Unit 2: Atomic Structure and Bonding

## Chapter 5: The Periodic Table and Chemical Trends

### Lesson 3: Predicting Properties Based on Periodic Trends

### 1. Big Idea:

- **Big Idea**: The periodic table is organized in such a way that the position of an element allows us to predict its properties, such as reactivity, atomic radius, and electronegativity.

### 2. Essential Questions

- **How can we predict the reactivity and properties of elements based on their positions on the periodic table?**

- Elements in the same group (vertical columns) tend to have similar properties because they have the same number of valence electrons. As you move across a period (horizontal rows), properties like atomic radius, ionization energy, and electronegativity change in predictable ways. These trends help us predict how an element will react with others, how stable it will be, and how its atoms behave.

### 3.1 Phenomenon-Based Learning

- **Unit Phenomenon**:

In northern countries, during winter, streets and roads are often covered in ice and snow. Road salt is spread to melt the ice and snow, making the streets safer. The salts used are typically sodium chloride, magnesium chloride, or calcium chloride. But why are these salts different? How can we predict how well each will work based on the elements they contain?

- **Chapter Phenomenon**:

Sodium chloride, magnesium chloride, and calcium chloride are salts used to melt ice. They all contain chloride, but the other element in each salt is different. Why do these salts behave differently? Can we predict these differences based on the location of sodium, magnesium, and calcium on the periodic table?

### 3.2 Lesson Phenomenon

- **Lesson Phenomenon**:

Imagine you have three different salts: sodium chloride (NaCl), magnesium chloride (MgCl₂), and calcium chloride (CaCl₂). You know that these salts can help melt ice on roads, but would they all work in the same way? Based on the position of sodium, magnesium, and calcium on the periodic table, can you predict which one might be the most effective?

### 4. Vocabulary

- **Atomic trends**: Patterns or tendencies in the properties of elements that can be observed as you move across a period or down a group in the periodic table.

- **Atomic radius**: The size of an atom, typically measured from the center of the nucleus to the outer boundary of the electron cloud.

- **Electron affinity**: The amount of energy released when an atom gains an electron.

- **Electronegativity**: A measure of how strongly an atom attracts electrons in a chemical bond.

- **Ionization energy**: The amount of energy required to remove an electron from an atom.

### 5. SMART Objectives

- **Recognize periodic trends**, such as ionization energy, electronegativity, and atomic radius.

- **Identify the factors** that cause these periodic trends.

- **Predict the properties** of an element, including reactivity and stability, based on its location in the periodic table.

### 6. Engage (Ignite)

**Phenomenon-Related Question**:

Have you ever wondered why different salts are used to melt ice on roads? Sodium chloride, magnesium chloride, and calcium chloride all contain chloride, but they don’t behave the same way. Can we figure out why by looking at the periodic table?

**Hands-On Experiment: Investigating Salt Solutions and Ice Melting**

In this experiment, different salts (NaCl, MgCl₂, and CaCl₂) will be tested to see how effectively they melt ice.

**Materials**:

- Small ice cubes

- Sodium chloride (NaCl)

- Magnesium chloride (MgCl₂)

- Calcium chloride (CaCl₂)

- Three small bowls

- A timer

- Thermometer

**Procedure**:

1. Place one ice cube in each of the three bowls.

2. Sprinkle a small amount of sodium chloride on the first ice cube.

3. Sprinkle a small amount of magnesium chloride on the second ice cube.

4. Sprinkle a small amount of calcium chloride on the third ice cube.

5. Start the timer and observe which ice cube melts the fastest.

6. Record the temperature of the melted water in each bowl after 5 minutes.

**Follow-Up Questions**:

1. Which salt melted the ice the fastest? Why do you think this is?

- **Answer**: Calcium chloride likely melts the ice the fastest because calcium has a higher charge (+2) compared to sodium (+1). This helps calcium chloride break down ice more quickly.

2. How does the position of sodium, magnesium, and calcium on the periodic table help explain the difference in their behavior?

