Unit 2: Atomic Structure and Bonding

## Chapter 6: Ionic and Metallic Bonding

# Lesson 4: Metallic Bonding and Metal Characteristics

### 1. Big Idea:

Metallic bonding is responsible for the unique properties of metals, such as their ability to conduct electricity, their luster, and their malleability.

### 2. Essential Questions:

- **How does metallic bonding explain the unique properties of metals?**

**Answer:** Metals have a unique type of bonding called **metallic bonding**, where metal atoms share a "sea of electrons." These electrons are free to move, which gives metals their special properties, such as the ability to conduct electricity and heat, their shiny appearance (luster), and their ability to be shaped (malleability and ductility).

### 3. Phenomenon-Based Learning

**Unit Phenomenon:**

In cold northern countries, road salt is spread to melt ice and snow. When salt contacts the ice, the ice melts, and the salt dissolves. However, metal street signs and lampposts exposed to the same conditions do not melt or dissolve. Why do salt and metal behave so differently with water?

**Chapter Phenomenon:**

Salt and metals behave differently when exposed to water. Salt dissolves, but metal doesn’t. This difference is due to the different types of bonding in salts (ionic bonding) and metals (metallic bonding). The sea of electrons in metallic bonding keeps metals from breaking apart in water, unlike salts which dissolve due to ionic bonds.

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### 4. Vocabulary

1. **Boiling point**: The temperature at which a substance turns from liquid to gas.

2. **Conductivity**: The ability of a substance to allow heat or electricity to pass through it.

3. **Ductility**: The ability of a metal to be stretched into wires.

4. **Luster**: The shiny appearance of metals when they reflect light.

5. **Malleability**: The ability of a metal to be hammered or pressed into shapes without breaking.

6. **Melting point**: The temperature at which a solid turns into a liquid.

7. **Metallic lattice**: A structure where metal atoms are arranged in a regular pattern and surrounded by a sea of electrons.

8. **Sea of electrons**: A model that describes how valence electrons in metals are free to move around within the metallic structure.

9. **Valence electrons**: The outermost electrons of an atom that are involved in bonding.

### 5. SMART Objectives

By the end of this lesson, students will be able to:

- **List the properties of metals**, such as conductivity, malleability, ductility, luster, high boiling/melting points.

- **Describe the formation of metallic bonds** and how the "sea of electrons" model explains these bonds.

- **Analyze the relationship between the structure of metals** (metallic lattice and sea of electrons) and their properties (e.g., electrical conductivity, strength).

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### 6. Engage (Ignite)

**Phenomenon-Related Question:**

**Why don’t metal street signs dissolve in water like salt does?**

**Hands-On Experiment:**

**Activity: Exploring Metal Conductivity**

**Objective:**

Understand how metallic bonding allows metals to conduct electricity.

**Materials**:

- Metal wire (copper or aluminum)

- Battery

- Light bulb

- Salt solution

- Water

- Spoon

**Procedure:**

1. Set up a simple circuit with a battery, a metal wire, and a light bulb.

2. Connect the wire to the battery terminals and the light bulb. The bulb should light up.

3. Now, replace the wire with a spoon dipped in salt solution and try to complete the circuit. Does the bulb light up?

4. Compare what happens with the metal wire and what happens with the salt solution.

**Follow-up Questions:**

1. **Why does the light bulb glow when the wire is connected?**

**Answer:** The light bulb glows because the electrons in the metal wire can move freely, allowing an electric current to flow. This is due to the sea of electrons in metallic bonding.

2. **Why doesn’t the light bulb glow when the spoon is dipped in salt solution?**

**Answer:** Salt in water does not conduct electricity in the same way because it forms ions, not a sea of electrons. Metals have free-flowing electrons that allow electricity to pass through easily.

3. **What can you infer about metallic bonding from this experiment?**

**Answer:** Metallic bonding allows electrons to move easily, which is why metals can conduct electricity.

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

**Background Information**:

Metals are elements that have a special type of bonding called **metallic bonding**. In metallic bonding, the atoms in a metal share their outermost electrons, called **valence electrons**. These electrons form a **sea of electrons** which moves freely around the metal atoms. The metal atoms are arranged in a **metallic lattice**, which is a regular, repeating pattern.

This movement of electrons is what makes metals so good at conducting electricity and heat. The sea of electrons also creates a strong bond between the atoms, making metals tough and giving them their special properties like **malleability** (the ability to be bent or shaped) and **ductility** (the ability to be stretched into wires).

