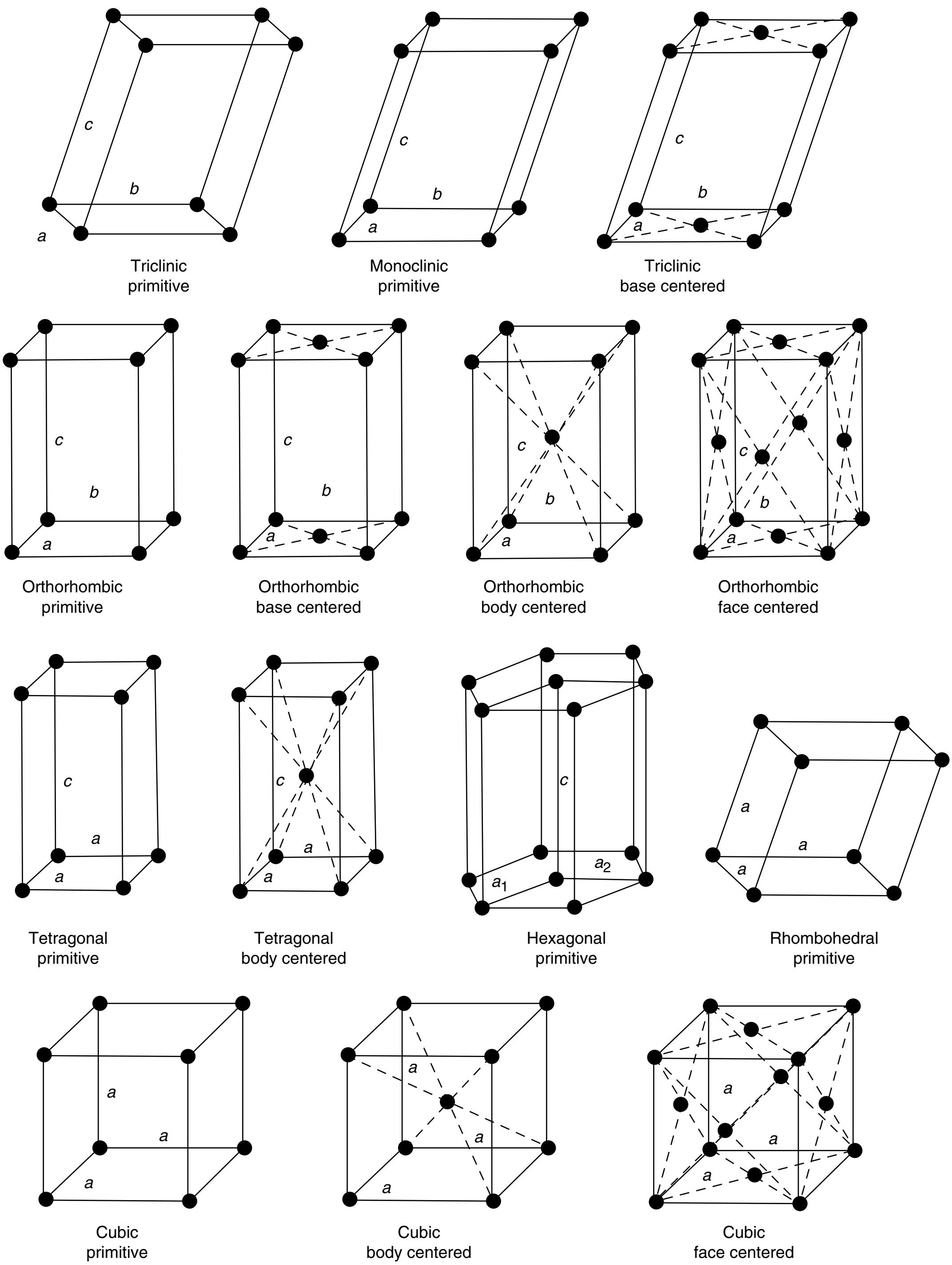


Fig. 1.8 The fourteen bravais lattices

planes repeat every other layer to give a stacking sequence of . . .*ABA*. . .. The number of atoms belonging to the hcp unit cell is six, and the atomic packing factor is 0.74. Note that this is the same packing factor as was obtained for the fcc structure. Also, the coordination number obtained for the hcp structure (CN = 12) is the same as that for the fcc structure. A basic rule of crystallography is that if the coordination numbers of two different unit cells are the same, then they both will have the same packing factors.

