Ib chemistry summary- by paul li and silvia riggioni

Science, Physics

IB Chemistry Summary- By Paul Li & Silvia Riggioni TABLE OF CONTENTS

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-2— IB Chemistry Summary- By Paul Li & Silvia Riggioni ATOMIC THEORY The Electromagnetic Spectrum - - - - - Electromagnetic radiation is a form of energy. The smaller the wavelength the higher he frequency => higher energy of the wave Radio waves, microwaves, infrared, visible light, ultraviolet, x-ray, gamma rays ï¿½ ï¿½ Î» Increases Velocity of waves = frequency x wavelength (c = V Ã— Î») Electromagnetic radiation comes in packages called quanta or photons. Atomic Emission Spectra - - - - - White light is made up of all the colours of the spectra. When it passes through a prism a continuous spectrum is obtained. When energy is applied to specific (individual) elements they emit a spectrum which only contains emissions of particular Î» s. A line spectrum is not continuous. Each element has its own characteristic line spectrum. Hydrogen spectrum- it consists of discrete lines that converge towards the high energy end of the spectrum. The lines converge as the shells are getting closer together. Energy levels increase because we get a higher frequency and a smaller wavelength. (E = Î·V) Explanation for the Emission spectra - - - - Electrons can only exist at specific energy levels When energy is supplied to an atom e- are exited from the lowest (ground) state to an exited state. When e- drop from a higher level to a lower level they emit energy (a photon of light). This energy corresponds to a particular Î» and shows up as a line spectrum. Jumps to the n= 1 have the highest V and the smallest Î» . n= x Continuum- e- shells merge- ionisation energy n= 5 Pfund lines n= 4 (N shell) Brackett lines n= 3 (M shell) Transitions causing Paschen lines. Infrared region. n= 2 (L shell) Transitions causing Balmer lines. Visible light region. n= 1 (K shell) Transitions causing Lymen lines. Ultraviolet region. - - The emission spectra can also be used to find the ionization energy. This is done using the Rydberg equation. First Ionization Energy ï¿½ the energy required to remove one electron from each atom of a mole of atoms in the gas state, to form one mole of cations in the gas phase, under s. t. p X (g) â†’ X (+g) + e âˆ’ Subatomic Particles Mass Proton Neutron Electron /a. m. u. 1 1 1/1840 Charge +1 0 -1 Found in atom Nucleus Nucleus Shells Nucleons -3- IB Chemistry Summary- By Paul Li & Silvia Riggioni - - - Almost all the mass of the atom is concentrated in the nucleus which has a very small radius. Much of the atom is empty space Electrons and protons are deflected by and electric field. Neutrons aren’t. Atomic Numbers - - It is the number of protons in the nucleus of an atom. Defines which element the atom belongs to and consequently its position in the Periodic Table. ï¿½ Z is the Atomic Number, X the symbol. ZX Mass Numbers - - It is the sum of the number of protons plus the number of neutrons in an atom or ion. A X ï¿½ A is the Atomic Mass, X the symbol. The relative atomic mass is the measure of the average mass, taking into account the various types of isotopes. E. g. RAM of chlorine is 35. 5, because Cl-35 is three times as abundant as Cl-37. Isotopes - - Two or more atoms of the same element which have the same number of protons (Z) but different number of neutrons. All isotopes react in the same way. However, the different masses will affect the physical properties such as density and the rate of diffusion of both elements and compounds. Calculating RAM by example — Lead (Pb) 204 206 Abundance in % 1. 5 Pb Pb 207 208 23. 6 Pb Pb 22. 6 52. 3 Average of 100 atoms: 1. 5 Ã— 204 23. 6 Ã— 206 22. 6 Ã— 207 52. 3 Ã— 208 = 306 = 4, 861. 6 = 4, 678. 2 = 1, 0878. 4 ========== 20, 724. 2 Ã· 100 --------------------207. 