- **Answer**: Sodium, magnesium, and calcium are in different groups on the periodic table. As you move down a group, atomic radius increases, and the ability to form strong ionic bonds changes, influencing how the salts interact with the ice.

3. Based on this experiment, how can you predict the effectiveness of these salts in real-world conditions?

- **Answer**: By knowing the trends of ionization energy and atomic radius, you can predict that calcium chloride, which is lower on the periodic table, will be more effective at melting ice than sodium chloride.

### 7. Pre-Explore (Direct Instruction)

**Background Information**:

The periodic table is organized so that elements with similar properties are grouped together. As you move across periods (rows) or down groups (columns), you can observe certain trends in how elements behave. These trends help us understand and predict properties like atomic radius, electronegativity, ionization energy, and reactivity.

- **Atomic Radius**:

Atomic radius decreases as you move across a period from left to right and increases as you move down a group. This happens because, across a period, protons are added to the nucleus, pulling the electrons closer. Down a group, more electron shells are added, making the atom larger.

- **Ionization Energy**:

Ionization energy is the energy needed to remove an electron from an atom. It increases as you move across a period because atoms become smaller and hold onto their electrons more tightly. It decreases as you move down a group because larger atoms have more electron shielding, making it easier to remove an electron.

- **Electronegativity**:

Electronegativity is the ability of an atom to attract electrons in a bond. Electronegativity increases across a period as atoms get smaller and want to gain electrons to complete their outer shell. It decreases down a group because larger atoms are less able to attract electrons.

**Discussion Questions**:

1. Why does the atomic radius decrease as you move across a period?

- **Answer**: It decreases because the nucleus gains more protons, which pull the electrons closer to the center.

2. Why do you think ionization energy increases across a period?

- **Answer**: Electrons are held more tightly as the nucleus gets stronger, making it harder to remove them.

### 8. Evaluate (Progress Check) - Pre-Explore

**Scaffolded Questions**:

1. How does the atomic radius change as you move down a group?

- **Answer**: The atomic radius increases as you move down a group because more electron shells are added, making the atom larger.

2. What happens to ionization energy as you go from left to right across a period?

- **Answer**: Ionization energy increases because the nucleus is stronger, and the electrons are closer, making them harder to remove.

3. Based on what you know, why might calcium chloride be more effective at melting ice than sodium chloride?

- **Answer**: Calcium has a higher charge (+2) compared to sodium (+1), which allows calcium chloride to more effectively break down the ice.

### 9. Explain (Lightbulb)

**Core Concept**:

The periodic table is more than just a list of elements. It is carefully organized to show trends, or patterns, in the properties of the elements. By understanding these trends, we can predict how elements will behave in chemical reactions and how their properties will change as we move across or down the table.

### Periodic Trends:

1. **Atomic Radius**:

Atomic radius is the size of an atom. As you move across a period (from left to right), the atomic radius decreases. This is because more protons are added to the nucleus, pulling the electrons closer. As you move down a group (from top to bottom), the atomic radius increases because more electron shells are added.

**Example**:

Sodium (Na) has a larger atomic radius than chlorine (Cl) because sodium is on the left side of the periodic table and chlorine is on the right.

**Sample Problem**:

Arrange the following elements in order of increasing atomic radius:

K, Li, Na

**Answer**: Li < Na < K (because atomic radius increases as you move down a group).

2. **Ionization Energy**:

Ionization energy is the energy required to remove an electron from an atom. As you move across a period, ionization energy increases because the electrons are held more tightly by the nucleus. As you move down a group, ionization energy decreases because the outer electrons are farther from the nucleus and shielded by inner electrons.

**Example**:

It is easier to remove an electron from sodium (Na) than from fluorine (F) because sodium has a lower ionization energy.

**Sample Problem**:

Which element has a higher ionization energy: Mg or Al?

**Answer**: Mg has a higher ionization energy than Al because it is farther to the left on the periodic table.

3. **Electronegativity**:

Electronegativity is a measure of how strongly an atom attracts electrons in a bond. It increases across a period because atoms become smaller and want more electrons to fill their outer shell. Electronegativity decreases down a group because the atoms are larger and less able to attract electrons.