Two lattice parameters, *c* and *a*, also shown in Fig. 1.10, are needed to completely describe the hcp unit cell. In an ideal hcp structure, the ratio of the lattice constants, *c*/*a*, is 1.633. In this ideal packing arrangement, the layer between the two basal planes in the center of the structure is located close to the atoms on the upper and lower basal planes. Therefore, any atom in the lattice is in contact with 12 neighboring atoms, and the coordination number is therefore CN = 12. It should be noted that there is often some deviation from the ideal ratio of $c/a = 1.633$. If the ratio is less than 1.633, it means that the atoms are compressed in the *c*-axis direction, and if the ratio is greater than 1.633, the atoms are elongated along the *c*-axis. In these situations, the hcp structure can no longer be viewed as truly being close packed and should be described as just being hexagonal. However, structures deviating from the ideal packing are still normally described as being hcp. For example, beryllium is described as having an hcp structure, but its *c*/*a* ratio of 1.57 is unusually low and causes some distortion in the lattice. This distortion and the unusually high elastic modulus of beryllium (3×10^5 MPa, or 42×10^3 ksi) result from a covalent component in its bonding. Contributions from covalent bonding are also present in the hcp metals zinc and cadmium, with *c*/*a* ratios greater than 1.85. This lowers their packing density to approximately 65%, considerably less than the 74% of the ideal hcp structure.

1.3.4 Body-Centered Cubic System

The body-centered cubic (bcc) system is shown in Fig. 1.11. The bcc system is similar to the simple cubic system except that it has an additional atom located in the center of the structure. Since the center atom belongs completely to the unit cell in question, the number of atoms belonging to the bcc unit cell is two. The coordination number for the bcc structure is eight, since the full center atom is in contact with eight neighboring atoms located at the corner points of the lattice. The atomic packing factor for the bcc structure is 0.68, which is less than that of the fcc and hcp structures. Since the packing is less efficient in the bcc structure, the closest-packed planes are less densely packed (Fig. 1.13).

Even though the bcc crystal is not a densely packed structure, it is the equilibrium structure of 15 metallic elements at room temperature, including many of the important transition elements. This is attributable to two factors: (1) Even though each atom has only eight nearest neighbors, the six second-nearest neighbor atoms are closer in the bcc structure

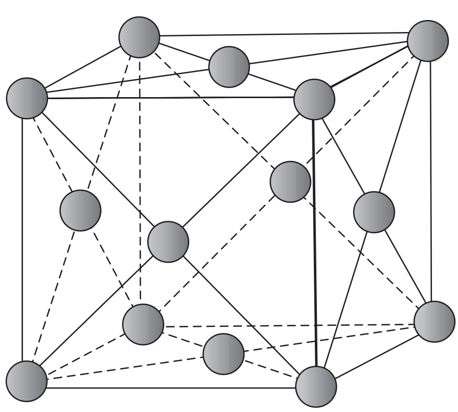
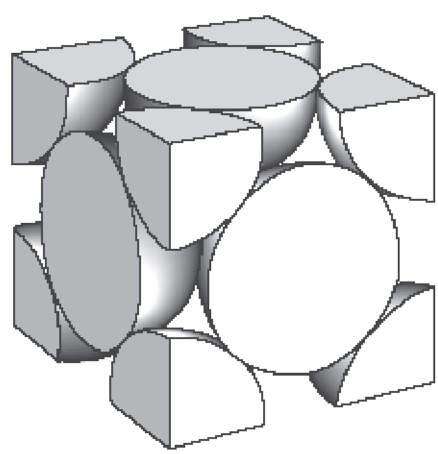


