Chemistry Year 1

SECTION

7

INTER ATOMIC BONDING

SYSTEMATIC CHEMISTRY OF THE ELEMENTS

Bonding

INTRODUCTION

Ionic bonds form when one atom gives an electron to another, creating solid crystals with high melting points that can conduct electricity in water; covalent bonds form when atoms share electrons, resulting in substances that can be gases, liquids, or solids with lower melting points and no electrical conductivity; metallic bonds occur when metal atoms share a pool of free electrons, making metals shiny, malleable, and good conductors of heat and electricity.

At the end of this section, you will be able to:

- Explain ionic bonding and its formation and state the properties of ionic compounds.
- Explain covalent bonding and its formation and state the properties of covalent compounds.

Key Ideas

- **Chemical bonds** are the forces that hold atoms together in a molecule or compound.
- **Lewis dot symbol** is a representation of an atom's valence electrons using dots around the symbol of the element.
- **Ionic bond** is a type of chemical bond formed between two atoms when one atom donates an electron to another, resulting in the formation of positively charged cations and negatively charged anions.
- **A covalent bond** is a chemical bond that involves the sharing of electron pairs between atoms.
- **Metallic bonds** are the forces that hold metal atoms together in a solid.
- **Electron affinity** is the energy change that occurs when an electron is added to a neutral atom in the gas phase to form a negative ion.

- **Large lattice energy** is the amount of energy released when ions in the gas phase come together to form a crystalline solid.
- **Polarising power** is the ability of a cation to distort the electron cloud of an anion.

CHEMICAL BONDING

Chemical bonds are defined as whenever two or more atoms are held strongly together. They do so in a particular way, which tends to give the constituent atoms specific whole number ratios. Gilbert Lewis, an American scientist, explained that atoms combine to achieve a more stable electron configuration. Maximum stability is attained when an atom loses, gains or shares electrons to achieve a similar electron configuration to that of a noble gas (inert gas).

In chemical bond formation, only the outermost shells of the constituent atoms come into contact. So, only the valence electrons are involved in bond formation. The **Lewis dot** symbol, which consists of the symbol of an element and a dot for one valence electron in an atom, helps to explain how atoms interact to form chemical bonds. Elements in the same group on the periodic table have the same number of valence electrons, hence similar Lewis' dot symbols, except for their atomic symbols. The constituent atoms in a molecule can donate, accept or share a certain number of electrons to form a specific type of chemical bond. The type of bond created depends on the electronegativity difference between the atoms involved.

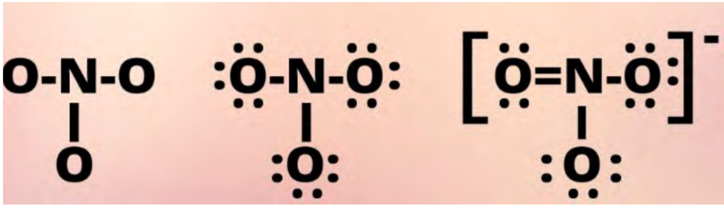


Fig. 7.1: A chemical bond.

Image source: https://sciencenotes.org/how-to-draw-a-lewis-structure/

Atoms connect through chemical bonds. Electronegativity strongly influences how atoms interact with each other and how they bond. The electronegativity difference between two bonded atoms determines the nature of the chemical bond that forms between them. If the electronegativity difference is large (***equal to or greater than approx. 1.7***), the bond that forms between the atoms will be ionic, and if it is small (***less than approx. 1.***7), a covalent bond will generally form.

Ionic Bond

This is an electrostatic force of attraction formed between a positive ion and a negative ion. Usually, the less electronegative element completely transfers its valence electron(s) to the more electronegative element. The bond is formed between a positive ion (cation) and a negative ion (anion). This exchange results in a more stable, noble gas electron configuration for both atoms involved. An ionic bond is based on attractive electrostatic forces between two ions of opposite charge.

