Summary and Notes for Exam Preparation: Periodicity

Periodicity of Physical Properties

1. **Periodic Patterns of Atomic Radii**
    - Definition: Atomic radius is the distance from the nucleus to the outermost shell of an atom.
    - Trends:
        - Across a period, atomic radius decreases due to increasing nuclear charge and number of protons, leading to stronger force of attraction between the nucleus and electrons.
        - Down a group, atomic radius increases due to increasing number of electron shells and the shielding effect of inner electrons, leading to a weaker force of attraction between the nucleus and outer electrons.
    - Example question: Explain the trend in atomic radii across Period 3 elements from sodium (Na) to argon (Ar).
2. **Periodic Patterns of Ionic Radii**
    - Definition: Ionic radius is the distance from the nucleus to the outermost shell of an ion.
    - Trends:
        - For cations, ionic radius decreases down a group due to the increasing nuclear charge and number of protons, leading to a stronger force of attraction between the nucleus and electrons.
        - For anions, ionic radius increases down a group due to the increasing number of electron shells and the shielding effect of inner electrons, leading to a weaker force of attraction between the nucleus and outer electrons.
    - Example question: Explain the trend in ionic radii of Group 2 elements from magnesium (Mg^{2+}) to barium (Ba^{2+}).
3. **Periodic Patterns of Melting Points and Electrical Conductivity**
    - Melting Point:
        - Trends:
            - Across a period, melting point generally increases due to the increasing metallic character and strength of metallic bonds.
            - Down a group, melting point generally increases due to the increasing strength of metallic bonds as a result of larger atomic/ionic radii.
        - Example: Explain the trend in melting points of Period 3 elements from sodium (Na) to argon (Ar).
    - Electrical Conductivity:
        - Trends:

            - Across a period, electrical conductivity generally decreases due to the increasing covalent character and decreasing number of delocalized electrons.
            - Down a group, electrical conductivity generally increases due to the increasing number of delocalized electrons.
        - Example: Explain the trend in electrical conductivity of Period 3 elements from sodium (Na) to argon (Ar).
4. **Periodic Patterns of First Ionisation Energies**
    - Definition: The first ionisation energy is the energy required to remove the outermost electron from an isolated gaseous atom to form a positive ion.
    - Trends:
        - Across a period, first ionisation energy generally increases due to the increasing nuclear charge and decreasing atomic radius, making it harder to remove an electron.
        - Down a group, first ionisation energy generally decreases due to the increasing atomic radius and shielding effect of inner electron shells, making it easier to remove an electron.
    - Example question: Explain the trend in first ionisation energies of Group 2 elements from beryllium (Be) to radium (Ra).

Example Questions with Detailed Explanations:

1. Explain the trend in atomic radii across Period 3 elements from sodium (Na) to argon (Ar). Explanation:
    - Across Period 3, the atomic radii decrease from sodium (Na) to argon (Ar).
    - This is because as you move across a period, the number of protons in the nucleus increases, leading to a stronger force of attraction between the nucleus and the electrons.
    - The increasing nuclear charge pulls the electrons closer to the nucleus, resulting in a decrease in atomic radius.
    - For example, sodium (Na) has 11 protons, while argon (Ar) has 18 protons. The greater nuclear charge in argon (Ar) pulls the electrons closer to the nucleus, giving it a smaller atomic radius compared to sodium (Na).
2. Explain the trend in ionic radii of Group 2 elements from magnesium (Mg^{2+}) to barium (Ba^{2+}). Explanation:
    - For cations in Group 2, the ionic radii increase down the group from magnesium (Mg^{2+}) to barium (Ba^{2+}).
    - This is because as you go down the group, the number of electron shells increases, leading to a greater shielding effect of the inner electron shells.
    - The increased shielding effect weakens the force of attraction between the nucleus and the outermost electrons, causing the ionic radius to increase.
    - For example, the Mg^{2+} ion has a smaller ionic radius compared to the Ba^{2+} ion due to the greater shielding effect in the larger barium atom.
3. Explain the trend in melting points of Period 3 elements from sodium (Na) to argon (Ar). Explanation:
    - Across Period 3, the melting points generally increase from sodium (Na) to argon (Ar).