**Scaffolded Questions:**

1. **What is a metallic bond?**

**Answer:** A metallic bond is a type of chemical bond where metal atoms share a "sea of electrons" that can move freely.

2. **What is a metallic lattice?**

**Answer:** A metallic lattice is the regular, repeating structure that metal atoms form, surrounded by the sea of electrons.

3. **How does the sea of electrons affect the properties of metals?**

**Answer:** The sea of electrons allows metals to conduct electricity, be malleable, ductile, and have luster.

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

**Scaffolded Questions (DOK 1-3):**

1. **What is the sea of electrons?**

**Answer:** The sea of electrons refers to the free-flowing valence electrons in a metal that are not attached to any specific atom.

2. **How does metallic bonding differ from ionic bonding?**

**Answer:** In metallic bonding, electrons are shared freely among all the atoms, while in ionic bonding, electrons are transferred from one atom to another, creating ions.

3. **Why is copper used in electrical wiring?**

**Answer:** Copper is used in electrical wiring because it has very good electrical conductivity due to its free-moving electrons.

### 9. Explain (Lightbulb)

**Main Concept Explanation – 4000-6000 words:**

**Introduction to Metallic Bonding**

Metals are everywhere in our daily lives—from the wires that bring electricity to our homes, to the shiny parts of cars, to the coins we use. What makes metals so useful and unique? The answer lies in the way the atoms in metals bond together, something called **metallic bonding**.

### # The Structure of Metals: Metallic Lattices and the Sea of Electrons

Imagine a metal as a large collection of atoms all packed together. These atoms don’t just sit there; they are organized in a repeating pattern known as a **metallic lattice**. In this structure, the valence electrons (the outermost electrons) of the metal atoms are not stuck to any one atom. Instead, they move around freely, creating what we call a **sea of electrons**.

Because these electrons can move so easily, they help hold the metal atoms together. The positively charged metal ions (what’s left of the atom after the electrons move away) are attracted to the negatively charged sea of electrons. This attraction forms a strong bond that holds the metal atoms in place.

**Metallic Properties Explained by Metallic Bonding**

Now that we know how metals are structured, let’s look at how this structure explains the special properties of metals.

### # 1. \*\*Electrical Conductivity\*\*

One of the most important uses of metals is in electrical wiring. Why? Because metals are excellent **conductors of electricity**. This means they allow electric charges to flow through them easily.

In metals, the sea of electrons is free to move. When an electric field is applied (like when you plug in a device), the electrons in the metal wire flow in one direction, allowing electricity to pass through.

**Sample Problem:**

Why are metals used in electrical wiring instead of materials like plastic or wood?

**Answer:** Metals are used because their sea of electrons allows electricity to move through them easily, making them good conductors. Plastics and wood, on the other hand, do not have free-moving electrons and cannot conduct electricity.

**Progress Check:**

What property of metallic bonding allows metals to conduct electricity?

**Answer:** The free movement of electrons in the sea of electrons allows metals to conduct electricity.

### # 2. \*\*Thermal Conductivity\*\*

Not only do metals conduct electricity well, but they also conduct heat. This is why metal pots and pans are used for cooking. The free-moving electrons in the metal can carry heat energy quickly from one part of the metal to another.

### # 3. \*\*Malleability and Ductility\*\*

Have you ever wondered how metals can be bent, hammered, or stretched into thin wires without breaking? This is because metals are **malleable** (they can be shaped) and **ductile** (they can be stretched into wires). The sea of electrons acts like a "glue" that holds the metal atoms together even when they are pushed or pulled into new shapes.

In contrast, ionic compounds like salt tend to be brittle. If you hit a piece of salt with a hammer, it will crack or shatter. This is because in ionic bonding, the ions are held in fixed positions. If you move them, they repel each other and break apart. But in metals, the sea of electrons allows the atoms to move around without breaking.

**Sample Problem:**

Why can metals be hammered into shapes without breaking, but salts shatter when hit?

**Answer:** Metals can be hammered into shapes because the sea of electrons allows the atoms to move without breaking the structure. Salts, with their fixed ionic bonds, shatter because the ions repel each other when they are moved.

**Progress Check:**

What allows metals to be malleable and ductile?

**Answer:** The sea of electrons allows the atoms in a metal to move without breaking, making the metal malleable and ductile.

