Week 4

BOND POLARITY

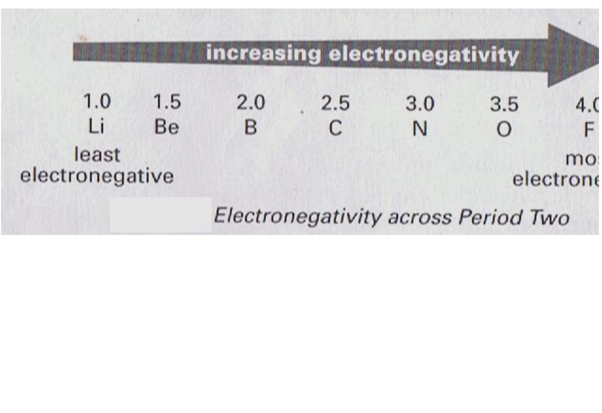
- The **electron pairs** shared between two atoms **are not necessarily shared equally.**
- The property of shared pair of electrons being closer to one atom than the other in a covalent bond is called **bond polarity.**

A covalent bond may be:

1. **Non-polar covalent bond** - when both atoms share electron pairs in a bond equally, i.e, when there are no charges present on atoms, the electron density is symmetrically distributed. Non-polar covalent bond occurs between atoms with identical electronegativity. In this way the electrons are evenly shared between two atoms forming the bond. For example, H-H bond in H2 molecule is a non-polar covalent bond making H2 a **non-polar molecule.**
2. **In a polar covalent bond** one atom is more electronegative than the other. For example, H-Cl bond in HCl molecules is a polar covalent bond making it a **polar molecule.**

Electronegativity (EN) is a measure of how strongly electrons are attracted to the nucleus of an atom.

- The Pauling electronegativity scale ranges from about 0.7 to 4.0. The following diagram shows the electronegativities of atoms of the second period:



Electronegativity across Period Two

- The least electronegative elements have the lowest values, while very electronegative elements have the highest values.
- The most electronegative atom of a bond has the greater pull on the shared electrons and will have a slight negative charge (shown by δ^-), because the shared electrons are held closer to it.
- The less electronegative atom has a slight positive charge, written δ^+.

<u>Features of the two types of Covalent Bonds</u>

<u>Non-polar</u>

- electrons are shared equally
- electrons usually found halfway between atoms
- occur between identical atoms e.g Cl_2, O_2, N_2

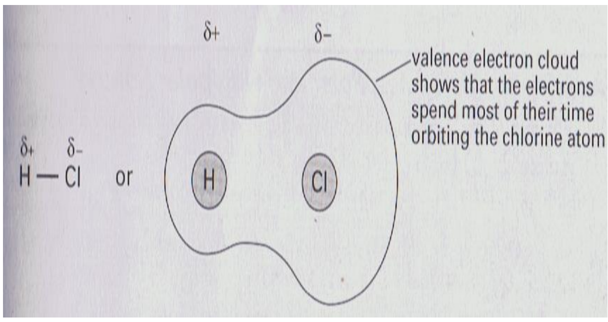
<u>Polar</u>

- electrons shared unequally
- electrons found closer to one atom than the other
- occur between different types of atoms e.g HCl, NH_3

Electronegativities can be used to **predict** which atom in a polar covalent bond will carry the slight negative charge and which the slight positive charge.

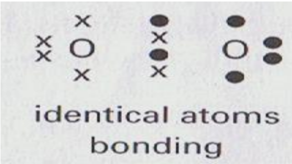
Example: Bonding electrons in a polar bond

Hydrogen chloride, HCl, contains polar covalent bonds because most of the time the bonding electrons are closer to the chlorine atom. The bond polarity can be shown by:

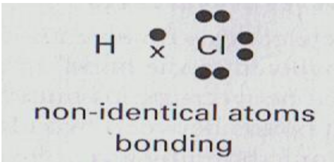


Bonding electrons are pulled closer to the chlorine because chlorine is more electronegative than hydrogen. (Chlorine has more protons than hydrogen, exerting a much greater pull on the bond electrons). The chlorine atom takes a sign.

Example: Polar and non-polar bonds



Oxygen, O_2, contains **non-polar covalent bonds** since the atoms involved in the bond are identical.



Hydrogen chloride, HCl, contains a **polar covalent bond**, since the atoms involved in the bond are not identical. The electrons are closer to the chlorine.

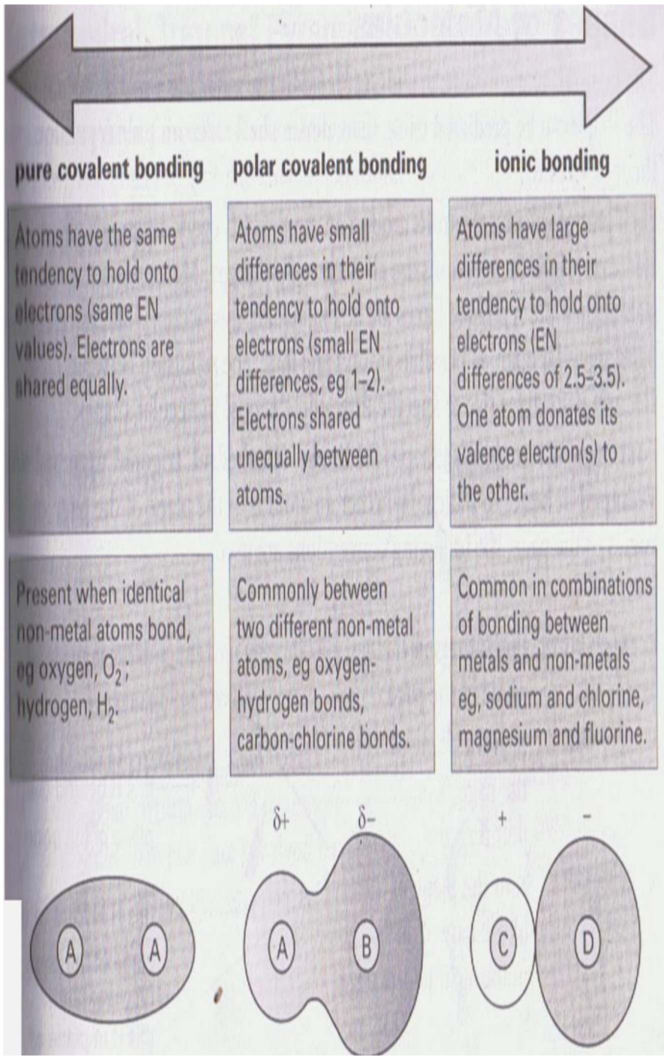
Example: Predicting polarity

$$\overset{\delta^-}{N} — \overset{\delta^+}{N}$$

- In a polar covalent bond between nitrogen and hydrogen, the bonding electrons will be closer to the nitrogen because nitrogen is more electronegative than hydrogen. The nitrogen atom will carry a slight negative charge.

Bonding as a Continuum

- Ionic bonds form when the electrons involved in the bond are transferred entirely to another atom involved in the bond.
- In covalent bonds, bonding electrons are not transferred between atoms but are shared between them.
- It is useful to consider covalent bonds between two identical atoms (forming a non-polar covalent bond) and ionic bonding as two extremes.
- In between are polar covalent bonds where the atoms in the bond are different.
- The electronegativity difference between atoms indicates where, in such a continuum, the bonding type lies, as the following diagram shows:



NOTES: The polarity of bond depends on electronegativity difference between bonded atoms. **The greater the electronegativity difference between bonded atoms, the greater the polarity, that is, the more polar the bond.**

Main Factors that affect the polarity of bond and overall molecule.

