On the occurrence of metallic character in the periodic table of the chemical elements

Friedrich Hensel¹, Daniel R. Slocombe² and Peter P. Edwards²

¹Fachbereich Chemie, Philipps-Universität Marburg, Hans-Meerwein-Strasse, Marburg 35032, Germany
²Inorganic Chemistry Laboratory, Department of Chemistry, University of Oxford, South Parks Road, Oxford OX1 3QR, UK

The classification of a chemical element as either ‘metal’ or ‘non-metal’ continues to form the basis of an instantly recognizable, universal representation of the periodic table (Mendeleeff D. 1905 The principles of chemistry, vol. II, p. 23; Poliakoff M. & Tang S. 2015 Phil. Trans. R. Soc. A 373, 20140211). Here, we review major, pre-quantum-mechanical innovations (Goldhammer DA. 1913 Dispersion und Absorption des Lichtes; Herzfeld KF. 1927 Phys. Rev. 29, 701–705) that allow an understanding of the metallic or non-metallic status of the chemical elements under both ambient and extreme conditions. A special emphasis will be placed on recent experimental advances that investigate how the electronic properties of chemical elements vary with temperature and density, and how this invariably relates to a changing status of the chemical elements. Thus, the prototypical non-metals, hydrogen and helium, becomes metallic at high densities; and the acknowledged metals, mercury, rubidium and caesium, transform into their non-metallic forms at low elemental densities. This reflects the fundamental fact that, at temperatures above the absolute zero of temperature, there is therefore no clear dividing line between metals and non-metals. Our conventional demarcation of chemical elements as metals or non-metals within the periodic table is of course governed by our experience of the nature of the elements under ambient conditions. Examination of these other situations helps us to examine the exact divisions of the chemical elements into metals and non-metals (Mendeleeff D. 1905 The principles of chemistry, vol. II, p. 23).
### 1. Metals and non-metals of the periodic table

Perhaps the most traditional and common classification of the chemical elements of the periodic table in textbooks and in chemistry classroom charts—and elsewhere [1]—is in terms of their metallic and non-metallic properties. Over a century ago, Mendeleeff noted ‘Many elements, although not all of them, have the peculiar lustre, opacity, malleability and the high thermal and electrical conductivities peculiar to the *metals* ... Those which do not possess the physical properties of metals are called *non-metals*’ [2]. Perhaps, the most outstanding physical property of a metal is its superb ability to transmit electricity. Metals are thus magnificent conductors of electricity and ultimately this is what distinguishes them from non-metals; the broadest, modern definition of metals and the metallic state is therefore of a substance transmitting electricity by electron transfer [3].

Even armed with room-temperature electrical conductivity data (as shown in figure 1), one can fairly easily identify elements of the periodic table for which the designation ‘metal’ or ‘non-metal’ is clearly appropriate. The difference in electrical conductivities between a metallic element and a non-metallic element (under ambient conditions) is itself a striking fact; between the most conductive elements (copper and silver) and the most resistive (sulfur, for example) it amounts to 23 orders of magnitude [5]. This range of over $10^{23}$ in electrical conductivity (or rather its inverse, resistivity, the quantity customarily used in the application of Ohm’s law) has been cited as one of the most widely ranging of all physical properties [5,6]. Furthermore—and remarkably—this range in the elements’ electrical property is even larger at low temperatures; the temperature dependence of the electrical resistivity of non-metals is in every way in direct contrast to that of metals [7]. Then, surely, only the most rigid purist would seriously question the ‘easy’ designation of mercury as a metal or sulfur, say, as a non-metal in the periodic table.

However, the very basis of such a conventional classification over time is undoubtedly influenced by our experience of the nature of the chemical elements *existing under ambient conditions (generally room temperature, room pressure) on our planet*. It has, however, been surmised...
very early (without a general proof) that metallic or non-metallic behaviour is not an inherent and unchanging property of any particular chemical element of the periodic table [3].

Once one begins to consider the extremes of temperature and pressure achievable in the laboratory, or the even more exotic conditions of geophysical and astrophysical relevance, many more varieties of physical behaviour become possible [7].

This view was beautifully illustrated by the notable discovery of a metallic form of phosphorus a century ago by Bridgman, who subjected white phosphorus at 200°C to pressures of 12–15 kbar [8]. This is most certainly the first experimental observation of the transition of an elemental solid from a non-metal to a metal.