242 ï¿½ RAM of Pb If the abundance is not given as a percentage then divide by the total abundance. Solutions - - - -4— Solute ï¿½ Solvent ï¿½ Solution ï¿½ Substance that is going to be dissolved. The liquid where we are going to dissolve the solute. Solute + Solvent IB Chemistry Summary- By Paul Li & Silvia Riggioni PERIODICITY Elements - - - - It is a pure substance which can't be made simpler by any chemical method and is made up of atoms, all of which have the same atomic number. A pure substance is made up of only 1 type of atomic number. Arranged in the periodic table by increasing atomic number. The majority are metals. Physical Properties - - A covalent radius is half the4 minimum distance between the nuclei of 2 atoms of the same element covalently bonded in a diatomic molecule. A Van der Waal's radius is half the minimum distance between the nuclei of two atoms in the same element, which are NOT chemically bonded. Covalent Radius - - Atomic radius: Ionic radius: Van der Waal's Radius radius of an atom. Half the distance between the nuclei of adjacent atoms. radius of an anion or cation. Atomic and Ionic Radii in the Periodic Table - - - - The Alkali Metals (Group I) o As you go down the group, both the atomic and the cationic radius increase. o Cations have a smaller atomic radius than the parent atom they come from because they have lost an outer shell electron. o The ions get smaller as they have the same electronic structure but a greater nuclear attration. The Halogens (Group VII) o As you go down the group, the anionic ionic radius increases. o Anions have a larger atomic radius than the parent atom they come from, because they gain an electron in the outer shell. The electrostatic repulsion makes the outer shell expand. o There are more e- for the same nuclear charge, so each is held less strongly and thus, can be further away from the nucleus for the same energy. Across a Period o The atomic radius decreases because the electrons are being pulled closer to the nucleus due to the increase in charge. E. g. Na ï¿½ Ar o The ions across the period have the same ionic structure, but an increase in the number of protons increases the electrostatic forces between the protons and electrons. Thus, the radius decreases. o Ions are isoelectronic, meaning that they have a similar electronic structure which resembles noble gas structures. E. g. Nitrogen’s is similar to Neon’s. Down a Group o Down any group the ionic radius increases as there are more electrons in more shells, which are further from the nucleus and with more shielding. E. g. Li+, Na+, K+ o As we go down the groups, the outermost electron is in a higher energy level, which is further from the nucleus, so the radius increases. -5- IB Chemistry Summary- By Paul Li & Silvia Riggioni Electronegativity (Pauling’s) - - - It is the ability of an atom in a covalent bond to attract electrons to itself. It is a relative measure; hence it doesn’t have any units. Electronegativity depends on 3 things: o Real nuclear charge, Z o Number of screening electrons (repulsion electrons) o Atomic Radius (Z âˆ’ Screening Electrons) Electronegativity = Atomic Radius If two elements have similar electronegativity, they bond covalently. If the difference is of 2ish or more, they will form ionic bonds. Electronegativity in the Periodic Table - - - - - - The Alkali Metals (Group I) o These elements have low values for electronegativity o As you go down the group, electronegativity decreases, even thought the number of protons is increasing. The Halogens (Group VII) o These elements have relatively high electronegativity values. o As you go down the group, electronegativity decreases, even thought the number of protons is increasing. Across a Period o Electronegativity increases across the Periodic Table from left to right. o This is due to the fact that the nuclear charge increases and the atomic radius decreases slightly. Down a Group o Electronegativity decreases because although Z increases, the effect is more than compensated by the increase in screening electrons and the diluting effect of atomic radii getting larger. o Another way of saying this is that the outer shell is further away from the nucleus and so the shielding increases. Noble Gases o These elements have no desire to gain electrons; therefore they do not have electronegativity values. o Argon and Xenon can form a bond. So these as the bottom of the group could have electronegativity values, but those at the top can’t. Halogens and Halides o Halides are the ions of the Halogens o The more electronegative the halogen elements are, the more able they are to pull the electron off the halide ions which are lower in the group than themselves. E. g. Cl2 (g) + 2 Br(âˆ’) â†’ 2Cl(âˆ’aq) + Br2