Ionic bonds can also be formed between species that are already ions rather than neutral atoms forming ions e.g. precipitation reactions.

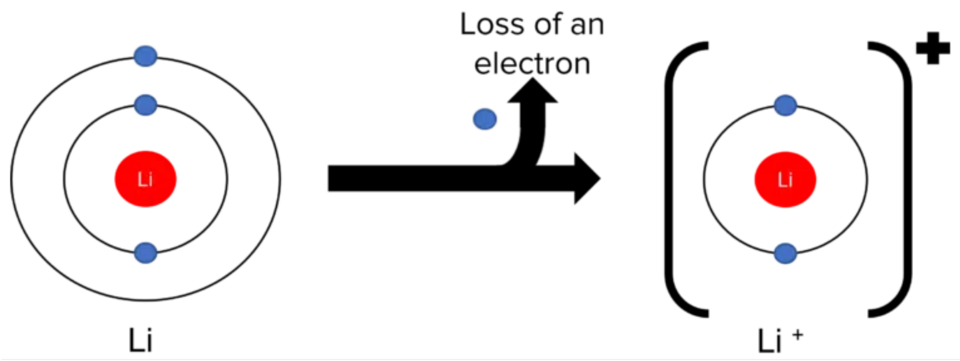
Cations and Anions

When an atom, typically a metal, loses an electron or electrons and becomes a positive ion, it is called a *cation*. When an atom, typically a non-metal, gains an electron or electrons and becomes a negative ion, it is called an *anion*.

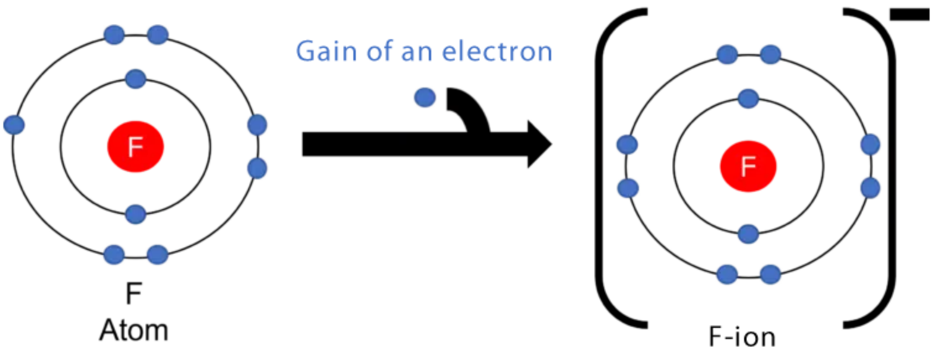
One example of an ionic bond is the formation of sodium fluoride, NaF, from sodium and fluorine atoms. In this reaction, the sodium atom loses its single valence electron to the fluorine atom, which has just enough space to accept it.

$Na \rightarrow Na^{+} + e^{-}$ ***and*** $F + e^{-} \rightarrow F^{-}$

The ions produced are oppositely charged and are attracted to one another due to electrostatic forces. $Na^{+} + F^{-} \rightarrow NaF$



