Chemistry 1 periodic trends worksheet

Name: __ Section: __ You should try to answer the question without speaking to your tutorial. If you're stuck, try asking another group for help. The size of the atom or ion and gravity between the core and the outer electrons play an important role in determining the chemistry of the element. Knowledge of trends in atomic and ionic sizes, ionics and affinity of electrons helps to understand the chemical behavior and nature of chemical bonds. Educational purpose Know periodic trends of atomic size, Ionic size, energy of ionization and affinity of electrons Understand the causes of metal, non-metallic and metallic nature Understand why some acids dissolve in water to make an acidic solution, while others dissolve in water to make major decisions Success criteria The ability to predict size differences between atoms and ions Be able to predict differences in energy ionization and electronic affinity between elements Be able to predict the relative reactivity of related elements Be able to predict the acid-alkaline nature of oxides Trends in the size of atoms are most important for understanding, because other trends can often be rationalized on this basis. The most commonly used measure of the size of an atom is its gluing atomic radius, also called the covalent radius (commonly given in picometer units (pm = 10-12 m) or Ångstroms (Å = 10-10 m) with 1 Å = 100 pm). The gluing atomic radius of the element is used as half the distance between the cores, when two identical atoms are fastened together. For example, the inter-nuclear distance in Cl2 is 198 p.m., so the atomic radius is taken before 99 p.m. or 0.99 Å. When two different atoms are fastened together, their dimensions can be changed. However, the sum of their individual atomic council is often used as the first approximation of the length of communication between two different atoms. Thus, the communication length of C-Cl in CCl4 is estimated as the amount of atomic carbon radius (0.77 Å) and chlorine (0.99 Å), which is 1.76 Å. The experimentally determined distance is 1,766 O. The following plot shows how atomic radiuses vary through the periodic table. Throughout the periodic table, the dimensions of atoms show the following trends, with many irregularities: Size increases down the group. The widest electrons are consistently extensive orbital as n increases. The size decreases over a certain period. Electrons are added to a single shell and do not protect each other very effectively from the growing nuclear charge. This causes all orbiters (including the most unpleasant orbital orbiters) to be stacked as \(Z\) increases. Transitional elements contract to the middle of the series and then expand to the end. Protection becomes more efficient after d under the shell halfway This causes the electrons that determine the size of the ns to move on towards the end of each row of transition. The Enthanides are slowly contracting throughout the series. This is the result of ineffective protection of 4f electrons below the size that determines electrons 6s. Since the filling of the enthanide occurs between the second and third transitional rows, the second and third transition elements in each period are almost the same size. For example, both Roux and Wasps have a 130 p.m. radius. Using only a periodic table, they predict the procedure for increasing the atomic radius for each of the following sets of elements: Ca, Mg, Be Ga, Br, Ge Al, Tl, Si Zn, Cd, Fe Experimentally defined B.C. communication distance in bismuth trichlorid is 2.48 Å. Given the tabulous value of 0.99 Å for the atomic radius of Cl, predict the atomic radius of Bi. If the experimentally defined distance of BI-I in the bismuth triyody is 2.81 O, it is almost impossible to predict the atomic radius of the I. Ionic radius for the ions, since it is almost impossible to experimentally determine where the density of electrons begins and the anion ends in any ionic pair. Tabulous values of ion dens represent the values that most consistently involve measured divisions between typical enanthia pairs in many binary ion compounds. Trends in ion sizes are very regular, almost without exceptions. Ion-radius of ions increases with negative charge and decreases with positive charge. Compare the radius (in picometers, PM) for the following isoelectronic species (isoelectronic means the same number of electrons) with ne configuration (ten electrons): O2–F– Na+ Mg2+ Al3+ 140 133 97 66 51 Note that for this series the size increases with negative charge and decreases with positive charge. In cases of anions, there are more electrons in the nucleus than protons. This means that each electron is kept about a core with less than one equivalent positive charge, so the entire distribution of electrons expands. The reverse occurs with cations. There are fewer electrons than protons, so each electron is kept about a core with more than one equivalent positive