- Molecular geometry
- Relative electronegativity of participating atoms
- Symmetry of molecules

Common Examples of Non-polar and Polar Molecules

Non-Polar Molecules	Polar Molecules
All noble gases,e.g, He, Ne, Ar, Kr….etc	All hydrogen halides e.g, hydrogen fluoride HF, hydrogen chloride HCl, Hydrogen bromide HBr, hydrogen iodide HI.
All halogens, e.g, Fluorine gas F_2, Chlorine gas Cl_2, liquid bromine Br_2, and solid iodine I_2	Water H_20
Gases such as hydrogen gas, oxygen gas, nitrogen gas and carbon dioxide gas	Ammonia NH_3
All hydrocarbons e.g, alkanes, alkenes and alkynes	Ethanol CH_3CH_2OH

VSEPR THEORY AND SHAPE MOLECULES

To determine the shapes of molecules, we must become familiar with the Lewis structure. Drawing a Lewis structure is the first step towards predicting the molecular geometry(shape of molecule). The Lewis Structure helps us identify the bonded pairs and the lone pairs.

Molecular geometry is the three-dimensional arrangement of atoms in a molecule. A molecule's geometry **affects** its physical and chemical properties such as melting point, boiling point, density and the types of reactions it undergoes.

Then, with Lewis structure, we apply the **Valence Shell Electron Pair Repulsion (VSEPR) theory** to determine the electron-group geometry and the molecular geometry.

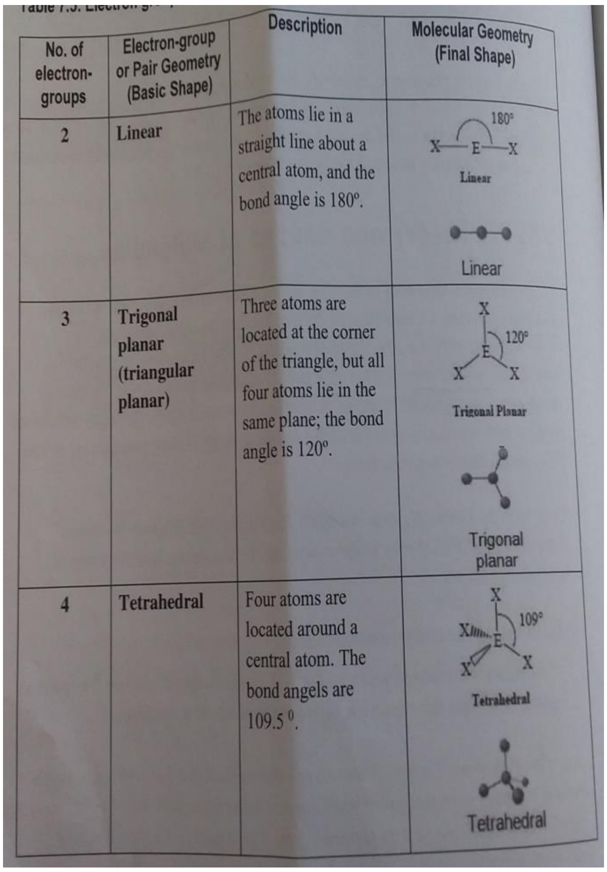
VSEPR theory is based on the idea that,

" valence shell electron pairs, being negatively charged, stay as far apart as possible so that the repulsion between them are at a minimum".

Electron-group is any collection of valence electron localised in a region around a central atom that exerts repulsion on other groups of valence electrons.

-It can be an electron-pair, lone pair, single unpaired electron, double bond or triple bond on central atom (double and triple bond treated like single bonds when counting electron group).

Table 7.5: Electron groups and Molecular Geometry



No. of electron-groups	Electron-group or Pair Geometry (Basic Shape)	Description	Molecular Geometry (Final Shape)
2	Linear	The atoms lie in a straight line about a central atom, and the bond angle is 180°.	Linear
3	Trigonal planar (triangular planar)	Three atoms are located at the corner of the triangle, but all four atoms lie in the same plane; the bond angle is 120°.	Trigonal planar
4	Tetrahedral	Four atoms are located around a central atom. The bond angels are 109.5 °.	Tetrahedral

NOTE:

- Electron-group geometry is determined by number of electron groups.
- Molecular geometry depends on not only the number of electron-groups but also the number of lone pairs.
- **When electron groups are all bonded pairs, molecular geometry (final shape) will be the same as the electron group geometry (basic shape).**

If central atom has both lone pairs (LP) and bonding pairs (BP), then 3 types of repulsive forces exist in such molecules. These are repulsions between bonding pairs, those between lone pairs, and those between a bonding pair and a lone pair. According to the VSEPR model, the repulsive forces decrease in the order summarised below.

Lone-pair (LP) Vs Lone-pair (LP) repulsions	>	Lone-pair (LP) Vs Bonding pair (BP) repulsions	>	Bonding-pair (BP) Vs Bonding pair (BP) repulsions

In VSEPR notation the molecule is represented by a formula using the letter **A** for the central atom, **X** for terminal atoms and **E** for lone pairs of electrons. **Table 7.6 a) and b)** below uses this notation and shows the geometry of various combinations, depending on the number of bonding pairs and lone pairs the molecule has.

Table 7.6 a): VSEPR Table (Electron-group geometry and Molecular geometry)

No. of Electron groups	Electron-group Geometry or Basic Shape (Bond Angle)	No. of Lone Pairs	VSEPR Notation	Molecular Geometry or Final Shape	Examples of Simple Molecules
2	Linear (180°)	0	AX_2	X–A–X Linear (180°)	$BeCl_2$
3	Trigonal planar (120°)	0	AX_3	Trigonal planar (120°)	BF_3
		1	AX_2E	Bent (< 120°)	SO_2
4	Tetrahedral (109.5°)	0	AX_4	Tetrahedral (109.5°)	CH_4
		1	AX_3E	Trigonal pyramidal (< 109.5°)	NH_3
		2	AX_2E_2	Bent or V-shaped (< 109.5°)	H_2O

Table 7.6 b): VSEPR Table (Electron-group/pair geometry and Molecular geometry)

Number of Electron Dense Areas	Electron-Pair Geometry	Molecular Geometry: No Lone Pairs	1 lone Pair	2 lone Pairs	3 lone Pairs
2	Linear	Linear			
3	Trigonal planar	Trigonal planar	Bent		
4	Tetrahedral	Tetrahedral	Trigonal pyramidal	Bent	

In VSEPR theory, pairs of electrons that surround the central atom of a molecule or ions are arranged as far apart as possible to minimise electron-electron repulsion.

STEPS TO PREDICT MOLECULAR GEOMETRY USING VSEPR THEORY

Step 1
- Draw the Lewis structure of the molecule or ion using the steps outlined in darwing lewis structures earlier in this sub-unit.

Step 2
- Count the number of electron-groups/pairs around the central atom (bonding pairs and lone pairs).
- Each of the following constitutes an electron-group:
  - a single, double or triple bond (multiple bonds count as one electron group)
  - a lone pair
  - an unpaired electron

Step 3
- Use **Table 7.5** to determine the **electron-group geometry (basic shape).**

Step 4
- Count the number of **bonding** and **non-bonding (lone)** electron pairs on the central atom and use **Table 7.6** to predict the **molecular geometry (final shape)** of the molecule.

Example 1

Predict the electron-group geometry (basic shape) and molecular geometry (final shape) of BCl_3.

SOLUTION:

Step1: Draw Lewis Structure of BCl_3.

