Chapter 4:

Periodic Table and Periodicity of Properties

Short Questions (Exercise)

2(i) Write the valence shell electronic configuration of an element present in the period and Group IIIA.

Period 2, Group IIIA: $2s^2\ 2p^1$ (e.g., Boron, B).
Period 3, Group IIIA: $3s^2\ 3p^1$ (e.g., Aluminium, Al).

2(ii) Define halogens.

Halogens are the elements of Group VIIA (Group 17) in the periodic table, including F, Cl, Br, I, and At. They are highly reactive non-metals and are known for forming salts with alkali metals.

2(iii) Which atom has a higher shielding effect, Li or Na?

Sodium (Na) has a higher shielding effect because it has more inner electron shells compared to Lithium (Li).

2(iv) Explain why Na has a higher ionization energy than K.

Sodium (Na) has a smaller atomic radius than Potassium (K), leading to stronger nuclear attraction for the valence electrons in Na. As a result, it requires more energy to remove an electron from Na than from K.

3. Arrange the elements in each group in order of decreasing ionization energy:

(a) Li, Na, K: Li > Na > K (Ionization energy decreases down the group because atomic size increases, reducing the nuclear attraction on valence electrons).
(b) Cl, Br, I: Cl > Br > I (Ionization energy decreases down the group for the same reason as above).

4. Arrange the elements in each group in order of decreasing shielding effect:

(a) Li, Na, K: K > Na > Li (Shielding effect increases down the group as the number of inner electron shells increases).
(b) Cl, Br, I: I > Br > Cl (Shielding effect increases down the group because of the additional inner shells).

5. Specify which of the following elements would you expect to have the greatest electron affinity: S, P, Cl.

Chlorine (Cl) has the greatest electron affinity because it is a halogen with a high tendency to gain electrons to achieve a stable octet.

6. Group the elements based on their electronic configuration into pairs that represent similar chemical properties:

Electronic configurations given:

- A = $1s^2\ 2s^2$
- B = $1s^2\ 2s^2\ 2p^6\ 3s^1$
- C = $1s^2\ 2s^2\ 2p^2$
- D = $1s^2\ 2s^2\ 2p^6$
- E = $1s^2\ 2s^2\ 2p^3$
- F = $1s^2\ 2s^1$

Grouping:

- A and F: Both belong to Group IA (alkali metals).
- B and D: Both represent elements in the noble gas configuration or adjacent to noble gases.
- C and E: Both belong to Group IVA and VA (p-block elements), respectively.

7. Arrange the elements in groups and periods for Z = 6.

Element: Carbon (C).

Group: IVA (14).

Period: 2.

8. For normal elements, the number of valence electrons of an element is equal to its group number. Find the group number for the following elements:

Li: Group IA (1 valence electron).

Cl: Group VIIA (7 valence electrons).

9. Write the valence shell electronic configuration for the following groups:

a. Alkali metals (Group IA): ns^1 (e.g., Li: $2s^1$, Na: $3s^1$).

b. Alkaline earth metals (Group IIA): ns^2 (e.g., Be: $2s^2$, Mg: $3s^2$).

c. Halogens (Group VIIA): ns^2np^5 (e.g., F: $2s^2\ 2p^5$, Cl: $3s^2\ 3p^5$).

d. Noble gases (Group VIIIA): ns^2np^6 (except He: $1s^2$) (e.g., Ne: $2s^2\ 2p^6$).

10. Write electron dot symbols for the atoms of the following elements:

Be: [Be]

K: [K] •

C: [C] ••

N: [N] ••••

I: [I] ••••••

11. Write the valence shell electronic configuration of the atoms of the following elements:

a. An element present in period 3 of Group VA: $3s^2\ 3p^3$ (e.g., Phosphorus, P).

b. An element present in period 2 of Group VIA: $2s^2\ 2p^4$ (e.g., Oxygen, O).

