CHAPTER 3

Periodic Table and Periodicity of Properties

SHORT Q/A

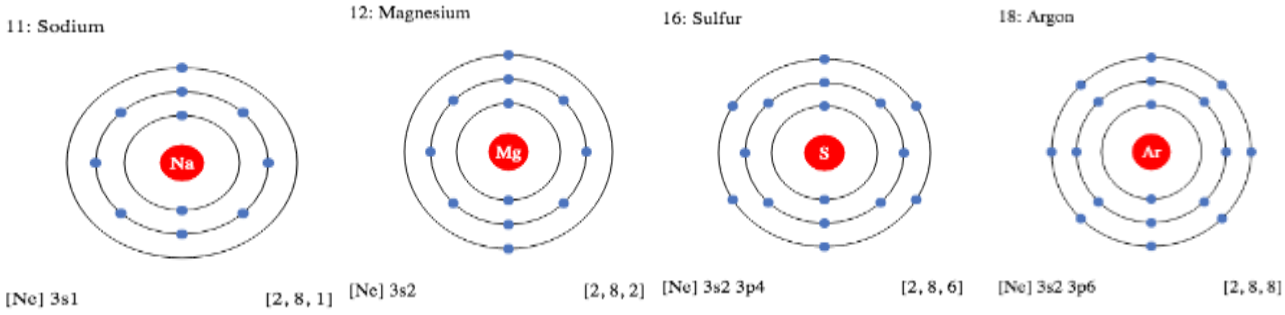
Q1. Which element in group 1A is not alkali metal and why?

Ans. Hydrogen of group 1A is not alkali metal because it does not form alkali when reacted with water as well it does not possesses the properties of metal such as ductility, conductivity, and shining surface. In addition Hydrogen is in gas form at room temperature.

Q2. Place the following elements in order of increasing ionization energy: Na, S, Mg, and Ar.

Ans. The increasing order of ionization energy of these elements are exhibited below:

$$Na < Mg < S < Ar$$



Q3. Name the group and state the group number of each of the following elements:

(a) K (b) Ne (c) Be (d) Cl (e) C

Ans.

K = Group 1A, Alkali metal

Ne = Group 8A, Noble Gas

Be = Group 2A, Alkaline Earth Metal

Cl =Group 7A, Halogen

C = Group 4A, Non-metal/ Carbon Family

Q4. Which element is the most electronegative among C, N, O, Br, and S? Which group does it belong to?

Ans. In the above elements C, N, O, Br, and S, Oxygen (O) is the most electronegative element because it belong to group 6A.

Q5. How do the first ionization energies of representative elements vary across a period and down a group?

Ans.

In periods:

The ionization energy of representative elements in periods from left to right increases due to the increase number of electrons in the outer most shell. The elements, which are on the extreme right of the period, have maximum ionization energy, and the elements which are on the extreme left of the period, have minimum ionization energy.

In Groups:

The ionization energy of representative elements in groups from top to bottom decreases due to the addition of shell. The elements at the top have maximum ionization energy, while the elements at the bottom have minimum ionization energy.

Q6. Which element is found in?

(a) Period2, Group 7 (b) Period 4, Group 3

(c) Period 5, Group 6 (d) Period 1, Group 8

Ans.

1. Period 2, Group 7 = Fluorine (F)
2. Period 4, Group 3 = Gallium (Ga)
3. Period 5, Group 6 = Tellurium (Te)
4. Period 1, Group 8 = Helium (He)

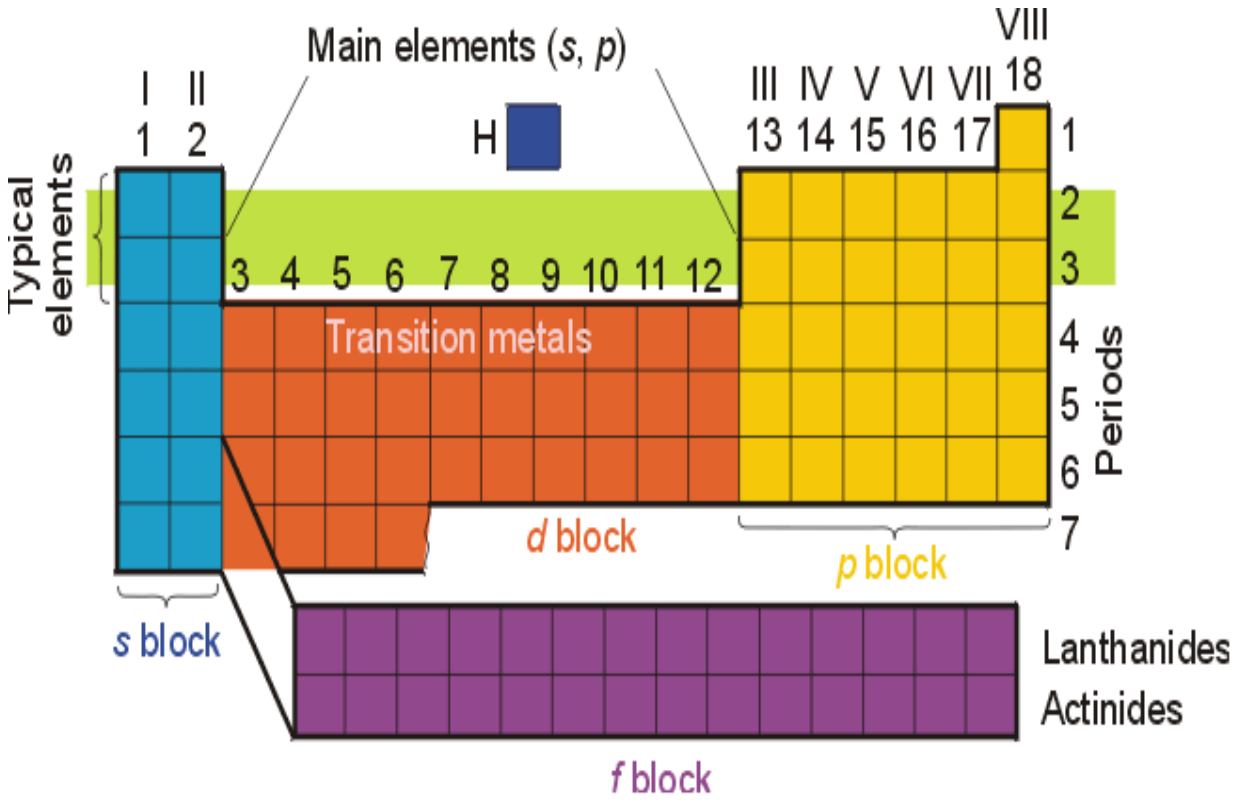
Q7. How you will differentiate between representative, and transition elements?

Ans.

Representative elements	Transition elements
• It belongs to Group A	• It belongs to Group B
• It has Fixed oxidation state	• It has variable oxidation sate
• It consists of metals, non-metals, and metalloid	• It consist of metal only
• It contain S, and P block elements	• It contains d, and f block elements

Q8. Make a general sketch of periodic table showing s, p, d, and f-block elements (without showing the symbols of the elements)

Ans.



Q9. Why the s-block elements have two groups only?

Ans. Elements in which the last electron enters the s-orbital are called s-block elements. Since s-subshell has only one orbital which can accommodate only two electrons, therefore there are only two groups of s-block elements, Group 1A elements and Group 2A elements.

Q10. What type of elements is Sulphur (S), a representative element, a transition element or lanthanide element?

Ans. We know that Sulphur belongs to Group A element, and Group A element exhibits representative elements, which means Sulphur (S) is a representative element.

Long Q/A

Q1. How modern periodic table is different from the Mendeleev's periodic table?

Ans.

MENDELEEV VS MODERN PERIODIC TABLE

Mendeleev's periodic table is, based on relation of properties of elements as dependent on the atomic weight of element. But Modern periodic table considers atomic number as the fundamental property that decides the properties of elements.

Modern periodic table does correct defects of the Mendeleev's periodic table. For example, in the Mendeleev's periodic table, in the element pairs, Argon-potassium, cobalt-nickel, tellurium-iodine and thorium and protactinium, elements with higher atomic mass precedes the element with lower atomic weight. Though it is the right places for them but is against the Mendeleev's periodic law.

These elements atomic number shows the reverse order compared to atomic mass. The supposed to be wrong positions in Mendeleev's table has the right explanation justifying their positions.

Uneven grouping of elements:

In Mendeleev's periodic table, coinage metals of copper, silver, and gold are, grouped together with very active alkali metals. Manganese metal was, grouped with halogens in the seventh group. The defects are rectified in the Modern periodic table.

Position of isotopes:

Isotopes with higher atomic weights could not, be accommodated in the Mendeleev's table. As isotopes have same atomic number with the stale atom they do not need any separate location in the modern periodic table.