2. The metal–non-metal transition in fluid chemical elements

(a) Concepts, models and criteria

This fundamental issue of the rigorous distinction between the metallic and non-metallic states of matter becomes even more important if we include consideration of the fluid state of the chemical elements. It has been more than 100 years since Strutt [9] and Bender [10] pointed out the interesting implications of an undoubted, density-induced metal to non-metal transition for the phase behaviour of certain fluid metals at high temperatures and pressures. Strutt noted ‘Mercury vapour is an insulator, while liquid mercury is a conductor. Since the liquid and saturated vapour are indistinguishable above the critical temperature, one or both of these must undergo a remarkable change of electrical properties as that temperature is approached’ [9]. Remarkably, Strutt found a reduction of close to eight orders of magnitude in electrical resistance for saturated mercury vapour at high temperatures and pressures as compared with its resistance at room pressure, even though experimentally he could not fully approach the critical temperature (1751 K) and pressure (1673 bar) at that time.

Such a metal to non-metal transition at finite temperature, in fact, raises fundamental questions of definition of both of these canonical states of matter [7]. As Landau & Zeldovich [11] and Mott (e.g. [12]) pointed out, there is a fundamental difficulty— theoretically and indeed experimentally—in drawing a distinction between a metal and a non-metal except at the absolute zero of temperature [7]. We now recognize that, at the absolute zero of temperature, a metal sharply and fundamentally differs from a non-metal with respect to its accessible spectrum of electronic energy levels.

The ground state of a metal comprises a continuous spectrum of electronic energy states. This explains the striking fact that in a metal the application of even the weakest electrical field produces an electrical current due to transitions to adjacent electronic levels in the continuum of electronic energy states. By contrast, the energy spectrum of a dielectric, i.e. a non-metal, is characterized by the existence of a finite electronic energy gap between levels. Thus, a rigorous criterion to distinguish between a metal and a non-metal can exist only at $T = 0$ K. There ... ‘... a metal conducts and a nonmetal doesn’t’ [13]. At any non-zero temperature, therefore, thermal excitation of conduction electrons, no matter how few, blurs the difference between metal and non-metal.

The most common modern—that is, quantum mechanical—description of metals and non-metals is that metals contain a partially filled conduction band, whereas non-metals have a completely filled valence band separated by a large electronic energy gap from an empty conduction band. The modern theory of electronic band structure provides a comprehensive picture of the behaviour of electrons in a solid; it derives completely from the underlying periodic lattice structure of atoms in crystalline metals or non-metals [14]. However, predicting, rationalizing and applying such concepts and band structures for amorphous or liquid substances require different models (figure 2).

The real difficulty—in either crystalline or amorphous substances—arises in the transition range where metallic properties evolve continuously into those characteristic of non-metals. Various approximate criteria exist for defining a boundary between metal and non-metal [15,16].
One of the earliest attempts—perhaps the earliest attempt—to explain the occurrence of metallic versus non-metallic behaviour of the chemical elements, and of course with that the first discussion of the metal–non-metal transition, can be traced back to the pioneering, pre-quantum-mechanical models of Goldhammer [17] and Herzfeld [18]. Herzfeld’s theory is based on the classical (Lorentz) oscillator model of an atom. Herzfeld pointed out that electrons localized around atomic nuclei constitute polarizable objects and that their internal dynamics in the dense assembly of the element leads to local corrections to the polarizing tendency of any external field impressed on the system.

He proposed that the corresponding element (viewed obviously as a collection of atoms) becomes metallic when the frequency of the oscillator placed in this dense, dielectric medium approaches zero; i.e. the valence electron is set free and the element—under those critical conditions—acquires metallic status.

For an isotropic substance, the Goldhammer–Herzfeld criterion is based upon the Lorenz–Lorentz or Clausius–Mosotti relation

$$\frac{n^2 - 1}{n^2 + 2} = \frac{\varepsilon - 1}{\varepsilon + 2} = \frac{4\pi \alpha}{3V} N_A = \frac{R}{V},$$

where \(n\) is the index of refraction, \(\varepsilon\) is the dielectric constant, \(\alpha\) is the atomic electronic polarizability, \(V\) is the molar volume, \(N_A\) is Avogadro’s number and \(R = \left(\frac{4}{3}\pi \alpha\right) N_A\) is the molar refractivity. Thus for \(R/V = 1\), the refractive index \(n\) or the dielectric constant \(\varepsilon\) diverge. This can only occur if electrons are no longer bound to individual atoms in the solid or liquid, and now become free, as in a metal.

