Worksheet periodic trends answer key

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The elements on the left side of a period tend to form more ionic bonds, while those on the right side form more covalent bonds. Key Terms boiling point: The temperature at which a liquid boils, with the vapor pressure equal to the given external pressure. melting point: The temperature at which the solid and liquid phases of a substance are in equilibrium; it is relatively insensitive to changes in pressure. The periodic table of elements has a total of 118 entries. Elements are arranged in a series of rows (periods) in order of atomic number so that those with similar properties appear in vertical columns. Elements in the same period have the same number of electron shells; moving across a period (so progressing from group to group), elements gain electrons and protons and become less metallic. This arrangement reflects the periodic recurrence of similar properties as the atomic number increases. For example, the alkali metals lie in one group (Group 1) and share similar properties, such as high reactivity and the tendency to lose one electron to arrive at a noble-gas electron configuration. Modern quantum mechanics explains these periodic trends in properties in terms of electron shells. The filling of each shell corresponds to a row in the table. In the s-block and p-block of the periodic table, elements within the same period generally do not exhibit trends and similarities in properties (vertical trends down groups are more significant). However, in the d-block, trends across periods become significant, and the f-block elements show a high degree of similarity across periods (particularly the lanthanides). If we examine the physical state of each element, we notice that on the left side of the table, elements such as lithium and beryllium are metallic solids, whereas on the right, nitrogen, oxygen, fluorine, and neon are all gases. This is because lithium and beryllium form metallic solids, whereas the elements to the right form covalent compounds with little intermolecular force holding them together. Therefore we can say that, in general, elements tend to go from solids to liquids to gases as we move across a given period. However, this is not a strict trend. Bonding As you move across a period in the periodic table, the types of commonly encountered bonding interactions change. For example, at the beginning of Period 2, elements such as lithium and beryllium form only ionic bonds, in general. Moving across the period, elements such as boron, carbon, nitrogen and oxygen tend to form covalent bonds. Fluorine can form ionic bonds with some elements, such as carbon and boron, and neon does not tend to form any bonds at all. Melting Points of the Halides Another physical property that varies across a period is the melting point of the corresponding halide. A halide is a binary compound, of which one part is a halogen atom and the other part is an element or radical that is less electronegative (or more electropositive) than the halogen, to make a fluoride, chloride, bromide, iodide, or astatide compound. Many salts are halides; the hal- syllable in halide and halite reflects this correlation. All Group 1 metals form halides that are white solids at room temperature. The melting point is correlated to the strength of intermolecular bonds within the element. First, we must analyze compounds formed from elements from Groups 1 and 2 (e.g., sodium and magnesium). To develop an understanding of bonding in these compounds, we focus on the halides of these elements. The physical properties of the chlorides of elements in Groups 1 and 2 are very different compared to the chlorides of the elements in Groups 4, 5, and 6. All of the alkali halides and alkaline earth halides are solids at room temperature and have melting points in the hundreds of degrees centigrade. For example, the melting point of sodium chloride (NaCl) is 808 °C. In contrast, the melting points of the non-metal halides from Periods 2 and 3, such as CCl4, PCl3, and SCl2, are below 0 °C, so these materials are liquids at room temperature. Furthermore, all of these compounds have low boiling points, typically in the range of 50 °C to 80 °C. Melting and boiling points of various halides Halide Melting Point (ºC) Boiling Point (ºC) LiCl 610º 1382º BeCl2 405º 488º CCl4 -23º 77º NCl3 −40º 71º OCl2 −20º 4º FCl −154º −101º NaCl 808º 1465º MgCl2 714º 1418º SiCl4 -68º 57º PCl3 −91º 74º SCl2 −122º 59º Cl2 −102º −35º KCl 772º 1407º CaCl 772º > 1600º The non-metal halide liquids are also electrical insulators and do not conduct electrical current. In contrast, when an alkali halide or alkaline earth halide melts, the resulting liquid is an excellent electrical conductor. This tells us that these molten compounds consist of ions, whereas the non-metal halides do not. This again demonstrates the type of bonding that these compounds exhibit: the left-most elements form more ionic bonds, and the further-right elements tend to form more covalent bonds. The physical properties (notably, melting and boiling points) of the elements in a given group vary as you move down the table. Describe the general trends of physical properties within a group on the periodic table. Key Takeaways Key Points The physical properties of elements depend in part on their valence electron configurations. As this configuration remains the same within a group, physical properties tend to remain somewhat consistent. The most notable within-group