    - This is because as you move across the period, the metallic character of the elements increases, leading to stronger metallic bonds.
    - Stronger metallic bonds require more energy to break, resulting in higher melting points.
    - For example, sodium (Na) has a lower melting point of 371 K compared to aluminum (Al) with a melting point of 932 K, due to the stronger metallic bonds in aluminum.
4. Explain the trend in electrical conductivity of Period 3 elements from sodium (Na) to argon (Ar). Explanation:
    - Across Period 3, the electrical conductivity generally decreases from sodium (Na) to argon (Ar).
    - This is because as you move across the period, the covalent character of the elements increases, leading to a decrease in the number of delocalized electrons available for conduction.
    - Delocalized electrons are responsible for the high electrical conductivity of metals, so a decrease in the number of delocalized electrons results in lower electrical conductivity.
    - For example, sodium (Na) is a good conductor with a conductivity of 0.218 S/m, while silicon (Si) and the subsequent elements are poor conductors due to their increased covalent character.
5. Explain the trend in first ionisation energies of Group 2 elements from beryllium (Be) to radium (Ra). Explanation:
    - Down Group 2, the first ionisation energies generally decrease from beryllium (Be) to radium (Ra).
    - This is because as you go down the group, the atomic radius increases due to the addition of more electron shells.
    - The increased atomic radius and shielding effect of the inner electron shells make it easier to remove the outermost electron, resulting in lower first ionisation energies.
    - For example, the first ionisation energy of beryllium (Be) is higher than that of barium (Ba) because the larger atomic radius and greater shielding effect in barium make it easier to remove an electron.

Periodicity of Chemical Properties

1. **Reactions of Sodium and Magnesium with Water**
    - Sodium (Na) reaction:
        - Sodium is a highly reactive metal that reacts violently with water.
        - The reaction produces hydrogen gas and a basic solution of sodium hydroxide (NaOH).
        - Balanced equation: $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
    - Magnesium (Mg) reaction:
        - Magnesium is also a reactive metal, but it reacts less violently with water compared to sodium.
        - The reaction produces hydrogen gas and a basic solution of magnesium hydroxide ($Mg(OH)_2$).
        - Balanced equation: $Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$

- Example question: Explain the differences in the reactions of sodium and magnesium with water.

2. **Oxides of Period 3 Elements**
   - Oxidation Numbers of Oxides:
     - The oxidation number of the central element in an oxide is equal to the charge on the oxide ion.
     - For example, in Na_2O, the oxidation number of Na is +1, and in Al_2O_3, the oxidation number of Al is +3.
   - Summary of the Acidic/Basic Nature of Period 3 Oxides:
     - Sodium oxide (Na_2O) and magnesium oxide (MgO) are basic oxides.
     - Silicon dioxide (SiO_2), phosphorus pentoxide (P_2O_5), and sulfur trioxide (SO_3) are acidic oxides.
     - Chlorine monoxide (Cl_2O) and argon oxide (Ar_2O_3) are not formed due to the high stability of the elements.
   - Example question: Explain the trend in the acidic/basic nature of the Period 3 oxides from sodium oxide (Na_2O) to sulfur trioxide (SO_3).
3. **Chlorides of Period 3 Elements**
   - Oxidation Numbers of the Period 3 Elements in their Chlorides:
     - The oxidation number of the central element in a chloride is equal to the charge on the chloride ion.
     - For example, in NaCl, the oxidation number of Na is +1, and in $AlCl_3$, the oxidation number of Al is +3.
   - Effect of Water on Chlorides of Period 3 Elements:
     - Sodium chloride (NaCl) and magnesium chloride ($MgCl_2$) are ionic compounds that dissolve in water to form neutral solutions.
     - Aluminum chloride ($AlCl_3$) and silicon tetrachloride ($SiCl_4$) react with water to form acidic solutions.
     - Phosphorus trichloride (PCl_3), sulfur dichloride (SCl_2), and chlorine monochloride (ClCl) hydrolyze to form acidic solutions.
   - The Chemical Bonding in Al_2Cl_6:
     - Aluminum chloride ($AlCl_3$) exists as a dimer, Al_2Cl_6, in the gaseous and solid states.
     - The bonding in Al_2Cl_6 is covalent, with the aluminum atoms sharing chlorine atoms to form a bridge-like structure.