Lithium cation



Fluorine anion

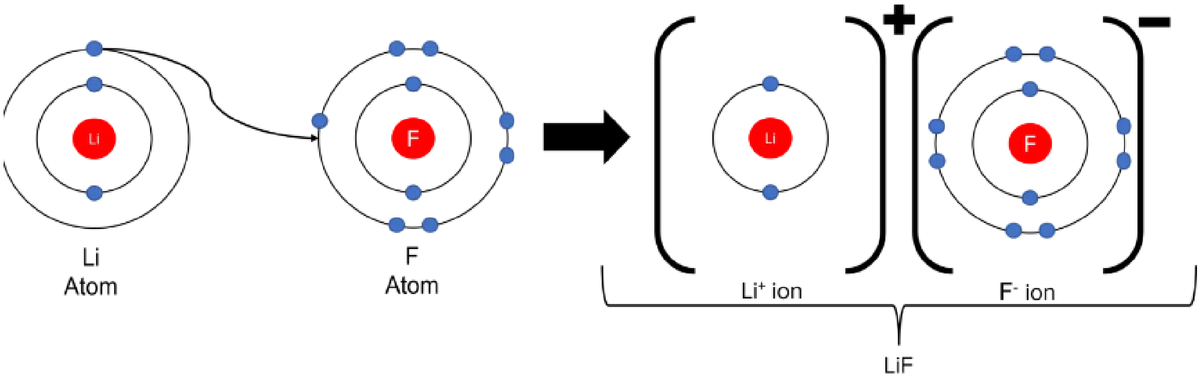


Fig 7.2: Formation of ionic compound **Lithium Fluoride (LiF).**

Image source: https://mmerevise.co.uk/gcse-chemistry-revision/ionic-bonding/

Factors that affect ionic bond formation

1. Low ionisation energy of the metallic element which forms the cation.
2. Large electron affinity of the non-metallic element which forms the anion.
3. Large lattice energy i.e., smaller size and higher charge of the ions.

The electronegativity difference of the elements affects the bonding of atoms. Elements with high electronegativity tend to form ionic bonds with elements of low electronegativity.

Properties of Ionic Compounds

1. Ionic compounds have high melting and boiling points.
2. Some dissolve in polar solvents like water.
3. They conduct electricity in the aqueous form or molten state (as the ions are free to move).
4. In their solid form, they serve as good insulators (as the ions cannot move).
5. They are hard and brittle in nature.

COVALENT BONDING

Covalent Bonds

This is a chemical bond that is formed when two atoms mutually share a pair of electrons. Covalent bonds are usually found between non-metal atoms. By doing so, the atoms attain a stable duplet or octet electronic configuration. In covalent bonding, overlapping of the atomic orbitals having one electron from each of the two atoms takes place, resulting in the sharing of the pair of electrons. Generally, the orbitals of the electrons in the valence shell of the atoms are used for electron sharing. For example, in a hydrogen molecule (H_2), a covalent bond is formed by the overlap of the two s-orbitals each containing one electron from each of the two H atoms of the molecule. Each electron in a shared pair of electrons is attracted to the nuclei of both atoms. Each H atom attains a $1s^2$ configuration. **H• + •H → H: H**

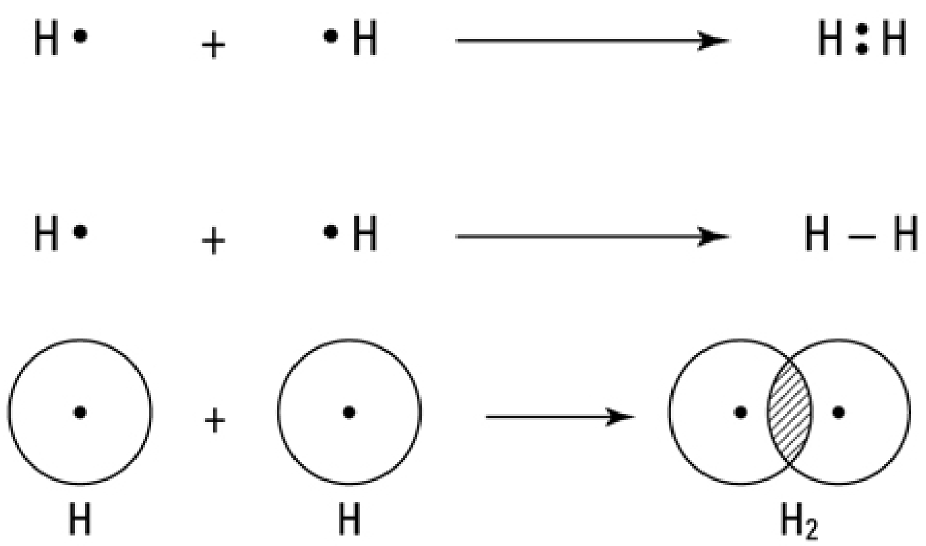


Fig 7.3: A covalent bond.