**Example**:

Fluorine (F) has the highest electronegativity of any element because it is small and highly reactive.

**Sample Problem**:

Which element is more electronegative: N or O?

**Answer**: Oxygen (O) is more electronegative than nitrogen (N) because it is farther to the right on the periodic table.

### Predicting Properties Based on Periodic Trends

By understanding these trends, we can predict the properties of elements based on their position in the periodic table. For example:

- **Reactivity**: Elements on the left side (like sodium) are more reactive because they want to lose electrons easily, while elements on the right side (like fluorine) are reactive because they want to gain electrons.

- **Stability**: Elements with full outer electron shells, like the noble gases (Group 18), are very stable and unreactive.

### Expand (Connect to Phenomenon)

Let’s revisit the experiment where we tested different salts to melt ice. Sodium chloride, magnesium chloride, and calcium chloride all behave differently because of the periodic trends of sodium (Na), magnesium (Mg), and calcium (Ca). Calcium has a higher charge than sodium, so calcium chloride works better at breaking apart the ice. This is an example of how understanding periodic trends, like ionization energy and atomic radius, helps us predict how elements will behave in real-life situations.

### 10. Evaluate (Progress Check) - Explain Section

**Question 1:**

What is an atom, and what are its basic parts?

**Answer:**

An atom is the smallest unit of matter. It contains three main parts: protons, which have a positive charge; neutrons, which have no charge; and electrons, which have a negative charge. Protons and neutrons are found in the nucleus, while electrons move around the nucleus in energy levels.

**Question 2:**

How does the periodic table organize elements, and what information can you get from it?

**Answer:**

The periodic table organizes elements by increasing atomic number (number of protons). It also groups elements with similar chemical properties into columns called "groups" or "families." From the periodic table, you can learn an element's atomic number, atomic mass, symbol, and general properties like whether it is a metal, nonmetal, or metalloid.

**Question 3:**

Why do atoms form bonds with other atoms, and what are the two main types of chemical bonds?

**Answer:**

Atoms form bonds to achieve a full outer energy level, which makes them more stable. The two main types of chemical bonds are ionic bonds, where atoms transfer electrons, and covalent bonds, where atoms share electrons.

### 11. Elaborate (Power Up) - Deeper Exploration

**Mini-task 1:**

Create a model that shows the difference between an ionic bond and a covalent bond.

**Answer:**

A model for an ionic bond could feature one atom losing an electron and another atom gaining it (e.g., sodium and chlorine forming NaCl). For a covalent bond, the model would show two atoms sharing electrons (e.g., two hydrogen atoms sharing electrons to form H₂).

**Mini-task 2:**

Consider water (H₂O). How do the bonds between the hydrogen and oxygen atoms affect the properties of water?

**Answer:**

Water molecules have polar covalent bonds, meaning the electrons are shared unequally between oxygen and hydrogen. Oxygen pulls the electrons closer, giving it a slight negative charge, while the hydrogen atoms have a slight positive charge. This polarity leads to hydrogen bonding between water molecules, making water cohesive, giving it a high boiling point, and allowing it to dissolve many substances.

**Mini-task 3:**

Can you think of another real-world example where the type of chemical bond affects the material's properties? Explain.

**Answer:**

In diamonds, carbon atoms bond covalently in a very strong, three-dimensional network. This covalent structure makes diamonds incredibly hard, giving them unique properties that make them useful in cutting tools and jewelry.

### 12. Final Evaluation

**Debate Question:**

Should we continue using fossil fuels as our main energy source, knowing their impact on the environment?

Arguments for continuing use:

- Fossil fuels are currently the most reliable and established energy source.

- Infrastructure for fossil fuels is already in place, making them cheaper in the short term.

Arguments for stopping:

- Burning fossil fuels leads to pollution and climate change, which threatens ecosystems and human health.

- Renewable energy sources like wind and solar are becoming more efficient and can reduce environmental harm.