A simple physical interpretation of the Goldhammer–Herzfeld criterion becomes apparent when one adopts the model of an isolated atom as a perfectly conducting sphere. It can be shown from electrostatics that the isolated atom’s electronic polarizability is \(\alpha = r^3\), where \(r\) is the atomic radius and the molar refractivity is then \(R = \left(\frac{4}{3}\pi N_A r^3\right)\); in reality, nothing more than a measure of the volume occupied by one mole of these perfectly conducting spheres. On the Goldhammer–Herzfeld view, metallization therefore occurs when the molar volume (the volume accessible to the atom at the specific conditions of temperature, pressure and elemental density) becomes equal to or less than \(R\), the molar refractivity; that is, the system of isolated conducting spheres becomes one large, macroscopic conductor of electricity.

The accuracy of the empirical Goldhammer–Herzfeld criterion with respect to the metallic character of the chemical elements of the periodic table at ambient conditions is nicely illustrated in figure 3. It is seen at a glance from figure 3 that, as one moves from left to the right across the periodic table, the elements transform from metal to non-metal around the IVA group just as the \(R/V\) values move below the dashed line at the critical \(R/V = 1\) threshold.
Figure 3. The metallization of chemical elements of the periodic table under ambient conditions imposed on the Earth. The figure shows the ratio \( R/V \) for elements of the s, p and d blocks of the periodic table. Here, \( R \) is the molar refractivity and \( V \) is the molar volume. The shaded circles represent elements for which both \( R \) and \( V \) are known experimentally. The open circles are for elements in which \( R \) is calculated and \( V \) is known experimentally. (Adapted from [19,20].)

Therefore, if we start with an element of the periodic table at low density—in the non-metallic state—and continually increase its density (by reducing the spacing between atoms in the solid or liquid), this leads ultimately to a dielectric catastrophe \( (\varepsilon = \infty) \) for \( R \geq 1 \) and the constituent atoms are no longer able to hold on to their valence electrons and metallic conductance will occur. Herzfeld showed that a large number of normally existing metals and non-metals across the periodic system do satisfy this condition. He also discussed—in 1927—the possible metallization of Xe and predicted the metallization volume \( 10.2 \text{ cm}^3 \text{ mol}^{-1} \) in close agreement with the modern experimentally observed value of \( 10.7 \text{ cm}^3 \text{ mol}^{-1} \) [21].

(b) Critical behaviour of fluid metals

As mentioned above, the metal–non-metal transition in elemental fluids is among the conceptually simplest examples of such an electronic phase transition from the metallic to the non-metallic state, and vice versa. A liquid metal near its triple point is in natural equilibrium with its corresponding low-density, non-metallic vapour. There is, therefore, a metal–non-metal transition that coincides exactly with the liquid–vapour transition, i.e. both the density and the electronic and molecular structures change when a metallic liquid vaporizes.

For example, liquid mercury and caesium approaching their triple points are considered as normal liquid metals with electronic properties typical of the condensed state. The remarkably small changes of basic properties of almost all metals, such as the electrical conductivity or magnetic susceptibility on melting [22], show that the electronic structure of the liquid metal is indeed quite similar to that of the crystalline solid, and both liquid and solid can be reasonably well described by the nearly free-electron approximation. This behaviour is usually rationalized by the fact that the short-range atomic correlations and the atomic density in the liquid near melting are closely similar to those of the crystals. Consequently, the liquid metallic phase is usually treated as a monatomic state, which reflects the solid structure being regarded as single screened metal ions each diffusively uncoupled from every other.
In contrast to melting, however, there are substantial, qualitative changes in the very nature of chemical bonding upon evaporation of a liquid metal. At sufficiently low densities, in the vapour phase, the valence electrons occupy spatially localized atomic or molecular orbitals. We take the example of caesium and mercury for illustration. In this vapour state, many caesium atoms form chemically bound dimers with a dissociation energy of 0.45 eV, whereas in the vapour phase of mercury, chemically bonded dimers do not form because the ground state of the mercury atom arises from a closed-shell electronic configuration, which by itself cannot form an appreciable chemical bond owing to the inert character of the 6s² shell. The interaction potential between two mercury atoms is thus generally regarded as acting between highly polarizable, closed-shell systems involving very little electronic density migration from the constituent partners, in a remarkably similar way to noble gas atoms. In this sense, mercury vapour has been denoted ‘pseudo-helium’ [23], while the vapour of the alkali metal caesium resembles hydrogen [24].

