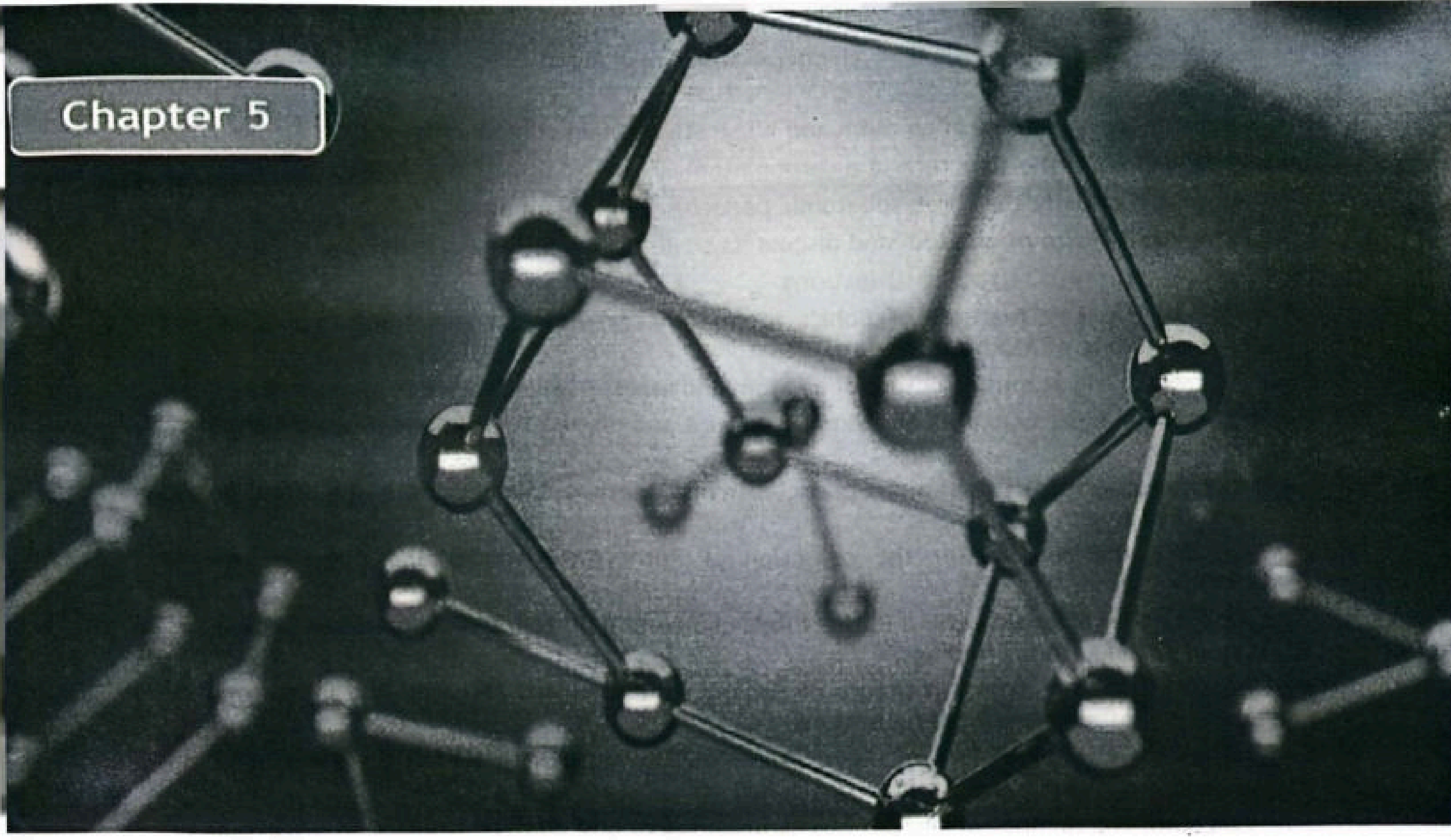


CHEMICAL BONDING

SLOs: After completing this lesson, the student will be able to:

1. Define Shells and Subshells
2. Determine the number of electrons in valence shell with help of periodic table.
3. Describe duplet and octet rule.
4. Identify elements of the periodic table and their ions using their atomic number and applying duplet and octet rule.
5. Justify why atoms form chemical bonds.
6. Describe the formation of an ionic bond.
7. Describe the characteristics of ionic compounds
8. Describe the formation of a covalent bond.
9. Describe the characteristics of covalent compounds.
10. Identify polar and nonpolar covalent compounds in heteroatomic and homoatomic molecules.
11. Differentiate between the ionic and the covalent compounds with examples.