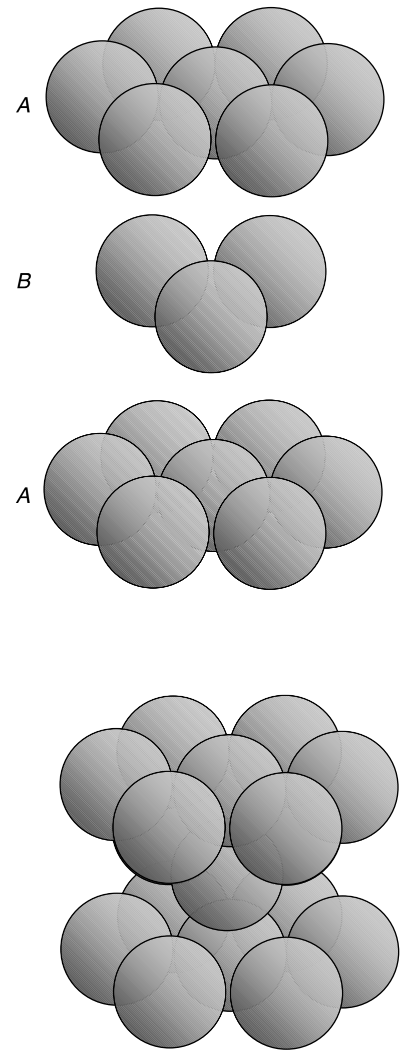
Fig. 1.9 Face-centered cubic structure



than in the fcc structure. Calculations indicate that these second-nearest neighbor bonds make a significant contribution to the total bonding energy of bcc metals; and, (2) in addition, the greater entropy of the less densely packed bcc structure gives it a stability advantage over the more tightly packed fcc structure at high temperatures. As a consequence, some metals that have the close-packed structures at room temperature transition to bcc structures at higher temperatures.

1.4 Slip Systems

Plastic deformation takes place by sliding (slip) of close-packed planes over one another. A reason for this slip plane preference is that



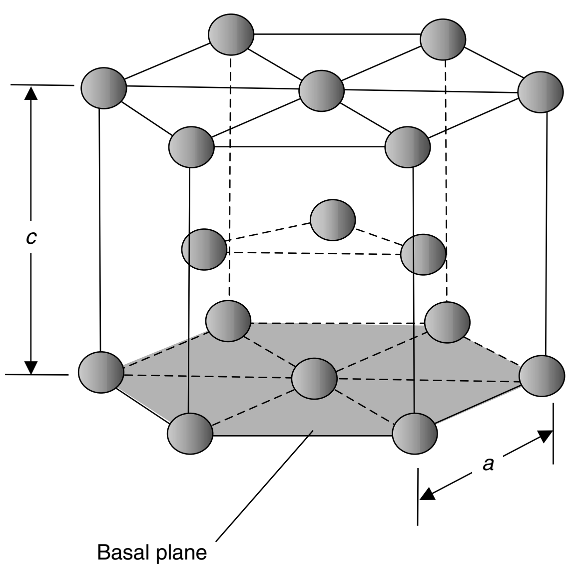


Fig. 1.10 Hexagonal close-packed structure

the separation between close-packed planes is greater than for other crystal planes, and this makes their relative displacement easier. Furthermore, the slip direction is in a close-packed direction. The combination of planes and directions on which slip takes place constitutes the slip systems of the material. In polycrystalline materials, a certain number of slip systems must be available in order for the material to be capable of plastic deformation. Other things being equal, the greater the number of slip systems, the greater the capacity for deformation. Face-centered cubic metals have a large number of slip systems (12) and are capable of moderate-to-extensive plastic deformation even at temperatures approaching absolute zero. A number of close-packed planes for the fcc, bcc, and hcp structures are illustrated in Fig. 1.14.