12. Copy and complete the following table:

Atomic Number	Mass Number	No. of Protons	No. of Neutrons	No. of Electrons
11	23	11	12	11
14	29	14	15	14
22	47	22	25	22
13	27	13	14	13

13. In which block, group, and period in the periodic table would you place each of the following elements with the given electronic configurations?

a. $1s^2\ 2s^2$:
- Block: s-block
- Group: IIA
- Period: 2

b. $1s^2\ 2s^2\ 2p^4$:
- Block: p-block
- Group: VIA
- Period: 2

c. $1s^2\ 2s^2\ 2p^6\ 3s^1$:
- Block: s-block
- Group: IA
- Period: 3

d. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$:
- Block: p-block
- Group: VA
- Period: 3

Answers to Think Tank Questions

14. What types of elements have the highest ionization energies and what types of elements have the lowest ionization energies? Argue.

Highest Ionization Energies:
Nonmetals, especially noble gases, have the highest ionization energies. This is because they have a stable electronic configuration, small atomic radii, and strong nuclear attraction for their valence electrons.

Lowest Ionization Energies:
Alkali metals (Group IA) have the lowest ionization energies due to their large atomic size and weak nuclear attraction for the single valence electron, making it easier to remove.

15(i). Match each ionization energy with one of the given electronic configurations. Give reasons for your choice.

Electronic configurations:
$1s^2\ 2s^2\ 2p^6$ and $1s^2\ 2s^2\ 2p^6\ 3s^1$.

Ionization energies: 2080 kJ/mol and 496 kJ/mol.

- $1s^2\ 2s^2\ 2p^6$: Ionization energy is 2080 kJ/mol.
Reason: This is the configuration of a noble gas (e.g., Neon, Ne), which has a stable octet and a strong nuclear attraction for electrons, requiring significantly high energy to remove an electron.

- $1s^2\ 2s^2\ 2p^6\ 3s^1$: Ionization energy is 496 kJ/mol.
Reason: This is the configuration of an alkali metal (e.g., Sodium, Na), where the outermost $3s^1$ electron is loosely bound, making it easier to remove.

15(ii). Use the second member of each group from Group IA, IIA, and VIIA to judge that the number of valence electrons in an atom of the element is the same as its group number.
- Group IA (e.g., Sodium, Na):
Configuration: $1s^2\ 2s^2\ 2p^6\ 3s^1$.
Sodium has 1 valence electron, matching its group number (1).

- Group IIA (e.g., Magnesium, Mg):
Configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2$.
Magnesium has 2 valence electrons, matching its group number (2).

- Group VIIA (e.g., Chlorine, Cl):
Configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$.
Chlorine has 7 valence electrons, matching its group number (7).

15(iii). Letter A, B, C, D, E, F indicates elements in the figure:
- A: Likely an alkali metal (e.g., Lithium, Li or Sodium, Na).
- B: Likely an alkaline earth metal (e.g., Beryllium, Be or Magnesium, Mg).
- C: Likely a p-block element (e.g., Carbon, C).
- D: Likely a halogen (e.g., Fluorine, F or Chlorine, Cl).
- E: Likely a noble gas (e.g., Neon, Ne or Argon, Ar).
- F: Likely a transition metal.

PROJECT

Step 1: Materials Needed

1. **Base Board**: Use a cardboard sheet, wooden board, or foam board as the base.
2. **Colored Papers or Paint:** To represent different element groups (e.g., alkali metals, transition metals, etc.).
3. **Small Boxes or Blocks**: For each element. You can use:
    - Cardboard cubes
    - Plastic blocks
    - Bottle caps or other small containers.
4. **Labels**: Printed or handwritten labels for element symbols, atomic numbers, and names.
5. **Glue or Adhesive**: To fix the blocks or cubes on the board.
6. **Markers or Pens**: To write or decorate.
7. **Optional**: LED lights to represent special properties like periods or groups, or clay for creative shaping.

Step 2: Plan the Layout

1. **Sketch the Periodic Table**: Draw the grid layout for periods and groups on the base board.
    - Include blocks: s-block, p-block, d-block, and f-block.
    - Make spaces for Lanthanides and Actinides below the main table.
2. **Color-Code Groups**: Assign specific colors to groups:
    - Alkali metals: Red
    - Alkaline earth metals: Orange
    - Transition metals: Yellow
    - Halogens: Green
    - Noble gases: Blue
    - Lanthanides and Actinides: Purple

Step 3: Create Element Blocks

1. **Prepare Blocks**:
    - Use small boxes, cubes, or clay molds for each element.

