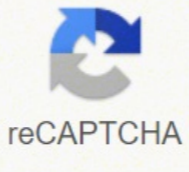




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Periodic Table of the Radioactive Elements

Electron Configurations in Groups

The Representative Elements

This figure shows a portion of the periodic table containing Groups 1A through 7A.

Elements in Groups 1A through 7A are often referred to as **representative elements** because they display a wide range of physical and chemical properties.

Note In moving down a group in the p-block, increasing ionization energies and decreasing bond strengths result from an inert-pair effect. Adam From Encyclopedia of Science and Technology < General Chemistry Descriptive chemistry is the study of the physical and chemical properties of elements and substances. As such, the bonding in such compounds has a significant covalent component, giving the compounds properties that can differ significantly from those expected for simple ionic compounds. The most important unifying principle in describing the chemistry of the elements is that the systematic increase in atomic number and the orderly filling of atomic orbitals lead to periodic trends in atomic properties. The chemistry of the second-period element of each group (n = 2; Li, Be, B, C, N, O, and F) differs in many important respects from that of the heavier members, or congeners, of the group. The semimetals lie along the diagonal line separating the metals from the nonmetals and exhibit intermediate properties. Ionization energies, the magnitude of electron affinities, and electronegativities generally increase from bottom to top. Answer nonmetal; C; semimetals; Si and Ge; metals; Sn and Pb Pb is most stable as M²⁺. Page ID49470 We will describe some of the chemistry of the representative elements, showing as we do how their properties may be rationalized on the basis of concepts and principles. This difference in behavior reflects the fact that within the same group of the periodic table, the relative energies of the n bond and the sigma (σ) bond differ. As shown by the M-Cl bond energies listed in Table (PageIndex{1}), the strength of the bond between a group 13 atom and a chlorine atom decreases by more than 30% from B to Tl. Similar decreases are observed for the atoms of groups 14 and 15. Because metals have relatively low electronegativities, they tend to lose electrons in chemical reactions to elements that have relatively high electronegativities, forming compounds in which they have positive oxidation states. The anomalous chemistry of second-period elements results from three important characteristics: small radii, energetically unavailable d orbitals, and a tendency to form pi (π) bonds with other atoms. For example, N is just above P in group 15: N₂ contains an N=N bond, but each phosphorus atom in tetrahedral P₄ forms three P-P bonds. Although Al and B both lie on the diagonal line, only B is a semimetal; the heavier elements are metals. As you read, you should try to see why a certain reaction occurs, as well as what actually happens, and instead of memorizing each specific equation, you should try to organize the chemistry of these elements according to the generalizations you have already learned. The same factors also lead to periodicity in valence electron configurations, which for each group results in similarities in oxidation states and the formation of compounds with common stoichiometries. Going down a group, the atoms generally became larger, and the overlap between the valence orbitals of the bonded atoms decreases. The diagonal line in Figure (PageIndex{1}) separates the metals (to the left of the line) from the nonmetals (to the right of the line). The energy of the 3d orbitals far exceeds the energy of the 2s and 2p orbitals, so using them in bonding is energetically prohibitive. One of the most dramatic differences between the lightest main group elements and their heavier congeners is the tendency of the second-period elements to form species that contain multiple bonds, predict which element differs the most from the others in its chemistry. Any thoughts from the collected masses on this topic? I don't really have a question (or a point, perhaps) but I am curious how others think about this topic. The most fundamental property leading to periodic variations is the effective nuclear charge (Z_{eff}). Arguments based on d-orbital availability and on the small size of the central atom, however, predict that coordination numbers greater than 4 are unusual for the elements of the second period, which is in agreement with experimental results. In group 13, we predict that the chemistry of boron will be quite different from that of its heavier congeners. C is most different. Increasing ionization energies and decreasing bond strengths lead to the inert-pair effect, which causes the heaviest elements of groups 13-17 to have a stable oxidation state that is lower by 2 than