Which electron affinity below would you predict to be the most exothermic?

Ionization Power and Electron Affinity Power Ionization First Energy required to remove one or more electrons from neutral atoms to form positively charged ions is a physical property that affects atomic chemical behavior. By definition, the element's first ionization energy is the energy needed to eliminate the outermost, or highest, electrons from the neutral atoms in the gas phase. The process by which hydrogen's first ioning power is measured will be represented by the following equations. H(g) H+(g) + e- Ho = -1312.0 kJ/mol Magnitud hydrogen's first ioning power can be brought into perspective by comparing it to the energy provided in chemical response. When we burn natural gas, about 800 kJ of energy is released per mole of methane used. CH4(g) + 2 O2(g) CO2(g) + 2 H2O(g) Ho = -802.4 kJ / mol Terrmite reaction, which is used to impeal iron rails, provides approximately 850 kJ of power per iron oxide mole used. Fe2O3(s) + 2 Al(s) Al2O3(s) + 2 Fe(l) Ho = -851.5 kJ/mol Hydrogen's first ionation power is the other half as far as the energy provided in any of these responses. The pattern in the First Ionization Energy the first ionization energy for helium is slightly less than twice the ionization energy for hydrogen because each electron in the helium senses the pulling power of two protons instead of one. Dia(g) He+(g) + e- Ho = 2372.3 kJ / mol It takes much lower power, however, to eject electrons from lithium atoms, which have three protons in their nucleus. Li(g) Li+(g) + e- Ho = 572.3 kJ / mol It can be explained by stating that the outer energy, or highest, electrons in lithium atoms are in orbit 2s. Since electrons in orbit 2s are already at higher energy than electrons in orbit 1s, it takes less energy to remove these electrons from atoms. The first ioning power for the main set element is given in the two numbers below. Two trends are clear from this data. In general, the first ionization energy increases when we go from left to right across the periodic table lines. The first ioning energy decreases when we go down the regular schedule column. The first trend is not surprising. We might expect the first ionization energy to become larger as we cross a row of periodical tables as the attraction between the nucleus and electrons becomes greater as the number of protons in the atomic nucleus becomes larger. The second trend is evident from the fact that the number of major quantum orbits holding the outer electrons becomes larger as we descend the periodic schedule column. Although the number of protons in the nucleus also becomes larger, electrons in smaller grafts and subshells tend to filter out electrons from some attraction In addition to that electrons ejected when the first ionization power is measured spend less time near the atomic nucleus, and therefore it takes less energy to remove these electrons from atoms. The exception to the First Ionation Powered General Pattern Below indicates the first ionation spirit for the elements in the second row of the periodic schedule. While there is a general trend towards the first increase in ionization power as we go from left to right across this line, there are two small barriers in this pattern. Boron's first ionation power was smaller than beryllium, and the first ionation energy of oxygen was smaller than nitrogen. This observation can be explained by looking at the electron configuration of these elements. Electrons are ejected when diioned beryllium atoms come from orbit 2s, but 2p electrons are ejected when boron diions. Be: [He] 2s2 B: [He] 2s2 2p1 Electrons are ejected when nitrogen and oxygen diions also come from orbit 2p. N: [He] 2s2 2p3 O: [He] 2s2 2p4 But there is an important difference in the way electrons are circulated in these atoms. Hund rules foresee that the three electrons in the 2p orbit of nitrogen atoms all have the same spin, but the electrons are paired in one of the 2p orbits on the oxygen atom. Hund's rules can be understood by assuming that electrons try to stay as far away as possible to minimize the rejection power between these zarahs. All three electrons in a 2p orbit on nitrogen therefore enter different orbits with their spins aligned in the same direction. In oxygen, two electrons must occupy one of the 2p orbits. The resurance between these electrons is minimized slightly by pairing the electrons. There are still some residual repulations between these electrons, however, which makes it a little easier to eject electrons from neutral oxygen atoms than we would expect from the number of protons in the atomic nucleus. Second, Third, Fourth, and Higher Ionation Powered Now you know that sodium forms Na + ions, magnesium forms Mg2+ ions, and aluminum forms Al3+ ions. But have you ever wondered why sodium doesn't form Na2+ ions, or even Na3+ ions? The answer can be obtained from data for both, third, and higher element-powered ioning. The first ionization power of sodium, for example, is the energy needed to remove a single electron from a neutral atom. Na(g) + Power Na +(g) + e- Second ionization power is the power needed to eliminate other electrons to form Na2+ ions in the gas phase. Na+(g) + Na2+(g) + e- The third ioning power can be represented by the following equations. Na2+(g) + Power Na3+(g) + e- The power required to form Na3+ ions in the gas phase is the first, second, and third ionation power elements. First, Second, Third, Third, Fourth Ionization Power of Sodium, Magnesium, and Aluminum (kJ/mol) It does not take much energy to remove a single electron from the sodium atom to form Na+ ions with an electron configuration containing cangkerang. However, once this is done, it takes nearly 10 times as much energy to break into this containing graft configuration to eliminate the second electron. Since