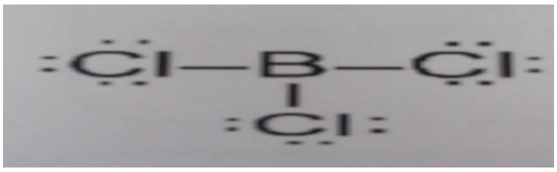
Valence electron count:

Boron (B) has 3 valence electrons × 1B atom = 3e's

Chlorine (Cl) has 7 valence electrons × 3Cl atoms = 21e's

Total Valence electrons = 24e's (21+3)

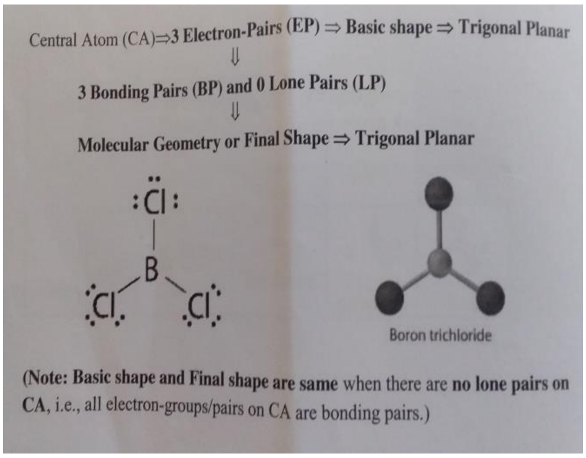
Lewis structure:



Step 2 and 3: Count the number of electron-groups/pairs around the central atom (bonding pairs and lone pairs) and using **Table 7.5** to determine the **electron group geometry (basic shape).**

Central Atom (CA)⟶3 Electron Pairs(EP) ⟶Basic Shape⟶ Trigonal Planar

Step 4: Count the number of **bonding** and **non-bonding** (lone) electron pairs on the central atom and using Table 7.6 to predict the **molecular geometry (final shape)** of the molecule.



Example 2

Predict the electron-group geometry (basic shape) and molecular geometry (final shape) of CH_4.

SOLUTION:

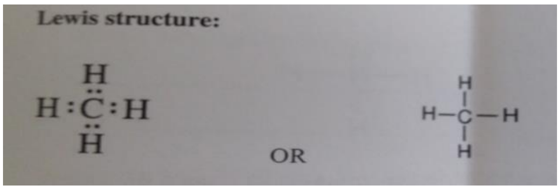
Step1: Draw Lewis Structure of CH_4.

Valence electron count:

Carbon (C) has 4 valence electrons × 1C atom = 4e's

Hydrogen (H) has 1 valence electrons × 4H atoms = 4e's

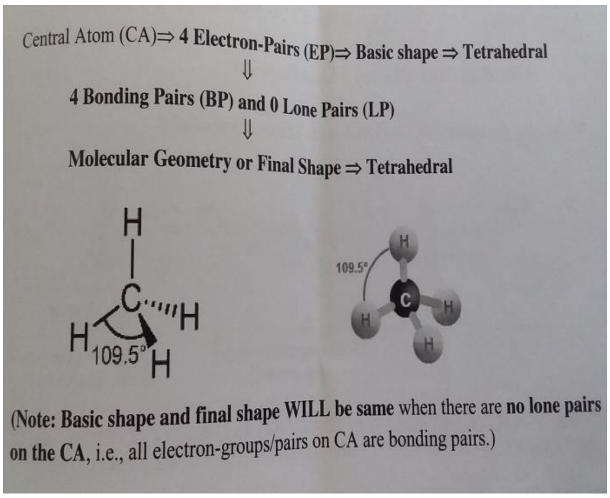
Total Valence electrons = 8e's (4+4)



Step 2 and 3: Count the number of electron-groups/pairs around the central atom (bonding pairs and lone pairs) and using **Table 7.5** to determine the **electron group geometry (basic shape).**

Central Atom (CA)⟶4 Electron Pairs (EP) ⟶Basic Shape⟶ Tetrahedral

Step 4: Count the number of **bonding** and **non-bonding** (lone) electron pairs on the central atom and using Table 7.6 to predict the **molecular geometry (final shape)** of the molecule.



EXERCISE 7: BOND POLARITY, VSEPR THEORY AND SHAPE MOLECULES

1. What is bond polarity?
2. Differentiate polar covalent bond from non-polar covalent bond.
3. The polarity of bond depends on what?
4. Give two examples of non-polar and polar molecules.
5. Explain VSEPR Theory in your own words.
6. Find out the electron-group geometry (basic shape) and molecular geometry (final shape) of NH_3.

MOLE CONCEPT AND AVOGADRO CONSTANT

The Mole (n)

In biology, moles are burrowing animals, but in chemistry the mole is a unit.

- The **mole** is a unit for the measure of the amount of a substance.
- It is **abbreviated as mol** and its symbol is **"n".**
- By definition, **a mole of any substance** contains the same number of elementary particles (atoms,molecules, or other particles) as there are atoms in exactly 12g of the ^{12}C isotope.

Avogadro's number (N_A) and Molar Mass

- The actual number of atoms in 12g of ^{12}C is determined experimentally. This number is called Avogadro's number (N_A) or Avogadro's constant, in honor of the Italian scientist Amedeo Avogadro.

Thus, Avogadro's number (N_A) = 6.02×10^{23}
1mol ^{12}C atom = 6.02×10^{23} ^{12}C atoms= 12.0g ^{12}C atoms
1mole of atom = 6.02×10^{23} atoms = atomic mass of atom in grams

Note:

-It is important to note that one mole of atoms contains 6.02×10^{23}atoms.
-One mole of molecules contains 6.02×10^{23} molecules.
-One mole of ions contains 6.02×10^{23}ions.
Thus, $\mathbf{N_A}$ can have any of these units;
molecules/mole
atoms/mole
ions/mole

- The mass (in grams) of I mol of substance is its molar mass (M), and is expressed in g/mol.
- Thus the atomic number of sodium(Na) is 23.0amu and its molar mass is 23.0g/mol; the atomic mass of chlorine (Cl) is 25.5amu and its molar mass is 35.5g/mol and so on.

Note:

1 mole of an element/atom = atomic mass of that element/atom in g = 6.02×10^{23}atoms

1 mole of a molecule = molecular mass of that molecule in g = 6.02×10^{23} molecules

Example 1:

Predict the mass of the following:

a. 1 mole of calcium (Ca) atoms.

b. 1 mole of water (H_2O) molecule.

Solution

a. 1 mole of Ca = 40.1 amu = 40.1 g

b. 1 mole of H_2O = 2 (H) + 1 (O) = 2 (1.0) + 1 (16.0) = 18.0 amu = 18.0 g

Mole-Mass relationship

The moles can be calculated using mass (in g) and molecular mass or molar mass (in g/mol), as stated in the formula below:

$$\text{no. of moles of an element} = \frac{\text{mass}}{\text{atomic mass}}, \quad n = \frac{m}{Ar}$$

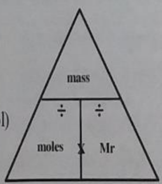
$$\text{no. of moles of a molecule} = \frac{\text{mass}}{\text{molecular mass}}, \quad n = \frac{m}{Mr} \quad \text{wher:}$$

n = no. of moles (mol)

m = mass in grams (g)

Ar = atomic mass in grams/mol (g/mol)

Mr = molecular mass or molar mass in grams/mol (g/mol)



$$n = \frac{m}{Mr} = \frac{g}{g/mol} = \cancel{g} \times \frac{mol}{\cancel{g}} = mol$$

Example 2:

Calculate the number of moles in:

a. 97.8g of K
b. 67.5g of $C_6H_{12}O_6$

Solution

a. $n = \frac{m}{Ar} = \frac{97.8g}{39.1g/mol}$ = **2.50mol K**
b. Mr ($C_6H_{12}O_6$)
= 6(C)+12(H)+6(O)
=6(12.0)+12(1.0)+6(16.0)
=180.0amu or g/mol

$n = \frac{m}{Ar} = \frac{67.5g}{180.0g/mol}$

= **0.375 or 0.38mol $C_6H_{12}O_6$**

Example 3:

Calculate the mass (in g) of 2.0moles of calcium carbonate ($CaCO_3$).

