Chapter 11 - the periodic table

Science, Physics

Chapter 11 - The Periodic Table 11-1 Organizing by Properties Dmitri Mendeleev (1834-1907) russian, published his element classification in 1869 - increasing molar mass - matching similar properties Mendeleev's periodic table included gaps - yet to be discovered - predicted the properties of missing elements repetition in properties of the elements was a fundamental pattern in nature - periodicity of the elements modern periodic law: the properties of the elements recur periodically when the elemetns are arranged in increasing order by their atomic numbers 11-2 The Periodic Table Today (Table 11-1 pg 307) helium, neon and argon do not form hydrides or fluorides - can be grouped the ratio to H and ratio to F are in groups with sequence 1, 2, 3, 4, 3, 2, 1 elements with same ratios mean they react similarly to H and F periodic table: vertical groups/chemical families groups 1, 2, 12, 14, 15, 16, 17 and 18 known as the representative elements elements in groups all react similarly to the same situations group 1: alkali metals (excluding H) - these metals react with water to form an alkali/basic solution group 2: alkaline earth metals group 17: halogens group 18: noble gases - nonreactive under most conditions group 3-12: transition metals - once believed that they behaved in a manner that is intermediate between the active metals and non metals horizontal rows: periods rare earth metals replaced with term inner transition metals or lanthanides series (57 to 71) actinide series (part of the inner trans. metals): elements 89 to 103 Patterns in Electrons Structure 11-3 The Periodic Table and Electron Configuration Li: 1s2 2s1 / K: 1s2 2s2 2p6 3s2 3p6 4s1 / Ar: 1s2 2s2 2p6 3s2 3p6 / Kr 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 (electron configurations) electrons in the outermost s and p orbitals are referred to as valence electrons each group is characterized by a similar outermost energy lvl configuration, that's why they react similarly transition metals in period 4: 3d orbital being filled one at a time transition metals in period 5: 4d orbital being filled one at a time 11-4 Electron Configuration and Chemical Behaviour Na+ ion and Ne atom are isoelectronic b/c they have the same electron configuration alkali metals: ion charge of 1+ alkaline metals: ion charge of 2+, isoelectronic with same element witb alkali metals halogens: ion charge of 1-, isoelectronic with the nearest noble gas metals lose electrons and form +ions / nonmetals gain (or share) and form - ions Periodic Trends 11-5 Atomic and Ionic Radii atomic radius: determine the distance between nuclei of metal atoms in a crystal / for elements in pure form as molecules, distance between nuclei for two atoms bonded together x ray diffraction technique covalent radius: half of measurement described above ionic radius: measure of the size of the electron probability volume for an ion atomic radii decrease ® in a period reason : as nuclear charge (protons in nucleus) increases, attraction is greater and atomic radius is smaller atomic/ionic radii increase ¯ in a group or family reason 1: # of electrons increase, therefore more orbitals further away from nucleus reason 2: shielding effect - inner electrons shield the valence electrons from the attractive force from nucleus - weakens the force between nucleus and valence electrons the positive ion formed when an atom loses its valence electron is smaller than the neutral atom if neutral atom gains, electron, bigger. b/c increase neg charge results in greater mutual repulsion among the electrons - also reduces the attractive force of the nucleus 11-6 Ionization Energy ionization energy: energy needed to remove an electron from a neutral gaseous atom Element (g) + ionization energy ® Ion+ (g) + e- determined in the 1920's by bombarding gaseous element samples with high energy electrons energy of bombarding electrons known precisely when electrons gain enough kinetic energy, atoms start to ionize first ionization energy - energy needed to eject the most weakly held electron to form pos ion low ionization energy: forms pos ion high ionization energy: may form neg ions or no ions at all IE1, IE2, IE3, IE4, IE5, is the energy needed to remove the outermost, next outermost, third, fourth and fifth outermost electrons from the atom, respectively (from valence to inner) IE increase ® in each period / alkali metal low, noble gas high noble gases the smallest = greater nuclear charge = high energy to rid electron IE2 for Li and He differs even though they both have same electron configuration reason: Li has nuclear charge of 3+ / He has nuclear charge of 2+ thus greater IE for Li problems 2. a) x= 1, y= 3 b) x= 3, y= 1 or 3 c) x= 2, y= 4 or 2 d) x= 3, y= 2 3. a) SrS b)GaF3 c) BeTe d)ClI e) AsBr3 7. 3+, 3