Materials having the bcc structure also often display 12 slip systems, although this number comes about differently than it does for the fcc lattice. A closest-packed bcc plane is defined by a unit cell edge and face diagonal. There are only two close-packed directions (the cube diagonals) in the closest-packed bcc plane, but there are six nonparallel planes of this type. Over certain temperature ranges, some bcc

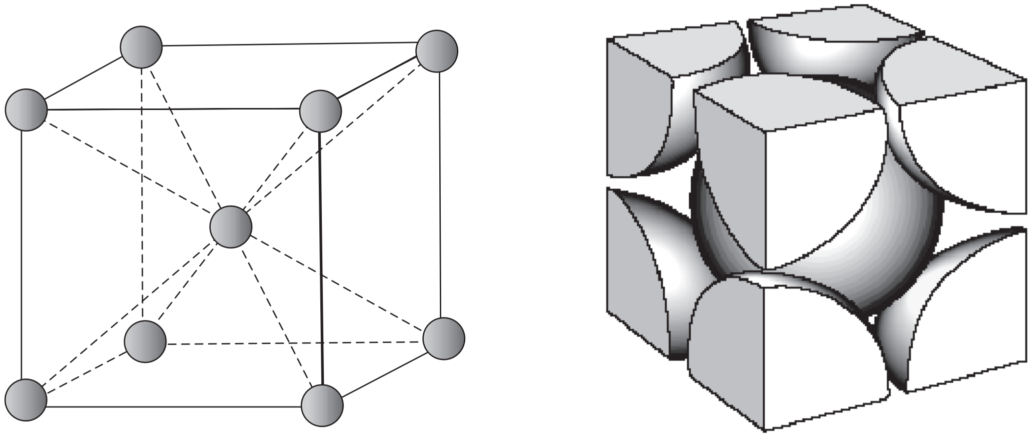


Fig. 1.11 Body-centered cubic structure

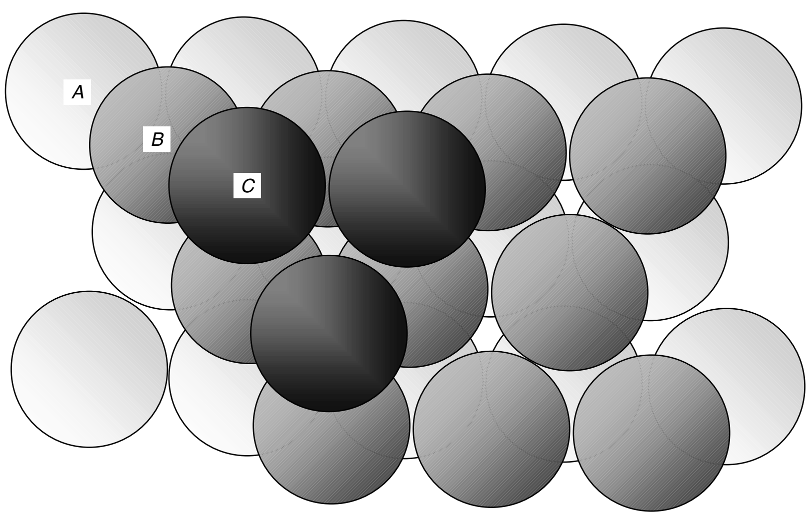


Fig. 1.12 Close packing of planes for the face-centered cubic structure

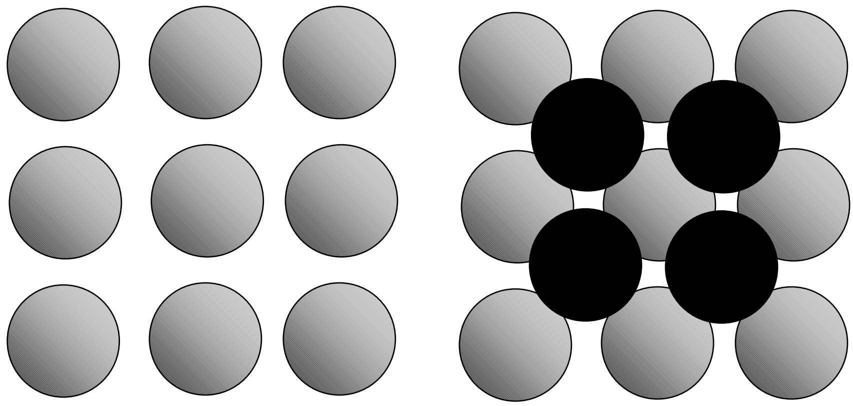


Fig. 1.13 Loose packing in body-centered cubic structure

metals display slip on other than close-packed planes, although the slip direction remains a close-packed one. Thus, bcc metals have the requisite number of slip systems to allow for plastic deformation. However, bcc metals often become brittle at low temperatures as a result of the strong temperature sensitivity of their yield strength, which causes them to fracture prior to undergoing significant plastic deformation.