Solution

Mr ($CaCO_3$)

= 1(Ca)+1(C)+3(O)

=1(40.1)+1(12.0)+3(16.0)

=100.1amu or g/mol

$n = \frac{m}{Ar}$=

m=n × Mr

=2.0mol × 100.1g/mol

= **200.2g $CaCO_3$**

The Mole Concept - Moles, Avogadro's number, Mass and Volume relationship

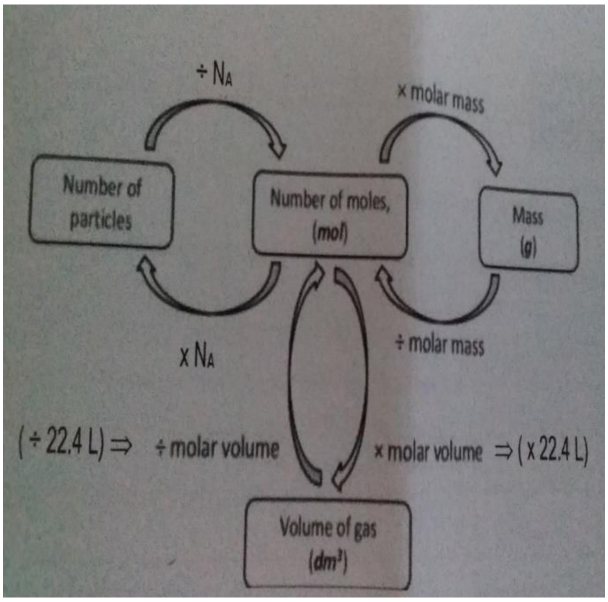
The mole may be expressed as:

Avogadro's number	Mass	Volume
▸ number of particles which is equivalent to Avogadro's number = 6.02×10^{23} ▸ **1 mol = 6.02×10^{23} atoms/molecules/ ions/particles/things**	▸ atomic or molecular mass or weight expressed in grams ▸ **1 mol element = Ar of element in g** ▸ **1 mol molecule = Mr of molecule in g**	▸ 1 mole of a gas at STP (standard temperature and pressure) occupies a volume of 22.4 L or dm^3 ▸ At STP, temperature = 0 °C (or 273 K) and pressure = 1 atmosphere (1 atm) ▸ **1 mol of a gas at STP = 22.4 L or 22.4 dm^3**

Example: 1 mole of methane (CH_4) gas at STP is equal to:

6.02×10^{23} molecules of methane gas	**16 grams** of methane gas (Note: Mr (CH_4) = 16 g/mol).	**22.4 L or dm^3** volume of methane gas

Mole Concept: Relationship between moles, number of atoms/molecules, mass and volume of gas at STP.



Example 4:

A student collects 0.25 mol of ethane gas at STP (C_2H_6). Determine the number of ethane molecules, the mass and volume of ethane gas (**Note:** Mr (C_2H_6) = 30.0 g/mol).

Solution

Molecules of ethane gas

1 mol = 6.02×10^{23} molecules
0.25 mol = x molecules

$$x = \frac{(0.25\ \text{mol})\ (6.02 \times 10^{23}\ \text{molecules})}{(1\ \text{mol})}$$
$$= 1.5 \times 10^{23} \text{ molecules of ethane gas}$$

Mass of methane

$$n = \frac{m}{Mr}\ ;\ m = (n)\ (Mr)$$
$$= (0.25\ \text{mol})\ (30.0\ \text{g/mol})$$
$$= 7.5\ \text{g}$$

Volume of gas

1 mol = 22.4 L or dm^3 at STP
0.25 mol = x L

$$x = \frac{0.25\ \text{mol} \times 22.4\ \text{L}}{1\ \text{mol}} = 5.6\ \text{L or dm}^3$$

Writing Chemical Formulas

A formula represents a relative number of atoms of different elements present in the substance. To write a formula, you need to analyse a chemical. However, we can take several points for granted, and concentrate on writing chemical formulas by a method called the **ion-charge method** (you may wish to "brush-up" your knowledge on what an **ion** is).

An **ion** is an atom or a group of atoms that has a net positive or negative charge. The number of positively charged protons in the nucleus of an atom remains the same during ordinary chemical changes (called chemical reactions), but negatively charged electrons may be lost or gained.

The loss of one or more electrons from a neutral atom results in a **cation,** an ion with a net positive charge. For example, a sodium atom (Na) can readily lose an electron to become a sodium cation, which is represented by Na^+:

Sodium atom (Na)	Sodium ion (Na^+)
11 protons ⇒ +11 11 electrons ⇒ -11 Overall charge = +11 + -11 = 0 charge	11 protons ⇒ +11 10 electrons ⇒ -10 (lost an electron) Overall charge = +11 + -10 = +1 charge

On the other hand, an **anion** is an ion whose net charge is negative due to an increase in the number of electrons. A chlorine atom (Cl), for instance, can gain an electron to become the chloride ion, Cl^- :

Chlorine atom (Cl)	Chloride ion (Cl^-)
17 protons ⇒ +17 17 electrons ⇒ -17 Overall charge = +17 + -17 = 0 charge	17 protons ⇒ +17 18 electrons ⇒ -18 (gained an electron) Overall charge = +17 + -18 = -1 charge

For the ion-charge method, we must know the charge of each ion, or the group of ions. **Table 4.1** shows the common ions, their charges and names. **Please note, it is very important to learn the names and charges of the common ions listed in the table below, as we will be using them regularly in our study of chemistry,**

Table 4.1: Common Ions and their Charges

COMMON IONS, THEIR CHARGES AND NAMES		
Positive ions (CATIONS)		
1+ charge	**2+ charge**	**3+ charge**
Sodium, Na^+ Potassium, K^+ Copper(I), Cu^+ Silver, Ag^+ Ammonium, NH_4^+	Magnesium, Mg^{2+} Barium, Ba^{2+} Calcium, Ca^{2+} Iron(II), Fe^{2+} Nickel(II), Ni^{2+} Copper(II), Cu^{2+} Zinc, Zn^{2+} Lead(II), Pb^{2+} Mercury(II), Hg^{2+}	Aluminium, Al^{3+} Chromium(III), Cr^{3+} Iron(III), Fe^{3+}
Negative ions (ANIONS)		**3- charge**
1- charge	**2- charge**	Phosphate, PO_4^{3-}
Fluoride, F^- Chloride, Cl^- Bromide, Br^- Iodide, I^- Hydroxide, OH^- Nitrate, NO_3^- Nitrite, NO_2^- Chlorate, ClO_3^- Hydrogen carbonate or bicarbonate, HCO_3^- Hydrogen sulfate or bisulfate, HSO_4^- Acetate, CH_3COO^-	Oxide, O^{2-} Peroxide, O_2^{2-} Carbonate, CO_3^{2-} Chromate, CrO_4^{2-} Dichromate, $Cr_2O_7^{2-}$ Sulfide, S^{2-} Sulfate, SO_4^{2-} Sulfite, SO_3^{2-}	

Example 1

To illustrate the use of this method, place each symbol side by side and under each write down the ionic charge, as shown below:

Sodium chloride

Na Cl … (symbols)

1 1 … (ionic charges)

Then, "cross – multiply" each symbol with the ionic charges

$$\begin{matrix} Na & & Cl \\ & \times & \\ 1 & & 1 \end{matrix} \quad = Na_1\,Cl_1 = NaCl$$

You have the formula for sodium chloride. Because we cannot divide the two numbers any further, NaCl becomes the simplest formula for sodium chloride (common salt).

Example 2

In this case, sodium is one element and can be represented by Na, but hydroxide contains two elements and is given on our chart as OH with a charge of one. Treat groups like OH as one ion.

Sodium hydroxide Nickel (II) nitrate - $Ni(NO_3)_2$

$$\begin{array}{ccc} Na & & OH \\ 1 & \times & 1 \end{array} = NaOH$$

Ca^{2+} PO_4^{-3} $= Ca_3(PO_4)_2$

More examples show **calcium hydroxide, aluminium nitrate and potassium chloride.**

Example 3 Ion charge method

Calcium hydroxide

$$\begin{array}{ccc} Ca & & OH \\ 2 & \times & 1 \end{array} = Ca_1(OH)_2 = Ca(OH)_2$$

superscript

Aluminium nitrate

$$\begin{array}{ccc} Al & & NO_3 \\ 1 & \times & 3 \end{array} = Al_1(NO_3)_3 = Al(NO_3)_3$$

subscript

Potassium chloride

$$\begin{array}{ccc} K & & Cl \\ 1 & \times & 1 \end{array} = K_1Cl_1 = KCl$$

Note:
1. Wherever the number "1" is involved, we leave it out. That is, when there is no number near a symbol, 1 is implied.
2. Parentheses ( ) are used if **a group** is taken more than once. Example, $Mg(OH)_2$ or $Al(NO_3)_3$.