No reasons were, offered for the triad elements of viii group. No such special grouping is given in the modern table.

Mendeleev's periodic table was for the arrangement of sixty-three elements known at that time.

Modern table accommodates all the 118- natural and synthetic elements.

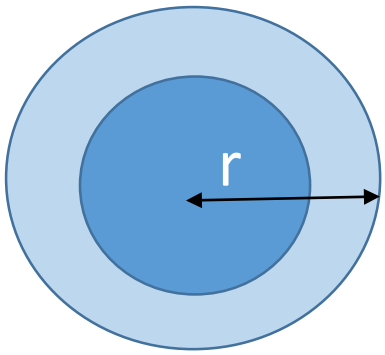
Atomic number is much fundamental property distinguishing each element and hence a better basic nature to decide the physical and chemical properties of elemental atoms than atomic weights.

Q2. Differentiate between atomic radii, and covalent radii. Explain the trends of atomic radius in the periodic table.

Ans.

Atomic Radii:

The distance between nucleus, and the valence shell (outer shell) of the atom is termed as atomic radius or radii. The atomic radius is represented by "r". The atomic radius is directly proportional with the number of shells. The atomic radius are expressed in the nanometer (1.0×10^{-9}) or the Pico meter (pm = 1.0×10^{-12}m).



Atomic Radius

Trends of Atomic radius in the periodic table:

a. Trends of Atomic radius in groups:

The atomic radius in periodic table increases from top to bottom, due to the addition of new shells in successive periods, and shielding effects of electrons

b. Trends of Atomic radius in periods:

The atomic radius in periods from left to right decreases, due to the addition of electrons in the same shell. As the number of electrons increases, the number of protons also increases, which increases the nuclear pull on the electrons, and attracts them nearer to the nucleus. Thus, the atomic radius decreases.

Covalent radii:

The one half of the distance between the nuclei of two similar atoms of the same molecule containing a single covalent bond. Therefore, the bond distance between the two atoms 'A', and 'B' is the average of the length 'A-A', and 'B-B'.

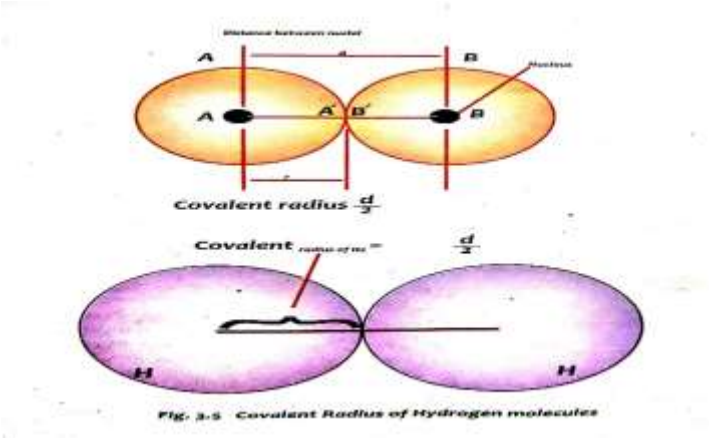


Fig. 3.5 Covalent Radius of Hydrogen molecules

Q3. What is electronegativity? Identify the most and least electronegative groups of elements in the periodic table. Why fluorine special in terms of electronegativity?

Ans. Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. In other words, we can say that the power of attraction for the shared pair of electron is called electronegativity.

Electronegativity is represented by E. N.

If two atoms have the same ability to attract shared pair of electron; we can say that they have the same electronegativity.

Pauling calculated the electronegativity values of the elements. For this purpose, he developed a scale from bond energies of diatomic molecules. On this scale the electronegativity of Fluorine is 4.0, which is the highest of all the periodic table

elements, and values range down to cesium and francium which are the least electronegative at 0.7.

Fluorine is the most electronegative element because it has 5 electrons in it's 2P shell. The optimal electron configuration of the 2P orbital contains 6 electrons, so since Fluorine is so close to ideal electron configuration, the electrons are held very tightly to the nucleus.
Electron Configuration: $1s^2 2s^2 2p^5$

Electronegativity: 3.98

Group: 17 (Halogens)

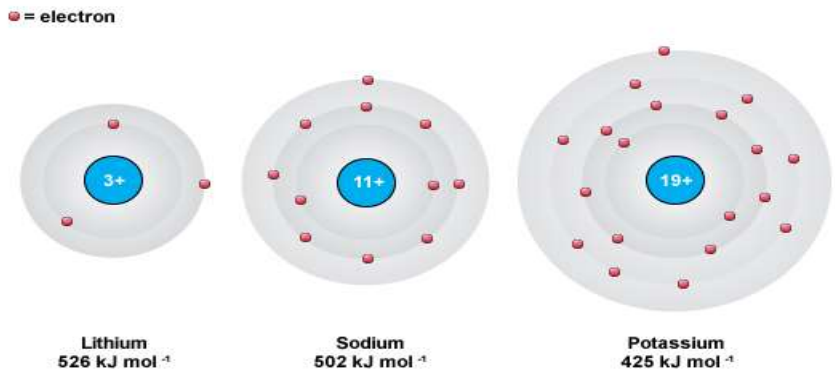
Atomic Number: 9

Q4. Define shielding effect, and draw it affects the ionization energy, electron affinity, and electronegativity.

Ans.

Shielding effect:

Electrons in an atom can shield each other from the pull of the nucleus. This effect, called the shielding effect, describes the decrease in attraction between an electron and the nucleus in any atom with more than one electron shell. The more electron shells there are, the greater the shielding effect experienced by the outermost electrons.



As you can see in the above figure, we have Lithium, Sodium and Potassium. Li possesses 2 shells, Na possesses 3 shells, and K possesses 4 shells, which means K possesses more shells than the rest of the elements, so its shielding effect will also be higher than the rest two. Hence in result, shielding effect is directly proportional

to the number of shells, as no. of shells increases the shielding effect also increases.

Ionization Energy:

Ionization Energy Trends in the Periodic Table. The ionization energy of an atom is the amount of energy required to remove an electron from the gaseous form of that atom or ion.

As you move down the group, the ionization energy decreases due to the size of atomic radius. The element at top possesses maximum ionization energy, while the elements at bottom possesses minimum ionization energy. However in periods, the ionization energy increases from left to right. The elements on extreme right possesses high ionization energy, while the elements on the extreme left possesses low ionization energy.

Electron affinity:

The amount of energy released when an electron is added to gaseous atom of an element in its outermost shell to form an anion is called electron affinity. It is represented by E.A. and expressed in KJ. Mol^{-1}.

In groups, as you move down the electron affinity decreases due to the addition of shells. The elements on top possesses maximum electron affinity, while the elements at the bottom possesses minimum electron affinity. However, in periods the electron affinity from left to right increases due to increase in nuclear charge, and decrease in the atomic radius.

Electronegativity:

The ability of an atom to attract the shared pair of electrons towards itself in a covalent bond is called electronegativity. In other words, we can say that the power of attraction for the shared pair of electron is called electronegativity.

There are various factors that affects electronegativity:

- Nuclear Charge
- Atomic radius
- Shielding effect
- Electronic configuration

In groups, from top to bottom the electronegativity decreases, due to previous mentioned factors, and in periods from left to right the electronegativity increases due to addition of electrons, and increase in nuclear charge.

Q5. Explain the following terms.

a. Periodicity of properties

In the context of chemistry and the periodic table, periodicity refers to trends or recurring variations in element properties with increasing atomic number. Periodicity is caused by regular and predictable variations in element atomic structure.

Mendeleev organized elements according to recurring properties to make a periodic table of elements. Elements within a group (column) display similar characteristics. The rows in the periodic table (the periods) reflect the filling of electrons shells around the nucleus, so when a new row begins, the elements stack on top of each other with similar properties. For example, helium and neon are both fairly unreactive gases that glow when an electric current is passed through them. Lithium and sodium both have a +1 oxidation state and are reactive, shiny metals.

b. Electron affinity

This is a measure of readily an atom accepts an electron. Electron affinity increases moving across a period and decreases moving down a group. Nonmetals usually have higher electron affinities than metals. The noble gases are an exception to the trend since these elements have filled electron valence shells and electron affinity values approaching zero. However, the behavior of the noble gases is periodic. In other words, even though an element group might break a trend, the elements within the group display periodic properties.

c. Modern periodic law

The modern periodic table is developed after the periodic law and a periodic table given by Mendeleev. In the latter part of the 18th century, Mendeleev made his periodic table. Scientists did not know about the internal structure of the atom back then.

The development of various atomic models and advances in quantum theory revealed that the atomic number is the most basic property of a chemical element.