changes in physical properties occur in Groups 13, 14, and 15, where the elements at the top are non-metallic, while the elements at the bottom are metals. The trends in boiling and melting points vary from group to group, based on the type of non-bonding interactions holding the atoms together. Key Terms physical property: Any property that is measurable whose value describes a physical system's state. malleable: Able to be hammered into thin sheets; capable of being extended or shaped by beating with a hammer or by the pressure of rollers. ductile: Capable of being pulled or stretched into thin wire by mechanical force without breaking. In chemistry, a group is a vertical column in the periodic table of the chemical elements. There are 18 groups in the standard periodic table, including the d-block elements but excluding the f-block elements. Each element within a group has similar physical or chemical properties because of its atom's outermost electron shell (most chemical properties are dominated by the orbital location of the outermost electron). Common Physical Properties A physical property of a pure substance can be defined as anything that can be observed without the identity of the substance changing. The observations usually consist of some type of numerical measurement, although sometimes there is a more qualitative (non-numerical) description of the property. Physical properties include such things as: Color Brittleness Malleability Ductility Electrical conductivity Density Magnetism Hardness Atomic number Specific heat Heat of vaporization Heat of fusion Crystalline configuration Melting temperature Boiling temperature Heat conductivity Vapor pressure Tendency to dissolve in various liquids These are only a few of the measurable physical properties. Within a group of the periodic table, each element has the same valence electron configuration. For example, lithium, sodium, potassium, rubidium, cesium, and francium all have a single electron in an s orbital, whereas every element in the group including fluorine has the valence electron configuration ns2np5, where n is the period. This means the elements of a group often exhibit similar chemical reactivity, and there may be similarities in physical properties as well. Boiling and Melting Points Before a discussion of the melting points of various elements, it should be noted that some elements exist in different forms. For example, pure carbon can exist as diamond, which has a very high melting point, or as graphite, whose melting point is still high but much lower than that of diamond. Different groups exhibit different trends in boiling and melting points. For Groups 1 and 2, the boiling and melting points decrease as you move down the group. For the transition metals, boiling and melting points mostly increase as you move down the group, but they decrease for the zinc family. In the main group elements, the boron and carbon families (Groups 13 and 14) decrease in their boiling and melting points as you move down the group, whereas the nitrogen, oxygen, and fluorine families (Groups 15, 16, and 17) tend to increase in both. The noble gases (Group 18) decrease in their boiling and melting points down the group. These phenomena can be understood in relation to the types of forces holding the elements together. For metallic species, the metallic bonding interaction (electron-sharing) becomes more difficult as the elements get larger (toward the bottom of the table), causing the forces holding them together to become weaker. As you move right along the table, however, polarizability and van der Waals interactions predominate, and as larger atoms are more polarizable, they tend to exhibit stronger intermolecular forces and therefore higher melting and boiling points. Metallic Character Metallic elements are shiny, usually gray or silver in color, and conductive of heat and electricity. They are malleable (can be hammered into thin sheets) and ductile (can be stretched into wires). Some metals, such as sodium, are soft and can be cut with a knife. Others, such as iron, are very hard. Non-metallic atoms are dull and are poor conductors. They are brittle when solid, and many are gases at STP (standard temperature and pressure). Metals give away their valence electrons when bonding, whereas non-metals tend to take electrons. A metal and a non-Metal: On the left is sodium, a very metallic element (ductile, malleable, conducts electricity). On the right is sulfur, a very non-metallic element. Metallic character increases from right to left and from top to bottom on the table. Non-metallic character follows the opposite pattern. This is because of the other trends: ionization energy, electron affinity, and electronegativity. You will notice a jagged line running through the periodic table starting between boron and aluminum – this is the separation between metallic and non-metallic elements, with some elements close to the line exhibiting characteristics of each. The metals are toward the left and center of the periodic table, in the s, d, and f blocks. Poor metals and metalloids (somewhat metal, somewhat non-metal) are in the lower left of the p block. Non-metals are on the right of the table. The electron configuration of a given element can be predicted based on its location in the periodic table. Predict the type of ions an element will form based on its position in the periodic table Key Takeaways Key Points The electron configuration of an element dictates the element's properties in a