Image source: https://www.dummies.com/article/academics-the-arts/science/chemistry/covalent-bonds-a-hydrogen-example-194230/

Covalent bonding between atoms that have electrons beyond the first shell involves only the unpaired valence electrons. In the covalent bond formation by two fluorine atoms in F_2, there are seven (7) valence electrons, which signifies that only one (1) unpaired electron exists on a fluorine atom

Only two (2) valence electrons participate in the formation of the covalent bond in F_2, leaving six (6) valence electrons in each atom not involved in forming the covalent bond. The pairs of electrons that are not used in bonding are called lone pairs. Therefore, F_2 has six lone pairs of electrons (3 lone pairs on each F atom).

Fig 7.4. Image source: Raymond Chang and Jason Overby, 2009

Types of Covalent Bonds

Pure covalent bond

When the atoms that form a covalent bond are identical, as in H_2, Cl_2 and other diatomic elements, then the electrons in the bond are shared equally. We refer to this as a pure covalent bond**.**

Electrons shared in pure covalent bonds have an equal probability of being near each nucleus.

Dative bond or coordinate covalent bonding

This is a type of covalent bond in which both electrons in the bond are donated by one atom. That is, the pair of electrons that are shared between the two atoms comes from only one of the bonding atoms. This type of bonding is different from a typical covalent bond, where each atom contributes one electron to form a shared pair.

Formation of dative bonds (coordinate covalent bonding)

a. Formation of Ammonium ion (NH_4^+). The reaction between an ammonia molecule (NH_3) and a proton (H^+) forms an ammonium ion (NH_4^+). $NH_3 + H^+ \rightarrow NH_4^+$. In ammonia, three hydrogen atoms combine directly with the nitrogen atom by normal covalent bonding. The nitrogen atom has a lone pair of electrons in its outermost shell. It acts as a lone pair donor, i.e. can combine with the hydrogen ion. It shares the lone pair of electrons with a proton from an acid to produce the ammonium ion, NH_4^+. The proton carries over its positive charge to give the ammonium ion, NH_4^+.

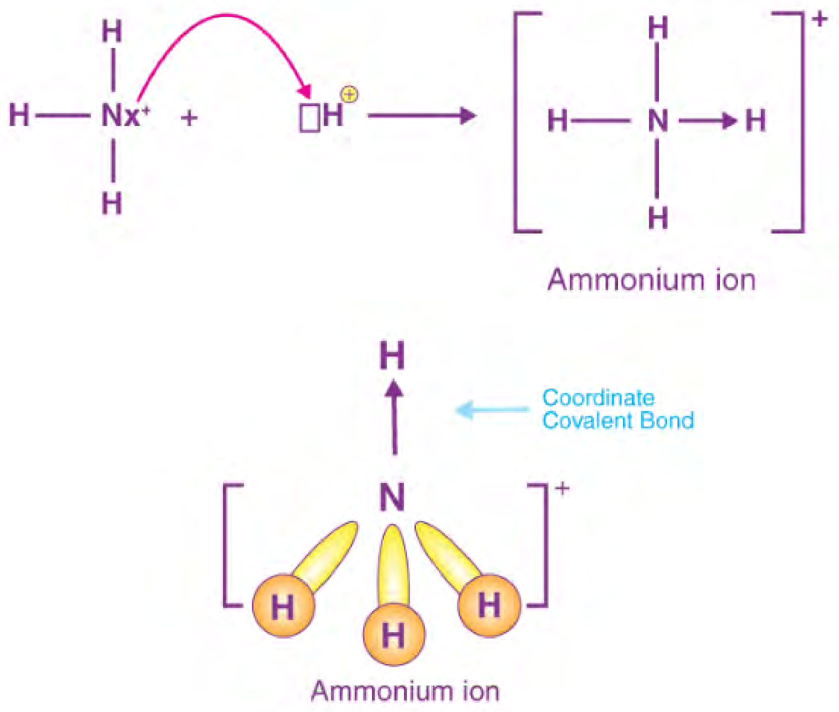


Fig 7.5: Formation of Ammonium ion.