However, this description applies only to conditions near the triple point, and whether the liquid–vapour phase change is always and necessarily accompanied by a metal–non-metal transition, even up to the liquid–vapour critical point, remains one of the major points at issue, dating back to the work of Landau & Zeldovich [11].

This almost certainly is not the case in mercury [25] but may be so in the alkali elements caesium, rubidium and potassium, for which extensive data exist [26]. Another point at issue is whether localized chemical bonding plays a significant role in liquid alkali metals as the density is lowered towards the critical density.

(c) The metallization of fluid hydrogen: the lightest alkali metal

Renewed interest in these questions is motivated by the exciting progress in the search for metallic character of the two most abundant chemical elements in our Universe, hydrogen and helium. Major advances have come from the shock wave experiments of Weir et al. [27] in the case of fluid hydrogen and from first-principles molecular dynamics calculation determined by Stixrude & Jeanloz [28] in the case of fluid helium.

Weir et al. [27] found a highly conducting state of fluid hydrogen (and also for deuterium) at high temperatures and in a pressure range above 1.4 Mbar. Their measurements at these extreme conditions showed electrical conductivity values of about 2000 Ω⁻¹ cm⁻¹, comparable, interestingly and importantly, with those for expanded supercritical fluid caesium and rubidium [29]. Weir et al. [27] interpreted their findings in terms of the existence of H₂ molecules within the highly conducting fluid phase. It has so far, however, not been possible to investigate the structural properties of fluid hydrogen at the extraordinarily high temperatures and pressures prevailing in the shock wave experiment in order to prove whether the pairing character persists in the highly conducting state.

However, given the close similarity in the atomic electronic structure of hydrogen and the alkali metals, an alternative, chemically intuitive approach is to compare and contrast the electrical and structural properties of hydrogen at high temperatures and high pressures with those of the experimentally accessible fluid alkali metals, which may serve as models for the metallization of shock-compressed hydrogen [30–32].

We therefore compare in figure 4 the measured electrical conductivity for compressed fluid hydrogen and fluid caesium at comparable temperatures (2000–4000 K) over a range of pressures.

Importantly, one can easily see that all three fluid elements reveal a continuous transition from a low-conductivity to a highly conducting state with increasing density. In the case of fluid hydrogen at these high temperatures, pressures of more than 1 Mbar are required to effect the transition at an atom density exceeding 3 × 10²³ cm⁻³. By contrast, the alkali metal caesium—recognized to be metallic under ambient conditions of temperature and pressure—reaches the conductivity of 2000 Ω⁻¹ cm⁻¹ already at the relatively moderate pressure of 100 bar and a concomitant atom density of ca 4 × 10²¹ cm⁻³. The large difference in the transition pressures indicates that the elemental density is the dominating factor for the transition to the metallic state. This
is obvious from figure 4, where it is clear that all three elements undergo a ‘density-induced’ continuous transition to the highly conducting state of 2000 \( \Omega^{-1} \text{cm}^{-1} \).

It is noteworthy—and important—that this conductivity value is comparable to that calculated on the assumption that high-temperature fluids metallize only, and remain metallic, when the mean free path of the itinerant conduction electrons becomes comparable to, or exceeds, the mean distance between the particles supplying the electrons—in this case, the constituent atoms of the various chemical elements [31,33].

It is therefore reasonable to speculate that the differences in elemental densities at the transition to metallic behaviour are related to the individual, characteristic atomic properties of the three elements. One such property, for example, is the static electronic polarizability \( \alpha \) of the isolated atom. The relatively small value of \( \alpha \) for atomic hydrogen (0.67 Å\(^3\)) then indicates from the Goldhammer–Herzfeld view that very high elemental densities are required for the transition to the metallic state of hydrogen. By contrast, the high polarizabilities of rubidium (47.3 Å\(^3\)) and caesium (59.7 Å\(^3\)) ensure that these elements are metallic at ambient conditions in the laboratory—which of course is common experience. The metallization densities predicted by the ‘Goldhammer–Herzfeld criterion’ indicated by vertical arrows in figure 4 are in remarkably good agreement with the observed onset of high electrical conductivity for all three chemical elements.