This led to the modification of Mendeleev's periodic law, which is today called as modern periodic law.

The atomic number is equal to the number of electrons or protons in a neutral atom. After knowing the fundamental unit of elements, scientists now had a clear idea about quantum numbers and electronic configuration of elements in the periodic table. After knowing the periodic law, chemists noticed that there is an analogy between the 94 naturally occurring chemical elements. This analogy made people more curious about the chemistry of these elements. Scientists made various artificial elements. A new periodic table was developed based on the modern periodic law by modifying the Mendeleev's periodic table.

Choose the correct option:

1(c), 2(d), 3(a), 4(d), 5(a), 6(d), 7(b), 8(c), 9(a), 10(d)

Best of luck

Reference: A text book of Chemistry Grade IX. Khyber Pakhtunkhwa Textbook Board Peshawar.

Chapter 04

Structure of molecules

Short Q/A

Q1. What is electron-sea model of metallic bonding?

Ans. **Metallic bond:**

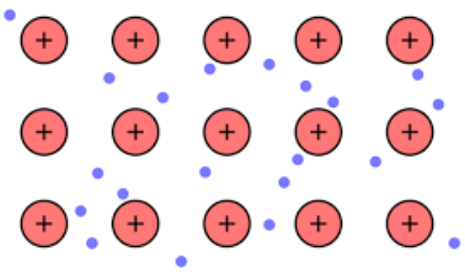
The chemical bonding that results from the attraction between metal positive ions and the surrounding sea of electrons is called metallic bonding.

The metallic bond can be explained on the basis of electron sea theory or electron gas theory:

Electron-sea model:

The nucleus of metal atom cannot hold the valence electron due to low ionization potential. These electrons move freely in vacant spaces present between the atoms. When atoms attract all electrons collectively and bound them together, a metal appears to have a sea of electrons in which all the nucleus of atoms are submerged. It results in the formation of metallic bond.

Metallic bond is neither ionic nor covalent because it is formed neither due to transfer nor due to sharing of valance electron between atoms.



Q2. Why most atoms are chemically bonded to other atoms in nature?

Ans. Every system in Universe tends to lower its energy in order to attain stability. Water flows from the higher level to the lower level. Similarly, electricity flows from higher potential to lower potential and the heat flows from a hot body to a cold body.

This happens because both water and electricity are trying to decrease their energy. Atoms in the same way have a tendency to decrease their energy.

For example:

The energy of the isolated hydrogen atom is higher than the energy of two hydrogen atoms because the combinations of atoms gives a stable molecule through emission of energy.

Atoms can decrease their energy by combining with the atoms and forms a chemical bond.

Q3. Identify and define the four major types of chemical bonding.

Ans. **Chemical Bond:**

The attractive force which holds the atom together to form a molecule or a compound is called chemical bond.

The four major types of chemical bonding are as follow:

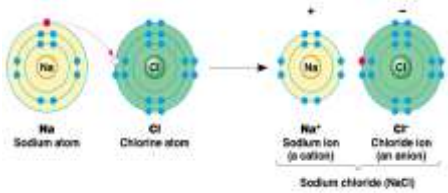
Types of chemical bond:

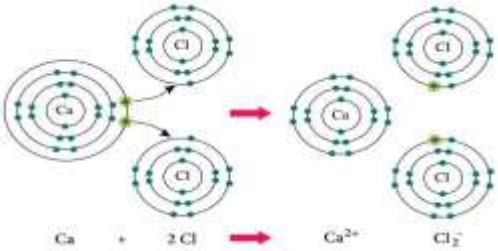
1. Ionic bond
2. Covalent bond
3. Co-ordinate or Dative bond
4. Metallic bond

Ionic bond

The bond which is formed by the complete transfer of one or more electrons from one atom to another is called ionic bond or electro covalent bond.

Examples: NaCl and $CaCl_2$

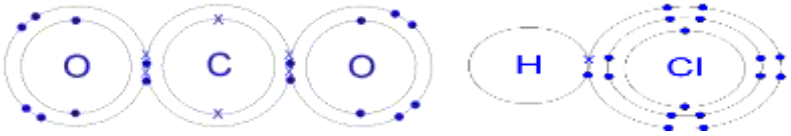




Covalent bond:

The chemical bond formed when two atoms are joined together by sharing of electrons or the bond which is formed by the mutual sharing of electrons is called covalent bond.

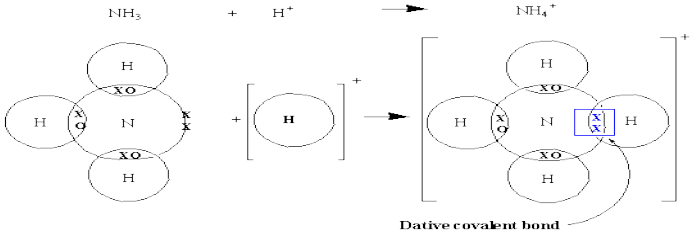
Examples: CO_2 and HCl



Co-ordinate covalent bond

The covalent bond which only one atom donates the shared pair of electrons is called the co-ordinate covalent bond or dative bond.

Example: $(NH_4)^+$



Metallic bond

The chemical bonding that results from the attraction between the metal positive ions and surroundings sea of electrons is called metallic bonding.

Example: Aluminum (Al) and Gold (Au).

Q4. Arrange the following from the strongest to the weakest attraction:

a. **Covalent bond**
b. **Dipole-dipole interaction**
c. **Hydrogen bond**
d. **Ionic bond**

Ans. **Arrangement from strongest to weakest**:

Ionic Bond > Covalent bond > Hydrogen bond > Dipole-dipole interaction.

a. **Ionic bond:** Ionic bond has the strongest attraction among atoms $_{(ions)}$ than other bonding. The attraction between the atoms $_{(ions)}$ is very strong because one of the atoms is positively charged while the other is negatively charged.
b. **Covalent bond:** Covalent bond is stronger than other bonds but weaker than ionic bond because it is formed by the mutual sharing of electrons among two atoms and these atoms are held together.
c. **Hydrogen bond:** Hydrogen bond is weaker than ionic and covalent bond but stronger than dipole-dipole interaction. Although bond is an attractive force between the higher electron deficient hydrogen atom and nearby highly electronegative atoms with lone pair of electrons such as F, O or N.
d. **Dipole-Dipole Interaction:** Dipole-dipole interaction is the weakest among all bonds. The attractive forces between the positive pole of one polar

molecule and negative pole of other polar molecule are called dipole-dipole interaction.

Q5. Why ionic compound are good electrolyte in molten and solution form and not in solid state?

Ans. Ionic compounds are good electrolytes in molten and solution form but not in solid state because.

- Electrical conductivity needs the movement of charged particles$_{(ions)}$
- In solid state, electrostatic force holds the ions together in crystals lattice. So, they are not free to move.
- In molten and solution form the ionic compound dissociates into ions which move freely and conduct electricity.

Q6. What type of elements tend to form the following type of bonding?

a. **Ionic**
b. **Covalent**
c. **Metallic**

Ans. The type of elements tend to form the following type of bonding are as follow:

a. **Ionic bond:** This type of bonding is always formed between a metal and non-metal. A metal always loses electrons to form a cation and non-metal gains electrons to form anion. The cation and anion attract each other and form an ionic bond.

   Example: Na^{+}———— Cl^{-}

   K^{+}———— Br^{-}

b. **Covalent bond:** This type of bond is formed between two non-metals by sharing of electrons. Sometimes it is formed between two similar elements and sometimes between different elements.

   Example: H——— Cl

   H_2, Cl_2 et.c

c. **Metallic bond:** This type of bond is between metal atoms. They are neither ionic nor covalent because they are neither formed due to transfer nor sharing of electrons. Atoms in metal are bonded with each other in such a way every atoms are attracted from all sides by other atoms.

   Example: Silver (Ag) and Tin (Sn)

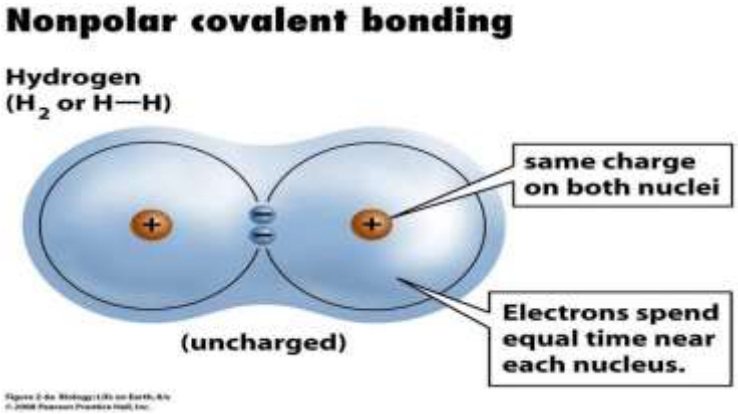
Q7.Give an example of non-polar bond? Give reason.

