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Stream 11-1

Chemistry Notes on Periodicity

3.1.1: Describe the arrangement of elements in the Periodic Table in Order of Increasing Atomic Number

Elements in the periodic table are arranged so that when they a read across the periods, the atomic number increases by one each time. This means that any element X that is directly to the right of element Y will have one more proton in its nucleus than X. This principle allows us to infer the properties of certain unknown elements.

3.1.2: Distinguish Between a Group and a Period

A *period* is a horizontal row on the periodic table:

 - Elements are arranged in 7 periods, where period 1 contains the elements He and H, and successive periods are corollary.

- The period number is equal to the number of energy levels the electrons of a given atom of any element in that period take up.

- Elements in the same period have different chemical properties because they have different numbers of valence electrons.

A *group* is a column on the periodic table:

- It is generally said that elements are ordered in 8 groups, where the area between groups 2 and 3, starting with period 4, houses transition elements. (Note Group 8 is sometimes referred to as group 0)

- The group number defines the number of valence electrons of each element listed in that group. For example, all elements in group 2 have two valence electrons (For this reason the chemical properties of elements in the same group are similar)

- Filing down a group, the number of electron orbitals inhabited by the atoms of successive elements increases by one. In other words, any element X, that is directly below element W, and in the same group will have one more occupied orbital that W.

3.1.3: Apply The Relationship between the Electron Arrangement of Elements and Their Position in the Periodic Table up to Z = 20

|  |  |  |  |
| --- | --- | --- | --- |
| Atomic Number (Z) | Period  | Group  | Electron Arrangement |
| 1 | 1 | 1 | 1 |
| 2 | 1 | 0 | 2 |
| 3 | 2 | 1 | 2.1 |
| 4 | 2 | 2 | 2.2 |
| 5 | 2 | 3 | 2.3 |
| 6 | 2 | 4 | 2.4 |
| 7 | 2 | 5 | 2.5 |
| 8 | 2 | 6 | 2.6 |
| 9 | 2 | 7 | 2.7 |
| 10 | 2 | 8 | 2.8 |
| 11 | 3 | 1 | 2.8.1 |
| 12 | 3 | 2 | 2.8.2 |
| 13 | 3 | 3 | 2.8.3 |
| 14 | 3 | 4 | 2.8.4 |
| 15 | 3 | 5 | 2.8.5 |
| 16 | 3 | 6 | 2.8.6 |
| 17 | 3 | 7 | 2.8.7 |
| 18 | 3 | 8 | 2.8.8 |
| 19 | 4 | 1 | 2.8.8.1 |
| 20 | 4 | 2 | 2.8.8.2 |

3.1.4: Apply the Relationship between the Number of Electrons in the Highest Occupied Energy Level for an Element and its Position in the Periodic Table

 The number of electrons in the highest occupied energy level is equivalent to the group number of that element, and the number of that energy orbital (i.e. Phosphorus’s outer energy level is 3) is equal to the period number of that particular element.

Part of 3.2.2: Nuclear Charge and Effective Charge, Trends in Atomic Radii and Ionic Radii

The **nuclear charge** of an atom is equal to the number of protons (hence the number of relative 1+ charges).

The **effective charge** of an atom is the portion of the nuclear charge that attracts the valence electrons of any given element, which are shielded by the magnetic repulsion of the inner electrons from the nuclear charge.

Trend for Atomic Radius

Since the radius of an atom cannot be rigidly defined because the atom is not a hard sphere, it is measured as half the distance between adjacent nuclei of two atoms of the same element.

 - Atomic radii increase down the group

 Atomic radii increase down a group because the number of electron orbitals increases and thus outer electrons are less affected by the nuclear charge and experience more electrostatic repulsion that pushes them apart and increases the radii of the atom.

 - Atomic radii decrease across the period

 Atomic radii decrease across the period because there are no new electron orbitals added and the nuclear charge increases relatively by 1+ with every added electron (as described by the Atomic number), thus, electrons are more influenced by the attracting positive charge of the nucleus and the atom becomes more compact.

Trends for Ionic Radius

Positive ions are always smaller than their parent atoms because when the atom of an element loses an electron, electron repulsion decreases and all atoms are more attracted by the nuclear charge. In some cases, the formation also causes the loss of an electron orbital.

Negative ions are always larger than their parent atoms because as they gain electrons, electron repulsion of the outer energy level increases and thus electrons push away from each other and the ionic radius increases.

**Ionic radii decrease across the periods groups 1 to 4 for positive ions**. The ions Na+, Mg 2+, Al 3+, Si 4+, all have the same electron arrangement, but their increasing nuclear charge draws electrons towards the nucleus and thus the ionic radius decreases.

**Ionic radii decrease across the period from groups 4 to 7 for negative ions**. Ions in the same group all have the same electron arrangement, but again, the increase in nuclear charge draws electrons closer to the center.

Positive ions are smaller than the negative ions because the former have two electron energy levels, and the latter have three.

The ionic radii increase down the group because the number of energy levels increases and when electrons are gained or lost, the radius of a previous ion is smaller than the one directly below it in the same group.

3.2.1, 3.2.2, 3.2.3: Define the Terms *First Ionization Energy* and *Electronegativity* + Trends (Including melting points for halogens and alkali metals, and trend in melting points across period three)

The first ionization energy of an atom is the measure of the attraction between the nucleus of a given atom and its outer electrons.

Strictly defined, the first ionization energy of an element is **the energy required to remove one mole of electrons from one mole of gaseous atoms**.

 - Ionization energies increase across a period as the increase in nuclear charge causes an increase in the attraction between out electrons and the nucleus, making electrons more difficult to remove.

 - Ionization energies decrease down a group because, as the electrons are removed from the furthest available energy level, the effective nuclear charge is about the same, due to the increasing distance between the outer electrons and the nucleus.