- Write or paste labels for each block:
  - Element symbol (e.g., H for Hydrogen)
  - Atomic number (e.g., 1 for Hydrogen)
  - Atomic mass (optional for advanced detail).
- Arrange blocks in rows and columns according to their period and group.

2. **Group Arrangement**:
   - Ensure elements are arranged in order of increasing atomic number.
   - Place Lanthanides and Actinides in separate rows below the main table.

Step 4: Assembly

1. Fix the blocks onto the base board with glue or double-sided tape.
2. Align them neatly within the grid you've drawn.
3. Use colors or markers to highlight the periodic trends, like reactivity, metallic character, or electron configuration.

Step 5: Enhance the Model

1. **Add Labels**: Indicate groups, periods, and blocks (s, p, d, f).
2. **Interactive Features**:
   - Add movable blocks to show element reactivity or bond formation.
   - Use LED lights for highlighting trends (e.g., noble gases glowing blue).
3. **Creative Elements**:
   - Use 3D printouts for special features like Hydrogen placement.
   - Add decorations around the edges.

Step 6: Presentation

1. Prepare a short explanation of:
   - Trends in the periodic table (e.g., ionization energy, electronegativity).
   - Group and period properties.
2. Explain the significance of color coding or added features.

Extera Short Questions (Topic Wise)

4.1: Periodic Table

1. **What is the periodic table?**
   The periodic table organizes elements based on increasing atomic number and similar chemical properties. Elements are arranged in periods (rows) and groups (columns).

2. **Who is credited with developing the periodic table?**
   Dmitri Mendeleev is credited with creating the first version of the periodic table by arranging elements by increasing atomic mass and grouping them based on properties.

3. **What are periods in the periodic table?**
   Periods are the horizontal rows in the periodic table, representing the number of electron shells in an atom.

4. **What are groups in the periodic table?**
   Groups are vertical columns where elements share similar chemical properties due to having the same number of valence electrons.

4.2: Group Number and Charge on an Ion

1. **How does the group number relate to valence electrons?**
   The group number for main group elements indicates the number of valence electrons. For example, Group 1 elements have 1 valence electron.

2. **What charge do Group 1 elements form?**
   Group 1 elements form +1 ions by losing one electron.

3. **What charge do Group 17 (halogens) elements form?**
   Halogens form -1 ions by gaining one electron to complete their octet.

4. **Why are noble gases (Group 18) unreactive?**
   Noble gases are stable because they have a full outer electron shell and usually do not form ions.

4.3: Periodicity of Properties

1. **What is periodicity?**
   Periodicity refers to the recurring trends in properties of elements, such as atomic radius and electronegativity, across periods and groups.

2. **Why does atomic radius decrease across a period?**
   Atomic radius decreases across a period because the increasing nuclear charge pulls electrons closer to the nucleus.

3. **Why does ionization energy increase across a period?**
   Ionization energy increases due to a stronger attraction between the nucleus and valence electrons, making it harder to remove an electron.

4. **What trend does electronegativity follow in the periodic table?**
   Electronegativity increases across a period and decreases down a group.

4.4: Characteristic Properties

1. **What is ionization energy?**
   Ionization energy is the energy required to remove an electron from an atom in its gaseous state.

2. **What is electron affinity?**
   Electron affinity is the energy change when an electron is added to a neutral atom, indicating its tendency to form anions.

3. **What is atomic radius?**
   Atomic radius is the distance from the nucleus to the outermost electron, which decreases across a period and increases down a group.

4. **What is metallic character?**
   Metallic character refers to an element's ability to lose electrons, which increases down a group and decreases across a period.

4.5: Transition Elements

1. **What are transition elements?**
   Transition elements are metals found in Groups 3-12 that have partially filled d-orbitals.

2. **Why are transition metals important?**
   Transition metals are important because they form colored compounds, act as catalysts, and exhibit variable oxidation states.

3. **What are examples of transition metals?**
   Iron (Fe), Copper (Cu), and Zinc (Zn) are common transition metals.