chemical reaction. Electron configurations vary regularly along the periodic table. The Aufbau principle determines the electron configuration of an element. The principle states that the lowest- energy orbitals are filled first, followed successively by higher-energy orbitals. Magnetism can result from unpaired electrons in a given ion of an element, depending on the spin states of the electrons. Key Terms electron configuration: The arrangement of electrons in an atom, molecule, or other physical structure, such as a crystal. The periodic table does more than just list the elements. The word "periodic" means that within each row, or period, the elements show a pattern of characteristics. This is because the elements are listed in part by their electron configuration. Blocking in the periodic table: The periodic table can be broken into blocks, corresponding to the highest energy electrons. The alkali metals and alkaline earth metals have one and two valence electrons (electrons in the outer shell), respectively; because of this, they lose electrons to form bonds easily and so are very reactive. These elements comprise the s block of the periodic table. The p block, on the right, contains common non-metals, such as chlorine and helium. The noble gases, in the column on the right, almost never react, since they have eight valence electrons forming a stable outer shell. The halogens, directly to the left of the noble gases, readily gain electrons and react with metals. The s and p blocks make up the main- group elements, also known as representative elements. The d block, which is the largest, consists of transition metals, such as copper, iron, and gold. The f block, on the bottom, contains rarer metals, including uranium. Elements in the same group or family have the same configuration of valence electrons, so they behave in chemically similar ways. Periodic table of the elements: This image is color-coded to show the s, p, d, and f blocks of the periodic table. Electron Configuration In atomic physics and quantum chemistry, the electron configuration is the distribution of electrons of an atom or molecule in atomic or molecular orbitals. For example, the electron configuration of the neon atom (Ne) is 1s2 2s2 2p6. According to the laws of quantum mechanics, a certain energy is associated with each electron configuration. Under certain conditions, electrons can move from one orbital to another by emission or absorption of a quantum of energy, in the form of a photon. Knowledge of the electron configurations of different atoms is useful in understanding the structure of the periodic table. The concept is also useful for describing the chemical bonds that hold atoms together. In bulk materials, this same idea helps explain the peculiar properties of lasers and semiconductors. The idea of an electron configuration was first conceptualized under the Bohr model of the atom, and it is still common to speak of "shells" and "subshells" despite the advances in understanding of the quantum-mechanical nature of electrons. Aufbau Principle The Aufbau principle (from the German Aufbau, meaning "building up, construction;" also called the Aufbau rule or building-up principle) is used to determine the electron configuration of an atom, molecule, or ion. The principle postulates a hypothetical process in which an atom is "built up" by the progressive addition of electrons. As electrons are added, they assume their most stable positions (electron orbitals) with respect to the nucleus and the electrons that are already there. According to the principle, electrons fill orbitals starting at the lowest available energy state before filling higher states (e.g., 1s before 2s). The number of electrons that can occupy each orbital is limited by the Pauli exclusion principle. If multiple orbitals of the same energy are available, Hund's rule states that unoccupied orbitals will be filled before occupied orbitals are reused (by electrons having different spins). Atomic orbitals ordered by increasing energy: Order in which orbitals are arranged by increasing energy according to the Madelung rule. Each diagonal red arrow corresponds to a different value of n + l. Magnetism Magnetism is a property of materials that respond to an applied magnetic field. Permanent magnets have persistent magnetic fields caused by ferromagnetism, the strongest and most familiar type of magnetism. However, all materials are influenced differently by the presence of a magnetic field. Some are attracted to a magnetic field (paramagnetism); others are repulsed by it (diamagnetism); still others have a much more complex relationship with an applied magnetic field (e.g., spin-glass behavior and antiferromagnetism). Substances that are negligibly affected by magnetic fields are considered non-magnetic, these are: copper, aluminum, gases, and plastic. Pure oxygen exhibits magnetic properties when cooled to a liquid state. The magnetic properties of a given element depend on the electron configuration of that element, which will change when the element loses or gains an electron to form an ion. If the ionization of an element yields an ion with unpaired electrons, these electrons may align the sign of their spins in the presence of a magnetic field, making the material paramagnetic. If the spins tend to align spontaneously in the absence of a magnetic field, the resulting species is termed ferromagnetic. Hierarchy for various types of magnetism: There are various types of magnetism identified to date that