it requires more energy to dispose of the second electron than is given in any chemical response, sodium can respond with other elements to form compounds containing Na+ ions but not Na2+ or Na3+ ions. The same pattern is noticed when magnesium-powered ioning is analyzed. The first ionization energy of magnesium is greater than sodium because magnesium has one more proton in its nucleus to hold on to electrons in orbit 3s. Mg: [Ne] 3s2 Mg's second ionization energy is greater than the first because it always takes more energy to remove electrons from positively charged ions than neutral atoms. The third ionization energy of magnesium is huge, however, because Mg2+ ions have an electron configuration containing cangkerang. The same pattern can be seen in aluminum ioning powered. The first aluminum ioning power is smaller than magnesium. The aluminum second ioning power is greater than the first, and the third ionation power is larger. Although it takes a huge amount of energy to eject three electrons from aluminum atoms to form Al3+ ions, the energy needed to break down into configurations containing Al3+ ion grafts is astronomical. Therefore, it would be a mistake to look for Al4+ ions as a chemical response product. Practice Issue 5:Foresees groups in a periodic schedule where elements with the following ioning power are most likely to be encountered. First IE = 786 kJ / mol 2nd IE = 1577 3rd IE = 3232 4th IE = 4355 5th IE = 16.36,36 4091 IE 6th = 19,784 Click here to check your answer to Practice Problem 5 Problem Practice 6:Use trend in powerful ionization of elements to explain the following observation. (a) The elements on the left of the periodic schedule are more likely than those on the right to form positive ions. (b) The maximum positive charge on ions is the same as the number of element sets Click here to check your answer to problem practice 6 Affinity Electron Ionization energy measures the tendency of neutral atoms to oppose electron loss. It takes a significant amount of energy, for example, to eject electrons from neutral fluorine atoms to form positively charged ions. F(g) F+(g) + e- Ho = 1681.0 kJ/mol Electron connection element is when neutral atoms in the gas phase get additional electrons to form negatively charged ions. Fluorine atoms in the gas phase, for example, provide energy when they get electrons to form fluoride ions. F(g) + + F-(g) Ho = -328.0 kJ / Electron mole connection is harder to measure than ionized power and is usually known to be less important numbers. The electron connection of the main group elements is shown in the numbers below. Several patterns can be found in this data. The electron connection usually becomes smaller as we go down the periodic schedule column for two reasons. First, the electrons added to the atom are placed in a larger orbit, where it spends less time near the atomic nucleus. Second, the number of electrons in atoms increases when we descend the column, so the regur outsizing power between the plus electrons and the electrons already in neutral atoms becomes greater. Electron connection data is complicated by the fact that the repulation between the electrons added to the atom and the electrons that already exist in the atom depends on the number of atoms. Between non-exemplary in the VIA and VIIA Groups, this resurance is the largest for the least atoms in this column: oxygen and fluorine. As a result, these elements have a smaller electron connection than the elements below in this column as shown in the numbers below. However, from that angle, the electron affinity decreases as we continue this column. At first glance, there appears to be no pattern in the electron connection across the periodic table lines, as shown in the numbers below. When these data are listed together with the electron configuration of these elements, however, they make sense. This data can be explained by stating that the electron connection is much smaller than the ionization power. As a result, elements such as helium, beryllium, nitrogen, and neon, which have an unusually stable electron configuration, have little connection to the additional electrons that no energy is given when the neutral atoms of these elements take up electrons. This configuration is so stable that it actually takes power to force one of these elements to take additional electrons to form negative ions. Electron Affinity and Electron Configuration for the First 10 Elements in the Affinity Electron Element Periodic Schedule (kJ/mol) Electron Configuration H 72.8 1s1 Dia <0 1s2 Li 1s2 Li59.8 [He] 2s1 Be <0 [He] 2s2 B 27 [He] 2s2 2p1 C 122.3 [He] 2s2 2p2 N & lt;0 [He] 2s2 2p3 O 141.1 [He] 2s2 2p4 F 328.0 [He] 2s2 2p5 Ne <0 [He] 2s2 2p6 Due to the Relative Size of Ionization Energies and Electron Affinities Students often believe that sodium responds with chlorine to form Na+ and Cl-ion because chlorine atoms like electrons are more than sodium atoms do. There is no doubt that sodium responds seriously with chlorine to form NaCl. 2 Na(s) + Cl2(g) 2 NaCl(s) Additionally, whichever NaCl in electric running is proof that this response product is salt, which contains Na+ and Cl-ion. Cl-ion. NaCl(s) H2O Na+(aq) + Cl-(aq) The only question is whether it is valid to assume that this reaction occurred because chlorine atoms such as electrons are more than sodium atoms. The first ionization energy for sodium is one and a half times greater than the electron connection for chlorine. Na: First IE = 495.8 kJ / mol Cl: EA = 328.8 kJ / mol Therefore, it requires more energy to eject electrons from neutral sodium atoms than is given when electrons are picked up by neutral chlorine atoms. We obviously need to find another explanation for why sodium responds with chlorine to form NaCl. Before we can do this, however, we need to know more about ionic compounds. The word majmuk.