Unit 2: Atomic Structure and Bonding

Chapter 6: Ionic and Metallic Bonding

# Lesson 4: Metallic Bonding and Metal Characteristics

1. Big Idea:  
- Metals have unique properties, such as malleability and conductivity, that are explained by metallic bonding and the structure of metals.  
  
 2. Essential Questions:  
\*\*1. How does metallic bonding explain the unique properties of metals?\*\*  
  
\*\*Answer:\*\*   
Metallic bonding occurs when valence electrons in metals are shared among a lattice of metal atoms. These electrons are not tied to any one atom and can move freely, creating what is known as a "sea of electrons." This sea allows metals to conduct electricity and heat, and it also gives metals their malleability, ductility, and luster. The structure of metallic bonds explains why metals can be bent or stretched without breaking and why they can carry electrical current efficiently.  
  
\*\*2. Why do metals differ from salts or ionic compounds in their behavior with water?\*\*  
  
\*\*Answer:\*\*   
Metals and salts react differently with water because of their distinct types of bonding. Salts are made of ions held together by ionic bonds, which can dissociate in water, allowing them to dissolve. Metals, on the other hand, are held together by metallic bonds, where a sea of electrons surrounds metal cations. This structure prevents metals from dissolving in water like salts. Instead, metals remain solid, resisting the interaction with water.  
  
 3. Phenomenon-Based Learning:  
\*\*Unit Phenomenon:\*\*   
In cold northern countries, road salt is spread to melt ice and snow on streets. As the 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 in the presence of water. Salt dissolves, but metal doesn’t. This happens because of the different types of bonding in salts (ionic) and metals (metallic). Metals stay intact due to their metallic bonds and sea of electrons, which prevent them from breaking apart in water.  
  
 4. Vocabulary:  
- \*\*Boiling Point:\*\* The temperature at which a substance changes from a liquid to a gas.  
- \*\*Conductivity:\*\* The ability of a material to allow the flow of electricity or heat.  
- \*\*Ductility:\*\* The ability of a material to be drawn into thin wires.  
- \*\*Luster:\*\* The shiny appearance of a metal when it reflects light.  
- \*\*Malleability:\*\* The ability of a material to be hammered or rolled into thin sheets without breaking.  
- \*\*Melting Point:\*\* The temperature at which a solid turns into a liquid.  
- \*\*Metallic Lattice:\*\* A regular, repeating arrangement of metal atoms in a crystal structure.  
- \*\*Sea of Electrons:\*\* A model in which valence electrons in a metallic bond move freely across the lattice of metal atoms.  
- \*\*Valence Electrons:\*\* The electrons found in the outermost shell of an atom that participate 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, and luster.  
- \*\*Describe the formation of metallic bonds\*\*, emphasizing the role of the "sea of electrons" and the metallic lattice structure.  
- \*\*Analyze the relationship between the structure of metals and their properties\*\*, explaining how metallic bonds contribute to properties such as ductility, conductivity, and melting points.  
  
 6. Engage (Ignite):  
Start the lesson by asking:   
\_"Why do metal street signs and lampposts, which are exposed to the same icy conditions as salt on the roads, not melt or dissolve?"\_  
  
\*\*Hands-On Experiment: Investigating the Conductivity of Metals\*\*   
Materials:   
- Copper wire   
- Steel nail   
- Light bulb   
- Battery   
- Saltwater solution   
- Plastic spoon   
- Wires with alligator clips   
  
\*\*Procedure:\*\*   
1. Connect the copper wire to the battery and the light bulb using the alligator clips. Observe the brightness of the light bulb.   
2. Replace the copper wire with a steel nail. Observe any changes in the bulb's brightness.   
3. Dip the ends of the wires into a saltwater solution and observe if the light bulb lights up.   
4. Finally, dip a plastic spoon into the saltwater solution and observe whether the light bulb lights up.   
  
\*\*Follow-Up Questions:\*\*   
1. Which materials conducted electricity well and why?   
2. Why do metals conduct electricity better than non-metals like plastic?   
3. How does the experiment help explain the role of the "sea of electrons" in metals?  
  
 7. Pre-Explore (Direct Instruction):  
Provide background information:   
Metals have unique properties that make them useful in many everyday applications. They are shiny (luster), can be shaped into wires (ductility), and can be hammered into sheets (malleability). These properties are due to the special way that atoms bond together in metals.  
  
