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Arrange the following elements from greatest to least tendency to accept an electron. cl s si na mg

Ionization energy and electron affinity The energy required to remove one or more electrons from atoms in an inability to form the first ionized energy-ionization ions is a physical property that affects the chemical action of atoms. By definition, the first ionization energy of an element is the outermost or highest energy in a natural atom on a gas, the energy required to remove electrons. The process of measuring the first ionization energy of hydrogen is shown by the following equation: H (g) H + (g) + e-Ho = -1312.0 kJ / mol the size of the first ionization energy of hydrogen can be brought into perspective as compared to the energy given in the chemical reaction. When natural gas is burned, about 800 kJ of energy is released per mole of methane consumed. CH4(g) + 2 O 2 (g) CO 2 (g) + 2 H 2 O (g) arc = -802.4 kJ / mol thermite reaction used for iron rail welding consumes about 850kJ of energy per iron oxide mole. Fe 2O3(s) + 2 Al2O3(s) + 2 Fe (l) Ho = -851.5 kJ / mol the first ionization of hydrogen is half as large again as the energy given in one of these reactions. The first ionization energy for the pattern helium of the first ionization energy is less than twice the ionization energy for hydrogen because each electron in helium feels the attractive force of two protons instead of one. He (g) that +(g) + e-Ho = 2372.3 kJ/mol it takes much less energy, however, to remove electrons from lithium atoms, that have three protons in the nucleus. Li (g) Li+(g) + e-Ho = 572.3 kJ/mol can be explained by pointing out that the outermost, or highest energy, electrons are in the 2s orbit to lithium atoms. Since electrons in 2s orbits are already at higher energies than electrons in 1s orbits, there is less energy to remove these electrons from atoms. The first ionization energy for the main group elements is provided in the two figures below. Two trends are evident in this data. Typically, the first ionization energy increases as it moves from left to right across rows of periodic tables. When you go down the heat of a periodic table, the first ionization energy decreases. The first trend is not surprising. As the number of protons in the nucleus of an atom grows, the force of attraction between the nucleus and the electron increases, so we can expect the first ionization energy to be greater as it travels across rows of periodic marks. The second trend comes from the fact that as we descend the heat of a periodic table, the main quantum number of orbits holding the outermost electrons becomes larger. Although the number of protons in the nucleus is also larger, electrons in smaller shells and subshells tend to inspect the outermost electrons from some of the forces of the nucleus's attraction. In addition Electrons are removed when the first ionization energy is measured and spend less time near the nucleus of the atom, so there is less energy to remove this electron from the atom. The figure below shows the exception to the general pattern of the first ionization energy shows the first ionization energy for the element in the second row of the periodic table. There is a general trend towards an increase in the first ionization energy as it moves from left to right across this row, but there are two minor reversals in this pattern. The first ionization energy of boron is less than beryllium, and the first ionization energy of oxygen is less than nitrogen. These observations can be explained by looking at the electronic composition of these elements. Electrons are removed when beryllium atoms are ionized, but 2p electrons are removed when boron is ionized. Be: [He] 2s2 B: [He] 2s2 2p1 electrons removed when nitrogen and oxygen were ionized also came from a 2p orbit. N: [He] 2s2 2p3 O: [He] 2s2 2p4, but there are important differences in how electrons are distributed in these atoms. Hund's rules predict that all three electrons in the 2p orbit of a nitrogen atom have the same spin, but the electrons are joined in one of the 2p orbits of oxygen atoms. Hund's rules are understandable, assuming that electrons should stay as far away as possible to minimize repulsion between these particles. So in a 2p orbit of nitrogen, the three electrons enter different orbits with spins aligned in the same direction. In oxygen, two electrons must account for one of the 2p orbits. The repulsion between these electrons is minimized to some extent by pairing the electrons. However, because there is still residual between these electrons, it becomes slightly easier to remove electrons from the atoms of atoms than you would expect from the number of protons in the nucleus of an atom. Second, third, fourth, and high ionization energy, so you know that sodium or + ions are formed, magnesium forms Mg2 + ions, and aluminum forms Al3 + ions. But have you ever wondered why sodium doesn't form Na2+ ions, even Na3+ ions? The answer can be obtained from the data on the second, third, and higher ionization energy of the element. The first ionization energy of sodium is, for example, the energy it takes to remove one electron from a erythy atom. Na(g) + Energy Na+(g) + e- The second ionization energy is the energy it takes to remove other electrons to form Na2+ ions on the gas. Na + (g) + energy Na 2 + (g) + e- the third ionization energy may be represented by the following equation: Na2+(g) + Energy Na3+(g) + e- The energy required to form Na3+ ions on the gas is the sum of the first, second, and third ionization energy of the element. First, second, third, The fourth ionization energy of sodium, magnesium and aluminum (kJ/mol) is a filling shell electron configuration that does not take much energy to remove a single electron from a sodium