charge. Thus, the entire distribution of electrons approaches the nucleus. The same trends can be seen in differently charged ions of the same element. For example, compare radiuses for \(Cu(I)\) and \(Cu(II)\) ions: \(\ce{Cu^{+}}\): 96 pm vs. \(\ce{Cu^{2+}}\): 18:00 Ionian radio will increase the group for single charge ions. In the graph below, look out for trends in sizes (rejoicing in picometers, PM) following ions with the same charge. Li+ 60 Be2+ 31 O2– 140 F– 136 Na+ 95 mg2+ 65 Al3+ 50 S2– 184 Cl– 181 K+ 133 Ca2+ 99 Ga3+ 62 Se2 – 198 BR– 195 Rb+ 148 Sr2+ 113 In3+ 81 Te2– 221 I– 216 Cs+ 16 Ba2+ 35 Tl3+ 9 For each of the following sets of atoms and organize members in order to increase the size. \(\ce{Se^{2-}, Te^{2-}, Se}\) \(\ce{Co^{3+}, Fe^{2+}, Fe^{3+}}\) \(\ce{Ca, Ti^{4+}, Sc^{3+} }\) \(\ce{S^{2-}, Se^{2–}, Ar}\) Ionization energy, I am the energy needed to remove the electron from the gase. The first energy of ionization, \(I_1\), refers to the removal of one electron from a neutral atom: \[X(g) \rightarrow X^+(g) + e^- \] from \(PE = I_1 > 0\) All ionization energies are positive because it takes energy to remove the electron from the atom nucleus attraction. Trends in the first ionization energies can be understood based on the size of the atoms. Valence electrons are closer to the nucleus to which they are attracted to a smaller atom; thus, more energy will be required to remove the electron by ionization. According to size trends, the first energy ionizations tend to increase over a certain period and decrease down the group. The following storyline shows a variation of the first ionization energies throughout the periodic table. Irregularities (runs in lines over a period) are the result of particularly stable electronic configurations either in a neutral atom (which led to slightly higher energy ionization) or +1 ion produced by ionization (which led to slightly lower energy ionization). The second and successful energy of ionization (I2, I3, ...), which remove additional electrons, require more and more energy, because each electron is removed from a smaller view with an increasingly positive charge (cation). For example, compare the first and second calcium ionization ionization: First ionization \(\ce{Ca(g) \rightarrow Ca+(g) + e-}\) \ (I_1 = 589.8\, kJ/mol\) Second ionization \(\ce{Ca+(g) \rightarrow Ca2+(g) + e- }\) \(I_2 = 1145.4\, kJ/mol\) Total (sum) \(\ce{Ca(g) \rightarrow Ca2+(g) + 2e- }\) \(I_1 + I_2 = 1735.2\, kJ/mol\) Based on the position in the periodic table, which element of the following pairs has a higher first energy of ionization? O vs Ne Ca vs Sr K vs Cr Br vs Sb In vs Sn First Ionisation energies tend to increase over a period. But in period 2, Be has a higher energy of the first ionization, than B, and N has a higher energy of the first ionization than O. Explain ,[Hint: Look at valence configurations, and recall that semi-filled and fully filled swaps have additional stability.] The affinity of electrons is determined by the overall reaction \[A(g) + e^– \rightrowar A^–(g)\] of \(PHP^o \equiv A\). In general, an additional electron is added to the lowest energy of the vacant orbiter in the valence sub-point, or in the next sub-point, if all the valence subcalps are already full. The first affinities of electrons can be either exothermic or endothermic, depending on whether the atom wants an additional electron or not. \(A < 0\): The reaction is exothermic and the atom wants electron \(A > 0\): the reaction is endothermic and does not want electron Although the first affinities of electrons can be either positive or negative, all the second and successful affinities of electrons are positive (endothermic). This is the result of repulsions between an additional electron and a negative ion, which was formed in the first process of electron affinity. For example, consider step-by-step formation \(\ce{O^{2-}(g)}\): First attachment \(\ce{O(g) + e^{-} \rightarrow O^{-}(g)}\) \(A_1 = -142 \,kJ/mol\) Second attachment \(\ce{O-(g) + e^{-} \rightarrow O^{2-}(g)}\) \(A_2 = +844\, kJ/mol\) Total (Sum) \(\ce{O(g) + 2e^{-} \rightarrow O^{2-}(g) }\) \(A_{tot} = +702\, kJ/mol\) Electron affinity data is limited, due to experimental difficulties, and values vary greatly depending on the data source. The following table of electron affinity data gives default values (kJ/mol). 