5.1 CHEMICAL BONDING (SHELLS AND SUBSHELLS)

Introduction

Chemical bonding is when atoms join together to make their outer shell of electrons stable. Only the outer electrons, called valence electrons, are involved in this joining. The inner electrons don't participatein bonding.

Do You Know?

Without chemical bonding, atoms would not be able to come together to form molecules and compounds, leading to the absence of everything we see around us.

Every atom has tendency to fill up its outer shell to be like a noble gas, because that makes it really stable. When different atoms come together, they make lots of different compounds that make up everything around us. **Chemical bonding** is attraction between two or more atoms which hold them together.

Bonding in compounds

The type and strength of chemical bonds present in a substance influence its boiling point, melting point, solubility, reactivity, and other properties.

Do You Know?

Atoms in molecules stick together because the nucleus and the electrons attract each other.

In the first shell around an atom, there can be up to two electrons. In the second shell, there can be up to eight. And for the first twenty elements, the third shell can also have up to eight electrons.

Shell

Electrons revolve around the nucleus in a specific circular path known as orbit or called as shells. Electrons are arranged in shells around the nucleus. The first shell closest to the nucleus is the lowest energy level. The further a shell is from the nucleus higher the energy level. Each shell can hold only the certain number of electrons.

The closest shell has a value of n=1.(n represents shell number)

The next shell has a value of n=2.

These are the rules.

1. The first shell can hold only 2 electrons. It fills first.
2. Second shell can hold 8 electrons. It fills next.
3. The third shell can hold 18 electrons. But it fills up to 8. The next 2 go into fourth shell. Then the rest of the third shell fills.

The maximum number of electrons possible in the first four energy levels are.

Table 5.1

n=	Shell	Maximum Number of Electron
1	1st Shell	2
2	2nd Shell	8
3	3rd Shell	18
4	4th Shell	32

Imagine electrons move into shells. They follow a rule called the $2n^2$ rule, which helps us figure out how many electrons can fit in each shell. We start filling the shells from the bottom, making sure each one is full before moving to the next. But when we reach the third shell, things get a bit tricky because each shell has smaller parts called subshells.

5. Subshells

Think of subshells as different rooms in a big house called an atom, where electrons like to hide Inside each room, there are smaller spots called orbitals where the electrons stay. There are four types of rooms: s, p, d, and f. Each room can hold a certain number of electrons, but they're all different.

Let's take Lithium, which has 3 electrons. The first two electrons move into the 1st shell, in the 1s room. So, it goes like this: $1s^2$. The last electron goes into the 2nd shell, in the 2s room. So, it's written as: $1s^2 2s^1$. That's Lithium's way of showing how its electrons are arranged.

Table 5.2

Shell	Subshell	Total Number of Electrons in Shell
1st Shell	1s	2
2nd Shell	2s, 2p	2+6=8
3rd Shell	3s, 3p, 3d	2+6+10=18
4th Shell	4s, 4p, 4d, 4f	2+6+10+14=32