4. **Why do transition metals form colored compounds?**
   The presence of partially filled d-orbitals allows electrons to absorb and emit light of specific wavelengths.

4.6: Lanthanides and Actinides

1. **What are lanthanides?**
   Lanthanides are elements from atomic numbers 57-71, known for their high reactivity and use in producing strong magnets.

2. **What are actinides?**
   Actinides are elements from atomic numbers 89-103, many of which are radioactive and used in nuclear applications.

3. **What is a common use of lanthanides?**
   Lanthanides are used in manufacturing strong permanent magnets and in phosphors for lighting.

4. **Why are actinides significant?**
   Actinides like Uranium and Plutonium are used as fuel in nuclear reactors.

4.7: Halogens

1. **What are halogens?**
   Halogens are nonmetals in Group 17, including Fluorine (F), Chlorine (Cl), Bromine (Br), and Iodine (I).
2. **Why are halogens reactive?**
   Halogens are highly reactive because they need only one electron to complete their octet.
3. **What are halogens used for?**
   Halogens are used in disinfectants, bleaching agents, and in producing salts like NaCl.
4. **Which halogen is the most reactive?**
   Fluorine is the most reactive halogen due to its high electronegativity.

4.8: Noble Gases

1. **What are noble gases?**
   Noble gases are inert elements in Group 18, including Helium (He), Neon (Ne), and Argon (Ar).
2. **Why are noble gases chemically inert?**
   Noble gases have a complete outer electron shell, making them stable and unreactive.
3. **What are noble gases used for?**
   Noble gases are used in lighting (neon lights), welding (argon), and as cryogenic coolants (helium).
4. **Why is helium preferred in balloons?**
   Helium is non-flammable and lighter than air, making it safe for use in balloons.

4.9: Comparison of General Physical Properties of Metals and Non-Metals

1. **How do metals and non-metals differ in conductivity?**
   Metals are good conductors of heat and electricity, while non-metals are poor conductors.
2. **What is the difference in malleability and ductility?**
   Metals are malleable and ductile, while non-metals are brittle and break under stress.
3. **What is the difference in appearance?**
   Metals are shiny (lustrous), while non-metals are dull in appearance.
4. **How do metals and non-metals differ in chemical reactivity?**
   Metals lose electrons to form cations, while non-metals gain electrons to form anions.

Extera Long Questions (Topic Wise)

1. How is the periodic table organized, and why is it important for understanding the properties of elements?

Organization of the Periodic Table

- **Rows and Columns:**
  The periodic table is organized into **periods** (horizontal rows) and **groups** (vertical columns). Periods represent the number of electron shells in an atom, while groups indicate elements with similar chemical properties due to having the same number of valence electrons.
  Example: Group 1 elements (alkali metals) all have one valence electron and react vigorously with water.

- **Increasing Atomic Number:**
  Elements are arranged in order of increasing atomic number, ensuring that properties repeat periodically.

- **Blocks:**
  The table is divided into blocks based on electron configuration:
  - **s-block:** Groups 1-2 (e.g., sodium and magnesium)
  - **p-block:** Groups 13-18 (e.g., oxygen and chlorine)
  - **d-block:** Transition metals (e.g., iron and copper)
  - **f-block:** Lanthanides and actinides

Significance of the Periodic Table

1. **Predicting Chemical Behavior:**
   The periodic table helps predict how elements react based on their position.
   Example: Elements in Group 17 (halogens) are highly reactive non-metals that readily form -1 ions.

2. **Understanding Periodic Trends:**
   Properties such as atomic radius, ionization energy, and electronegativity follow predictable trends across periods and groups.
   - Atomic radius decreases across a period and increases down a group.
   - Ionization energy increases across a period and decreases down a group.

3. **Identifying Families:**
   Groups like alkali metals, halogens, and noble gases have distinct properties that are useful in industrial and scientific applications.
   Example: Noble gases are inert and used in lighting and welding.

4. **Applications in Chemistry and Industry:**
   The periodic table is a foundational tool for understanding chemical reactions, material

properties, and designing new compounds.
Example: Transition metals like platinum are used as catalysts in industrial processes.

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2. What are the key trends in periodicity, and how do they influence the behavior of elements?