In metals, atoms are arranged in a regular pattern called a \*\*metallic lattice\*\*. The outer electrons (valence electrons) of metal atoms are not tightly held and can move freely throughout the lattice. This free movement of electrons is called the \*\*sea of electrons\*\*. The positive metal atoms are surrounded by this sea, and the attraction between the positive atoms and the free-moving electrons holds the structure together.  
  
This structure explains metal properties:  
- \*\*Conductivity:\*\* The sea of electrons can carry electric current easily.  
- \*\*Malleability and Ductility:\*\* The atoms in the lattice can slide past each other without breaking the bonds, allowing the metal to be shaped without breaking.  
- \*\*Luster:\*\* The free electrons can reflect light, giving metals their shiny appearance.  
  
 8. Evaluate (Progress Check) - Pre-Explore:  
1. What is the metallic lattice, and how does it differ from the lattice found in ionic compounds?  
2. How does the sea of electrons explain the ability of metals to conduct electricity?  
3. Why are metals malleable, while ionic compounds are typically brittle?  
  
 9. Explain (Lightbulb):  
 \*\*Formation of Metallic Bonds and Their Role in Metal Properties\*\*:  
Metals are made up of atoms that are closely packed in a regular pattern, called a \*\*metallic lattice\*\*. The unique thing about metal atoms is that they have a few electrons in their outermost shell, called \*\*valence electrons\*\*. In a metal, these valence electrons are not bound tightly to any one atom. Instead, they can move freely throughout the entire structure. This results in what’s known as a \*\*sea of electrons\*\*.  
  
The \*\*sea of electrons\*\* is a key feature of metallic bonding. The electrons can move around, but the metal atoms stay in place, forming a lattice of positive ions. The attraction between the positive ions and the negative sea of electrons holds the metal together. This is why metals are strong and have high melting and boiling points.  
  
 \*\*How the Structure of Metals Explains Their Properties\*\*:  
The structure of metal atoms in a metallic lattice, combined with the sea of electrons, gives rise to several important properties:  
  
- \*\*Malleability and Ductility\*\*: Metals can be bent or hammered into sheets (malleability) or drawn into wires (ductility). This is because the metal atoms can slide past each other when a force is applied, without breaking the bonds that hold them together. Unlike ionic compounds, where moving atoms would cause the bonds to break, metallic bonds allow atoms to shift while staying bonded.  
  
- \*\*Conductivity\*\*: Metals are excellent conductors of electricity and heat. This happens because the electrons in the sea of electrons can move freely. When an electric current is applied, the electrons flow through the metal, carrying the charge. Similarly, heat is transferred because the electrons can move and transfer energy throughout the metal.  
  
- \*\*Luster\*\*: Metals have a shiny appearance because the free electrons in the sea of electrons can absorb and reflect light. When light hits the surface of the metal, the electrons reflect it, making the metal appear shiny.  
  
- \*\*Melting and Boiling Points\*\*: Metals generally have high melting and boiling points. This is because the metallic bonds are strong, and it takes a lot of energy to break the bonds and turn the metal into a liquid or gas.  
  
  
  
 \*\*Sample Problem:\*\*  
\*\*Question 1:\*\*   
Why do metals conduct electricity better than ionic compounds when they are solid?  
  
\*\*Explanation:\*\*   
In metals, the sea of electrons allows free movement of electrical charge, even in solid form. In contrast, ionic compounds only conduct electricity when dissolved in water or melted because their ions are locked in place in the solid state.  
  
  
  
 10. Post-Explore (Elaborate and Extend):  
\*\*Further Explanation:\*\*   
In the hands-on experiment, you saw that metals like copper and steel allowed electricity to pass through, lighting the bulb. This is because the sea of electrons allows charges to move easily through the metal. The saltwater solution conducted electricity as well since ionic compounds dissolve in water to release charged ions. However, the plastic spoon did not conduct electricity because it does not have free-moving charges like the sea of electrons in metals.  
  
Metals and salts both have lattice structures, but the way they bond is different. Metals have metallic bonds with free-moving electrons, while salts (ionic compounds) have ionic bonds that only conduct electricity when dissolved or melted.  
  