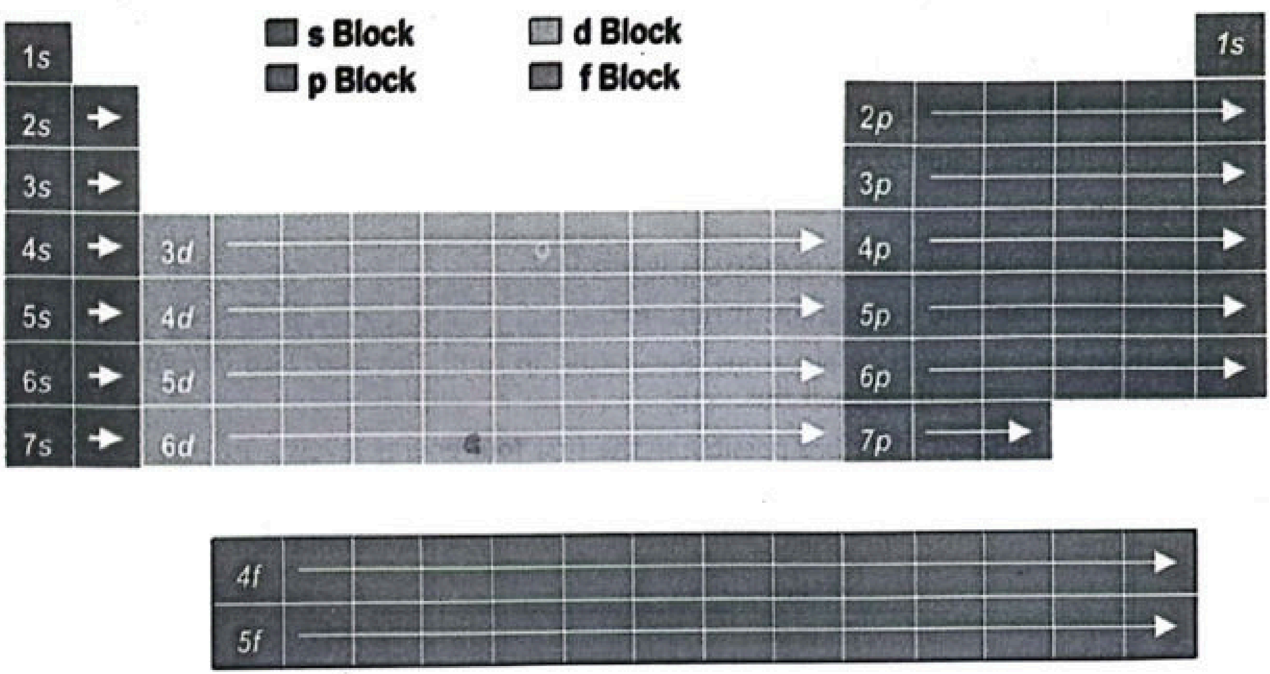


Fig. 5.1: Blocks of Periodic Table

5.2. NUMBER OF ELECTRONS IN VALENCE SHELL

Imagine the world of atoms as a big house with different rooms. The first shell (closest to the nucleus) can hold two electrons. The second shell can hold 8 electrons. The third shell can hold 32 electrons. Within the shells, electrons are further grouped into subshells of four different types, identified as s, p, d, and f in order of increasing energy. This grouping helps us see which elements act similarly because they share the same electron preference.

Additional information

S block and p block elements are called representative elements. d block elements are transition elements f block elements are called inner transition elements. In p block element helium is an exception. Helium has two electrons in s orbital what it is placed in the block due to the similar properties of noble gas elements

The first shell only has the s subshell ⟹⟹ 2 electrons.

The second shell has the s and p subshells ⟹⟹ 2 + 6 = 8 electrons.

The third shell has the s, p, and d subshells ⟹⟹ 2 + 6 + 10 = 18 electrons.

The fourth shell has the s, p, d, and f subshell ⟹⟹ 2 + 6 + 10 + 14 = 32 electrons.

Table 5.3 Electrons in the Valence Shell of Fist 20 Elements

Atomic Number	Element Name	K	L	M	N	O
1	Hydrogen (H)	1s1				
2	Helium (He)	1s2				
3	Lithium (Li)	1s2	2s1			
4	Beryllium (Be)	1s2	2s2			
5	Boron (B)	1s2	2s2	2p1		
6	Carbon (C)	1s2	2s2	2p2		
7	Nitrogen (N)	1s2	2s2	2p3		
8	Oxygen (O)	1s2	2s2	2p4		
9	Fluorine (F)	1s2	2s2	2p5		
10	Neon (Ne)	1s2	2s2	2p6		
11	Sodium (Na)	1s2	2s2	2p6	3s1	
12	Magnesium (Mg)	1s2	2s2	2p6	3s2	
13	Aluminum (Al)	1s2	2s2	2p6	3s2	3p1
14	Silicon (Si)	1s2	2s2	2p6	3s2	3p2
15	Phosphorous (P)	1s2	2s2	2p6	3s2	3p3
16	Sulfur (S)	1s2	2s2	2p6	3s2	3p4
17	Chlorine (Cl)	1s2	2s2	2p6	3s2	3p5
18	Argon (Ar)	1s2	2s2	2p6	3s2	3p6
19	Potassium (K)	1s2	2s2	2p6	3s2	3p6, 4s1
20	Calcium (Ca)	1s2	2s2	2p6	3s2	3p6, 4s2

A **period** is the horizontal row of the periodic table. All elements in a row have the same number of *electron shells*. If we know the period of an element, then we can predict the number of electrons shell of a neutral atom.

Periodic table Method for Calculation of Electrons in Valence Shell in an Atom

For a neutral atom the number of valence electrons is equal to the atom's main group number. The main group number for an element can be found. Every atom likes to keep a certain number of electrons around, and we call these "valence electrons." For most atoms in the main groups of the periodic table, the number of valence electrons matches the group number. So, if an atom is in group 1, it has 1 valence electron; if it's in group 6, it has 6 valence electrons, like oxygen.