(aq) aq o The chlorine molecule is reduced- it is the oxidising agent. The bromide ions are oxidised, they're the reducing agents. Ionisation Energy - First Ionization Energy ï¿½ the energy required to remove one electron from each atom of a mole of atoms in the gas state, to form one mole of cations in the gas phase, under s. t. p X (g) â†' X (+g) + eâˆ' - âˆ†H is positive. - Across a Period - In every period the noble gas has the highest value. - Going across the Periodic Table, the general trend is a rise in âˆ†H1st . This is because the outer electrons are going into the same electron shell, but the nuclear charge in increasing. This means they are held more tightly and thus, more energy is required to remove them. -6— IB Chemistry Summary- By Paul Li & Silvia Riggioni - Although the general trend is up, elements in the 2nd and 5th groups have a higher than normal value for âˆ†H1st because the ones in Group II have a full s shell while those in Group V have a half full p shell. This means that we see a drop in the 3rd and 6th elements. Down a Group - The ionisation energy decreases as it is easier to pull out the electron because: o The outer electron is further away from the nucleus. This outweighs the increase in Z, thus the electrostatic forces fall (since they are inversely proportional to the square of the distance). o The number of electrons in the inner shells of the atom increases, thus increasing shielding. These two factors make the energy required to pull one electron off smaller. Successive Ionization Energies - - 1. 2. 3. These are the ionisation energies required to remove more than one electron from a single atom. Electrons are removed from the outside outwards. Special rules exist for the d shell. The energy increases as each electron is removed from the same shell, for the ion has become more positive (and smaller, so there are 2 reasons why the electrostatic forces get larger). There is a large increase in the energy as a new shell of electrons is broken. This is because electrons are closer to the nucleus and there are fewer electrons shielding electrons. The step increases on the energy change increases as we get closer to the nucleus. Log Scale of Energy Energy Required to to Remove the Electron Electron Successive Ionisation Energies Number of Electrons Stripped Tips for drawing: - The electronic structure backwards - Each “ flat" set gets slightly steeper each time. - The “ steps" get slightly bigger each time. Melting Points The m. p. t depends on the type of bonding. The bonding depends on the arrangement of the outer electrons. Going across a period shows a behaviour going from metallic to non metallic. Having one, two or three electrons in the outer shell encourages metallic behaviour. When we get to the half filled shell we get a giant covalent structure. The next elements are gasses, being made of simple covalent molecules. The eighth element forms a monoatomic gas. Melting Points Log m. p. t - - - 1 2 3 4 5 6 7 8 Number of Electrons in Outer Shell -7- IB Chemistry Summary- By Paul Li & Silvia Riggioni - - The Alkali Metals (Group I) o As we go down the group the m. p. t decreases. Also, metals are softer. o All the elements have a body centred cubic structure made up of cations and delocalised electrons. As we go down, the radius of the ions increases and so the cations are further away from the delocalised electrons, therefore the attraction is weaker. The Halogens (Group VII) o As we go down the group, the m. p. t increases. o The effect is due to increased Van der Waal’s forces between the molecules- the higher the number of electrons, the higher the possible forces between the temporary dipoles. Chemical Properties - The Alkali Metals (Group I) o All are soft and have low densities. o All have coloured flame tests. o They’re very electropositive: like to make cations easily. o Behave as very strong reducing agents, thus oxidising themselves. o Metal (I) hydroxides are very soluble. E. g. KOH o Metal (I) hydroxides are very strong bases. E. g. 2 H 2O + 2e âˆ’ â†’ H 2 + 2OH âˆ’ o Reactivity increases down the group, due to the easiness with which the electron can be lost from the metal. (Ionization energy decreases). o Reaction with water: 2 Li(s) + 2 H 2O(l) â†’ 2 LiOH (aq) + H 2(g) 2 Li(s) + 2 H (+aq) + 2OH (âˆ’aq) â†’ 2 Li(+aq) + 2OH (âˆ’aq) + H 2 (g) Removing the Spectator Ions: 2 Li(s) + 2 H (+aq) â†’ 2 Li(+aq) + H 2 (g) o - Reaction with Halogens: ï¿½ All Group I metals react with halogens to form alkali metal halides. ï¿½ 2 Na + Cl2 â†’ 2 NaCl ï¿½ Alkali metal + Halogen ï¿½ Alkali metal Halides The Halogens (Group VII) o They are the salt makers. o As we go down the group, atomic, ionic and covalent radii increase. Ionisation energies for this group are the second highest in the periodic table. o As we go down the group, the reactivity of the halogen decreases, but the halide’s increases. ï¿½halogen molecule becomes a stronger oxidising agent. o As we go up the group ï¿½halide ion becomes a weaker reducing agent. o The further up the group the halogen, the better the oxidising agent. The better down, the better the reducing agent. Use of Standard Electrode Potentials - Electrode potentials are a measure of how strong oxidising or reducing an agent is. They predict whether a redox reaction is energetically favourable. The hydrogen half-cell was used as the base for all the S. E. P. s EÎ¸ Product ï¿½ Reactant +1. 09 Br2 + 2e âˆ’ â‡” 2 Br(âˆ’) â†� â†� aq This is the anticlockwise rule â†’ â†’ +2. 97 F2 + 2e âˆ’ â‡” 2 F(âˆ’) aq Reactant ï¿½ Product â†“ - - -8— Place the most negative value at the top Place the most positive value at the bottom. â†‘ IB Chemistry Summary- By Paul Li & Silvia Riggioni Test for Halide Ions 1. 2. Add nitric acid. This will prevent other pseudo halides from reacting with the silver ions in the silver nitrate, forming a white p. p. t. E. g. 2 Ag + + CO32 âˆ’ â†’ Ag 2CO3 , which is a white p. p. t. If there is no noticeable change, add silver nitrate. If chlorine ions are present, and instant whit p. p. t. of silver chloride will be produced. Ag (+aq) + Cl(âˆ’aq) â†�â†’ AgCl(s) photons 3. 4. In the presence of light, the white p. p. t will decompose and turn grey. 2 AgCl â†’ 2 Ag + Cl2 The silver chloride p. p. t. dissolves easily in ammonium hydroxide (ammonia). In the case of Bromide Ions, a cream p. p. t. is produced which also decomposes with light and turns grey. However, it doesn’t dissolve in ammonia. In the case of Iodide ions, a yellow p. p. t is produced which does NOT decompose with light. Trends across the Third Period Na Mg Bonding Metallic Metallic Al Si P S Cl Ar Metallic Giant Covalent Simple or Giant Simple Covalent Simple Covalent Mono-Atomic Structure Body Centred Cubic Hexagonal Closepacking Face-centred Cubic Diamond Type Trigonal Piramidal 8-atoms in a ring Cl= Cl Mono-Atomic Conductivity Good Good Melting Point High High Reaction w/ Water Basic Basic Good Semi-conductor Insulator Insulator Insulator Insulator Higher Highest Low Low Low Lowest Amphoteric Amphoteric Acid Acid Acid - BONDING Ionic Bonding - When are two elements likely to combine to form a Binary Compound which is ionic in nature? M + x â†’ M + xâˆ’ M + Y2 â†’ M 2 + + 2Y âˆ’ 2 M + 3Z â†’ M 3+ Z 2 âˆ’ M X, Y or Z - - - - - - Likely to be on the left side of the P. T Likely to be metal Likely to have a low electronegativity Small First Ionization Energy (small 2nd , 3rd also) Forms cations isoelectronic to a noble gas Good Reducing Agent - - - - - - - An ionic bond is formed by the transfer of electrons to form 2 oppositely charged particles (ions) which are held together by electrostatic forces. Likely to be on the left side of the P. T Likely to be a non-metal Likely to have a high electronegativity Large 1st Ionization Energy. Forms anions isoelectronic to a noble gas