  
  
 Sample Problem:   
\*\*Question 2:\*\*   
Why can metals be hammered into thin sheets without breaking, while salts (ionic compounds) tend to shatter when struck?   
  
\*\*Answer:\*\*   
In metals, the atoms can slide past each other without breaking the metallic bonds, thanks to the sea of electrons. In salts, the ions are held together by ionic bonds, and if the layers shift, like charges repel each other, causing the salt to shatter.  
  
  
  
 Conclusion:  
The unique properties of metals—such as being good conductors, malleable, ductile, and shiny—can all be explained by their structure and the formation of metallic bonds. The \*\*sea of electrons\*\* allows metals to conduct electricity, be stretched into wires, and hammered into sheets. Understanding metallic bonding helps explain why metals behave differently from other types of materials, like salts, in different environments.  
  
 10. Evaluate (Progress Check) - Explain  
  
\*\*Scaffolded Questions:\*\*  
1. \*\*(DOK 1 - Recall):\*\* What is the atomic number, and how does it relate to the elements on the periodic table?  
 - \*\*Answer:\*\* The atomic number is the number of protons in the nucleus of an atom. It determines the identity of the element and its position on the periodic table. For example, hydrogen has an atomic number of 1.  
  
2. \*\*(DOK 2 - Skill/Concept):\*\* Why do atoms form bonds with other atoms? Provide an example.  
 - \*\*Answer:\*\* Atoms form bonds to achieve a full outer shell of electrons, which makes them more stable. For example, sodium and chlorine form an ionic bond, where sodium gives up an electron to chlorine, resulting in a stable arrangement for both atoms.  
  
3. \*\*(DOK 3 - Strategic Thinking):\*\* How does the concept of electronegativity help us understand the type of bond formed between two atoms?  
 - \*\*Answer:\*\* Electronegativity refers to an atom's ability to attract shared electrons. If two atoms have a large difference in electronegativity, they are likely to form an ionic bond. If the difference is small, they will form a covalent bond. For example, the bond between hydrogen and oxygen in water is polar covalent due to a difference in electronegativity.  
  
  
  
 11. Elaborate (Power Up)  
  
\*\***Mini-Tasks:\*\***  
1. \*\*(DOK 2 - Skill/Concept)\*\*: Predict how the properties of an element change as you move across a period in the periodic table.  
 - \*\*Answer:\*\* As you move across a period from left to right, the atomic radius decreases, ionization energy increases, and electronegativity generally increases. This is because the number of protons increases, pulling electrons closer to the nucleus.  
  
2. \*\*(DOK 3 - Strategic Thinking)\*\*: Compare and contrast ionic and covalent bonds in terms of how they form and the properties of the resulting compounds.  
 - \*\*Answer:\*\* Ionic bonds form when one atom donates an electron to another, resulting in charged ions that attract each other. Ionic compounds tend to have high melting points and conduct electricity when dissolved in water. Covalent bonds form when atoms share electrons, and covalent compounds tend to have lower melting points and do not conduct electricity in water.  
  
3. \*\*(DOK 3 - Strategic Thinking)\*\*: Design an experiment to test how the structure of a molecule affects its boiling point. What would you expect to find?  
 - \*\*Answer:\*\* An experiment could involve comparing the boiling points of molecules with different structures, such as water (H₂O) and methane (CH₄). Since water has hydrogen bonds (a strong intermolecular force), it has a higher boiling point than methane, which only has weaker London dispersion forces.  
  
  
  
 12. Final Evaluation  
  
**\*\*Debate Question:\*\***  
- \*\*Should we focus more on developing new materials through synthetic chemistry, or should we prioritize finding natural alternatives?\*\*  
 - \*Arguments for synthetic chemistry\*: Synthetic materials can be tailored for specific purposes, often have better durability, and can be mass-produced. They also reduce the strain on natural resources.  
 - \*Arguments for natural alternatives\*: Natural materials are often more environmentally friendly, biodegradable, and sustainable. They reduce pollution and the reliance on non-renewable resources.  
  
\*\***Multiple-Choice Questions:\*\***  
  
1. \*\*Which bond is most likely between two atoms with a large difference in electronegativity?\*\*  
 - A) Covalent Bond   
 - B) Ionic Bond   
 - C) Metallic Bond   
 - D) Hydrogen Bond   
 - \*\*Correct Answer:\*\* B) Ionic Bond   
 - \*\*Explanation:\*\* A large difference in electronegativity causes one atom to transfer electrons to another, forming an ionic bond.  
  