Another property common to compressed fluid hydrogen and the expanded fluid alkali metals points to their chemical similarity in the periodic table of the chemical elements. The characteristic atomic property, for example, is the radius of the principal maximum in the electron charge density of the \( ns \) valence orbital, \( a_{H^*} \), which enters into the venerable Mott criterion for the onset of the metal–non-metal transition, viz.

\[
n_{c}^{1/3} a_{H^*} = 0.26,
\]

and \( n_{c} \) is the critical carrier density for the transition to the metallic state at \( T = 0 \) K [34,35].

It is instructive to plot the measured electrical conductivity as a function of the product—the Mott scaling parameter, \( n^{1/3} a_{H^*} \)—where \( n \) is the valence electron density.

We include here also the data for expanded Hg as well as data for expanded rubidium, caesium and compressed hydrogen [30–32].
Figure 5. The transition to the metallic state for high-temperature \((T > 1000\, \text{K})\) fluid caesium, rubidium, mercury and hydrogen: the dependence of the electrical conductivity on the scaling parameter \(n^{1/3}a_{H^*}\). The dotted line drawn at \(n^{1/3}a_{H^*} = 0.38\) indicates the common metallization condition for these chemical elements. To the left of the metallization condition, we have the non-metallic form of the four elements (non-metallic fluids); to the right, we have the corresponding metallic state (the metallic fluids). Above the metallization threshold, we anticipate conduction based on the theory put forward by Ziman [36].

One sees here (figure 5) a clear scaling behaviour for all of hydrogen, the alkali metals, Rb and Cs, and Hg, and all show the density-induced transition to the conducting metallic state (recall the high temperatures of these studies). The change in the slope at the conductivity value of approximately \(2000\, \Omega^{-1}\, \text{cm}^{-1}\) (i.e. when the mean free path of the conduction electrons becomes comparable to the mean distance between the scattering centres) leads one to conclude that all four of these high-temperature fluids acquire metallic properties at a constant value (approx. 0.38) of the scaling parameter \(n^{1/3}a_{H^*}\). Under the appropriate experimental conditions, one is therefore led naturally to identify all three fluid elements, hydrogen, rubidium and caesium, as metallic, with hydrogen now assuming the position of the lightest alkali metal of group IA of the periodic table.

(d) Molecular dimers: their role in the metallization of hydrogen and the heavier alkali elements

The nature of molecular hydrogen at high pressures has been extensively investigated. Labet et al. [37] examine computationally the transition from paired to monatomic hydrogen. They suggest the possibility of a microscopically non-crystalline phase of hydrogen at elevated pressures, one in which there is a substantial range of roughly equi-enthalpic geometries available to the system.

Weir et al. [27] assumed for fluid hydrogen that the tendency of the structure to dimerize persists up to the densities and temperatures prevailing in the shock experiment. Then the question emerges whether dimers and other clusters can be detected by structural measurements in expanded fluid alkali metals. Naturally, to pose this question is to ask further ‘What is the relation of dimer formation to the metal–non-metal transition?’

The underlying view [38,39] is that the transition in hydrogen and the alkali metals may begin with metallization of the diatomic system by an overlap of the molecular valence and conduction bands (figure 2) and this is followed at higher density by a gradual transition into the monatomic, metallic state. Thus, the dimer molecules present in the vapour may well survive
Figure 6 shows the transition to the metallic state for the high-temperature conditions of the chemical elements helium and mercury, plotting the dependence of the electrical conductivity on the Mott scaling parameter, $n^{1/3}a_H^*$ [44]. The shaded region indicates the transition region to the passage to a low-density liquid. Interesting insights regarding up to what density remnants of the diatomic unit, present in the dilute vapour phase, can survive have been obtained from measurements of the dynamic structure factor of fluid rubidium expanded over a relatively wide density range [39]. At a scaled density $n^{1/3}a_H^* = 0.41$, the observed dynamic structure factor is in accord with molecular dynamics calculations [40] employing effective pair potentials derived by pseudopotential theory, typical for monatomic metals. By contrast, at a lower scaled density of $n^{1/3}a_H^* = 0.37$, drastic changes in the dynamic structure factor are observed and a well-defined excitation peak appears at about $\hbar\omega = 3.2$ meV with the maximum intensity at a momentum transfer near $1\,\text{Å}^{-1}$. This observation was interpreted as an optic-type mode [41] in which two atoms tend to move in opposite directions.