**Electronegativity** is the ability of an atom to attract electrons in a covalent bond.

 - Electronegativity increases from left to right across a period due to the increasing nuclear charge, resulting in an increased attraction between the nucleus and bonded electrons.

 - Electronegativity decrease down the group due to the increased distance, and consequently lower attraction between the outer energy level electrons and the nucleus.

3.2.4: The most electronegative element is obviously at the top right of the periodic table, while the least electronegative element is at the bottom left of the periodic table. (4.0 – Flourine, 0.7 – Cesium)

Note: Ionization Energies are properties of Gaseous atoms, while electronegativity is a property of an atom in a molecule whose values are derived empirically.

**Melting points decrease down group one (alkali metals).** These elements have metallic structures which are held together by attractive forces between delocalized electrons and positively charged ions. Because the distance between the positive ions and out delocalized electrons increase with the increasing number of orbitals, the melting points decrease.

**Melting points increase down group seven (halogens).** These elements have molecular structures held together by van der Waals (intermolecular) forces. The magnitude of the forces increases with the increasing number of electrons in the molecule.

**Melting points rise across the period 3 and reach a maximum at group 4, and a minimum at group 0.** At the left of the period are metallic elements and going across the period, the strength of the metallic bonding increases as there is an increase in the nuclear charge (increasing atomic number). The **number of mobile valence electrons increases (one valence electron added on each time)**, and the **atomic radius decreases**. These two factors yield higher melting points because the metal cat-ions are held together more tightly.

Following the metals, at the center of the period, the **giant covalently bonded structures** including C, and Si, where every atom is bonded to all other atoms of the same element around it by a very strong **covalent bond.**

Following the giant covalently bonded structures, come the elements that have **simple non-polar molecular structures** and only **weak van der Waal’s forces** exist between them (N, P). These melting point are low, and depend on the mass and size of the molecules (P4, S8, Cl2)

Noble gases exist as single atoms with even weaker van der Waal’s forces and hence have very low melting points.

3.3.1: Similarities and Differences of Chemical Properties of Elements in the same Group

Elements in the same group have similar chemical properties as they have the same number of valence electrons.

 Group 0: The Noble Gases

- They are colorless gases.

- They are monatomic.

- They are unreactive

Noble gases have the highest ionization energies because they have a stable octet, turning them into negative ions would require adding extra electrons to an empty outer shell, that would experience negligible nuclear force.

Elements in groups 1 to 3 lose electrons to adopt the arrangement nearest noble gas with lower atomic number. For example Boron would lose three electrons to adopt the electron configuration of Neon.

Elements in groups 4 to 7, gain electrons and adopt the arrangement of the next noble gas to the right of them. For example, Bromine, would become 1- charged in order to adopt the configuration of Krypton.

 Group 1: Alkali Metals

Physical Properties include:

 - Good conduction of electricity

 - Low density

 - High malleability

Chemical Properties include:

 - High reactivity

They form single charged ions because they have low ionization energies and their outer shell electrons can be easily lost. Going down the group, their ionization energy decreases while their atomic radius increases and thus they become more reactive. **Reactivity increases down the group**

Reactions of some alkali metals with water:

 Lithium reacts slowly, and keeps its shape.

 Sodium reacts so vigorously that the heat produced melts the unreacted metal which forms a small ball and moves around on the surface of the metal.

 Potassium reacts most vigorously to produce sufficient heat to ignite the hydrogen produced. This produces a lilac colored flame and move excitedly on the waters surface.

 Group 7: Halogens

Group seven elements exist as diatomic molecules.

Physical Properties:

 - The are colored

 - They show gradual change from gases (F2 and Cl2) to liquid (Br2) to solid (I2 and At2)

 **The reactivity decreases down the group** as the atomic radius increases and the attraction for outer electrons decreases.

 Reaction with group 1 metals:

 Halogens react with group one metals to form halides. The electrostatic forces of attraction between oppositely charged ions pull them together. The most vigorous reaction takes place between elements that are furthest apart on the periodic table. Therefore the reaction between Fr, the least electronegative, and thus most reactive metal, and Fluorine, the most electronegative and most reactive halogen, would form the halide: FrF

 It is possible to determine the relative reactivity of two halogens by a displacement reaction in which a solution of one halogen is introduced to a solution of one halide. If the halogen is more reactive then the halogen in the compound, it will “steal” the electron donated by the alkali metal.

 Halogens form insoluble salts with silver. Adding a solution containing a halide to a solution containing silver ions produces a precipitate which can be used to identify the halide.

3.3.2: Discuss the Changes in Nature, from Ionic to Covalent and from basic to acidic, of the oxides across period three

Ionic compounds are generally formed between metal and non-metal elements and thus the oxides of Na to Al have **giant ionic structures**.

Covalent compounds are formed between non-metals so the oxides of phosphorus, sulfur, and chlorine are **molecular covalent**. Silicon, is the only metalloid in group three, that has a **giant covalent structure.**

**The ionic character of the oxides in group three decreases from left to right** as the elements in the group approach the electronegativity of Oxygen which is 3.5. The ionic characters of the oxides increase down the group because the electronegativity decreases.

Acid-Base Properties of Period Three Oxides

 - Metallic elements form ionic oxides are basic.

 - Non-metallic elements form covalent oxides that are acidic.

 - Both **Aluminum and Silicon** form oxides that are amphoteric (meaning they react with both acids and bases, having the properties of both.

**Basic oxides** react with acids to form ***salt and water***.

**Basic oxides** dissolve in water to produce **alkaline solutions.** (these have the OH- ion.)

**Acidic oxides** react with water to produce acidic solutions.

**Amphoteric oxides** behave like bases when they react with acid, and behave like acids when they react with bases.