Remember: This method has the following limitations:

- There are many covalent compounds whose chemical formulae cannot be written using the ion-charge method.
- For example, the chemical formula for butane is C_4H_{10} but obviously we cannot use our ion-charge method for butane. Chemical formulae for compounds like these are determined by what we call **"combustion analysis"** (burning and analyzing the weight percentages of the products).
- It is possible to write down a chemical formula, and then learn that such a compound does **not** exist.

3. Similarly, compounds formed with the positive ammonium ion NH_4^+, are not binary. All binary compounds end in **– ide**, but not all compounds with **– ide** endings are binary. For example, NaOH, $Ca(CN)_2$ and NH_4Cl have names that end in **–ide**, but they are not binary compounds.

Refer to **Table 4.1** for a list of **common cations and anions, their charges and names.**

Table 4.2 shows the examples of common binary compounds.

Table 4.2: Common Binary Compounds

Chemical Formula	Name
NO	nitrogen oxide
NO_2	nitrogen dioxide
N_2O	dinitrogen monoxide or dinitrogen oxide
N_2O_3	dinitrogen triodixide
N_2O_5	dinitrogen pentoxide
SO_2	sulphur dioxide
SO_3	sulphur trioxide
CCl_4	carbon tetrachloride
CO	carbon monoxide
CO_2	carbon dioxide
NaCl	sodium chloride
NaBr	sodium bromide
NaI	sodium iodide
LiCl	lithium chloride
Na_2O	sodium oxide
Li_2O	lithium oxide
$MgCl_2$	magnesium chloride
$FeCl_2$	ferrous chloride
$FeCl_3$	ferric chloride
FeO	iron(II) oxide or ferrous oxide
Fe_2O_3	iron(III) oxide or ferric oxide
CuO	copper(II) oxide
Cu_2O	copper(I) oxide
$ZnCl_2$	zinc chloride

$Cu(OH)_2$ - copper (II) hyd...

Naming Chemical Formula

Many chemical compound names consist of two words. The words are derived from the ions that combine to form the compound. Those compounds that consist of only 2 elements are called **binary compounds.** Some elements form more than one covalent compound with another element. For example, 2 compounds made of sulphur and oxygen could be: SO_2 and SO_3.

The names of binary compounds usually end with **"-ide"**. This is from the second combining element. For example: bromine – bromide or sulphur – sulphide and so forth. Binary compounds can be either ionic or covalent.

Covalent binary compounds (e.g. N_2O) are named by following these steps:

1. The first word of the name is made up of (a) a prefix indicating the number of atoms of the first element appearing in the formula, if more than one; and (b) the name of the first element in the formula (e.g, di-nitrogen, for "N_2" of N_2O).
2. The second word of the name is made up of (a) a prefix indicating the number of atoms of the second element appearing in the formula, if more than one compound of these two elements exists: (b) the root of the name of the second element and (c) the suffix **"- ide"**, which means that only the two elements named are present (e.g. mono-oxide for "O" of N_2O).

Thus the name of the compound with the chemical formula, **N_2O is dinitrogen monoxide.**

Note:
The Prefixes:

mono – meaning one
di – meaning two
tri – meaning three
tetra – meaning four
pent – meaning five

1. The prefix **'mono'** can be omitted if **only one** of the combining elements is made of one atom. For example,
   N_2O - Could either be dinitrogen monoxide or dinitrogen oxide.
   CO - Will be carbon monoxide and not carbon oxide because carbon and oxygen are both one (mono).
2. A few negative polyatomic ions (ions made up of two or more atoms) have names with an **"-ide"** suffix, for example, hydroxide ion, OH^- and cyanide ion, CN^-. Thus, the compounds LiOH and KCN are named lithium hydroxide and potassium cyanide, respectively.

Compounds formed with either of these polyatomic ions are not binary. These and a number of other such ionic compounds are called **ternary compounds,** meaning **compounds consisting of three elements.**

Exercise 8: Writing and Naming Chemical Formula.

1. Write down the chemical formula for each of the following compounds:

a. Sodium carbonate
b. Potassium sulfate
c. Calcium nitrate
d. Aluminium hydroxide
e. Lithium oxide

2. State the charges of the following ions with their chemical formula:

a. Ammonium
b. Silver
c. Calcium
d. Magnesium
e. Carbonate

Week 5 and Week 6
SOLIDS AND RELATED PROPERTIES

States of matter and particle theory

Learning Outcomes:

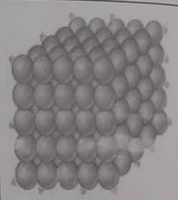
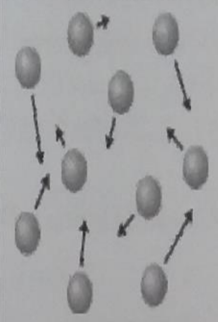
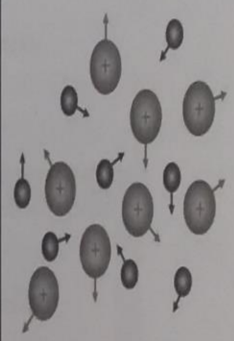
i.	Define matter.	1
ii.	Define particles of matter.	1
iii.	List the 3 different states of matter.	2
iv.	Identify the state of matter from a given set of data or illustrations.	1
v.	Describe the characteristics of particles at a given state of matter. *(in terms of arrangement, force between particles, particle energy & movement)*	2
vi.	Name, State or Identify the process of changing the state of matter from one form to another. *(include melting, vapourisation, condensation, freezing or solidification, sublimation)*	1
vii.	Describe the nature of particles during the process of change in the states of matter.	2
viii.	Differentiate between the physical properties of any two states of matter. *(Physical properties include: melting point, boiling point, freezing point, compressibility, movement, electrical conductivity)*	3
ix.	Account for the differences in the physical properties of any two states of matter of the same substance.	3

What is Matter?

Matter is known as anything that occupies space and has mass. Matter includes things we can see and touch (such as water, trees and earth), as well as anything we cannot (such as air).

Matter is made up of particles called atoms. The arrangements of these particles are usually divided into four major states i.e. solid, liquid, gas and plasma.

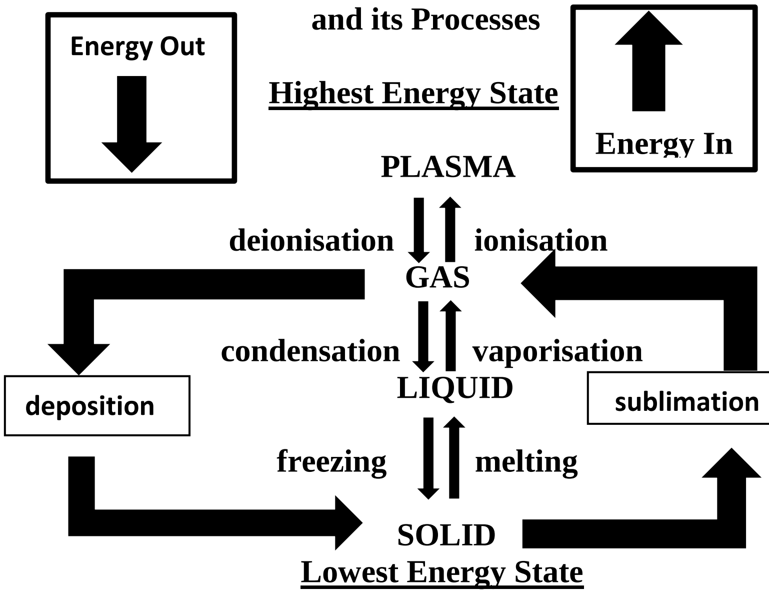
Table 1: State of Matter Particle Arrangement