can be organized in a hierarchy. Applications of Magnetism A lodestone, or loadstone, is a naturally magnetized piece of the mineral magnetite (Fe3O4). Ancient people first discovered the property of magnetism in lodestone. Pieces of lodestone, suspended so they could turn, were the first magnetic compasses, and their importance to early navigation is indicated by their very name, which in Middle English means "course stone" or "leading stone." Lodestone is one of only two minerals that is found naturally magnetized; the other, pyrrhotite, is only weakly magnetic. Atomic radii decrease from left to right across a period and increase from top to bottom along a group. Predict the relative atomic sizes of the elements based on the general trends in atomic radii for the periodic table. Key Takeaways Key Points The atomic radius of a chemical element is a measure of the size of its atoms, usually the mean or typical distance from the nucleus to the boundary of the surrounding cloud of electrons. Since the boundary is not a well-defined physical entity, there are various non-equivalent definitions of atomic radius. The periodic trends of the atomic radii (and of various other chemical and physical properties of the elements) can be explained by the electron shell theory of the atom. Key Terms quantum theory: A theory developed in early 20th century, according to which nuclear and radiation phenomena can be explained by assuming that energy only occurs in discrete amounts called quanta. electron shell: The collective states of all electrons in an atom having the same principal quantum number (visualized as an orbit in which the electrons move). noble gas: Any of the elements of Group 18 of the periodic table, being monatomic and (with very limited exceptions) inert. In chemistry, periodic trends are the tendencies of certain elemental characteristics to increase or decrease as one progresses along a row or column of the periodic table of elements. The atomic radius is one such characteristic that trends across a period and down a group of the periodic table. Periodic trends: A graphic showing overall periodic trends in the periodic table. Meaning of the Atomic Radius The atomic radius of a chemical element is a measure of the size of its atoms, usually the mean or typical distance from the nucleus to the boundary of the surrounding cloud of electrons. Since the boundary is not a well-defined physical entity, there are various non-equivalent definitions of atomic radius. Depending on context, the term atomic radius may apply only to isolated atoms, or also to atoms in condensed matter, covalently bound in molecules, or in ionized and excited states. The value of an atomic radius may be obtained through experimental measurements or computed with theoretical models. Under some definitions, the value of a radius may depend on the atom's state and context. For our purposes, we are generally looking at atoms in their elemental state. Sizes of atoms and their ions in picometers (pm): Red numbers are ionic radii of cations, black numbers are for neutral species, and blue numbers are for anions. Atomic radii vary in a predictable and explicable manner across the periodic table. Radii generally decrease from left to right along each period (row) of the table, from the alkali metals to the noble gases; radii increase down each group (column). The radius increases sharply between the noble gas at the end of each period and the alkali metal at the beginning of the next period. These trends of the atomic radii (and of various other chemical and physical properties of the elements) can be explained by the electron shell theory of the atom. Radii measurements provided important evidence for the development and confirmation of quantum theory. Explanation of the General Trends The way atomic radius varies with increasing atomic number can be explained by the arrangement of electrons in shells of fixed capacity. Shells closer to the nucleus—those with a smaller radius—are generally filled first, since the negatively charged electrons are attracted by the positively charged protons in the nucleus. As the atomic number increases along a row of the periodic table, additional electrons are added to the same, outermost shell. The radius of this shell gradually contracts as the attraction between the additional electrons and the nucleus increases. In a noble gas, the outermost shell is completely filled. Therefore, the additional electron of next alkali metal (one row down on the periodic table) will go into a new outer shell, accounting for the sudden increase in the atomic radius. Atomic number to radius graph: A chart showing the atomic radius relative to the atomic number of the elements. The increasing nuclear charge is partly counterbalanced by the increasing number of electrons, a phenomenon that is known as shielding; this explains why the size of atoms usually increases down each column. Underlying causes of the periodic trends in atomic radius also have an impact on other chemical and physical properties of the elements. Similarly charged ions tend to decrease in size across a period (row) and increase in size down a group (column). Identify the general trends of the ionic radius size for the periodic table. Key Takeaways Key Points The ionic radius is the distance between the nucleus and the electron in the outermost shell of an ion. When an atom loses an electron to form a cation, the lost electron no longer contributes to shielding the other electrons from the charge