Key Periodic Trends

1. **Atomic Radius:**
    - **Trend:** Decreases across a period due to increasing nuclear charge, pulling electrons closer to the nucleus. Increases down a group due to added electron shells.
    - **Example:** Sodium has a larger atomic radius than chlorine in the same period, while cesium is larger than sodium in the same group.
2. **Ionization Energy:**
    - **Trend:** Increases across a period because of stronger nuclear attraction, making it harder to remove an electron. Decreases down a group due to increased distance from the nucleus.
    - **Example:** Helium has the highest ionization energy, while francium has one of the lowest.
3. **Electronegativity:**
    - **Trend:** Increases across a period as atoms more strongly attract electrons in a bond. Decreases down a group as the atom's size increases.
    - **Example:** Fluorine is the most electronegative element, while cesium is one of the least.
4. **Reactivity:**
    - Metals: Reactivity increases down a group as electrons are lost more easily.
    - Non-metals: Reactivity decreases down a group as electrons are harder to gain.
    - **Example:** Potassium is more reactive than lithium among alkali metals, while fluorine is more reactive than iodine among halogens.

Influence on Element Behavior

- **Chemical Bonding:**
  Trends like electronegativity and ionization energy determine how elements form bonds.
  Example: Chlorine forms ionic bonds with sodium (NaCl) and covalent bonds in Cl_2 molecules.
- **Industrial Applications:**
  Reactivity trends guide the use of elements in manufacturing and reactions.
  Example: Reactive alkali metals are used in batteries, while noble gases are used in inert environments.

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3. What are the characteristics and significance of transition elements, lanthanides, and actinides?

Transition Elements

1. **Characteristics:**
   - Found in Groups 3-12 (d-block) with partially filled d-orbitals.
   - Exhibit high melting points, malleability, and electrical conductivity.
   - Show variable oxidation states and form colored compounds.
     Example: Iron (Fe) forms Fe^{2+} and Fe^{3+} ions, giving compounds like rust and ferric chloride.
2. **Applications:**
   - Used as catalysts in industrial processes.
   - **Example:** Platinum and palladium are used in catalytic converters to reduce emissions.
   - Essential in the production of alloys like stainless steel.

Lanthanides

1. **Characteristics:**
   - Elements 57-71, part of the f-block.
   - Known as rare earth elements with high reactivity and magnetic properties.
   - Often form +3 ions.
2. **Applications:**
   - Used in strong permanent magnets, phosphors in screens, and rechargeable batteries.
     Example: Neodymium is used in making powerful magnets for wind turbines and electric vehicles.

Actinides

1. **Characteristics:**
   - Elements 89-103, part of the f-block.
   - Most are radioactive and exhibit multiple oxidation states.
   - Include naturally occurring elements like uranium and thorium.
2. **Applications:**
   - Used as fuel in nuclear reactors.
     Example: Uranium-235 is used for energy production in nuclear power plants.
   - Plutonium-239 is used in nuclear weapons.

Significance of Transition Elements, Lanthanides, and Actinides

- Transition elements are crucial for industrial applications and biological systems (e.g., hemoglobin with iron).
- Lanthanides and actinides play a key role in advanced technologies like renewable energy and medical imaging.

4. How do periodic trends such as atomic radius, ionization energy, and electronegativity vary across periods and groups in the periodic table?

Atomic Radius

- **Definition:** The atomic radius is the distance from the nucleus to the outermost electron shell.
- **Trend Across a Period:** Atomic radius decreases across a period due to an increase in nuclear charge, which pulls electrons closer to the nucleus.
  Example: Lithium (Li) has a larger atomic radius than fluorine (F) in Period 2.
- **Trend Down a Group:** Atomic radius increases down a group because additional electron shells are added, increasing the distance between the nucleus and the outer electrons.
  Example: Sodium (Na) has a smaller atomic radius than potassium (K) in Group 1.

Ionization Energy

- **Definition:** The energy required to remove an electron from an atom in the gaseous state.
- **Trend Across a Period:** Ionization energy increases because of greater nuclear charge, which holds the electrons more tightly.
  Example: Fluorine (F) has a higher ionization energy than lithium (Li).
- **Trend Down a Group:** Ionization energy decreases because electrons are further from the nucleus and experience less nuclear attraction.
  Example: Cesium (Cs) has a lower ionization energy than sodium (Na).