Ans. **Non-polar bond:**

The bond which is formed by the mutual sharing of electrons between atoms, having similar electro negativities is called non-polar bond or non-polar covalent bond.

Example.

An example of a non-polar covalent bond is the bond between two hydrogen atoms because they equally share the electrons. Also they have same electronegativity.



Q8. Predict the bond type ionic, polar, non-polar covalent bond in each of the following.

a. **$CaCl_2$**
b. **H_2O**
c. **CO_2**
d. **C_2H_4**

Ans. **$CaCl_2$:** Contains ionic bond.

Ionic bond is present in Calcium Chloride. Calcium loses two electrons and form C^{+2} ion. Two chloride atoms gain electrons and form Cl^{-2} ions.

Both ions attract each other and form an ionic bond.

H_2O: Contains polar covalent bond.

Polar Covalent bond is present in water Hydrogen is less electro negative than Oxygen that's why the shared pair of electron is attracted more towards Oxygen and it becomes partial negative while Hydrogen becomes partial positive and forms polar covalent bond.

CO_2: Contains polar covalent bond.

Polar covalent bond is formed in Carbon-Dioxide (CO_2). Oxygen is more electronegative than Carbon. Therefore, shared pair of electron is attracted more towards Oxygen than Carbon, due to which Oxygen becomes partial negative to form polar covalent bond.

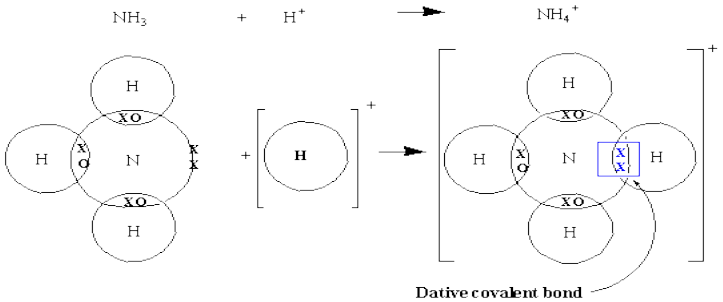
C_2H_4: Contains non-polar covalent bond.

Non-polar covalent bond is present in ethane. Carbon and Hydrogen have nearly similar electronegativity.

So, C-H bond is considered as non-polar covalent bond.

Q9: Give the reason why Dative bond is always polar?

Ans. In co-ordinate covalent bond the shared pair of electrons is donated by one of the bonded atoms. The atom which donates pair of electrons called donor atom and possesses partial positive charge. The atom which accepts pairs of electron is called acceptor and possesses partial negative charge. As the two bonded atoms consists of positive and negative poles so it is said to be polar covalent bond.



Q10. Why ionic compounds are good conductor as compared to covalent compounds?

Ans. Ionic compounds are good conductor as compared to covalent compounds because:

Ionic compounds:

Ionic compounds have ions which are free to move when melted or dissolved in water. This free movement of ions is responsible for good conductivity of ionic compounds.

Covalent compounds:

Covalent compounds have bonds where electrons are shared between atoms. Due to the sharing of electrons, they exhibit characteristic physical properties that

include lower melting points and electrical conductivity compared to ionic compounds.

Q11. Define Hydrogen bond, also explain properties of hydrogen bond.

Ans. **Hydrogen bond:**

A hydrogen bond is the attractive force between the highly electron deficient hydrogen atom and nearby highly electro negative atom with lone pair of electrons such as F, O or N.

Properties of hydrogen bond:

1. Hydrogen bond is stronger than dipole-dipole forces but weaker than covalent bond. It is about twenty times weaker than covalent bond and ten times stronger than dipole-dipole interaction.
2. Hydrogen bond is directional.
3. Hydrogen bond forms long chain and helps in the formation of network of molecules.

Long Q/A

Q1. Define the octet rule in term of noble gas, configuration and stability.

Ans. **The octet rule**: refers to the tendency of atoms to prefer to have eight electrons in the *valence shell* (outer orbital). When atoms have fewer than eight electrons, they tend to react and form more stable compounds.

Noble gases or Inert gases:

The elements in group 8 of the periodic table such as: Helium (He), Neon (Ne) and Argon (Ar) are known as noble gases.

They are also called inert gases because they are very stable and rarely take part in chemical reactions to forma compound. Their stability is due to their completely filled outer most shells.

Octet rule in terms of Noble gas, configuration and stability:

All shells with eight electrons are called an octet and are very stable.

Except for Helium, that has two electrons in its shell, all other noble gases and their outer most shells are filled with eight electrons.

Thus, when atoms take part in a chemical reaction, they tend to combine in ways to complete eight electrons in their outer most shells, to attain electronic configurations of the noble gases.

Q2: (a) What is the main distinguish between ionic and covalent bonding?

Ans.

Ionic Bond	Covalent Bond
• It is formed by the complete transfer of electrons	• It is formed by the sharing of electrons.
• Ionic bond has electronegativity difference greater than 1.7	• Covalent bond has electronegativity difference is less than 1.7
• Compound of ionic bond are good conductor in molten state	• The compound of covalent bond are bad conductor
• Compound of ionic bond are hard	• Compound of covalent bond are soft
• Ionic bond is strong	• Covalent bond is weak
Example, NaCl, $CaCl_2$ etc	Example, HCl, O_2 etc

(b). How is E.N used in determining the ionic or covalent character of the bonding between two elements?

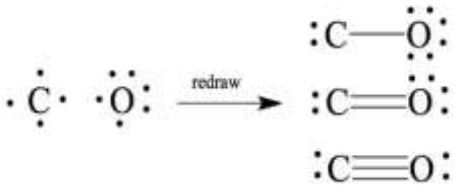
Ans. Sharing of electrons in the formations of covalent bonds and the complete transfer of electrons from one atom to the other in the formation of ionic bonds depend upon the difference in electronegativity of the bonding atoms.

If this difference is more than 1.7, electron transfer will occur and the ionic bond will form. While if their difference is less than 1.7, electrons will be shared and covalent bond will form.

Q3. Draw the Lewis structure of the following?

a. **CO**
b. **HCl**
c. **SO_2**
d. **CCl_4**
e. **BF_3**
f. **NH_3**

Ans.



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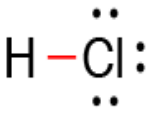
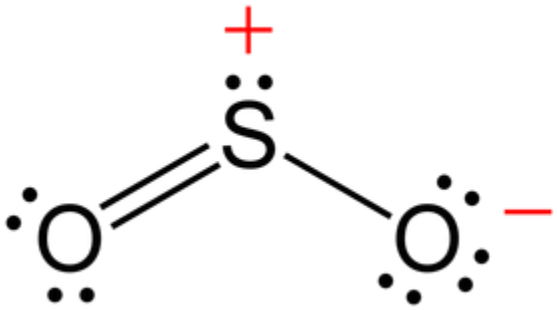
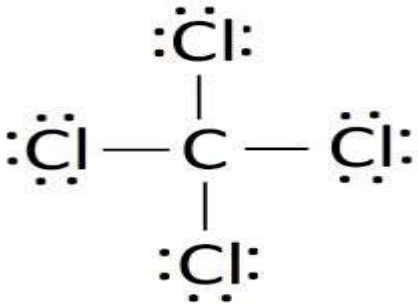
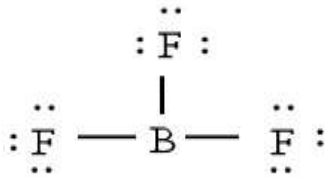
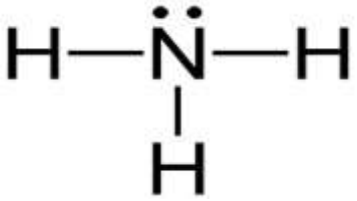










Q4. (a).What is the meaning of term polar, as applied to chemical bonding?

(b). Distinguish between polar and non-polar covalent bonds.

Ans. (a). A **polar** bond is a covalent bond between two atoms where the electrons forming the bond are unequally distributed. This causes the molecule to have a slight electrical dipole moment where one end is slightly positive and the other is slightly negative.

The more electronegative will be partial negative charged and the less electronegative atom will be partial positive charged.

Example:

In case of HCl, Hydrogen is less electronegative than Chlorine. So, it is polar molecule and has a polar bond.

Ans. (b). **Polar covalent bond:**

1. The covalent bond formed by the mutual sharing of electrons between atoms having different electro negativities is called polar covalent bond.
2. Both atoms do not attract the shared pair of electron equally.
3. Opposite poles are created in bonds.
4. One atom is partially positive charged and other is partially negative charged.
5. HCl has a polar covalent bond.