Image source: https://byjus.com/chemistry/ammonium-ion/

b. Formation of a hydroxonium ion (H_3O^+). $H_2O + H^+ \rightarrow H_3O^+$. In a water molecule, two hydrogen atoms share two pairs of electrons with an oxygen atom by normal covalent bonding. The oxygen atom in a water molecule has two lone pairs of electrons in its outermost shell. It shares this with a proton (hydrogen ion, H^+) from an acid to produce the hydroxonium ion H_3O^+. The positive charge on the hydrogen ion is carried over to give the positively charged hydroxonium ion, H_3O^+.

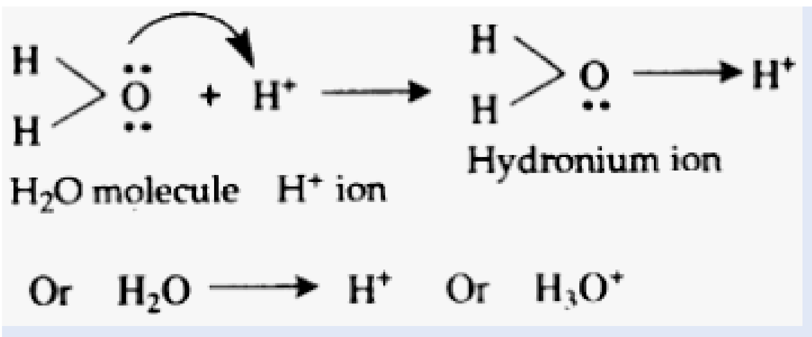


Fig. 7.6: Formation of hydroxonium ion.

Image source: https://www.sarthaks.com/274945/explain-the-formation-of-h3o-and-nh4-ion

Polar covalent bond

This is a type of chemical bond formed between two atoms, where the electrons are shared unequally. In this type of bond, the bonding electrons are more attracted to one atom than the other, giving rise to a shift of electron density toward that atom. The atom that attracts the electrons more strongly acquires a partial negative charge whilst the other acquires a partial positive charge. For example, the electrons in the H–Cl bond of a hydrogen chloride molecule spend more time near the chlorine atom than near the hydrogen atom.

Whether a bond is non-polar or polar covalent is determined by a property of the bonding atoms called electronegativity. Electronegativity is a measure of the tendency of an atom to attract electrons (or electron density) towards itself. It determines how the shared electrons are distributed between the two atoms in a bond. The more strongly an atom attracts the electrons in its bonds, the larger its electronegativity. Electrons in a polar covalent bond are shifted toward the more electronegative atom; thus, the more electronegative atom is the one with the partial negative charge. The greater the difference in electronegativity, the more polarised the electron distribution and the larger the partial charges of the atoms.

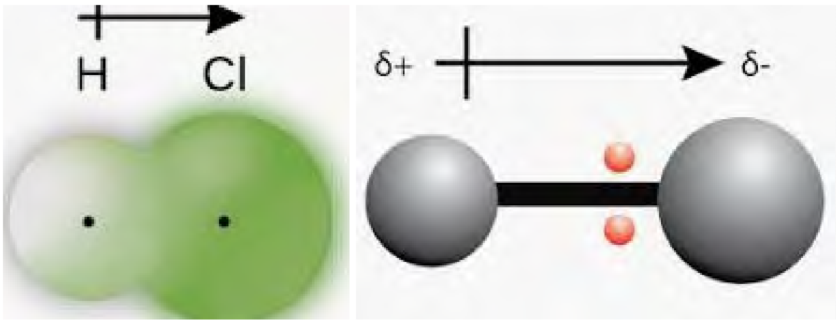


Fig. 7.7: Formation of HCl.