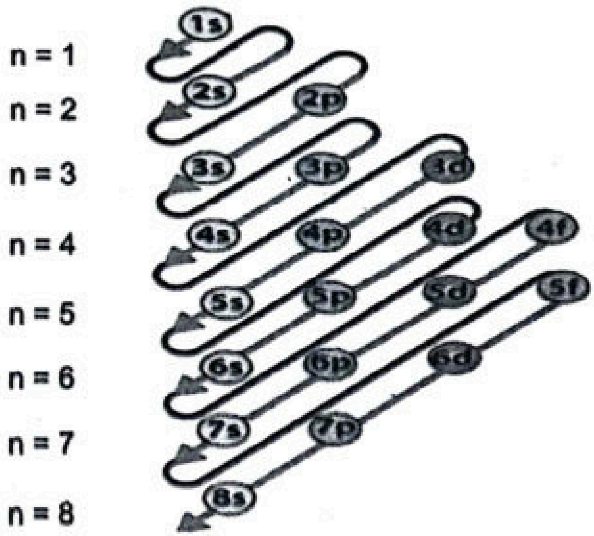


Fig. 5.2: Filling of Electrons in Sub Shell

But there's a special group called the transition metals that don't follow this rule. So, for those, we just have to look at the periodic table.

Group 1: 1 valence electron

Group 2: 2 valence electrons

Group 3: 3 valence electrons

Group 4: 4 valence electrons

Group 5: 5 valence electrons

Group 6: 6 valence electrons

Group 7: 7 valence electrons

Group 8: 8 valence electrons

For example, Sodium is in group 1, so it has 1 valence electron, while Oxygen is in group 6, so it has 6 valence electrons

5.3. DEFINITION OF DUPLET AND OCTET RULE

Atoms have rules they like to follow to feel balanced. One rule is the octet rule, which means atoms prefer to have up to eight electrons in their outer shell.

Then there's the duplet rule, which is simpler. It says some atoms are stable with just two electrons in their outer shell.

Hydrogen, Helium, and Lithium are special because they follow the duplet rule. Hydrogen can share or gain an electron to reach two, while Lithium is okay with losing one to get to two. They're like players following their own special rules to find balance.

5.4 APPLICATIONS OF DUPLET RULE

A Hydrogen atom has reached its stability by filling its outermost orbital with two electrons. When the element has two electrons in its outermost orbital instead of eight electrons it does not affect the stability of the individual element because it has 1s orbital which can maximum hold two electrons in shell.

The atomic number of hydrogen atom is one, so two hydrogen atoms share its one electron each to obtain its duplet state by forming a hydrogen molecule.

By using this rule, we can understand how different atoms bond. For instance:

Group I metals give away 1 electron to have 8.

Group II metals give away 2 electrons to have 8.

Group VI non-metals gain 2 electrons to have 8.

Group VII halogens gain 1 electron to have 8.

Sometimes, there are exceptions to the octet rule

Some molecules have an odd number of electrons.

Some atoms don't need 8 electrons; they're okay with fewer.

And some atoms can even have more than 8 electrons in their outer shell.

Understanding these rules helps us predict how atoms will bond together, like in carbon dioxide (CO_2), where carbon and oxygen share electrons to reach stability.

Octet Rule Examples

A few examples which follow the octet rule are:

CO_2, NaCl, MgO

1.CO_2

Carbon contains four electrons In Its outermost shell. Also, carbon should have four electrons to complete its octet when it is combined with two molecules of oxygen. Here each carbon atom requires two electrons to complete Its octet Carbon and oxygen share their outermost electron and form CO, which further completes the octet.

2. NaCl

Chlorine contains seven electrons In Its outermost shell and requires only one electron to complete its octet whereas sodium contains one electron in its outermost shell. Both sodium and chlorine complete their octet by forming Sodium Chloride (NaCl).

Exceptions to the Octet Rule

Many elements do not follow the octet rule. Some of the exceptions to the octet rule are given below:

An electron or molecule which contains unpaired electrons In Its outermost shell or valence shell is considered a free radical. These electrons are less stable and do not obey the octet rule.

Elements like hydrogen, lithium, helium do not obey the octet rule because they can only lose or gain one electron to become stable

5.5. WHY ATOM FORM CHEMICAL BOND.

All neutral atoms are unstable, except for the noble gases in group 8. All neutral atoms, except the noble gases, form bonds in order to become stable and try to reach the most stable state (having lowest energy) that they can. Atom gets stable when their valence shell is filled with electron or they satisfy the octate rule.