the maximum predicted for their respective groups. Consequently, the elements in the upper right of the periodic table are the smallest and most electronegative; the elements in the bottom left are the largest and least electronegative. As we begin our summary of periodic trends, recall that the single most important unifying principle in understanding the chemistry of the elements is the systematic increase in atomic number, accompanied by the orderly filling of atomic orbitals by electrons, which leads to periodicity in such properties as atomic and ionic size, ionization energy, electronegativity, and electron affinity. There appear to be two major reasons for the inert-pair effect: increasing ionization energies and decreasing bond strengths. In contrast, atomic size decreases from left to right and from bottom to top. I think we can all agree that the descriptive chemistry of the elements (I am thinking mostly of the d block, but some aspects of the p block as well) is: important hard or boring to teach This may be a generational thing, and maybe I'm young(er), and the pendulum is way on one side, but when I read books like Cotton & Wilkinson or Greenwood & Earnshaw, where they say that facts are more important than theories, I tend to disagree. Some of my coworkers, however, firmly believe that facts trump theories. Given the limited time we all have to teach inorganic chemistry, one has to make choices about content to include or discard. I have tended to spend a lot (my students might say a LOT) of time on MO theory, while paying only the barest lip service to descriptive chemistry. I have, once or twice, taught units on halogens and noble gas chemistry (since those are at least things the students have not seen, especially noble gas chemistry), but these days, I have added units on the "more important" fields of organometallics or solid state or bioinorganic. I put a little descriptive chemistry in here and there, where it fits (VSEPR is a good place to at least mention a little, for example, why you can form IF₇ but only ClF₅ or BrF₅) and I put quite a bit in my acid/base unit (that I teach in Analytical). Recall that these trends are based on periodic variations in a single fundamental property, the effective nuclear charge (Z_{eff}), which increases from left to right and from top to bottom in the periodic table. The inert-pair effect should be most important for the heaviest element (Tl), so it is most likely to form compounds in an oxidation state that is lower by 2. The very small cations derived from second-period elements have a high charge-to-radius ratio and can therefore polarize the filled valence shell of an anion. In contrast, an Si=Si bond, with less p-orbital overlap between the valence orbitals of the bonded atoms because of the larger atomic size, is only about 40% stronger than an Si-Si bond. Explanation of these properties is organized according to the concept of periodicity, with each subsequent section corresponding to one of the eight groups of representative elements. Exercise (PageIndex{1}) Based on the positions of the group 14 elements C, Si, Ge, Sn, and Pb in the periodic table and the general trends outlined in this section, classify these elements as metals, semimetals, or nonmetals. As you study the elements, you will find the periodic table very helpful. It is therefore substantially more difficult than expected to remove these ns²electrons, as shown in Table (PageIndex{1}) by the difference between the first ionization energies of thallium and aluminum. Solution Group 13 spans the diagonal line separating the metals from the nonmetals. Because Tl is less likely than Al to lose its two ns² electrons, its most common oxidation state is +1 rather than +3. Within the s and p blocks, similarities between elements in different groups are most marked between the lightest member of one group and the element of the next group immediately below and to the right of it. Contributors Was this article helpful? Conversely, nonmetals have high electronegativities, and they therefore tend to gain electrons in chemical reactions to form compounds in which they have negative oxidation states. Thus BF₃ forms only the four-coordinate, tetrahedral BF₄⁻ ion, whereas under the same conditions AlF₃ forms the six-coordinate, octahedral AlF₆³⁻ ion. Consequently, electron configurations with more than four electron pairs around a central, second-period element are simply not observed. Because d orbitals are never occupied for principal quantum numbers less than 3, the valence electrons of second-period elements occupy 2s and 2p orbitals only. These elements exhibit similar electronegativities and charge-to-radius ratios. The inert-pair effect refers to the empirical observation that the heavier elements of groups 13-17 often have oxidation states that are lower by 2 than the maximum predicted for their group. Then use the