Non-Polar covalent bond:

1. The covalent bond formed by the mutual sharing of electrons between two atoms having same electro negativities is called non-polar covalent bond.
2. Both atoms attract shared pair of electrons equally.
3. Opposite poles are not created in these bonds.
4. No partial positive or negative charge appears on the atoms.
5. H_2 has a non-polar covalent bond.

Q5. Explain why most metals are malleable and ductile but ionic crystals are not?

Ans. Metals are malleable, which means they can be beaten into sheets and foils.

Metals are ductile, which means they can be drawn into wires. Metals are malleable and ductile but ionic crystals are not because:

In metals, the metallic bond is present which has weak intermolecular forces. Due to which electrons are free to move around.

Therefore, when we bend the metal, it can easily change its shape. On the other hand, ionic bond crystals are very strong because of their ionic bond. Due to strong bonding it is hard to break the bond, also difficult to change its shape.

Q6. Properties of ionic compounds, covalent compounds, and metallic compounds.

Ans. **Properties of Ionic compounds:**

1. Solid at room temperature.
2. Have sharp melting and boiling point.
3. Soluble in polar solvents like water.
4. Good electrolytes in molten or solution form.
5. Have reactions in molten state or in solution form.
6. Compounds are composed of cations and anions in crystalline form.

Properties of Covalent compound:

1. Geometrical shape of molecules.
2. Polarity and intermolecular forces among molecules.
3. Bond type, whether single, double or triple.
4. Covalent compounds have low melting and boiling points.
5. The bonds in covalent compound are directional
6. Covalent compounds are non-electrolytes in their solution form.
7. The crystals of covalent compounds are composed of molecules.
8. Reactions of covalent compounds are slower than the ionic compounds.

Properties of Metallic compounds:

1. All metals are solid at room temperature and pressure except Mercury (Hg).

2. Metals are malleable; they can be beaten into sheets and foils.
3. Metals are good conductor of heat and electricity.
4. Metals are lustrous; they have shiny surfaces.
5. Metals are sonorous; they produce specific ringing sounds when struck.
6. Metals are ductile; they can be drawn into wires.

Q7. What are the types of covalent bond?

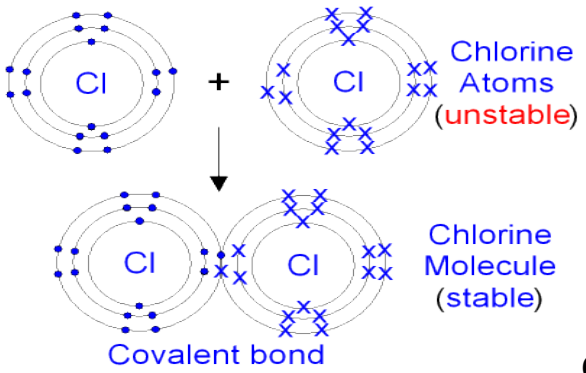
Ans. The types of covalent bond are as follow:

1. Single covalent bond
2. Double covalent bond
3. Triple covalent bond

Single covalent bond:

The bond in which two atoms share only one pair of electrons is known as a single covalent bonds.

For example: Cl_2



CL — CL

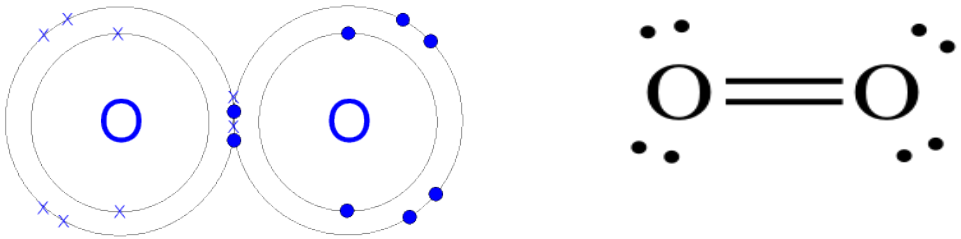
- Chlorine is formed from two chlorine atoms. The electronic configuration of Chlorine atom is (2, 8, 7), as it has seven electrons in its outer most shells or

valence shell. To attain stability it needs one electron to complete its valence shell, and for that chlorine makes single covalent bond with the chlorine atom to attain one electron each and make a compound of CL_2. As shown in the diagram.

Double Covalent bond:

The bond in which two atoms share two electrons each to form two pairs of electrons is called double covalent bond.

For example: O_2

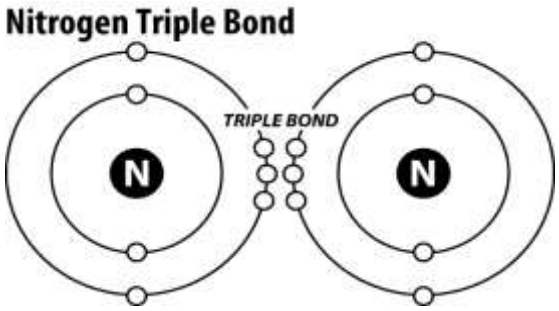


- Oxygen molecule is formed from two oxygen atoms. The electronic configuration of Oxygen atom is (2, 6). An oxygen atom has six electrons in its valence shell. To attain stability it needs two electrons to complete its outer most shell. For that its makes a double covalent bond with oxygen atom to attain two electrons each for to complete its valence shell. As shown in the figure above.

Triple Covalent bond:

The bond in which two atoms share three electrons each, to form three pairs of electrons is called triple covalent bond.

For example: N_2



$$N \equiv N$$

- Nitrogen molecule is formed from two nitrogen atoms. The electronic configuration of Nitrogen atom is (2, 5). A nitrogen atom has five electrons in its valance shell. To attain stability or to complete its outer most shell it needs three electrons. For that its makes triple covalent bond with nitrogen atom to achieve three electrons each to complete its valance shell. As shown in the figure above

Choose the correct answer.

1.(**d)** 2.(**c)** 3.(**c)** 4.(**b)** 5.(**c)** 6.(**c)** 7.(**b)** 8.(**c)** 9.(**d)** 10.(**b)**

Best of luck

Reference: A text book of Chemistry Grade IX. Khyber Pakhtunkhwa Textbook Board Peshawar.

Chapter: 05

Numerical Questions

1. **Calculate the initial volume of a sample of gas at 1.20 atm. If its volume is changed to 70.4 cm^3 as its pressure is changed to 3 atm. at constant temp.**

   Given data:

   Initial Pressure, P_1= 1.20 atm.

   Finial Pressure, P_2= 3atm.

   Final Volume, V_2= 70.4 cm^3

   Required:

   Initial Volume, V_1=?

   Formula:

   $P_1 V_1 = P_2 V_2$

   $V_1 = P_2 V_2 / P_1$

   Putting all the values:

   $$V_1 = \frac{3atm \times 70.4\ cm^3}{1.20atm}$$

   $$V_1 = \frac{211.2\ cm^3}{1.20}$$

   V_1 = 176 cm^3

   Initial volume, V_1 = 176 cm^3

2. **Calculate the final pressure of a sample of gas that change at constant temp 14.3 dm^3 from 7.55 dm^3 at 828 torr.**

   Give data:

   Initial Volume, V_1 = 7.55 dm^3

Initial Pressure, $P_1 = 828$ torr

Finial Volume, $V_2 = 14.3\ dm^3$

Required:

Final Pressure, $P_2 = ?$

Formula:

$P_1 V_1 = P_2 V_2$

$P_2 = P_1 V_1 / P_2$

Putting all the values

$$P_2 = \frac{828 \times 7.55}{14.3}$$

$P_2 = 437.16$ torr

3. **Calculate the final volume at 302k of 5.41 dm^3 sample of gas originally at 353k, if the pressure does not change.**

Given data:

Initial volume, $V_1 = 541\ dm^3$

Initial Temperature, $T_1 = 302k$

Final Temperature, $T_2 = 353k$

Required:

Final Volume, $V_2 = ?$

Formula:

$V_1 / T_1 = V_2 / T_2$

$V_2 = V_1 / T_1 \times T_2$

Putting all the values

$$V_2 = \frac{5.41 \times 353}{302}$$

$V_2 = 6.323\ dm^3$

4. **Calculate initial volume at 0^0 C of a simple of gas that is changed to 731 cm^3 by cooling -14^0 C at constant pressure.**

   Given data:

   Initial Temperature, $T_1 = 0 + 273 = 273k$

   Final Temperature, $T_2 = -14 + 273 = 259k$

   Final Volume, $V_2 = 731\ cm^3$

   Required:

   Initial Volume, $V_1 = ?$

   Formula:

   $V_1 / T_1 = V_2 / T_2$

   $V_1 = V_2 / T_2 \times T_1$

   Putting all the values

   $V_1 = \frac{731 \times 273}{259}$

   $V_1 = 770.513\ cm^3$

5. **A sample of a gas at room temperature occupies 0.80 dm^3 at 1.5 atm. What will be its volume when the pressure of a gas is raised to 2.1 atm?**

   Given data:

   Volume, $V_1 = 0.80\ dm^3$

   Pressure, $P_1 = 1.5$ atm

   Pressure, $P_2 = 2.1$ atm

   Required:

   Volume, $V_2 = ?$

   Formula:

   $P_1 V_1 = P_2 V_2$

   $V_2 = P_1 V_1 / P_2$

   Putting all the values

$V_2 = \frac{1.5 \times 0.80}{2.1}$

$V_2 = 0.57\ dm^3$

6. **Calculate the final volume at 319^{0c} of a sample of a gas original 5.13 dm^3 at 171^{0c}, if the pressure does not change.**

Given data:

Initial Volume, $V_1 = 5.43\ dm^3$

Initial Temperature, $T_1 = 319^{0c} = 319+273= 592k$

Final Temperature, $T_2 = 171^{0c} = 171+273= 444k$

Required:

Final Volume, $V_2 =?$

Formula:

$V_1 / T_1 = V_2 / T_2$

$V_2 = V_1 / T_1 \times T_2$

Putting all the values

$V_2 = \frac{5.13 \times 444}{592}$

$V_2 = 3.8475\ dm^3$

Or

$V_2 = 4\ dm^3$

Physical States of Matter

Short Q/A

Q1. Can you give reason why it takes longer time to cook at high altitude?

Ans. The reason why it takes longer time to cook at high altitude is as follow:

Reason:

Boiling point of a liquid depends upon the external pressure, with the increase in external pressure, the boiling point also increases. Similarly, the decrease in external pressure causes decrease in the boiling point.

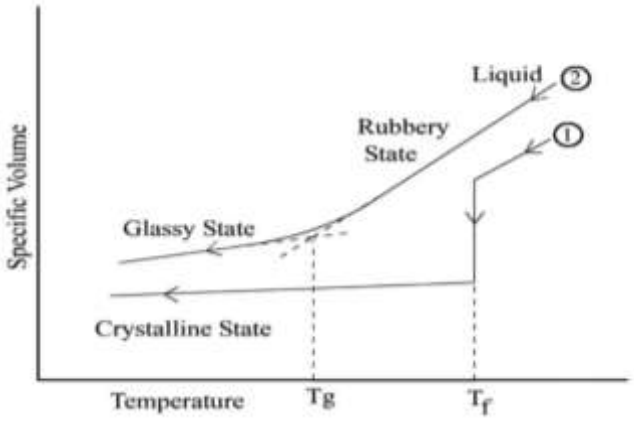
It takes longer time to cook at higher altitudes because the external pressure is lower there i.e. the atmospheric pressure is less than 760mm of Hg. The decrease in external pressure causes decrease in boiling point.

Since the boiling point of liquid depends on the external pressure. Therefore, food will take more time to cook at the high altitude due to low boiling points.

Q2. Glass softens over wide range of temperature. Ice melts at a specific temperature. Explain the reason for this difference.

Ans. The reason for their difference is as follow:

Reason: Glass softens over a wide range of temperature while ice melts at a specific temperature because glass is an example of amorphous solid and they do not have sharp melting point rather they melt over a range of temperature. While ice is an example of crystalline solids and has sharp, fixed or specific melting point.



Q3. Explain why it happens that on a hot summer day, when there is sweat on the body of a person, one feels cool under fast moving fan.

Ans. When there is sweat on the body of a person, he/she feels cool under the moving fan because:

Explanation:

During the hot summer days, we perspire as our body attempts to maintain our body temperature by releasing sweat through the pores of our skin, and when we sit under the fan, the sweat on our body starts to evaporate due to the air directed upon us, making our body cool.

This is due to the reason that when evaporation occurs, the liquid (sweat in this case) absorbs heat from the surroundings in order to evaporate. Thus, it takes a lot of heat from your body as well. Your skin thus becomes cool.

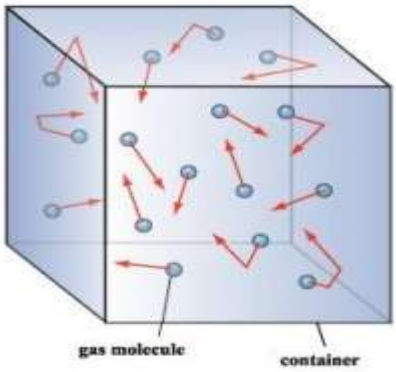
Q4. Why are the densities of gases lower than that of liquids?

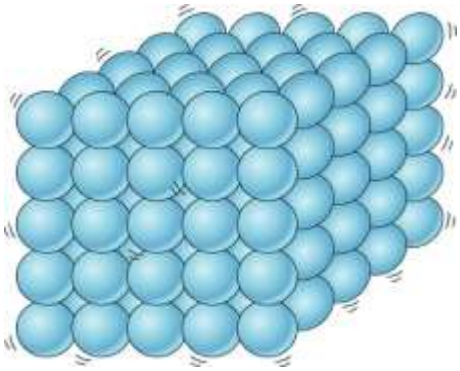
Ans. The densities of gases is lower than that of liquids because:

Reason:

According to the formula ($d = \frac{m}{v}$), density is inversely proportional to volume. A gas usually has much lower density than a solid or liquid. The molecules or atoms in a gas are much further apart than in a solid or a liquid. Gas molecules or atoms are usually flying around at very high speeds, occasionally bouncing off each other or the walls of the container the gas is in.

So, that is why the density of gas is very low, as compared to the same amount if liquid.





Gas Molecules *Solid Molecules*

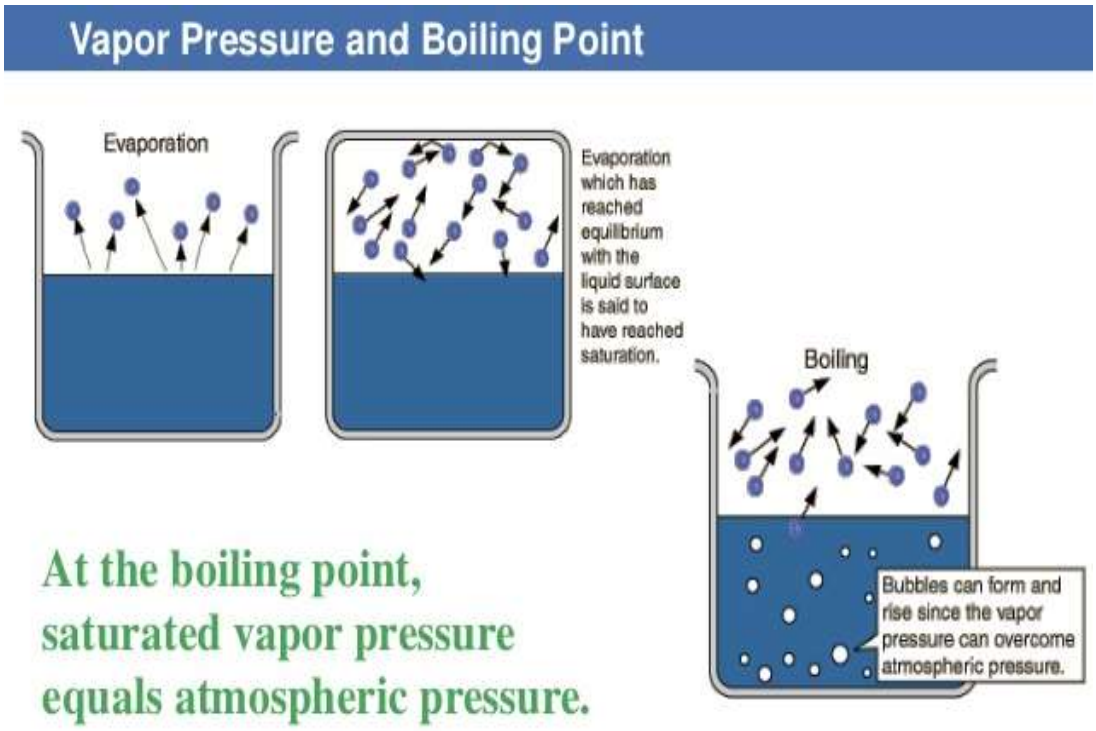
Q5. What is the relationship between atmospheric pressure and boiling point of a liquid?

Ans. The relationship between atmospheric pressure and boiling point of a liquid is as follow:

Relationship:

The higher the pressure the higher is the boiling point of a liquid. To explain this simply: as you increase the temperature of a liquid, its molecules start moving faster and faster.