Image source: https://socratic.org/questions/how-does-a-covalent-bond-become-polar

Factors That Affect the Formation Of Covalent Bonds

1. Number of valence electrons
2. High ionisation energy
3. Comparable electron affinity
4. Comparable electronegativities
5. Atomic size
6. High nuclear charge and small internuclear distance.

Properties Of Covalent Compounds

Compounds that contain covalent bonds exhibit different physical properties than ionic compounds. Covalent compounds;

1. Exist in three states of matter namely solid, liquid and gas.
2. Do not conduct electricity in a solid, molten or aqueous state.

Polarisation

During the formation of an ionic compound or ionic molecules, two oppositely charged ions (cations and anions) must come closer to each other. During this process, the cation attracts the electron charge cloud of the outermost shell of the anion toward itself. Therefore, the symmetrical shape of the anion gets distorted, deformed or polarised.

The phenomenon of the distortion of the symmetrical shape of the electron cloud of an anion by the nearby cation is called polarisation of anion.

There is also the chance of the polarisation of a cation by an anion. However, due to the smaller size of the cation, its electron cloud is strongly held to the nucleus, and the shape of the cloud is not distorted to an appreciable extent. Hence, the polarisation of a cation by an anion is not generally considered.

Polarising Power

The ability of a cation to polarise (distort) an anion is called its polarising power or polarisability.

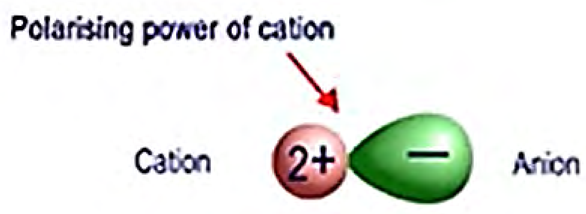


Fig. 7.8: Polarising power of a cation

Factors affecting polarising power

The factors affecting the magnitude of a cation's polarising power are listed below:

a. *Magnitude of positive charge on the cation*

The greater the charge on the cation, the more strongly it attracts the outermost shell electron cloud of an anion toward itself and polarises the given anion easily. Therefore, the polarising power of a cation is directly proportional to the magnitude of the positive charge on it. E.g. $Na^{+} < Mg^{2+} < Al^{3+}$. If the same element has a different positive charge, the higher positive charge has greater power of polarisation. E.g., $Sn^{4+} > Sn^{2+}$.

b. *Size of cation*

The smaller the size of the cation, the more strongly it attracts the outermost shell electron cloud of an anion towards itself, hence, the greater its polarising ability. In other words, with the decreasing size of the cation, the polarising power of the cation increases. Thus, the polarising power of the cation is inversely proportional to the size of the cation. Example: $Li^{+} > Na^{+} > K^{+} > Rb^{+}$

Fig. 7.9: Factors affecting polarising power.

Polarisability of anions

The tendency of an anion to get polarised by a cation is called its polarisability. The factors affecting the polarisability of an anion are ***magnitude*** of the ***negative charges*** on the anion and the size of the anion.

a. *Magnitude of negative charges on anion*

The higher the negative charge on the anion, the more easily its outermost electron cloud is attracted by cations, hence the Polarisability of an anion is directly proportional to the magnitude of the negative charges on it. E.g., $C^{4-} > N^{3-} > O^{2-} > F^{-}$

b. *Size of the anion*

The larger the size of the anion, the more easily its outermost shell electron cloud is attracted by the cation towards itself, hence, the greater the polarisability of an anion. Thus, the polarisability of an anion is directly proportional to the size of an anion.

METALLIC BONDING

This is a type of chemical bond that occurs in metals. It is an electrostatic force of attraction between the fixed positive metal ions and the delocalised electrons around the cations. The lattice structure of the atoms is held together by a sea of delocalised valence electrons. The delocalised electrons can also be referred to as a sea of electrons or mobile electrons.