State	Particle Arrangement	
Solid	▸ Atoms are very tightly packed together in a fixed position. ▸ Solids retain their shape. ▸ Cannot be compressed to bring atoms any closer together. Even if high pressure is used, this would not be possible. ▸ The only result would be that the atoms held in fixed positions would be distorted or damaged. ▸ Some examples of solids include tables, chairs, and books.	 (Source of image: https://chemstuff.co.uk/academic-work/year-7/particle-model-of-solids-liquids-and-gases/)
Liquid	▸ Atoms are less tightly but closely packed and not fixed in any position. ▸ Atoms of liquids are able to move and roll over each other. ▸ Liquids do not have a fixed shape and take up the shape of the container they are placed in. ▸ For instance, liquids placed in a square container take the shape of a square, but when placed in bottle, it takes the shape of the bottle.	 (Source of image: https://chemstuff.co.uk/academic-work/year-7/particle-model-of-solids-liquids-and-gases/)
Gas	▸ Particles in gases are far apart and are in constant motion. ▸ Particles usually take the shape of the containing vessel. In fact, the particles spread to occupy the whole vessel. ▸ This can easily be seen using coloured gases such as chlorine gas (yellowish green) and nitrogen dioxide gas (brownish). Some colourless gases include oxygen, hydrogen and carbon monoxide gas.	 (Source of image: https://chemstuff.co.uk/academic-work/year-7/particle-model-of-solids-liquids-and-gases/)
Plasma	▸ Plasma is not a common state of matter here on Earth, but may be the most common state of matter in the universe. ▸ Plasma has neither a definite volume nor a definite shape. ▸ It is often seen in ionized gases, but it is distinct from a gas. Unlike gases, plasmas are electrically conductive, produce magnetic fields and electric currents. ▸ Positively charged nuclei swim in a "sea" of freely-moving disassociated electrons, where this "sea of electron" allows matter in the plasma state to conduct electricity. ▸ The plasma may be formed by heating and ionizing a gas. Heating matter to high temperatures causes electrons to leave the atoms, resulting in the presence of free electrons. This creates a partially ionized plasma. Stars are essentially superheated balls of plasma. ▸ Examples of plasma include stars, lightning, electric sparks, fluorescent lights, neon signs and **plasma televisions. They are called "plasma" displays because they use small cells containing electrically charged ionized gases, which are plasmas.**	 Source of image: https://tetronics.com/our-technology/what-is-plasma/)

Changes in States of Matter

From our basic knowledge of science we are fully aware that the states of matter change from one to the other when we simply apply heat or carry out the cooling process.

Figure 1: Changes in the State of Matter and its Processes



Note:

Figure 1 highlights processes that are commonly known to us as they are a part of everyday lives. For example, melting and freezing can easily change a solid substance to liquid and vice versa as evident when we freeze water to get ice on a hot day or how the ice melts when exposed to the heat from the sun.

HOW CAN WE EXPLAIN THESE CHANGES IN THE PHASES/STATES WHEN WE APPLY OR REMOVE HEAT FROM A SUBSTANCE?

-The answer lies in the study of heating and cooling curves.

Heating and Cooling Curves

-Prior to learning about heating and cooling curves it is vital that we have an understanding of what **system** is.

-**System** is a **closed imaginary vessel or container that is isolated from its surroundings or environment.**

-A system will change when it absorbs or loses heat.

-Adding and removal of heat from system and how it changes in states can be illustrated using **heating and cooling curve.**

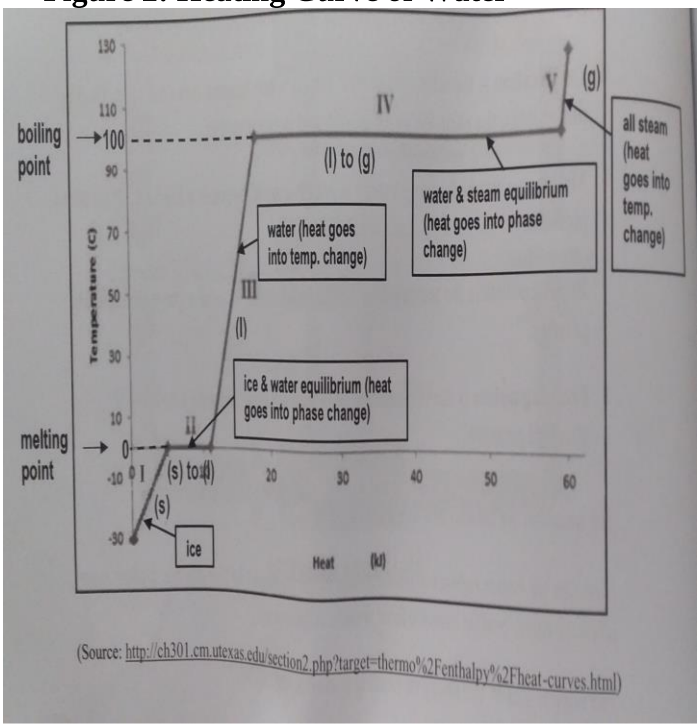
WHAT ARE HEATING AND COOLING CURVE?

It is a simple line graph that shows the phase changes a given substance undergoes with increasing or decreasing temperature. (Temperature, y-axis and heat/time, x-axis)

HEATING CURVE OF WATER

-express the absorbing of heat into system that changes states from solid to liquid to gas.

Figure 2: Heating Curve of Water



-System (water) begins in solid state. i.e. ice and as heat flows into system from surroundings the ice melts and becomes liquid. As more heat flows into the system, liquid water will vaporise into gaseous state.

Figure 3: Stages in the Heating Curve of Water

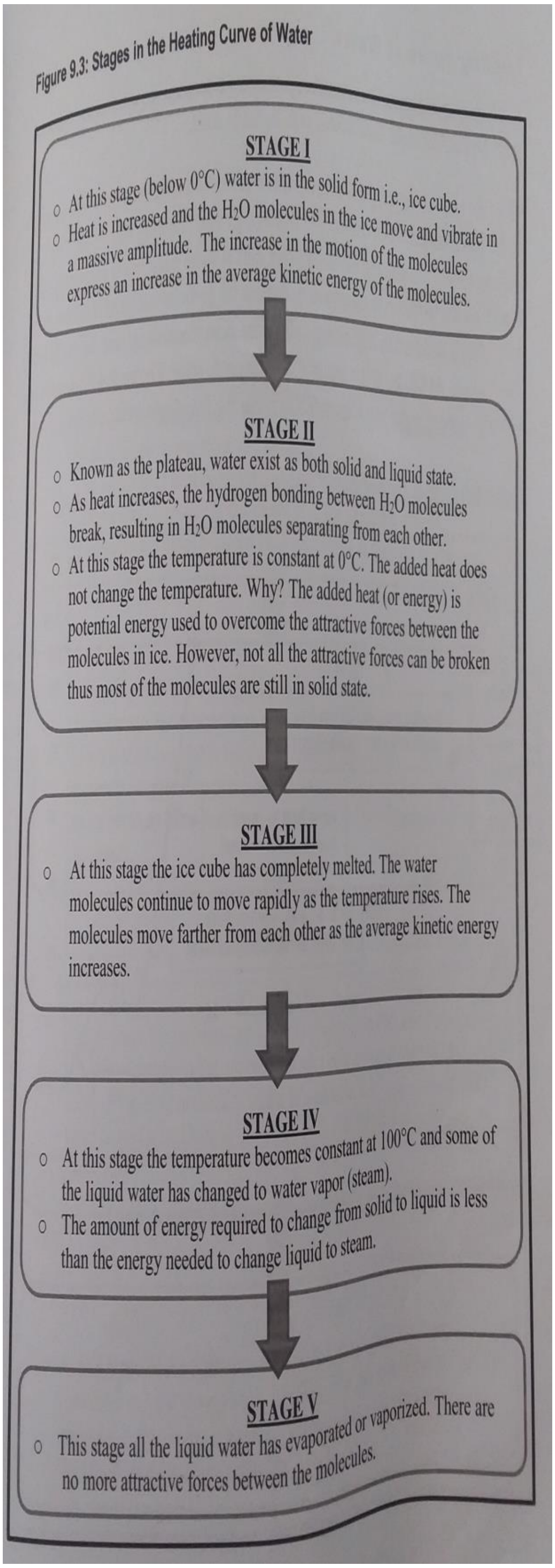


Figure 9.3: Stages in the Heating Curve of Water

STAGE I

- At this stage (below 0°C) water is in the solid form i.e., ice cube.
- Heat is increased and the H_2O molecules in the ice move and vibrate in a massive amplitude. The increase in the motion of the molecules express an increase in the average kinetic energy of the molecules.