A time is reached when the vapor pressure of liquid becomes equal to the atmospheric pressure and the liquid begins to boil which shows direct relationship between atmospheric pressure and the boiling point of a liquid. So, when the atmospheric pressure increases the boiling point will also increase. Similarly, when the atmospheric pressure decreases the boiling point will also decrease.



Q6. Why a gas is compressible but a solid is not compressible? Give reason?

Ans. The reason for why gas is compressible but solid is not compressible:

Reason:

Gases are highly compressible. This is because the gas molecules have large empty spaces as compared to the solids in which the particles are closely packed together due to the strong intermolecular forces. So, when pressure is applied to the gases, the distance between the molecules decreases, its volume decreases as well and the gases get compressed while solids do not get compressed because the particles in solids are fixed and closely packed together.

Long Q/A

Q1: Define Boyle's law and verify it experimentally.

Ans. **Boyles' Law:**

A law stating that the pressure of a given mass of an ideal gas is inversely proportional to its volume at a constant temperature.

Mathematical form:

$V \alpha 1/P$

$V = K_b \times 1/P$

$PV = K_b$

Where K_b is called constant for Boyle's law.

When the volume of a given mass of gas is changed from V_1 to V_2 and the pressure is changed from P_1 to P_2, then Boyles's law equation can be written as,

$P_1V_1 = P_2V_2 = K_b$ (at constant temperature)

Where, P_1= Initial pressure

P_2= Final pressure

V_1= Initial volume

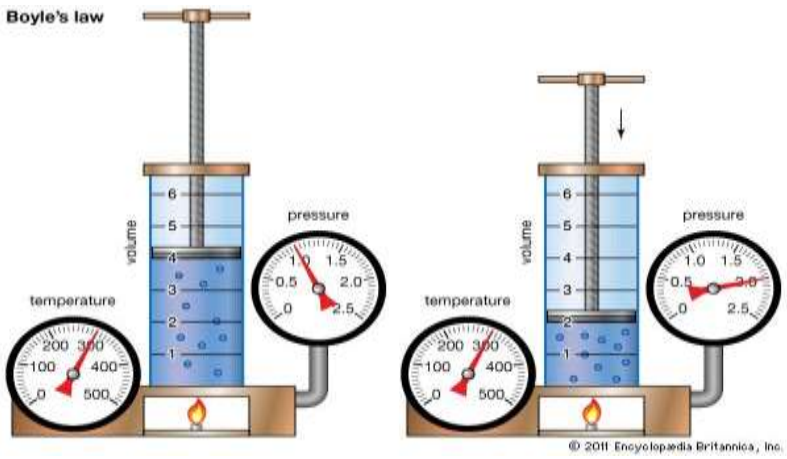
And V_2= Final volume.

According to the above equation, the Boyle's law can also be defined as,

The product of pressure and volume of given mass of gas remains constant provided the temperature is constant.

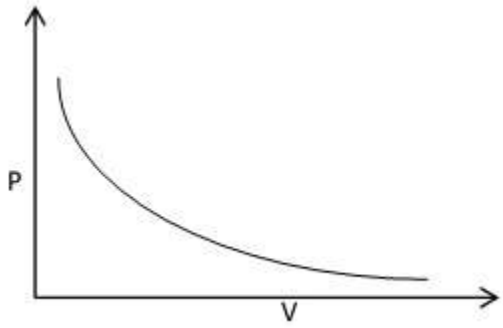
Experimental Verification of Boyle's law:

A certain mass of gas is enclosed in the cylinder. The volume of the gas is changed by increasing and decreasing the pressure. The volume at various pressure is noted. In each case, the product of pressure and volume remains constant at constant temperature and is found according to the Boyle's law, as shown in the figure:



Graphical representation:

If we plot the values of pressure 'P' and volume 'V', curve line is obtained which shows that the volume is inversely proportional to the pressure, as show in the figure:



Q2: a) What is the difference between evaporation and boiling?

Ans. The difference between evaporation and boiling is as follow:

Evaporation	Boiling
1) The spontaneous change of liquid into gaseous state is called evaporation	1) The conservation of liquid to gaseous state at boiling point us called boiling
2) Evaporation occurs at all temperature	2) Boling point occurs only at boiling point
3) It is a spontaneous process	3) It is not a spontaneous process
4) It produce cooling.	4) It does not produce cooling
5) It is surface phenomena	5) It is a bottom phenomena
6) External atmospheric pressure has no effect on evaporation	6) It increases with the increase of atmospheric pressure

b) Differentiate between effusion and diffusion

Ans. The difference between diffusion and effusion is as follow:

Diffusion	Effusion
1) The random mixing of gas molecule is called diffusion	1) The escape of gas molecule through a small hole one by one is called effusion
2) In diffusion different gas molecule take place	2) In effusion one kind of gas molecule take place
3) In diffusion collision occur between gas molecule	3) There is no collision between gas molecule

c) Differentiate between condensation and evaporation.

Ans. The difference between condensation and evaporation is as follow:

Condensation	Evaporation
1) The conversion of vapor back to liquid us called condensation	1) The conversion of liquid into vapor is called evaporation
2) During condensation heat is loss by liquid molecule	2) The liquid molecule gains heat and results in evaporation

Q3: Define 'Allotropy'? And give examples of allotropic forms of carbon.

Ans. **Allotropy:**

The existence of an element in more than one crystalline forms is called allotropy. The different forms are called allotropic forms.

For example:

1. The existence of an element in more than one physical state or form, such as Carbon (Kajol, Soot, Diamond, Graphite etc.)
2. The existence of two or more kinds of molecules of an element. In this case, each molecule has different number of atoms such as allotropes of Oxygen are Oxygen (O_2) and Ozone (O_3).

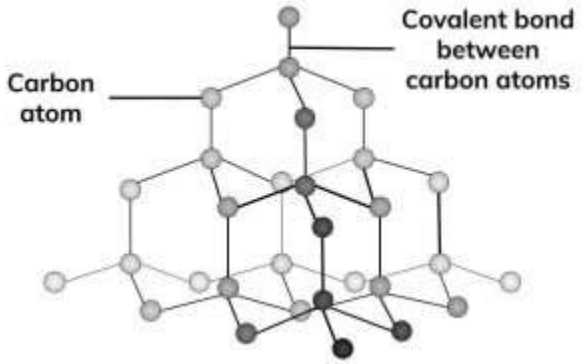
"The temperature at which at which one allotrope changes into another allotropic form is called ***Transition temperature****".*

Allotropic forms of Carbon:

Crystalline Carbon exists in three allotropic forms:

1. Diamond
2. Graphite
3. Bucky Ball

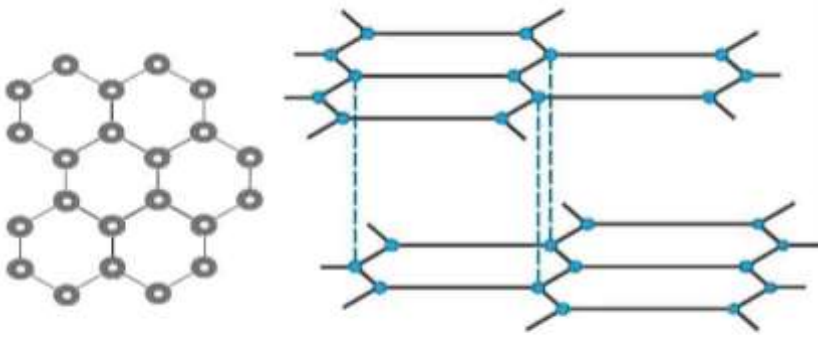
Diamond: It is the crystalline form of Carbon. In diamond, Carbon exists in cubic form. Each Carbon atom is tetrahedral bonded by four covalent bonds with other Carbon atoms. Covalent bond is very strong, so the diamond is very hard and has high melting point. Diamond is a bad conductor for electricity because all four valence electrons are used in formation of covalent bonds, which are tightly held with each other.



Graphite:

In graphite, each carbon atom is covalently bonded to three other carbon atoms rather than to four atoms as in diamond. Carbon exists in hexagonal form of sheet/ layers. These sheets linked with each other Carbon atoms by weak attractive forces. These sheets slide over each other.