Here are some important things to remember about bonding:

1. Only the outermost electrons, called valence electrons, join in bonding. The inner electrons stay out of it. Every atom wants to be like a noble gas, which has a full outer shell, because that makes it super stable.
2. The first shell can hold up to 2 electrons, the second shell up to 8, and for the first 20 elements, the third shell can also hold up to 8.
3. When atoms bond, they can do it in different ways. In an ionic bond, one atom gives away an electron to another atom, making them both stable.
4. In a covalent bond, atoms share electrons, which helps them reach maximum stability. And there are other types of bonds too.

We can use the periodic table to figure out how atoms might bond. On the right side of the table are the noble gases, like helium and neon. They're already super stable because they have full outer shells, so they don't usually bond with other atoms.

5.6. FORMATION OF IONIC BONDS

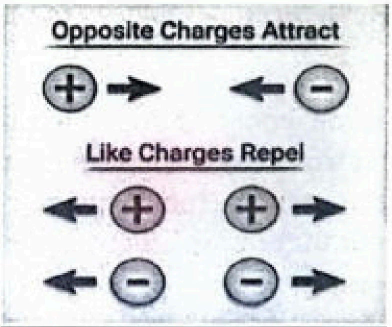
Ionic bonds are formed between ions, which are atoms with a charge. When a metal loses an electron, it becomes a positively charged ion called a cation. And when a non-metal gains that electron, it becomes a negatively charged ion called an anion. This exchange makes both atoms feel more stable.

Ionic compounds form solid structures with high melting points, and they can dissolve in water.

For example, sodium and fluorine combine to make sodium fluoride, NaF. Sodium loses its electron to fluorine, which gains it. The oppositely charged ions are attracted to each other, forming an ionic bond.

A **cation** is formed when a metal ion loses a valence electron while an **anion** is formed when a non-metal gains a valence electron. They both achieve a more stable electronic configuration through this exchange.

Ionic solids form crystalline lattices, or repeating patterns of atoms, with high melting points, and are typically soluble in water.



Key Terms
cation: A positively charged ion.
anion: A negatively charged ion.

How ions are formed

Ions are formed when the atoms lose or gain electrons in order to fulfil the octet rule and have full outer valence electron shell.

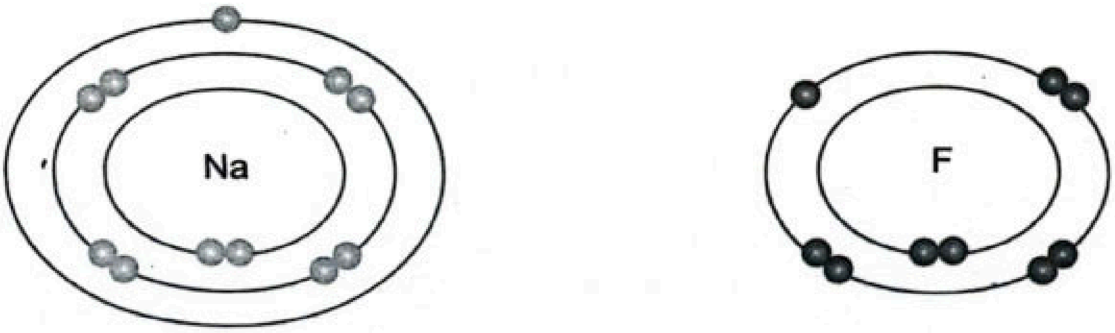


Fig. 5.3: Formation of Sodium F

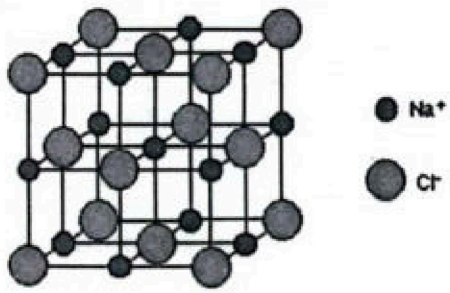


Fig. 5.4: Sodium Chloride Lattice

Ionic compounds are special because they form neat structures called lattices. These compounds are solid and look like crystals most of the time. They're also tough to melt because they have high melting points.

When you put these compounds in water, most of them dissolve easily, and the dissolved ions can conduct electricity. Take table salt, NaCl, for example. When it dissolves in water, it breaks into sodium and chloride ions, which can conduct electricity.

The charges on these ions are decided by how many electrons they lose or gain to become stable. And when they combine to make a compound, they do it to balance out the charges.

In the names and formulas of ionic compounds, the cation comes first, followed by the anion.