Metallic Bonds

Metallic bonds are formed by the process of metal atoms transferring their outermost electrons to an electron sea, which is the collection of shared electrons that surround the positively charged atomic cores of the metal atoms. When heated, the metal atoms release their outermost electrons, which then move freely throughout the lattice. Since metal atoms have very low electronegativity, they tend to lose their valence electrons, readily forming positive ions in the process. These positive ions are then surrounded by a cloud of delocalised electrons, which form the metallic bond that holds the ions together in a regular lattice structure.

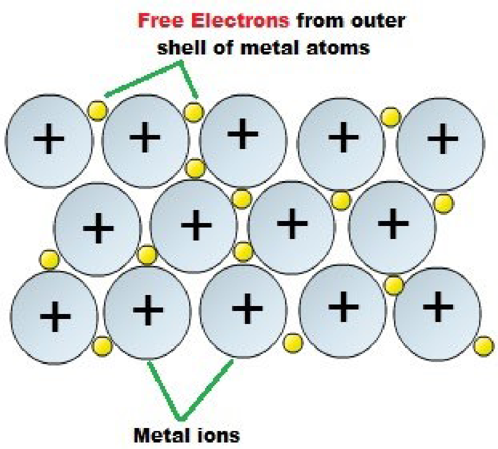


Fig. 7.10: A metallic bond

The strength of the metallic bond depends on factors such as the number of valence electrons in the metal atom, the size of the atoms and the proximity of the atoms in the lattice. Metallic bond strength increases as the number of valence electrons increases. Thus, aluminium is harder than magnesium, which is in turn harder than sodium because 3, 2, and 1 valence electrons are attracted by the fixed positive lattice points respectively. As the atomic size of these atoms reduces with increased nuclear attraction, their melting points also increase accordingly.

Properties of Metals

Malleable

Think of a lump of clay. You can flatten it with your hands or a rolling pin into a thin sheet. Metals are kind of like that. Malleable means a metal can be hammered or pressed into different shapes, especially thin sheets, without breaking.

This is why we can make things like aluminium foil, car bodies, and metal sheets for roofs. If metals were brittle (like glass), they would just shatter when you tried to flatten them.

Gold is extremely malleable. You can hammer it into incredibly thin sheets called gold leaf, which are used for decoration.

Ductile

Ductile means a metal can be stretched into thin wires. This is how we make electrical wires, which are essential for carrying electricity to our homes and devices. Copper is very ductile, which is why it's used so much for electrical wiring.

Good conductors of heat and electricity

1. **Heat:** If you put a metal spoon in a hot cup of tea, the spoon quickly gets hot too. This is because metals allow heat to flow through them easily. "Good conductor of heat" means heat travels through the metal quickly. This is why we use metal pots and pans for cooking. They heat up quickly and evenly.
2. **Electricity:** Electricity is the flow of electrons. Metals have free electrons that can move easily, allowing electricity to flow through them. "Good conductor of electricity" means electricity travels through the metal easily. This is why we use metal wires to carry electricity.

Silver and copper are excellent conductors of both heat and electricity.

Activity 7.1: Malleability Demonstration (Aluminium Foil Shaping)

Materials needed:

- Aluminium foil
- A small hammer (or a heavy spoon)
- A flat surface (like a cutting board)

Procedure:

1. Take a piece of aluminium foil.
2. Try to shape it with your hands. Notice how easily it bends and changes shape.
3. Place the foil on a flat surface.
4. Gently tap the foil with the hammer (or spoon). Observe how the foil flattens and spreads out without breaking easily.

Activity 7.2: Ductility Demonstration (Copper Wire Stretching)

Materials:

- Thin copper wire (easily found in some discarded electronics or crafting supplies)
- Two pairs of pliers (optional, for better grip)

Safety Note: Be careful not to cut yourself with the wire.

Procedure:

1. Hold each end of the copper wire firmly.
2. Gently and slowly pull on the wire, trying to stretch it.
3. Observe how the wire becomes thinner and longer before it eventually breaks.
4. If you have pliers, grip each end of the wire with the pliers, and pull.