2. \*\*What happens to atomic size as you move down a group in the periodic table?\*\*  
 - A) It decreases   
 - B) It increases   
 - C) It stays the same   
 - D) It fluctuates   
 - \*\*Correct Answer:\*\* B) It increases   
 - \*\*Explanation:\*\* As you move down a group, more electron shells are added, making the atom larger.  
  
3. \*\*Which of the following molecules is polar?\*\*  
 - A) O₂   
 - B) CO₂   
 - C) CH₄   
 - D) H₂O   
 - \*\*Correct Answer:\*\* D) H₂O   
 - \*\*Explanation:\*\* Water (H₂O) has a bent shape and an uneven distribution of electrons, making it polar.  
  
4. \*\*Which of the following best describes the role of valence electrons in bonding?\*\*  
 - A) They are shared between atoms in covalent bonds.   
 - B) They are transferred from one atom to another in ionic bonds.   
 - C) Both A and B   
 - D) Neither A nor B   
 - \*\*Correct Answer:\*\* C) Both A and B   
 - \*\*Explanation:\*\* Valence electrons can be shared in covalent bonds or transferred in ionic bonds.  
  
**\*\*Long-Answer Questions:\*\***1. \*\***Explain how the structure of a molecule can affect its physical properties, such as melting point, boiling point, and solubility.\*\*  
 - \*\*Answer:\*\*** The structure of a molecule affects the type of intermolecular forces present. Molecules with strong intermolecular forces, like hydrogen bonds, have higher melting and boiling points. Polar molecules, like water, are also more likely to be soluble in polar solvents, whereas non-polar molecules dissolve better in non-polar solvents.  
 **2. \*\*Describe how ionic and covalent compounds differ in their electrical conductivity.\*\*  
 - \*\*Answer:\*\*** Ionic compounds conduct electricity when dissolved in water or melted because the ions are free to move and carry charge. Covalent compounds generally do not conduct electricity in solid or liquid form, as they do not have charged particles capable of conducting electrical current.  
  
3. \*\***How does the periodic trend of electronegativity explain the type of bond formed between sodium (Na) and chlorine (Cl)?\*\*  
 - \*\*Answer:\*\*** Sodium has a low electronegativity, while chlorine has a high electronegativity. This large difference causes sodium to transfer its electron to chlorine, forming an ionic bond. Sodium becomes a positive ion (Na⁺) and chlorine becomes a negative ion (Cl⁻).  
  
4. **\*\*Provide an example of a real-world application where understanding molecular polarity is important.\*\*  
 - \*\*Answer**:\*\* One example is in the design of medicines. Polar molecules are often more soluble in blood, which is mostly water (a polar solvent). This affects how drugs are absorbed and transported in the body. Understanding polarity helps chemists design drugs that can be more easily delivered to the target areas in the body.  
  
  
  
 13. Extend (Beyond the Lesson)  
  
**\*\*Additional Tasks:\*\***- \*\*Task:\*\* Research the role of chemistry in environmental sustainability. How can chemistry help in developing renewable energy sources, such as solar panels or biofuels? Write a short report on your findings.  
 - \*Purpose\*: This task encourages students to apply their knowledge of chemistry to environmental issues, linking classroom learning to real-world applications.  
  
**\*\*Readings:\*\***  
- \*\*Reading Assignment:\*\* "The Chemistry of Climate Change" – an article explaining the role of greenhouse gases, such as CO₂, and how chemical reactions contribute to global warming.  
 - \*Purpose\*: This reading will help students understand how chemical reactions in the atmosphere impact the environment.  
 **\*\*Spaced Practice Activity:\*\***- \*\*Activity:\*\* Over the next week, keep a journal of daily items you use (e.g., plastics, medicines, fuels). Identify the type of chemical bonds present in these items and how they affect their physical properties.  
 - \*Purpose\*: This activity encourages students to revisit core concepts regularly, reinforcing their understanding of chemical bonds and physical properties.  
  
These tasks, readings, and activities help students build on foundational knowledge and think critically about how chemistry is applied in the real world.