Depending on their *c*/*a* ratio, polycrystalline hcp metals may or may not have the necessary number of slip systems to allow for appreciable plastic deformation. The ideal hcp structure has only three slip systems, because there is only one nonparallel close-packed plane in it (the basal plane, which contains three nonparallel close-packed directions). However, three slip systems are insufficient to permit polycrystalline plastic deformation, and so hcp polycrystals for which slip is restricted to the basal plane are not malleable. When *c*/*a* is less than the ideal ratio, basal planes become less widely separated, and other planes compete with them for slip activity. In these instances, the number of slip systems increases, and material ductility is beneficially affected. In addition, polycrystalline hcp metals can also deform by a mechanism called twinning, which is covered in Chapter 2, "Crystalline Imperfections and Plastic Deformation," in this book.

The methods for identifying crystalline planes and crystalline directions can be found in Appendix C, "Crystallographic Planes and Directions."

1.5 Allotropy

Depending on pressure and temperature, many metals can exist in more than one crystalline form, a phenomenon known as allotropy. The important metal iron undergoes a series of allotropic transformations during heating and cooling, as shown in Fig. 1.15. Note that an allotropic transformation is a solid-state phase transformation and, as such, occurs at a constant temperature during either heating or cooling.

Under equilibrium cooling conditions, the solidification of pure iron from the liquid occurs at 1540 °C (2800 °F) and forms what is called delta iron (δ_{Fe}), which has a bcc structure. Delta iron is then stable on further cooling until it reaches 1395 °C (2541 °F), where it undergoes a transformation to an fcc structure called gamma iron (γ_{Fe}). On still further cooling to 900 °C (1648 °F), it undergoes yet another phase transformation, transforming from the fcc structure back to the bcc structure, called ferrite iron (α_{Fe}) to distinguish it from the higher-temperature delta iron. This last transformation, $\gamma_{Fe} \rightarrow \alpha_{Fe}$, is extremely important because it forms the basis for the hardening of steel. Note that the $\gamma_{Fe} \rightarrow \alpha_{Fe}$ transformation occurs at 900 °C (1648 °F) on cooling, somewhat lower than the 910 °C (1673 °F) transformation temperature on heating. This temperature differential is known as the temperature hysteresis of allotropic phase transformation, and its magnitude increases with increases in the cooling rate. The temperatures (designated *A*) associated with heating contain the subscript "c," which is