STAGE II

- Known as the plateau, water exist as both solid and liquid state.
- As heat increases, the hydrogen bonding between H_2O molecules break, resulting in H_2O molecules separating from each other.
- At this stage the temperature is constant at 0°C. The added heat does not change the temperature. Why? The added heat (or energy) is potential energy used to overcome the attractive forces between the molecules in ice. However, not all the attractive forces can be broken thus most of the molecules are still in solid state.

STAGE III

- At this stage the ice cube has completely melted. The water molecules continue to move rapidly as the temperature rises. The molecules move farther from each other as the average kinetic energy increases.

STAGE IV

- At this stage the temperature becomes constant at 100°C and some of the liquid water has changed to water vapor (steam).
- The amount of energy required to change from solid to liquid is less than the energy needed to change liquid to steam.

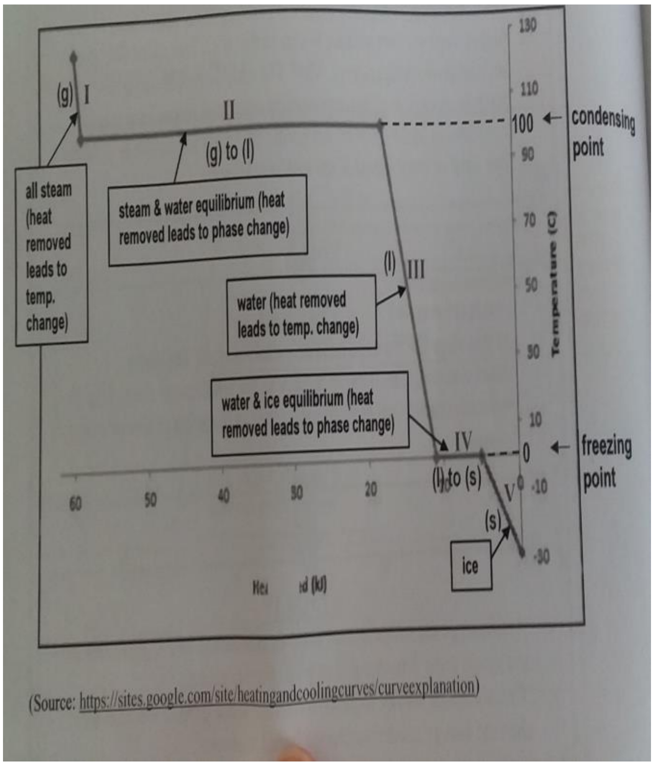
STAGE V

- This stage all the liquid water has evaporated or vaporized. There are no more attractive forces between the molecules.

COOLING CURVE OF WATER

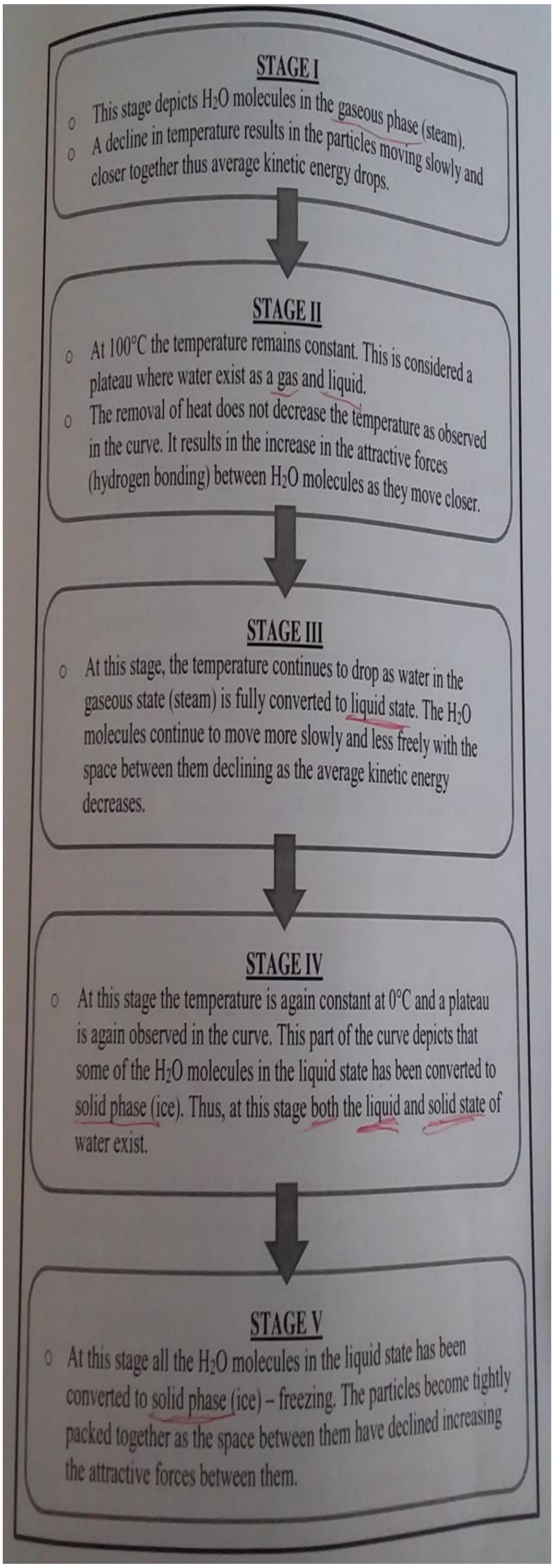
Cooling curve's describes the reversible process of a heating curve where a system loses heat and changes state from the gas to liquid to solid.

Figure 4: Cooling Curve of Water



-When system in gaseous state cooled it will lose energy thus resulting in the particles moving slowly and coming closer together thus at some point gas is converted to liquid. As the temperature continues to decline the particles continue to move even more slowly and tend to lose energy until there is no more energy for the particles to move freely. At this point the liquid state changes to solid.

Figure 5: Stages of Cooling Curve of Water



STAGE I

- This stage depicts H_2O molecules in the gaseous phase (steam).
- A decline in temperature results in the particles moving slowly and closer together thus average kinetic energy drops.

STAGE II

- At 100°C the temperature remains constant. This is considered a plateau where water exist as a gas and liquid.
- The removal of heat does not decrease the temperature as observed in the curve. It results in the increase in the attractive forces (hydrogen bonding) between H_2O molecules as they move closer.

STAGE III

- At this stage, the temperature continues to drop as water in the gaseous state (steam) is fully converted to liquid state. The H_2O molecules continue to move more slowly and less freely with the space between them declining as the average kinetic energy decreases.

STAGE IV

- At this stage the temperature is again constant at 0°C and a plateau is again observed in the curve. This part of the curve depicts that some of the H_2O molecules in the liquid state has been converted to solid phase (ice). Thus, at this stage both the liquid and solid state of water exist.

STAGE V

- At this stage all the H_2O molecules in the liquid state has been converted to solid phase (ice) – freezing. The particles become tightly packed together as the space between them have declined increasing the attractive forces between them.

Heating and Cooling Curves of Naphthalene

Just like water, naphthalene as a system has heating and cooling curves that also can be utilised to represent the changes in matter.