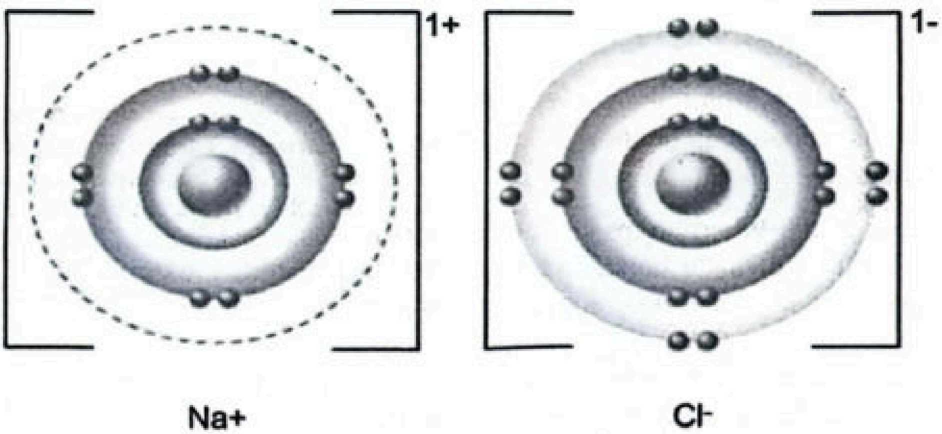


Fig. 5.5: Formation of Sodium Chloride

The ionic compound comprises ions.
These ions are held together by attractive forces among the opposite charges.
An example of an ionic compound is Sodium chloride.

5.7. PROPERTIES OF IONIC COMPOUND

1. Ionic compounds have high melting as well as boiling points.
2. They are hard and brittle in nature.
3. They are good **insulators** when solid
4. They conduct electricity when dissolved in water.
5. Ionic compounds are soluble in Water.
6. Ionic compounds have unique properties. They all form crystals.

5.8. FORMATION OF COVALENT BONDS

When non-metal elements come together, they can form covalent bonds. This happens when they share pairs of electrons. They do this because they want to have the same kind of electron setup as noble gases, which makes them stable. Unlike ionic bonds, no ions are made in covalent bonding.

Interesting fact

Covalent bonds are like a partnership in a competition - electrons are shared, an atoms perform a synchronised routine to create a stable molecule.

These bonds are held together by something called Van der Waals forces, which are pretty weak and depend on how close the atoms are to each other. These forces disappear quickly if the atoms move apart.

Atoms in molecules stick together because the nucleus and the electrons attract each other.

There are three types of covalent bonds: single, double, and triple. The simplest example is a hydrogen molecule. A single hydrogen atom is unstable because it only has one electron. But when two hydrogen atoms share their electrons, they both end up with two, which makes them stable. That's how they make a molecule of H2.

Types of Covalent Bond

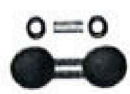
Covalent bonds are when atoms share electrons. There are different kinds depending on how many pairs of electrons they share.

Carbon is an incredible element. Arrange carbon atoms in one way, and they become soft, graphite. In other arrangement, the atoms form diamond, one of the hardest materials in the world.

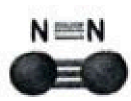
1. Single Covalent Bond: When atoms share one pair of electrons, it's called a single bond. Examples are H_2 (hydrogen gas), Cl_2 (chlorine gas), and CH_4 (methane gas).
2. Double Covalent Bond: If they share two pairs of electrons, it's called a double bond. Examples include O_2 (oxygen gas) and CO_2 (carbon dioxide).
3. Triple Covalent Bond: When they share three pairs of electrons, it's called a triple bond. Examples are N2 (nitrogen gas) and C_2H_2 (acetylene).

There are also two other types of covalent bonds based on whether the electrons are shared equally or not:

Single bond



Double bond



Triple bond

Polar Covalent Bond: If the atoms have a big difference in how much they attract electrons, the bond can be polar. One atom becomes slightly negative and the other slightly positive. Examples are H_2O (water) and HCl (hydrochloric acid).

Nonpolar Covalent Bond: When there's no difference in how much the atoms attract electrons; the bond is nonpolar. Electrons are shared equally. Examples include H_2 (hydrogen gas) and O_2 (oxygen gas).

These different types of bonds help atoms stick together in different ways!