of the nucleus; consequently, the other electrons are more strongly attracted to the nucleus, and the radius of the atom gets smaller. When an electron is added to an atom, forming an anion, the added electron repels other electrons, resulting in an increase in the size of the atom. The trend observed in size of ionic radii is due to shielding of the outermost electrons by the inner-shell electrons so that the outer shell electrons do not "feel" the entire positive charge of the nucleus. Key Terms cation: A positively charged ion, as opposed to an anion. ion: An atom or group of atoms bearing an electrical charge, such as the sodium and chlorine atoms in a salt solution. anion: A negatively charged ion, as opposed to a cation. In chemistry, periodic trends are the tendencies of certain elemental characteristics to increase or decrease along a period (row) or group (column) of the periodic table of elements. Ionic radius (rion) is the radius of an ion, regardless of whether it is an anion or a cation. Although neither atoms nor ions have sharp boundaries, it is useful to treat them as if they are hard spheres with radii. In this way, the sum of ionic radii of a cation and an anion can give us the distance between the ions in a crystal lattice. Ionic radii are typically given in units of either picometers (pm) or Angstroms (Å), with 1 Å = 100 pm. Typical values range from 30 pm (0.3 Å) to over 200 pm (2 Å). Trends in Ionic Radii Ions may be larger or smaller than the neutral atom, depending on the ion's charge. When an atom loses an electron to form a cation, the lost electron no longer contributes to shielding the other electrons from the charge of the nucleus; consequently, the other electrons are more strongly attracted to the nucleus, and the radius of the atom gets smaller. Similarly, when an electron is added to an atom, forming an anion, the added electron repels other electrons, resulting in an increase in the size of the atom. The ionic radius is not a fixed property of a given ion; rather, it varies with coordination number, spin state, and other parameters. For our purposes, we are considering the ions to be as close to their ground state as possible. Nevertheless, ionic radius values are sufficiently transferable to allow periodic trends to be recognized. Sizes of atoms and their ions: Relative sizes of atoms and ions. The neutral atoms are colored gray, cations red, and anions blue. As with other types of atomic radii, ionic radii increase upon descending a group and decrease going across a period. Note that this only applies if the elements are the same type of ion, either cations or anions. For example, while neutral lithium is larger than neutral fluorine, the lithium cation is much smaller than the fluorine anion, due to the lithium cation having a different highest energy shell. The ionization energy tends to increase as one moves from left to right across a given period or up a group in the periodic table. Recognize the general periodic trends in ionization energy. Key Takeaways Key Points The ionization energy is the energy required to remove an electron from its orbital around an atom to a point where it is no longer associated with that atom. The ionization energy of an element increases as one moves across a period in the periodic table because the electrons are held tighter by the higher effective nuclear charge. The ionization energy of the elements increases as one moves up a given group because the electrons are held in lower-energy orbitals, closer to the nucleus and therefore are more tightly bound (harder to remove). Key Terms ionization energy: The energy needed to remove an electron from an atom or molecule to infinity. The ionization energy of a chemical species (i.e., an atom or molecule) is the energy required to remove electrons from gaseous atoms or ions. This property is also referred to as the ionization potentia and is measured in volts. In chemistry, it often refers to one mole of a substance (molar ionization energy or enthalpy) and is reported in kJ/mol. In atomic physics, the ionization energy is typically measured in the unit electron volt (eV). Large atoms or molecules have low ionization energy, while small molecules tend to have higher ionization energies. The ionization energy is different for electrons of different atomic or molecular orbitals. More generally, the nth ionization energy is the energy required to strip off the nth electron after the first n-1 electrons have been removed. It is considered a measure of the tendency of an atom or ion to surrender an electron or the strength of the electron binding. The greater the ionization energy, the more difficult it is to remove an electron. The ionization energy may be an indicator of the reactivity of an element. Elements with a low ionization energy tend to be reducing agents and form cations, which in turn combine with anions to form salts. Ionization energy: This graph shows the first ionization energy of the elements in electron volts. Moving left to right within a period or upward within a group, the first ionization energy generally increases. As the atomic radius decreases, it becomes harder to remove an electron that is closer to a more positively charged nucleus. Conversely, as one progresses down a group on the periodic table, the ionization energy will likely decrease since the valence electrons are farther away from the nucleus and experience greater shielding. They experience a weaker attraction to the positive charge of the