The measured excitation energy of 3.2 meV is surprisingly close to that calculated for Rb$_2$ molecules on a simple cubic lattice [39].

Thus, the results of the calculation and the experiment support the view that rubidium, which is stable as a diatomic molecule at low elemental density, may undergo a density-induced molecule-to-metal transition at a scaled density lower than 0.38, which precedes the occurrence of monatomic Rb, which is analogous to the situation that is suggested to occur in shock-compressed hydrogen in the transition to the conducting metallic phase of hydrogen.

It has been discussed very early (specifically for mercury [42,43]) that there exists a major difference in the way the metallic state develops in divalent systems like mercury and helium and monovalent systems like alkali metals and hydrogen. As is well known, an expanded divalent metal such as mercury is predicted by the Bloch–Wilson band model to undergo a metal–semiconductor transition when the s valence and the p conduction bands no longer overlap. In a crystal, a real gap appears and widens as the density decreases. Mott proposed, specifically for mercury, that the disordered fluid retains features of the crystal’s electronic structure, i.e. a minimum in the density of states at the chemical potential of the electrons develops with decreasing density. As the results of the recent work of Stixrude & Jeanloz [28] for helium are completely consistent with Mott’s scenario developed for mercury, it is certainly tempting to compare experimentally determined properties of mercury with the calculated data for helium at scaled densities $n^{1/3}a_H^*$ with $a^*$ now the s-orbital radius of the isolated helium or mercury atoms, respectively.

Figure 6. The calculated conductivity values of He (at 10 000 K) and the experimentally determined conductivity values of Hg (at 1800 K) as a function of the scaling parameter, $n^{1/3}a_H^*$. The shaded region indicates the transition region to metallic behaviour. This transition range is indicated as the onset of band-structure effects [28,42].
metallic behaviour. It is not, however, trivial for the experimentalist to determine the metallization condition in this instance, since the common (and often used) dependence of the sign of the conductivity change with changing temperature is not obeyed by divalent metals. The sign of the conductivity dependence is always positive at constant volume [36]. This is observed here for He and Hg, in contrast to the alkali metals at higher densities.

Stixrude and Jeanloz [28] argue that the band gap is closed, as s–p hybridization causes the valence band to broaden with increasing temperature. Indeed, it is suggested that the valence band broadens by a factor of 2 from 0 to 50 000 K. The calculated scaled gap closing density of mercury is 0.329 at 10 000 K, which agrees remarkably with the measured scaled closing gap density 0.333 [44].

In Figure 6, the shaded region indicates the transition region to metallic behaviour. This transition range is indicated as the onset of band-structure effects [28,42] (pseudo-gaps or minima in the density of states, i.e. Mott’s famous g-factor). For the lower limit of conductivity (about 200–250), both Mott and Stixrude assume that metallization occurs close to densities and temperatures of gap closing. The upper limit (about 2000) is the ubiquitous Ioffe–Regel limit (i.e. for a scaled density 0.358 in the case of Hg). Above this density, Hg is a nearly free-electron metal [25,36,42]. In the shaded range of figure 6, band-structure effects modify the Ziman formula via Mott’s g-value. Stixrude et al. claim that in the case of He the value of g is everywhere less than unity.

Figure 6 is also illuminating, as it reflects the transition to—and from—the metallic state in helium, conventionally a non-metal, and mercury (‘pseudo-helium’), conventionally a metal in the periodic table of the chemical elements (figure 1).

3. Concluding remarks

The development of the periodic table of the chemical elements is one of the greatest achievements in science. The underlying periodic law is now universally recognized as the consequence of the periodic variation in electronic configuration; its most fundamental ramification leads to either metals or non-metals in the periodic table. It is remarkable to reflect that in the pre-quantum-mechanical descriptions of Goldhammer [17] and Herzfeld [18], one can understand, and rationalize, the metallic or non-metallic status of the chemical elements of the periodic table under both ambient and extreme conditions. The very title of Herzfeld’s [18] paper—‘On atomic properties which make an element a metal’—beautifully emphasizes the pivotal role of atomic properties in dictating the nature—metal or non-metal—of a chemical element in the condensed state (either amorphous or crystalline). Thus, the instantly recognizable ‘diagonal relationship’ separating metals from non-metals in the periodic table is now rationalized and forms the cornerstone of our conventional, universally recognized classification of the chemical elements of the periodic table [1].

References


13. Mott NF. 1996 Communication to P. P. Edwards, cited in [7].