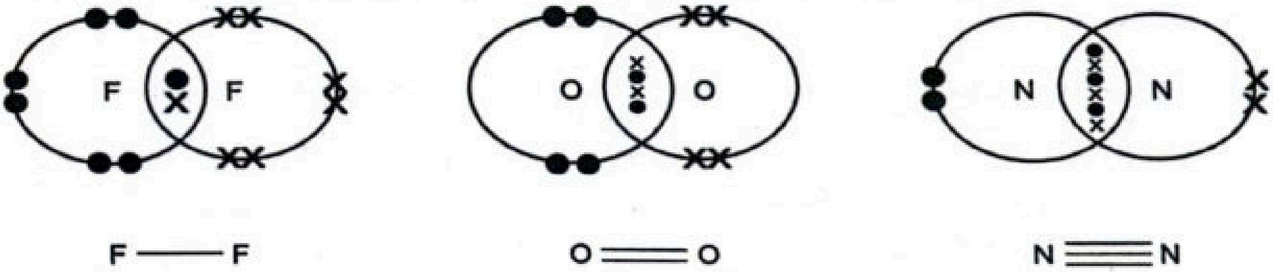


Fig. 5.6: Examples of Non Polar Molecules

5.9. PROPERTIES OF COVALENT COMPOUNDS:

1. The covalent compounds exist as gases or liquids or soft solids.
2. The melting and boiling points of covalent compounds are generally low.
3. Covalent compounds are insoluble in water but dissolve in organic solvents.
4. They are non-conductors of electricity in solid, molten, or aqueous state.
5. They are soft and flexible.

5.10. POLAR AND NON POLAR BONDS

Some molecules are made of only one type of element. These are called **homonuclear molecules.** The most common ones are diatomic, which means they have two atoms of the same element. Examples include hydrogen (H_2), oxygen (O_2), nitrogen (N_2), and the halogens. There are also triatomic homonuclear molecules like ozone (O_3), and even tetratomic ones like arsenic (As_4) and phosphorus (P_4).

On the other hand, there are molecules made of different elements. These are called **heteronuclear molecules.** One example is hydrochloric acid (HCl). When atoms of different elements bond together, the molecule might become polar because of their

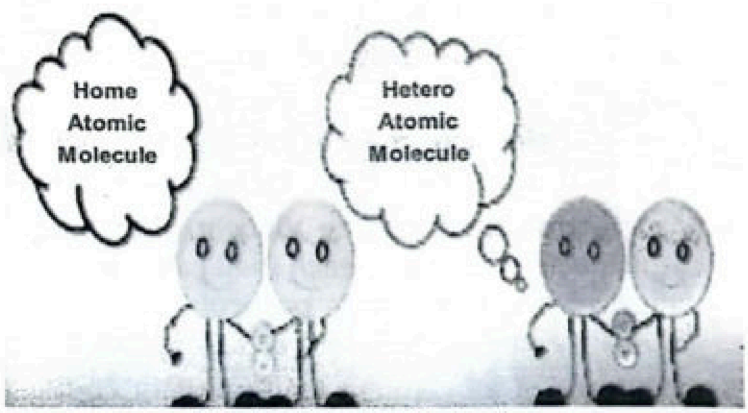


Fig. 5.7: Types of Molecules

Funny Fact

Homonuclear molecules are like identical twins they are made-up of the atoms of the same element. heteronuclear molecules are atoms of the different elements come together form unique compounds creating a molecule for example oxygen and hydrogen team up to form water.

differences in electronegativity. These molecules can have different numbers of atoms, like diatomic, tri-atomic, tetra-atomic, and even more.

5.11. DIFFERENCE BETWEEN COVALENT AND IONIC COMPOUNDS

	Covalent compounds	Ionic compounds
1	Covalent compounds are formed by sharing of electrons between 2 atoms	Ionic compounds are formed by complete transfer of electrons.
2	They exist in all three states solid liquid and gases.	They mostly exist in solid-state.
3	Covalent compounds have low melting and boiling points.	They have higher melting and boiling point.
4	They are generally insoluble in water.	They are soluble in water.
5	They are do not conduct electricity in the molten state as well as in aqueous solution.	onic compounds are good conductor of electricity in the molten state as well as in aqueous solution.
6	Weak van der Waal forces (intermolecular forces) are there.	Strong bonds with electrostatic forces of attraction.
7	Electron orbits overlap.	Electron orbits are separate.
8	Covalent bonds are formed among nonmetals.	Ionic bonds are formed between metals and nonmetals.
9	Relatively soft.	Hard and Brittle.

SUMMARY

1. A chemical bond is a connection between atoms.
2. Atoms can achieve stable configuration by forming chemical bonds with other atoms.
3. Only the electrons present in the outer shell, also known as the valence electrons, take part in the formation of chemical bonds.
4. An electron shell may be thought of as an orbit followed by electrons around an atom nucleus. Because each shell can contain only a fixed number of electrons, each shell is associated with a particular range of electron energy, and thus each shell must fill completely before electrons can be added to an outer shell.
5. The atom of an element having duplet configuration has only two electrons in the outermost shell, whereas the atom of an element having octet configuration has 8 electrons in its outermost shell.
6. Ionic bond is the electrostatic attraction between the positively charged nuclei of atoms and the electrons.
7. Ionic compounds have high melting and boiling points, are brittle in nature, have a lattice structure and can conduct electricity when in solution.