nucleus. Ionization energy increases from left to right in a period and decreases from top to bottom in a group. Rationale for the Periodic Trends in Ionization Energy The ionization energy of an element increases as one moves across a period in the periodic table because the electrons are held tighter by the higher effective nuclear charge. This is because additional electrons in the same shell do not substantially contribute to shielding each other from the nucleus, however an increase in atomic number corresponds to an increase in the number of protons in the nucleus. The ionization energy of the elements increases as one moves up a given group because the electrons are held in lower-energy orbitals, closer to the nucleus and thus more tightly bound (harder to remove). Based on these two principles, the easiest element to ionize is francium and the hardest to ionize is helium. Periodic trends in ionization energy – YouTube: This video explains the periodic trends in ionization energy....periodicity. The periodic table is arranged in a manner to show trends in the characteristics of the elements. The electron affinity of the elements generally increases across a period and sometimes decreases down a group in the periodic table. Recognize the general periodic trends for electron affinity. Key Takeaways Key Points The electron affinity of an atom or molecule is the propensity for that particle to gain an electron. This is an exothermic process for all non-noble gas elements. There are general trends in electron affinity across and down the periodic table of elements. Electron affinity generally increases across a period in the periodic table and sometimes decreases down a group. These trends are not necessarily universal. The chemical rationale for changes in electron affinity across the periodic table is the increased effective nuclear charge across a period and up a group. Key Terms electron affinity: The electron affinity of an atom or molecule is defined as the amount of energy released when an electron is added to a neutral atom or molecule to form a negative ion. electronegativity: The tendency of an atom or molecule to attract electrons to itself. The electron affinity (Eea) of a neutral atom or molecule is defined as the amount of energy released when an electron is added to it to form a negative ion, as demonstrated by the following equation: [latex]X(g) + e^- \rightarrow X^{-}(g)[/latex] Electron affinity is measured for atoms and molecules in the gaseous state only, since in the solid or liquid states their energy levels would be changed by contact with other atoms or molecules. Robert S. Mulliken used a list of electron affinities to develop an electronegativity scale for atoms by finding the average of the electron affinity and ionization potential. A molecule or atom that has a more positive electron affinity value is often called an electron acceptor; one with a less positive electron affinity is called an electron donor. Together they may undergo charge-transfer reactions. To use electron affinities properly, it is essential to keep track of the sign. For any reaction that releases energy, the change in energy (ΔE) has a negative value, and the reaction is called an exothermic process. Electron capture for almost all non-noble gas atoms involves the release of energy and therefore is an exothermic process. Periodic Properties: Part 4, Ionic Charges, Ionization Energy, Electron Affinity – YouTube: We conclude our discussion of periodic properties by wrapping up the prediction of ionic charges of the transition metals, ionization energies, and electron affinity. Confusion may arise in mistaking Eea for ΔE. The numbers listed in tables of Eea are all positive because they are magnitudes; the values of Eea in a table of electron affinities all indicate the amount of energy released when an electron is added to an element. Because the release of energy is always an exothermic event, these all correspond to negative values of ΔE (indicating an exothermic process). Periodic Trends in Electron Affinity Although Eea varies greatly across the periodic table, some patterns emerge. Generally, nonmetals have more positive Eea than metals. Atoms, such as Group 7 elements, whose anions are more stable than neutral atoms have a higher Eea. The electron affinities of the noble gases have not been conclusively measured, so they may or may not have slightly negative values. Chlorine has the highest Eea while mercury has the lowest. Eea generally increases across a period (row) in the periodic table, due to the filling of the valence shell of the atom. For instance, within the same period, a Group-17 atom releases more energy than a Group-1 atom upon gaining an electron because the added electron creates a filled valence shell and therefore is more stable. A trend of decreasing Eea down the groups in the periodic table would be expected, since the additional electron is entering an orbital farther away from the nucleus. Since this electron is farther away, it should be less attracted to the nucleus and release less energy when added. However, this trend applies only to Group-1 atoms. Electron affinity follows the trend of electronegativity: fluorine (F) has a higher electron affinity than oxygen (O), and so on. The trends noted here are very similar to those in ionization energy and change for similar (though opposing) reasons. Electron affinities in the periodic table: This table shows the electron affinities in kJ/mol for the elements in the periodic table.