atom to form Na+ ions. However, once this is done, removing the second electron requires nearly 10 times the energy to break into this filled shell configuration. Because it takes more energy to remove a second electron than is given in a chemical reaction, sodium contains Na+ ions, but can react with other elements to form compounds containing Na2+ or Na3+ ions. Similar patterns are observed when analyzing the ionization energy of magnesium. Magnesium's first ionization energy is more than sodium because it has one more proton in its nucleus to hold electrons in 큼 orbit. Mg: [Ne] 3s2 Mg's second ionization energy is always higher than the first because it takes more energy to remove electrons from positively charged ions than 큼 atoms. However, because Mg2+ ions have a filled shell electron composition, magnesium's third ionization energy is enormous. The same pattern can be seen in the ionization energy of aluminum. Aluminum's first ionization energy is smaller than magnesium. Aluminum's second ionization energy is greater than the first, and the third ionization energy is 큼 energy. Removing three electrons from an aluminum atom to form Al3+ ions requires a significant amount of energy, but the energy required to break into the filled shell composition of al3+ ions is astronomical. Thus, it will be a mistake to find Al4+ ions as a product of chemical reactions. Exercise 5: Predict the group in a periodic table where the elements with the next ionization energy are most likely. Primary IE = 786 kJ/mol 2th IE = 1577 3RD IE = 3232 4th IE = 4355 5TH IE = 16,091 6th IE = 19,784 Exercises 5 Exercises 6:0 Check the element's ionized energy trends to explain the next observation. (a) Elements on the left side of the periodic table are more likely to form positive ions than elements on the right. (B) the maximum positive potential of the ions is the same as the munitions of the element, where the practice problem 6 to determine the answer to the electron affinity ionization energy to measure the tendency of the heavy atoms to resist the loss of electrons. For example, removing electrons from adelectoms of atoms of atoms requires a significant amount of energy to form positively charged ions. F(g) F+(g) + e-Ho = 1681.0 kJ/mol When atoms obtain additional electrons on a gas, the electron affinity of the element obtains extra electrons to form negatively charged ions. For example, fluoro atoms on a gas emit energy when they get electrons to form fluoro ions. F(g) + F-(g) Ho = -328.0 kJ/mol electron affinity is more difficult to measure than ionized energy and is generally known as less important. The electron affinity of the main group elements is shown in the figure below. You can find multiple patterns in this data. Electronic affinity usually gets smaller when you go down a column in a periodic table for two reasons. First, the electrons added to the atoms are placed in a larger orbit that spends less time near the nucleus of the atom. Second, the number of electrons in atoms increases as we go down the column, so electrons are added and the repulsion between electrons already present in neutral atoms increases. Electron affinity data is complicated by the fact that the repulsion between atoms already present in atoms and electrons added to electrons depends on the volume of the atoms. Among the via groups and non-metals of VIIA, this force of repulsion is the largest for very small atoms in this column: oxygen and fluororus. As a result, these elements have a smaller electron affinity than the elements below this column, as shown in the figure below. But at that point, electronic preferences decrease as we keep this column down. At first glance, as shown in the figure below, there appears to be no pattern in the electronic affinity across the rows of periodic tables. However, it makes sense if such data is listed with the electronic configuration of these elements. This data can be explained by pointing out that electron affinity is much smaller than ionized energy. As a result, elements such as helium, beryllium, nitrogen and neon with unusually stable electron compositions have a small affinity for extra electrons, where energy is not given when neutral atoms of these elements pick up electrons. This configuration is so stable that it actually takes energy to get one of these elements to choose extra electrons to form the cathode. 전자 선호도 및 전자 구성 주기표 요소 전자 선호도 (kJ/mol) 전자 구성 H 72.8 1s1 그는 <0 1s2 Li 59.8 [그는] 2s1 < 2s2 B 27 [그는] 2s2 2p1 C 122.3 [그는] 22 [그는] 222.3 [그는] 22 [그는] 22 [그는] 22 [그는] 222.3 [그는] 22 [그는] 222.3 [그는] 22 [그는] 22 [그는] 222.3 [그는] 22 [그는] 22 [그는] 22 < [그는] 222.3 [그는] 22 [그는] 222[22] [그는] 22 [그는] 222[22] s2 2p3 O 141.1 [그는] 2s2 2p4 F 328.0 [그는] 2s2 2p5 Ne <0 [그는] 2s2 2p6 이온화 에너지와 전자 Affinities의 상대적 크기의 결과 학생들은 종종 나트륨이 염소와 반응하여 Na+와 Cl-Ions를 형성한다고 Believe. There is no doubt that sodium reacts violently with chlorine to form NaCl. 2 Nas) + Cl2 (g) 2 NaCl(s) In addition, the ease with which NaCl's solution in water is transporting electricity is a testament to the fact that the product in this reaction is salt containing Na+ and Cl-ions. NaCl(s) H2O Na+(aq) + Cl- (aq) The only question is whether it is legal to assume that this reaction occurs because chlorine atoms like electrons over sodium atoms. The first ionization energy for sodium is one and a half times more than electron affinity for chlorine큼 chlorine. Na: Primary IE = 495.8 kJ/mol Cl: EA = 328.8 kJ/mol Therefore, more energy is needed to remove electrons from a sodium atom than is given when capturing electrons by a chlorine atom. We will obviously have to find another explanation as to why sodium reacts with chlorine to form NaCl. Before we can do this, however, we need to know more about the chemistry of ion compounds. Compound.

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