**Assessment Questions:**

**Multiple Choice Questions:**

1. **What determines the identity of an atom?**

a) The number of electrons

b) The number of protons

c) The number of neutrons

d) The number of energy levels

**Correct Answer:** b) The number of protons

Explanation: The number of protons, also known as the atomic number, defines the type of atom.

2. **Which of the following is an example of an ionic bond?**

a) H₂

b) NaCl

c) CO₂

d) O₂

**Correct Answer:** b) NaCl

Explanation: NaCl (sodium chloride) forms when sodium donates an electron to chlorine, creating an ionic bond.

3. **What type of bond involves the sharing of electrons?**

a) Ionic bond

b) Metallic bond

c) Covalent bond

d) Hydrogen bond

**Correct Answer:** c) Covalent bond

Explanation: Covalent bonds involve atoms sharing electrons to achieve stability.

4. **Which property of water is due to its polar covalent bonds?**

a) Water's ability to dissolve many substances

b) Water's solid state being denser than its liquid state

c) Water's lack of surface tension

d) Water's inability to conduct electricity

**Correct Answer:** a) Water's ability to dissolve many substances

Explanation: The polarity of water allows it to interact with and dissolve various substances.

**Long-answer Questions:**

1. **Explain how the structure of the periodic table helps predict the chemical properties of elements.**

**Answer:**

The periodic table is organized in a way that groups elements with similar properties together. Elements in the same group (vertical columns) have the same number of valence electrons, which largely determine an element's chemical behavior. For example, alkali metals in Group 1 have one valence electron, making them highly reactive. As you move across a period (horizontal row), elements gain more protons and electrons, changing their properties like reactivity and electronegativity.

2. **Describe the process of ionic bonding and give an example.**

**Answer:**

Ionic bonding occurs when one atom transfers one or more electrons to another atom. This happens because one atom (usually a metal) wants to lose electrons to achieve a full outer shell, while the other atom (usually a nonmetal) wants to gain electrons. For example, in sodium chloride (NaCl), sodium donates one electron to chlorine, forming a positively charged sodium ion (Na⁺) and a negatively charged chloride ion (Cl⁻). These opposite charges attract, forming an ionic bond.

3. **Why is water considered a polar molecule, and how does this affect its properties?**

**Answer:**

Water is a polar molecule because the oxygen atom pulls the shared electrons closer to itself, creating a partial negative charge on the oxygen and a partial positive charge on the hydrogens. This polarity allows water molecules to form hydrogen bonds with each other, which gives water its unique properties such as high surface tension, the ability to dissolve many substances, and a relatively high boiling point for a small molecule.

4. **Compare and contrast covalent and ionic bonds in terms of electron movement and bond strength.**

**Answer:**

In covalent bonds, atoms share electrons, leading to relatively strong bonds because both atoms achieve a stable electron configuration by sharing. In ionic bonds, one atom transfers electrons to another, creating oppositely charged ions that are held together by electrostatic forces. Ionic bonds are generally strong in solid form but can break apart in water, while covalent bonds tend to remain intact in water. Covalent bonds also allow for molecules with specific shapes, while ionic compounds form crystal structures.

### 13. Extend (Beyond the Lesson)

**Additional Tasks:**

1. **Research Project:**

Pick an element from the periodic table and create a report about its discovery, uses, and how its atomic structure relates to those uses. For example, how does the structure of carbon explain why it can form both diamonds and graphite?

2. **Real-World Challenge:**

Look at the materials used in a smartphone. Identify at least two metals used and explain why their chemical properties make them suitable for use in electronic devices.

**Spaced Practice Suggestions:**

1. **Periodic Table Practice:**

Revisit the periodic table weekly by choosing a new group (such as halogens or noble gases) and writing about their properties, uses, and reactivity.

2. **Bonding Comparisons:**

Create a Venn diagram comparing ionic, covalent, and metallic bonds, focusing on how they form, their properties, and their everyday applications.

By revisiting these concepts over time, students will reinforce their understanding and be able to apply their knowledge in various contexts.