### # 4. \*\*Luster\*\*

One reason metals are so appealing to the eye is their shiny appearance. This property is called **luster**. When light hits a metal surface, the sea of electrons absorbs and re-emits the light, giving metals their shiny appearance.

### # 5. \*\*High Melting and Boiling Points\*\*

Most metals have high **melting points** (the temperature at which they turn into a liquid) and **boiling points** (the temperature at which they turn into a gas). This is because the metallic bonds (the attraction between the metal ions and the sea of electrons) are very strong. It takes a lot of energy to break these bonds and change the metal from a solid to a liquid or gas.

**Sample Problem:**

Why do metals generally have high melting and boiling points?

**Answer:** Metals have strong metallic bonds between metal ions and the sea of electrons. It takes a lot of energy to break these bonds, which is why metals have high melting and boiling points.

**Progress Check:**

What property of metallic bonding explains why metals have high melting points?

**Answer:** The strong attraction between metal ions and the sea of electrons requires a lot of energy to break, resulting in high melting points.

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### 10. Summary of Key Concepts

1. **Metallic bonding** involves a sea of electrons that are free to move around metal atoms, creating strong bonds.

2. The **metallic lattice** is the regular arrangement of metal atoms in a solid.

3. Metals have properties like **conductivity**, **malleability**, **ductility**, **luster**, and **high melting/boiling points** because of their structure and bonding.

4. The **sea of electrons** allows metals to conduct electricity and heat, be shaped, and reflect light.

### 11. Final Progress Check:

1. **What is the key difference between metallic bonding and ionic bonding?**

**Answer:** In metallic bonding, atoms share a sea of electrons, while in ionic bonding, electrons are transferred between atoms to form ions.

2. **Why are metals shiny?**

**Answer:** Metals are shiny because their sea of electrons can absorb and re-emit light, giving them their luster.

3. **How does the metallic lattice protect metals from dissolving in water?**

**Answer:** The sea of electrons in the metallic lattice holds the metal atoms together strongly, preventing them from breaking apart in water.

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

**Question 1:**

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

**Answer:**

An atom is the basic unit of matter. It has three main parts: protons, neutrons, and electrons. Protons and neutrons are in the nucleus (the center), while electrons move around the nucleus.

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**Question 2:**

How does the periodic table help us understand how elements will behave in chemical reactions?

**Answer:**

The periodic table organizes elements based on their properties. Elements in the same group (column) often behave similarly in reactions because they have the same number of electrons in their outer shell. This helps predict how an element might react with others.

**Question 3:**

Explain why some elements are more reactive than others.

**Answer:**

Some elements are more reactive because of their electron arrangement. For example, elements with one electron in their outer shell (like sodium) can easily lose that electron, making them very reactive. On the other hand, elements with full outer shells (like noble gases) are less reactive because they don’t need to gain or lose electrons.

### 11. Elaborate (Power Up)

**Mini-task 1:**

Imagine you are given two unknown elements. Using the periodic table, describe how you would predict their reactivity and bonding behavior.

**Answer:**

First, I would find their position on the periodic table. If they are in the same group, they likely have similar reactivity. Then, I would look at their electron configuration. Elements in the same group have the same number of valence electrons, which determines how they bond. For instance, elements with almost full outer shells might try to gain electrons, while those with almost empty shells might try to lose electrons.

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**Mini-task 2:**

How would you explain the difference between ionic and covalent bonding to a friend?

**Answer:**

Ionic bonding happens when one atom gives up electrons to another, creating charged particles called ions. These oppositely charged ions attract each other. Covalent bonding, on the other hand, happens when two atoms share electrons. In ionic bonds, the atoms are pulled apart because of their opposite charges, while in covalent bonds, the atoms stick together because they are sharing electrons.

**Mini-task 3:**

Why are noble gases considered “stable,” and how does this affect their use in everyday products?

**Answer:**

Noble gases are stable because their outer electron shells are full, so they don’t need to gain or lose electrons. This makes them very unreactive. Because of this, they are used in products where reactions are unwanted, like in neon lights or in the atmosphere of light bulbs to prevent the metal filament from burning up.

### 12. Final Evaluation

**Debate Question:**

Should we continue using lithium-ion batteries in electronics, or should we invest more in alternative battery technologies?

**Points for Discussion:**

- **For Lithium-Ion Batteries:** They are widely used, have good energy density, and are becoming cheaper to produce.