1A 2A 3A 4A 5A 6A 7A H -73 Li -60 be +100* B -27 C -122 N +9* O -1 141 F -328 Na -53 Mg +30* Al -44 Si -134 P -72 S -200 Cl -348 K -48 Ca ? Ga -30 Ge -120 How -77 Se -195 Br -325 Rb -47 Sr? In -30 Sn -121 Sb -101 Te -190 I -295 Cs -45 Ba ? Tl -30 Pb -110 Bi -110 Po -183 At -270 * Estimated values In general, gross trends in the first affinity of electrons: Electron affinity tends to become more negative over a period. This is the result of gradually smaller atomic sizes, placing the added electron closer to the nucleus. Electron affinity tends to become less negative (or positive) down the group. The added electron is placed in orbital, which is gradually farther from the nucleus. Nonmetals tend to have a high (negative) affinity of electrons, and metals tend to have low (slightly negative or positive) affinity electrons. This is consistent with the tendency of nonmetalists to form aions, while metals resist the formation of negative ions. Noble gases have positive affinities of electrons. The added electron is placed in an orbital orbit of higher energy over a very stable closed shell configuration. The combination of high protective charge of nuclear weapons and violation of stable configuration softens the formation of onion. Based on electronic configurations or any other relevant considerations, explain the differences in the affinity of electrons of the following pairs of species. Na, A1 < 0; Mg, A1 > 0 N, A1 > 0; O, A1 < 0 Br, A1 << 0; Kr, A1 > 0 O, A1 < 0; O–, A2 > 0 Chemical behavior of elements is a consequence of their electronic structure. Unsurprisingly, the similarities in the behaviour of elements in the same group are the result of their presence of the same kind of valence configuration. Differences are often a matter of degree, reflecting changes in gravity between the core and the valence of electrons (i.e. effective nuclear charge). For example, \(\ce{Li}\) and \(\ce{Na}\) react with water to the formation of aqueous solutions of ionic hydroxides and hydrogen gas: \[\ce{2 Li(s) + 2 H2O(l) \rightarrow LiOH(aq) + H2(g)}\] \[\ce{2 Na(s) + 2 H2O(l) \rightarrow 2 NaOH(aq) + H2(g)}\] The reaction with lithium is very gentle, but with sodium it usually causes burning hydrogen to explode in the air. The difference is clear by comparing smaller size and higher lithium energy ionization with larger sizes and less sodium ionization energy. In short, the energy reserve is less in the case of sodium, and the exothermic output of energy reaction is greater. As we have seen, the elements can be classified as metals, nonmetallics and metaloids. Each category has a characteristic range of ionization energies and electron affinity. Metals – low energy ionization, low affinity of electrons (negative or poorly positive) Nonmetals – high energy ionization, high affinity of electrons (negative) Metaloids - intermediate ionization energy and affinity of electrons Due to relatively low energy ionization, metals tend to form cations, and when combined with nonmetalizing substances. For example, when metals are combined with oxygen, they form ionic oxides. \[\ce{4 Fe(s) + 3 O2 \rightarrow 2 Fe2O3(s)}\] Metal oxides tend to dissolve in water to form a hydroxide ion (\(\ce{OH^{-}})). Therefore, metal oxides are the main ones. \[\ce{Na2O(s) + H2O(l) \rightarrow 2 NaOH(aq)}\] This is the result of hydrolysis of ion \(\ce{O^{2-}}\): \[\ce {O^() 2-}(aq) + H2O(l) \rightarrow 2 OH^{-}(aq)}\] Nonmetalis have ionization energy too high for stable formation of this, but their favorable affinity of electrons, usually will lead to the formation of anony. In combination with metal products or complex products, the resulting compounds are ionic. However, when nonmetals are combined, the compounds are molecular. Unlike metals, nonmetal oxides are acidic. Molecular oxides do not have oxide ions, and oxygen is too tightly tied to hydrolysis in the ion of hydroxide in water. Instead, nonmetal oxides tend to add both oxygen and hydrogen from the water to make molecular acid. \[\ce{CO2(g) + H2O(l)} \rightarrow \underbrace{\ce{H2CO3(aq)}}_{text{carbonic acid}} \] \[\ce{N2O5(s) + H2O(l)} \rightarrow \underbrace{ \ce{2 HNO3(aq). }} _{\text{nitric acid}}\] \[\ce{P4O10(s) + 6 H2O(l)} \rightarrow \underbrace{ \ce{4 H3PO4(aq)}}_{text{phosphoric acid}} \] Metallic properties fall between metals and non-metals. To behave like metals, metals such as \(\ce{Si}\), \(\ce{Ge}\), As you would like to form very highly charged replications: (\(\ce{Si^{4+}}\), \(\ce{Ge^{4+}}\), \(\ce{As^{5+}}\)). Such small and highly charged notchs will have a strongly attractive effect on electrons around any eions with which they can be combined. This will lead to a significant distribution of the density of electrons between them. As we will see, electron density metabolism is usually associated with molecule formation. So of metaloids are generally not ionic, and are generally not as basic as would be the case with real metals. They may not have an acid-meadow in water (essentially inert) or may be acidic (such as molecular oxides). \[\underbrace{ \ce{SiO2(s) + H2O(l) \rightarrow } \, \text{no reaction}}_{text{inert behavior}}\] \[\underbrace{ \ce{As2O5(s) + 3 H2O(l) \rightarrow 2 H3AsO4(aq)}_{text{acid oxide behavior}}\] Calcium is usually less reactive than potassium, but more reactive, For each of the following oxides, specify whether it is ionic or molecular and whether it is acidic or basic. , which is expected between each oxide and water: \(\ce{SO2(g)}\), \(\ce{CaO(s)}\), \(\ce{Li2O(s)}\), \(\ce{SeO3(s)}\) and \(\ce{P4O6(s)}\). \(\ce{P4O6(s)}\).