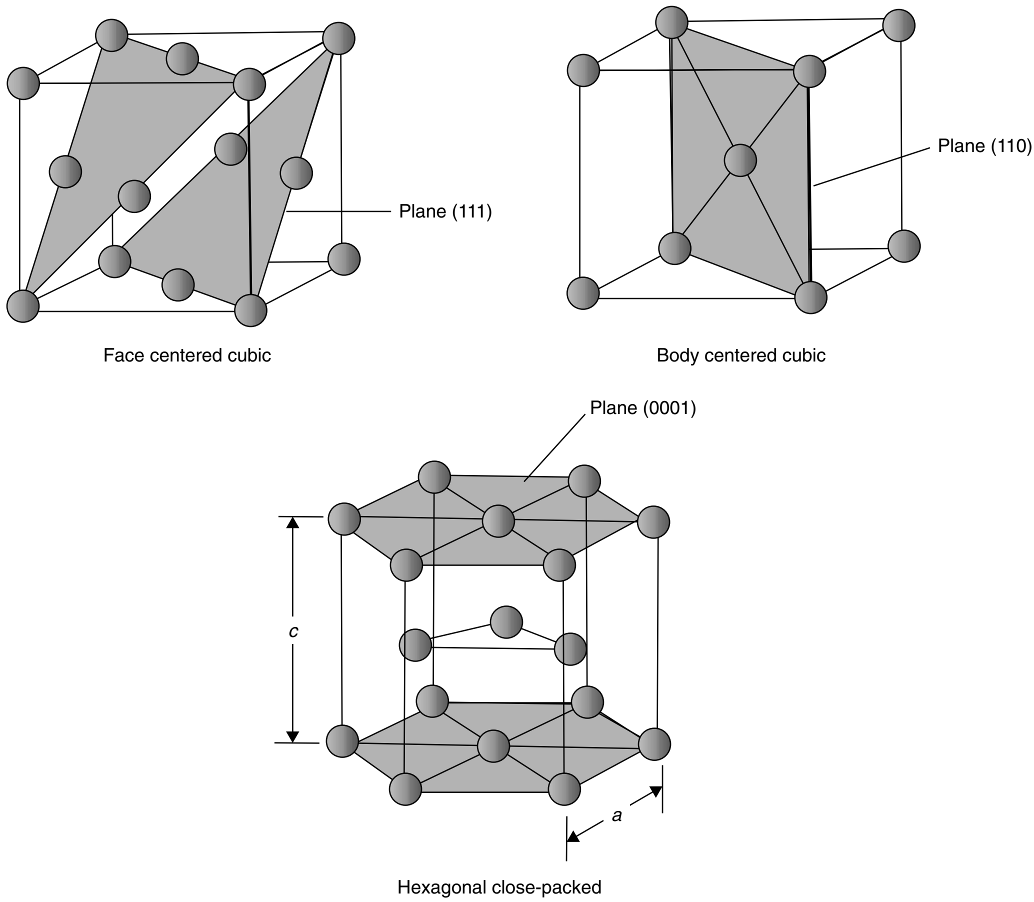


Fig. 1.14 Close-packed planes

French for *chauffage*, meaning heating, while the ones for cooling have the subscript "r" for the French *refroidissement*, meaning cooling.

Many other metals, as well as some non-metals, also exhibit allotropic transformations. For example, titanium, zirconium, and hafnium exhibit a transition from an hcp structure to bcc on heating. Note that in each case, a close-packed structure is stable at room temperature, while a looser packing is stable at elevated temperatures. While this is not always the case, it is a trend experienced with many metals.

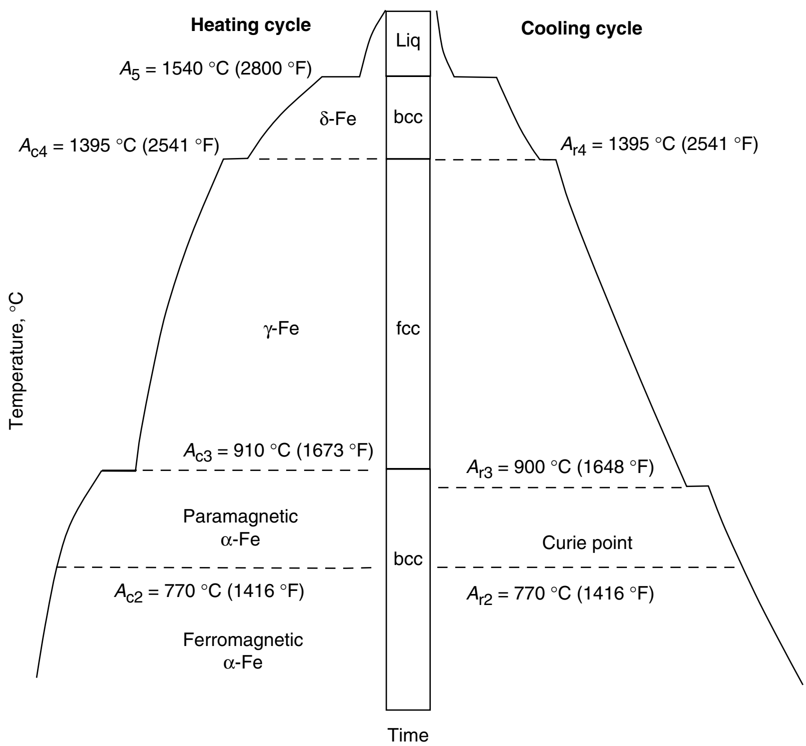


Fig. 1.15 Allotropic transformations in pure iron. Source: Ref 2

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