Example Questions with Detailed Explanations:

1. Explain the differences in the reactions of sodium and magnesium with water.
   Explanation:
   - Sodium (Na) is a highly reactive metal that reacts violently with water, producing hydrogen gas and a basic solution of sodium hydroxide (NaOH).
   - The reaction is highly exothermic, and the hydrogen gas produced often ignites, leading to a dramatic reaction.
   - In contrast, magnesium (Mg) is also a reactive metal, but it reacts less violently with water compared to sodium.
   - The reaction of magnesium with water produces hydrogen gas and a basic solution of magnesium hydroxide ($Mg(OH)_2$).

    - The reaction of magnesium is less exothermic than the reaction of sodium, and the hydrogen gas produced does not typically ignite.
    - The differences in reactivity can be attributed to the relative ease with which sodium and magnesium lose their valence electrons to form cations.
    - Sodium has a lower ionization energy and is more electropositive, making it more reactive than magnesium.
2. Explain the trend in the acidic/basic nature of the Period 3 oxides from sodium oxide (Na_2O) to sulfur trioxide (SO_3). Explanation:
    - Sodium oxide (Na_2O) and magnesium oxide (MgO) are basic oxides, meaning they react with water to form basic (alkaline) solutions.
    - This is because sodium and magnesium have low ionization energies and readily lose their valence electrons, forming positively charged cations (Na^+ and Mg^{2+}) in the oxides.
    - As you move across Period 3, the elements become more electronegative, and their oxides become more acidic in nature.
    - Silicon dioxide (SiO_2), phosphorus pentoxide (P_2O_5), and sulfur trioxide (SO_3) are acidic oxides, meaning they react with water to form acidic solutions.
    - This is because silicon, phosphorus, and sulfur have higher ionization energies and are more electronegative, forming oxyanions (SiO_4^{4-}, PO_4^{3-}, and SO_4^{2-}) in their oxides.
    - The trend from basic to acidic oxides across Period 3 is due to the increasing electronegativity and covalent character of the elements.
3. Explain the effect of water on the chlorides of the Period 3 elements. Explanation:
    - Sodium chloride ($NaCl$) and magnesium chloride ($MgCl_2$) are ionic compounds that dissolve in water to form neutral solutions.
    - This is because sodium and magnesium are electropositive metals that readily form stable cations (Na^+ and Mg^{2+}) in their chlorides, and the chloride ions (Cl^-) do not hydrolyze.
    - Aluminum chloride ($AlCl_3$) and silicon tetrachloride ($SiCl_4$) react with water to form acidic solutions.
    - This is because aluminum and silicon are more electronegative and form covalent chlorides. When these chlorides react with water, they undergo hydrolysis, producing acidic solutions.
    - Phosphorus trichloride (PCl_3), sulfur dichloride (SCl_2), and chlorine monochloride ($ClCl$) also hydrolyze in water, forming acidic solutions.
    - The acidic nature of these solutions is due to the formation of hydrogen ions (H^+) as a result of the hydrolysis reactions.
4. Describe the chemical bonding in aluminum chloride (Al_2Cl_6). Explanation:
    - Aluminum chloride ($AlCl_3$) exists as a dimer, Al_2Cl_6, in the gaseous and solid states.
    - The bonding in Al_2Cl_6 is covalent, with the aluminum atoms sharing chlorine atoms to form a bridge-like structure.
    - Each aluminum atom is bonded to three chlorine atoms, and the remaining chlorine atoms are shared between the two aluminum atoms, forming the bridge.

- This bridge-like structure is more stable than individual $AlCl_3$ molecules due to the additional covalent bonds and the sharing of electron density between the aluminum atoms.
- The covalent bonding in Al_2Cl_6 is in contrast to the ionic bonding found in other Period 3 chlorides, such as sodium chloride ($NaCl$) and magnesium chloride ($MgCl_2$).

These example questions cover the key concepts related to the periodicity of chemical properties, including the reactions of sodium and magnesium with water, the acidic/basic nature of Period 3 oxides, the oxidation numbers and hydrolysis of Period 3 chlorides, and the bonding in aluminum chloride. The explanations provided highlight the underlying reasons for the observed chemical behavior of these elements and their compounds.