Electronegativity

- **Definition:** The ability of an atom to attract shared electrons in a bond.
- **Trend Across a Period:** Electronegativity increases due to increasing nuclear charge and smaller atomic radius.
  Example: Oxygen (O) is more electronegative than carbon (C).
- **Trend Down a Group:** Electronegativity decreases as atoms become larger and the nucleus is less able to attract bonding electrons.
  Example: Fluorine (F) is more electronegative than iodine (I).

Significance of Trends

- These trends help predict chemical reactivity and bonding behavior. For instance, highly electronegative elements like fluorine are strong oxidizing agents, while low ionization energy metals like sodium are highly reactive.

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5. What are the properties and industrial applications of transition metals?

Properties of Transition Metals

1. **Variable Oxidation States:**
   Transition metals can lose different numbers of electrons from their d-orbitals, forming ions with multiple oxidation states.
   Example: Iron forms Fe^{2+} and Fe^{3+}.

2. **Formation of Colored Compounds:**
   Transition metals form compounds that are often brightly colored due to d-d electronic transitions.
   Example: Copper sulfate ($CuSO_4$) is blue, while potassium permanganate ($KMnO_4$) is purple.

3. **Catalytic Properties:**
   Transition metals and their compounds act as catalysts in industrial and biological reactions.
   Example: Platinum is used in catalytic converters to reduce emissions.

4. **High Melting and Boiling Points:**
   These metals have strong metallic bonds, contributing to their high melting and boiling points.

Industrial Applications

1. **Construction and Manufacturing:**
   - Iron is used to make steel for construction.
   - Titanium is used in aerospace engineering for its strength and lightness.

2. **Chemical Catalysts:**
   - Nickel is used in hydrogenation of vegetable oils.
   - Vanadium pentoxide (V_2O_5) is used in the production of sulfuric acid.

3. **Electronics and Electrical Equipment:**
   - Copper is used for electrical wiring due to its high conductivity.
   - Gold is used in microchips for its excellent corrosion resistance.

4. **Medical and Biological Use:**
   - Cobalt is used in vitamin B12, which is essential for red blood cell production.
   - Transition metals like platinum are used in chemotherapy drugs.

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6. What are the differences in physical and chemical properties between metals and non-metals, and how are they used in industry?

Physical Properties

1. **Metals:**
   - Good conductors of heat and electricity.

- High malleability and ductility.
- Shiny appearance (lustrous).
- High melting and boiling points.
  Example: Copper is used in wiring because of its excellent conductivity.

2. **Non-Metals:**
   - Poor conductors of heat and electricity (insulators).
   - Brittle and not malleable.
   - Dull in appearance.
   - Lower melting and boiling points.
     Example: Sulfur is used in rubber production for its insulating properties.

Chemical Properties

1. **Metals:**
   - Lose electrons to form cations during reactions.
   - React with oxygen to form basic oxides.
     Example: Magnesium reacts with oxygen to form magnesium oxide (MgO).
2. **Non-Metals:**
   - Gain electrons to form anions during reactions.
   - React with oxygen to form acidic oxides.
     Example: Sulfur reacts with oxygen to form sulfur dioxide (SO_2).

Industrial Applications

1. **Metals:**
   - Used in construction (steel, aluminum).
   - Electrical and thermal applications (copper, gold).
   - Catalysts in chemical reactions (platinum, iron).
2. **Non-Metals:**
   - Used as insulators (rubber, sulfur).
   - Components in fertilizers (phosphorus, nitrogen).
   - Essential in making plastics (carbon, hydrogen).

Significance

The contrasting properties of metals and non-metals allow for their complementary use in industries. For example, metals like aluminum are used for structural strength, while non-metals like nitrogen are vital for agriculture.