Activity 7.3: Heat Conductivity Demonstration (Spoon in Hot Water)

Materials:

- A metal spoon (preferably a stainless steel or silver spoon)
- A plastic or wooden spoon (for comparison)

- A mug or heat-safe container
- Hot water (**be careful!**)

Procedure:

1. Pour hot water into the mug.
2. Place the metal spoon and the plastic/wooden spoon into the hot water, with the handles sticking out.
3. Wait for a minute or two.
4. Carefully touch the handles of both spoons.
5. Notice the temperature difference**.**

Safety notes:

- **Teacher supervision is recommended, especially when working with hot water or tools.**
- **Always be careful when handling any type of wire, especially when handling any electricity. Never use household current to test conductivity.**

REVIEW QUESTIONS

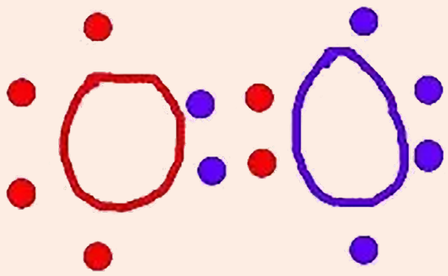
Review Question 7.1

1. In which of these compounds is the ionic bond strongest?
   $NaCl$, $MgCl_{2,}$ $AlCl_3$
2. Explain why sodium chloride does not conduct electricity when solid.
3. Why do ionic compounds have high boiling and melting points?
4. Use Lewis dot structures to explain the nature of chemical bonding found in oxygen gas (O_2) and identify the number of bonds formed in the molecule

Review Question 7.1

1. $AlCl_3$
2. At normal temperatures, sodium chloride is solid, and ionic compounds do not conduct electricity in the solid state because their ions are fixed in position and cannot move but in aqueous solution and in a molten state, they conduct electricity as ions are free to move.
3. The ionic compounds have high melting and boiling points because they are made up of positive and negative ions which have the strong force of attraction between the oppositely charged ions. So, a lot of heat energy is required to break this strong force of attraction or the ionic bond making ionic compounds have high melting points.
4. Oxygen has an atomic number of 8. It has 6 valence electrons. According to the octet rule, atoms strive to achieve a stable octet (8 valence electrons) configuration. In the Lewis dot structure of O_2, each oxygen atom shares 2 electrons with the other, forming a total of 2 covalent bonds. This allows each oxygen atom to achieve a stable octet configuration, with 8 valence electrons.

   The Lewis dot structure shows 2 lone pairs of electrons on each oxygen atom, in addition to the 2 shared electrons in the covalent bonds. This gives each oxygen atom a total of 8 valence electrons, satisfying the octet rule. The sharing of electrons in the covalent bonds and the presence of lone pairs allows the oxygen atoms to reach a stable, octet configuration, which is the preferred state according to the octet rule.



O_2 molecule

EXTENDED READING

Go online and read about how to use the Lewis electron dot to explain bonding. You may also watch YouTube videos by clicking on the links below:

- (https://www.youtube.com/watch?v=cIuXl7o6mAw)
- https://www.kau.edu.sa/Files/0053615/Files/21715_CHEM%20110%20UniChem%20Chapter%205.pdf
- https://www.lwtech.edu/campus-life/learning-lab/resources/docs/lwtech-learning-lab-science-lewis-dot-structure.pdf
- https://www.youtube.com/watch?v=cIuXl7o6mAw
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GLOSSARY

- **Polarisation** in chemistry refers to the distortion of the electron cloud of an atom or ion by a nearby positive charge.
- **Malleable** is the property of a material, especially metals, that allows it to be hammered or rolled into thin sheets without breaking.
- **Ductile** is the property of a material that allows it to be stretched into a wire without breaking.
- **Ionisation energy** is the amount of energy required to remove an electron from an atom or ion in its gaseous state.
- **Polarisability** is the tendency of an atom or ion's electron cloud to be distorted by an external electric field or nearby cation.

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