trends discussed in this section to compare their relative stabilities and chemical reactivities. Among the main group elements, the lightest member of each group exhibits unique chemistry because of its small size resulting in a high concentration of charge, energetically unavailable d orbitals, and a tendency to form multiple bonds. The second-period elements ($n = 2$) in each group exhibit unique chemistry compared with their heavier congeners because of their smaller radii, energetically unavailable d orbitals, and greater ability to form π bonds with other atoms. It is not surprising that they tend to exhibit properties and reactivities intermediate between those of metals and nonmetals. In fact, theoretical descriptions of the bonding in molecules such as SF₆ have been published without mentioning the participation of d orbitals on sulfur. A C=C bond, for example, is approximately 80% stronger than a C-C bond. C and P are most similar in chemistry. Due to their small radii, second-period elements have electron affinities that are less negative than would be predicted from general periodic trends. Figure [\PageIndex{1}](#)): Summary of Periodic Trends in Atomic Properties. Because of the position of the diagonal line separating metals and nonmetals in the periodic table, the chemistry of groups 13, 14, and 15 is relatively complex. Dean, Lange's Handbook of Chemistry, 15th ed. Consequently, bond strengths tend to decrease down a column. predict which element of group 14 will be chemically most similar to a group 15 element. Because of the smaller atomic size, simple binary ionic compounds of second-period elements also have more covalent character than the corresponding compounds formed from their heavier congeners. For example, although an oxidation state of +3 is common for group 13 elements, the heaviest element in group 13, thallium (Tl), is more likely to form compounds in which it has a +1 oxidation state. Consequently, compounds that contain both multiple and single C to C bonds are common for carbon, but compounds that contain only sigma Si-Si bonds are more energetically favorable for silicon and the other third-period elements. Because Al is the second member of group 13, we predict that its chemistry will be most similar to that of Be, the lightest member of group 2. Thus the +1 oxidation state is predicted to be most important for thallium. predict which element of another group will exhibit chemistry most similar to that of Al. Given: positions of elements in the periodic table Asked for: classification, oxidation-state stability, and chemical reactivity Strategy: From the position of the diagonal line in the periodic table separating metals and nonmetals, classify the group 13 elements. Because the elements of groups 13, 14, and 15 span the diagonal line separating metals and nonmetals, their chemistry is more complex than predicted based solely on their valence electron configurations. Key Takeaway The chemistry of the third-period element in a group is most representative of the chemistry of the group because the chemistry of the second-period elements is dominated by their small radii, energetically unavailable d orbitals, and tendency to form π bonds with other atoms. The ionization energies increase because filled ($n - 1$)d or ($n - 2$)f subshells are relatively poor at shielding electrons in ns orbitals. Consequently, the elements of the third period ($n = 3$: Na, Mg, Al, Si, P, S, and Cl) are generally more representative of the group to which they belong. When an electron is added to such a small atom, increased electron-electron repulsions tend to destabilize the anion. All five elements in group 13 have an ns²np¹ valence electron configuration, so they are expected to form ions with a +3 charge from the loss of all valence electrons. predict which element forms the most stable compounds in the +2 oxidation state. predict which element forms the most stable compounds in the +1 oxidation state. (New York: McGraw-Hill, 1999). Note In contrast to the chemistry of the second-period elements, the chemistry of the third-period elements is more representative of the chemistry of the respective group. Thus the two electrons in the ns subshell experience an unusually high effective nuclear charge, so they are strongly attracted to the nucleus, reducing their participation in bonding. As an example, LiCl, which is partially covalent in character, is much more soluble than NaCl in solvents with a relatively low dielectric constant, such as ethanol ($\epsilon = 25.3$ versus 80.1 for H₂O). The net effect of these two factors—increasing ionization energies and decreasing bond strengths—is that in going down a group in the p-block, the additional energy released by forming two additional bonds eventually is not great enough to compensate for the additional energy required to remove the two ns² electrons. You may recall that the role of d orbitals in bonding in main group compounds with coordination numbers of 5 or higher remains somewhat controversial. Moreover, the small sizes of these elements prevent them from forming compounds in which they have more than four nearest neighbors. Example [\PageIndex{1}](#)) Based on the positions of the group 13 elements in the periodic table and the general trends outlined in this section, classify these elements as metals, semimetals, or nonmetals. Representative Elements Group 1 (1A): Alkali Metals Group 2 (2A): Alkaline Earth Metal Group 13 (3A): Boron Family Group 14 (4A): Carbon Family Group 15 (5A): Nitrogen Family Group 16 (6A): Oxygen Family Group 17 (7A): Halogens Group 18 (8A): Noble Gases Hydrogen Transition Metals Group 11 (1B): Coinage Metals Group 12 (2B): Zinc Family Group 8/9/10 (8B): Platinum Family Inner Transition Metals Lanthanoids Actinoids (including Nuclear Chemistry) Synthetic Elements Learning Objectives To know important periodic trends in several atomic properties. The semimetals lie along the diagonal line dividing metals and nonmetals. The most important periodic trends in atomic properties are summarized in Figure [\PageIndex{1}](#)). Table [\PageIndex{1}](#)): Ionization Energies (I) and Average M-Cl Bond Energies for the Group 13 Elements Element Electron Configuration I (kJ/mol) I1 + I2 + I3 (kJ/mol) Average M-Cl Bond Energy (kJ/mol) B [He] 2s²2p¹ 801 6828 536 Al [Ne] 3s²3p¹ 578 5139 494 Ga [Ar] 3d¹⁰4s²4p¹ 579 5521 481 In [Kr] 4d¹⁰5s²p¹ 558 5083 439 Tl [Xe] 4f¹⁴5d¹⁰6s²p¹ 589 5439 373 Source of data: John A. At this point it is useful to look and consolidate some of the general trends observed for the representative elements.

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