MCQS

1. **Who is credited with developing the first periodic table?**
   a) J.J. Thomson
   b) John Dalton
   c) Dmitri Mendeleev ☑
   d) Ernest Rutherford

2. **What does a group in the periodic table represent?**
   a) Number of electron shells
   b) Number of valence electrons ☑
   c) Number of neutrons
   d) Atomic number

3. **Which group contains the alkali metals?**
   a) Group 2
   b) Group 18
   c) Group 7
   d) Group 1 ☑

4. **What is the group number for halogens?**
   a) 1
   b) 2
   c) 16
   d) 17 ☑

5. **Which element is a noble gas?**
   a) Oxygen
   b) Chlorine
   c) Neon ☑
   d) Sodium

6. **Why do noble gases rarely react?**
   a) They lack neutrons
   b) They are very small
   c) They have a full outer electron shell ☑
   d) They lack protons

7. **What charge does an ion formed by Group 1 elements have?**
   a) +1 ☑
   b) -1
   c) +2
   d) -2

8. **Which element is in Group 2 and forms a +2 ion?**
   a) Calcium (Ca) ☑
   b) Sodium (Na)

c) Chlorine (Cl)
d) Fluorine (F)

9. **What is periodicity in the periodic table?**
   a) Elements arranged by mass
   b) Recurring trends in properties ☑
   c) Random arrangement of elements
   d) Similar properties in periods

10. **Which property increases across a period?**
    a) Atomic radius
    b) Metallic character
    c) Electronegativity ☑
    d) Ion size

11. **What happens to atomic radius across a period?**
    a) Increases
    b) Decreases ☑
    c) Remains constant
    d) First increases, then decreases

12. **What is ionization energy?**
    a) Energy released during a reaction
    b) Energy absorbed by the nucleus
    c) Energy required to remove an electron ☑
    d) Energy lost when forming a bond

13. **Which element is a transition metal?**
    a) Sodium
    b) Magnesium
    c) Iron ☑
    d) Neon

14. **Why do transition metals form colored compounds?**
    a) Their electrons are stable
    b) Their d-orbitals are partially filled ☑
    c) They have no neutrons
    d) They form gases

15. **What is a common property of transition elements?**
    a) Variable oxidation states ☑
    b) Always stable
    c) Low melting points
    d) No metallic properties

16. **Which element is a lanthanide?**
    a) Uranium
    b) Neodymium ☑
    c) Plutonium
    d) Thorium

17. **What are lanthanides known for?**
    a) Radioactivity
    b) Magnetic properties ☑
    c) Bright colors
    d) Heavy mass

18. **Which element is an actinide?**
    a) Titanium
    b) Beryllium
    c) Uranium ☑
    d) Zinc

19. **What is a common use of actinides?**
    a) Nuclear energy ☑
    b) Jewelry making
    c) Conductors in electronics
    d) Making plastics

20. **Which group is the most reactive nonmetals?**
    a) Group 1
    b) Group 2
    c) Group 18
    d) Group 17 ☑

21. **Which halogen is the most reactive?**
    a) Bromine
    b) Chlorine
    c) Fluorine ☑
    d) Iodine

22. **What are noble gases used for?**
    a) Cooking
    b) Lighting and welding ☑
    c) Fertilizers
    d) Batteries

23. **Which noble gas is used in balloons?**
    a) Helium ☑
    b) Neon
    c) Argon
    d) Krypton

24. **Which property is higher in metals than in non-metals?**
    a) Brittle nature
    b) Insulating ability
    c) Electrical conductivity ☑
    d) Electronegativity

25. **Which property is characteristic of non-metals?**
    a) Malleability

b) Shiny appearance

c) High electronegativity ☑

d) Good electrical conductivity

26. **What type of bond do metals typically form?**

a) Metallic bonds ☑

b) Covalent bonds

c) Ionic bonds

d) Hydrogen bonds

27. **What is the most malleable metal?**

a) Iron

b) Platinum

c) Copper

d) Gold ☑

28. **What is the least reactive noble gas?**

a) Neon

b) Helium ☑

c) Argon

d) Krypton

29. **What property increases down a group in halogens?**

a) Electronegativity

b) Ionization energy

c) Atomic radius ☑

d) Reactivity

30. **What distinguishes metals from non-metals?**

a) Metals gain electrons

b) Metals lose electrons to form cations ☑

c) Metals have low density

d) Metals have high electronegativity