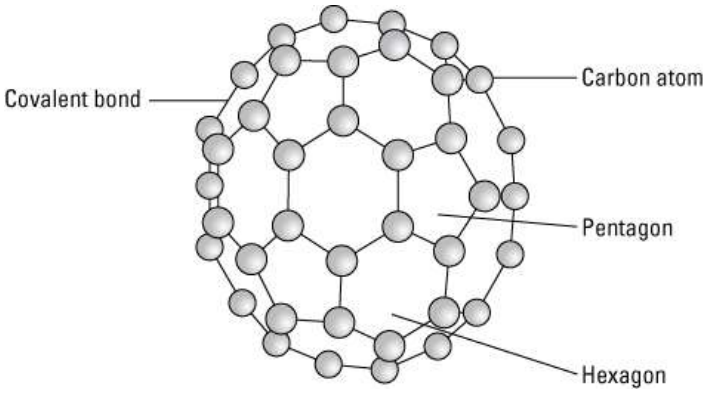
Graphite is soft. It is used as lubricant in heavy machinery. It is good conductor of electricity because it has free electron available.



graphite

Bucky Ball:

Bucky ball is the crystalline allotropic form of carbon. It is recently discovered in 1985. Bucky ball consist of twenty to hundred carbon atoms. In Bucky Ball, the atoms are arranged in a hollow cage like structure. The Carbon atoms linked with each other and adopt the shape of football In Bucky Ball, Carbon atoms joined together making pentagonal, hexagonal, etc. structures. Bucky Ball are used as semi-conductors, super conductors and lubricants.



Q4. What are the solids? Differentiate between amorphous solids and crystalline solids.

Ans. **Solids:**

In solid state of matter, the particles are closely packed together in a fixed pattern. In solids there occurs a strong force of attraction between the solid particles, which hold them firmly together, so that they cannot leave there position. Solid particles pass only the vibrational motion. Hence, solid cannot be fused like gas and liquids.

Typical properties of solids:

Some typical properties of solid state are:

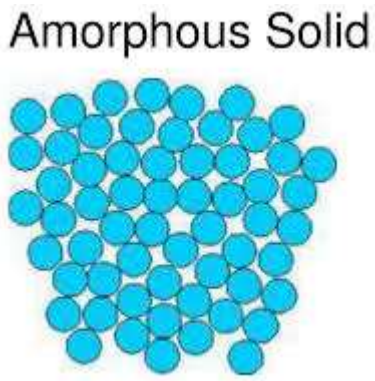
1. **Volume and shape:**
   Solids have strong inter molecular forces present between their particles, thus having a definite shape and a definite volume.
2. **Melting point:**
   The temperature at which the solids start to melt and exists in dynamic equilibrium with liquid state is called melting point.
3. **Rigidity:**
   The particles in solid are fixed and closely packed. The particles in solid neither move nor slides over their mean position. Therefore, the solids are rigid in their structure. The solids resists the deforming forming force due to hard structure and strong intermolecular force.
4. **Density:**
   Solids are denser than liquids and gases. In solids, the particles are closely packed together and have no empty spaces between the particles. There mass per unit volume is greater. Therefore, they have higher densities as compared to the other two states of matter.

Types of solids:

Solids can be classified into two types, based on the arrangements of particles. There are amorphous solids and crystalline solids.

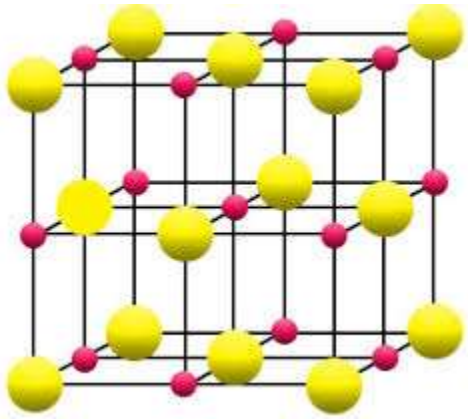
1. **Amorphous solids:**
   Amorphous means shape less. Solids at which the particles are not regularly arranged or in which the particles are not properly arrange in three dimensions are called amorphous solids. In simple words, we can say that

amorphous solids are one that lacks ordered arrangements of its particles. Amorphous solids are hard like true solids but they do not have sharp melting point. They melt over a range of temperature. For example, glass, wax, butter, plastic etc.



2. **<u>Crystalline solids:</u>**

Crystalline solids are the solids in which particles are arranged in regular three dimensional pattern. They have definite surfaces or faces. Each face has definite angel which the other face. Pure crystalline solids have sharp melting point. For example Sodium Chloride (NaCl), Naphthalene ($C_{10}H_8$) etc.



Q5. Define Charles's law and verify it graphically and diagrammatically?

Ans. **Charles's law:**

This law states that, the volume of a given mass of a gas is directly proportional to the absolute temperature at constant pressure.

Mathematically the law can be expressed as:

$V \alpha T$ (Constant Pressure)

$V = K_c T$

$V/T = K_c$

Where K_c is called constant of Charles's law.

When the volume is changed from V_1 to V_2 by changing the temperature form T_1 to T_2, then the relationship can be written in the following form,

$V_1/ T_1 = V_2/T_2 = K_c$

Where T_1= Initial temperature

T_2= Final temperature

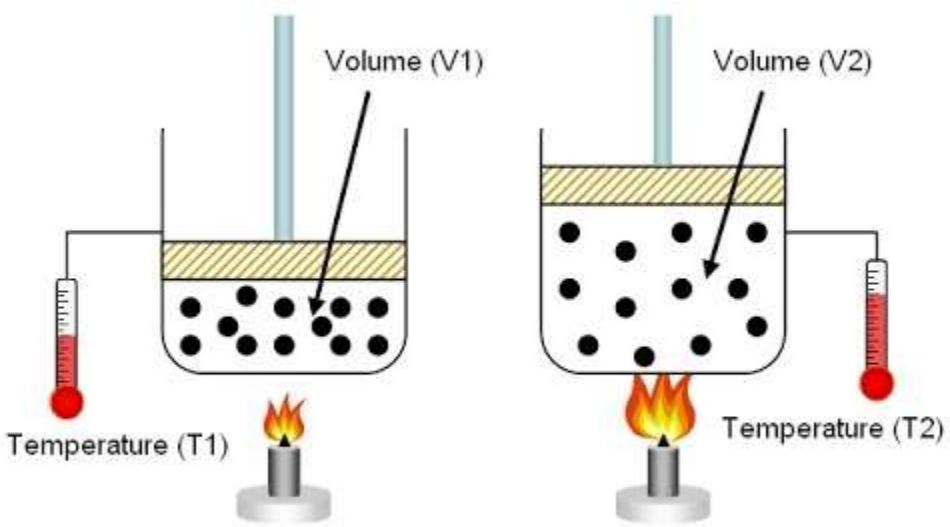
V_1= Initial volume

V_2= Final volume

From the above equation, the Charles's law can be defined as, the ratio between volume and absolute temperature of the given mass of a gas is constant at constant pressure.

Experimental verification of Charles's law:

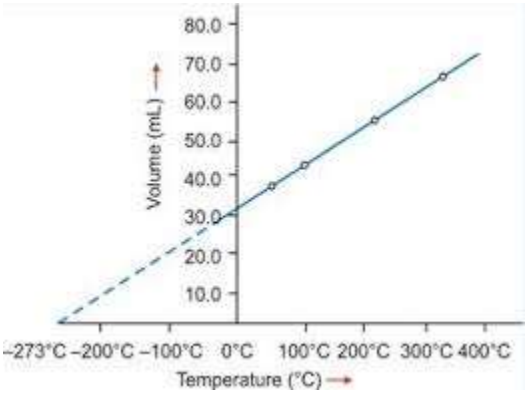
The apparatus used for the experimental verification of the Charles's law consists of a cylinder. The cylinder has a piston. The walls of the cylinder are heat insulator while the base of the cylinder is heat conductor. When the cylinder is heated at constant pressure, the piston moves upward and the volume will increase. It is noted from the various



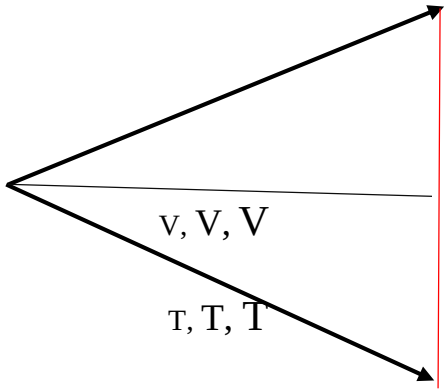
observations that the ratio between volume and absolute temperature remains constant. This verifies Charles's law.

Graphical Representation:

If the values of volume 'V' is plotted against the temperature 'T', a straight line is obtained, which shows that the volume is directly proportional to the absolute temperature.



The Charles's law can be diagrammatically represented as,



Exercise:

***Choose the correct option*:**

1. (c) 3. (c) 5. (b) 7. (c) 9. (c)
2. (a) 4. (b) 6. (c) 8. (c) 10. (a)

Best of luck

Reference: A text book of Chemistry Grade IX. Khyber Pakhtunkhwa Textbook Board Peshawar.