8. Ionic solids are arranged in a crystal lattice structure.
9. Covalent bonds form when electrons are shared between atoms and are attracted by the nuclei of both atoms.
10. In nonpolar covalent bonds, the electrons are shared equally. In polar covalent bonds, the electrons are shared unequally, as one atom exerts a stronger force of attraction on the electrons than the other.
11. Homo-atomic molecules or homonuclear molecules are molecules composed of only one type of element. The most familiar homonuclear molecules are diatomic, meaning they consist of two atoms, though not all diatomic molecules are homonuclear.
12. Hetero-atomic molecules or heteronuclear molecules, are molecules composed of more than one type of element, for example, HCl.

EXERCISE

Section I: Multiple Choice Questions

Select the correct answer:

1. The region around the nucleus which can be filled with one or two electrons is
   A) field B) axis C) zone D) atomic orbital
2. The electronic configuration of atom which possess lowest energy is the most
   A) stable B) unstable C) transitional D) ductile
3. The reason of bonding in ionic compounds is -------------
   A) Sharing of electrons B) Repulsion force
   C) Attraction force D) None of these
4. Ionic compounds have______ melting points due to ______ Ionic bonds.
   A) High, weak B) High, strong
   C) low, strong D) low, weak
5. A gas having double bonds is:
   A) Carbon monoxide B) Oxygen
   C) Carbon dioxide D) None of these
6. Aluminium tends to lose:
   A) 2 electron B) 4 electron
   C) 1 electron D) 3 electron
7. In covalent compounds, the bond is formed due to the..................
   A) sharing of electrons B) donation of electrons
   C) high electronegativity of atoms D) high electron affinity of atoms

8. A molecule of........ Contains a triple bond
   A) sulphur B) nitrogen
   C) carbon D) oxygen
9. Identify the molecule with a single covalent bond.
   A) CO_2 B) CO C) Cl_2 D) N_2
10. A polar covalent bond will be formed in which one of these pair of atoms:
   A) HF B) H_2 C) Cl_2 D) O_2

Section II: Short Response Questions

1. Explain why atoms form bonds.
2. Describe how a sodium atom gains a stable outer shell.
3. Analyze why the atoms of group 0 elements are unreactive.
4. Define the terms: 1. ions, 2. covalent bond, 3. ionic bond.
5. Explain why covalent compounds are generally gases, liquids, or soft solids.
6. List 5 examples of covalent bonds.
7. Identify four properties of covalent compounds.
8. Name the three types of covalent bonds.
9. Define a covalent bond and describe how it is formed.
10. Provide two examples of each: homoatomic molecules and heteroatomic molecules.
11. Explain what homoatomic molecules are.
12. Determine the charge of an atom that has lost 3 electrons.

Activity:
Students will explore and understand the different types of chemical bonding (ionic, covalent, and metallic) by constructing models and analyzing the properties of various compounds.
Materials Needed:
1. Colored balls or beads (representing different atoms, e.g., red for oxygen, white for hydrogen, yellow for sodium, green for chlorine, etc.)
2. Toothpicks or thin wires (to connect the atoms, representing bonds)
3. Periodic table (for reference)
4. Large chart paper or poster board
5. Markers
6. Labels or sticky notes

Water, salt, sugar, metal pieces (for observation of physical properties)

Section III: Extensive Response Questions

1. Predict the physical properties of covalent bonds and relate these properties to their structure and bonding.
2. Predict the physical properties of ionic bonds and relate these properties to their structure and bonding.

3. Describe the formation of covalent bonds between non-metallic elements, such as Calcium Chloride, oxygen, ammonia, and carbon dioxide, using dot and cross diagrams.
4. Discuss the features of covalent and ionic bonds.
5. Differentiate between the properties of covalent and ionic compounds.
6. Explain why ionic compounds are good conductors of electricity in aqueous solution or in a fused state.
7. Compare homoatomic and heteroatomic molecules.
8. Name the types of chemical bonds and explain which one is the strongest.
9. Determine whether $FeSO_4$ is a heteroatomic or homoatomic molecule.
10. Complete the following Tables.

Ionic compound	Positive Ion	Negative Ion	Formula
Potassium bromide			
	Na^+	F^-	
			MgO

Covalent molecule	Formula	Number and type of atoms presents
Hydrogen gas		
	CH_4	
		One Nitrogen and three Hydrogen