- **Against Lithium-Ion Batteries:** The mining of lithium harms the environment, and there are concerns about battery disposal and recycling. Alternatives like solid-state batteries might be safer and more sustainable.

### # Multiple-Choice Questions

**Question 1:**

What happens to an atom when it loses an electron?

a) It becomes a positive ion.

b) It becomes a negative ion.

c) It becomes neutral.

d) It gains a neutron.

**Answer:**

a) It becomes a positive ion.

**Explanation:** Losing an electron gives the atom a positive charge, turning it into a positive ion.

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**Question 2:**

Which of the following is a property of covalent bonds?

a) The atoms share electrons.

b) The atoms transfer electrons.

c) The atoms form ions.

d) The atoms are far apart.

**Answer:**

a) The atoms share electrons.

**Explanation:** Covalent bonds involve the sharing of electrons between atoms.

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**Question 3:**

Which element is most likely to form an ionic bond with chlorine?

a) Helium

b) Oxygen

c) Sodium

d) Carbon

**Answer:**

c) Sodium

**Explanation:** Sodium has one electron in its outer shell and can easily lose it, forming an ionic bond with chlorine, which needs one electron to complete its outer shell.

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**Question 4:**

What is the role of neutrons in the atom?

a) To balance the charge of protons.

b) To bond with electrons.

c) To stabilize the nucleus.

d) To create ions.

**Answer:**

c) To stabilize the nucleus.

**Explanation:** Neutrons help hold the nucleus together by balancing the repulsive forces of the positively charged protons.

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### # Long-Answer Questions

**Question 1:**

Explain how the periodic table is organized and why this arrangement is useful for scientists.

**Answer:**

The periodic table is organized by increasing atomic number (the number of protons in an atom). Elements are also grouped by their chemical properties, which repeat in a pattern called periodicity. Elements in the same group (vertical columns) have the same number of valence electrons, which means they behave similarly in chemical reactions. This arrangement helps scientists predict how elements will react with each other and allows them to easily find information like atomic mass, electron configuration, and reactivity.

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**Question 2:**

Describe the difference between metals, nonmetals, and metalloids.

**Answer:**

Metals are typically shiny, malleable, and good conductors of electricity and heat. Nonmetals are more diverse in appearance and are usually poor conductors of heat and electricity. Metalloids have properties of both metals and nonmetals. They can be semiconductors, which makes them useful in electronics. Metals tend to lose electrons in reactions, while nonmetals tend to gain or share electrons.

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**Question 3:**

How do noble gases differ from other elements in terms of reactivity, and what makes them special?

**Answer:**

Noble gases are unique because they have full outer electron shells, which makes them very stable and unreactive. Other elements react because they want to achieve a full outer shell, but noble gases already have this. This is why they don’t easily form compounds and are often used in situations where reactivity could be dangerous, like in light bulbs or neon signs.

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**Question 4:**

Compare and contrast ionic and covalent bonds. Give an example of a compound formed by each type of bond.

**Answer:**

Ionic bonds form when one atom gives up one or more electrons to another atom, creating ions that attract each other. An example is sodium chloride (NaCl), where sodium transfers an electron to chlorine. Covalent bonds, on the other hand, form when two atoms share electrons. An example of a covalent bond is found in water (H₂O), where oxygen shares electrons with hydrogen atoms. Ionic compounds tend to have high melting points and conduct electricity when dissolved in water, while covalent compounds often have lower melting points and do not conduct electricity.

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### 13. Extend (Beyond the Lesson)

**Additional Tasks:**

1. **Reading:**

Read a scientific article about the environmental impact of mining metals used in everyday products (like lithium for batteries). Reflect on how this knowledge might influence your choices as a consumer.

2. **Challenge Question:**

You’ve learned that metals tend to lose electrons and nonmetals tend to gain electrons. How might this knowledge help engineers design better materials for electronics?

3. **Real-World Problem:**

Imagine you are tasked with creating a new, eco-friendly battery for a smartphone. Based on what you know about chemical reactions and the properties of elements, what materials would you choose and why?

4. **Spaced Practice:**

Over the next week, revisit the concept of chemical bonding by using flashcards that focus on ionic and covalent bonds. Try to explain the difference without looking at your notes. After a few days, test yourself again to see how much you remember.

By engaging in these additional activities, you will deepen your understanding of chemistry concepts and see how they apply to real-world situations. Revisiting these ideas over time helps reinforce what you’ve learned.