Figure 6: Heating Curve of Naphthalene

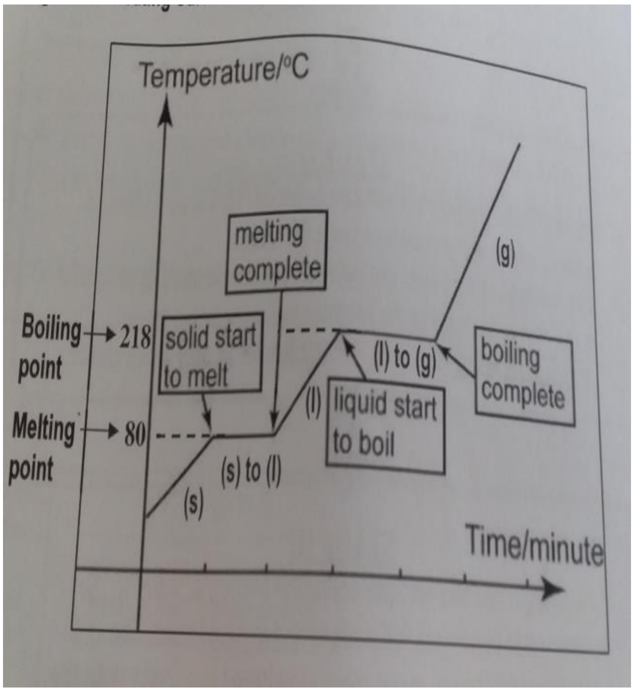
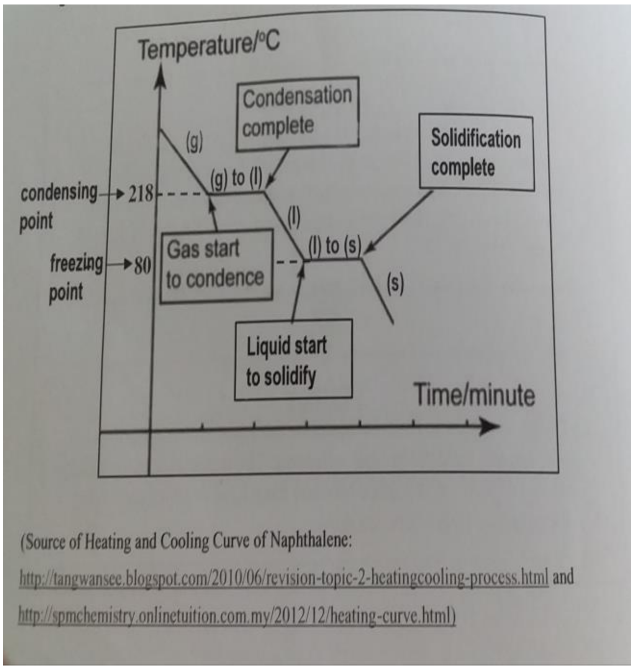


Figure 7: Cooling Curve of Naphthalene



Exercise 9: States of Matter

1. Fill in the blanks (Note: Some of the words can be used more than once.)

Fixed	particles	shape
compressed	Gas	moving

Solids have a ____________ ____________ and its _______ cannot be _________ so easily ________ . Whereas, liquid _______ tend to move ______are known to be further apart and are constantly.

2. Complete the table.

State Change	**Process Name**
a)Solid $\longrightarrow$ Gas	
b)	Vaporisation
c)Liquid $\longrightarrow$ Solid	
d)	Melting
e) Gas $\longrightarrow$ Solid	
	Condensation

3.Explain, what are heating and cooling curves?

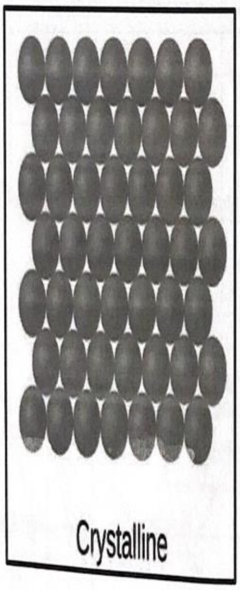
4.The following statements are based on the heating and cooling curves of water. Classify the statements as either true or false by placing a tick.

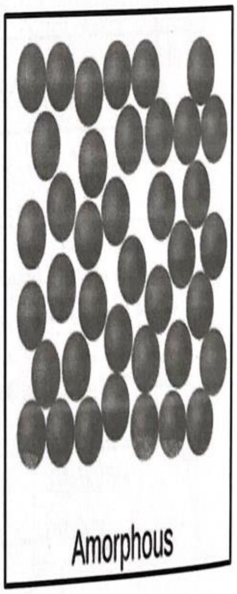
Statement	**True**	**False**
a)The heating curve of water begins with the rapid movement and vibration of solid water particles when heat energy is released from the system.		
b)In stage II of the heating curve of water both solid and liquid water exist at a constant temperature.		
c)The cooling curve of water begins with water in the liquid state.		
d)In the final stage of the heating curve of water, the particles tend to move more slowly.		

Crystalline and Non-Crystalline Solids

i. **Crystalline (regular) solids** – the particles are organised in a pattern that is regular, repeating three-dimensional structure. Some common examples of crystalline solids are diamond, sugar, table salt, ice, iodine sulphur, and metals like copper, iron, magnesium and many more.

ii. **Non-crystalline (irregular) solids (or Amorphous solids)** – the particles are organised randomly and do not have a definite pattern in their structure. Some common examples of amorphous solids are glass, candle wax and plastics.

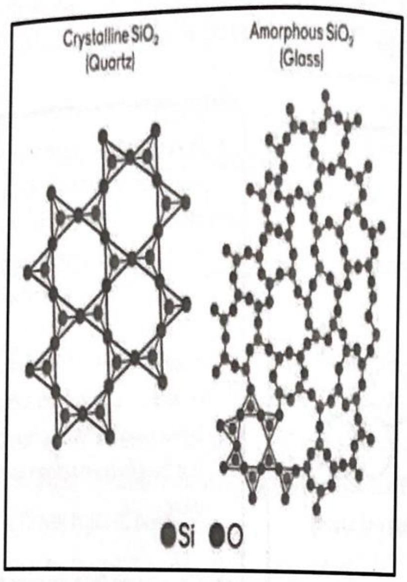




(Source of images: https://chem.libretexts.org/Textbook_Maps/General_Chemistry_Textbook_Maps/Map%3A_Chemistry_(OpenSTAX)/10%3A_Liquids_and_Solids/10.5%3A_The_Solid_State_of_Matter)

Some substances can occur in both crystalline and amorphous forms. A familiar example is silicon dioxide (SiO_2), which occurs naturally as quartz sand with a crystalline structure. When it is melted and quickly cooled, however, it forms glass, which is amorphous, refer to **Figure 9.8**. Glass is also rigid, but, unlike quartz, it breaks into random fragments because its structure does not contain well-defined planes. Many products are made today with safety glass, which has either been heat treated so that it will break into small, granular chunks or layered with plastic so that pieces will stick together if the glass breaks. Common uses for safety glass include car windows, skylights, and shower doors.

Figure 8: Silicon Dioxide in Crystalline and Amorphous Forms



(Source: https://www.learner.org/courses/chemistry/text/text.html?dis=U&num=Ym5WdElURS9OQ289&sec=YzJWaklUQS9NaW89)

In this unit we will be focusing more on crystalline solids.

Crystalline Solids

Hundreds of years ago, people utilised the regular shapes of crystals to provide them with clues of the arrangement of atoms in substances. However, during that time there was a lack to no systematic way of finding out until the technique of **x-ray diffraction analysis** was discovered and was used to determine the positions of atoms in the crystals.

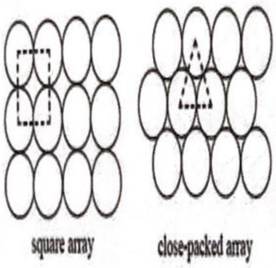
What does the structure of a crystal look like?

Crystalline solids are considered to be **three dimensional arrangement of individual atoms, ions or whole molecules that display a repeating pattern.** These atoms or ions or molecules that crystals are composed of are called **lattice points**. It is these points that form layers which in turn result in crystalline solids. We call the array of points that define the position of particles in the crystal structure a **crystal lattice. Figure 9.9** display the processes or steps involved in the formation of crystalline solids.

Figure 9: Processes in the Formation of Crystalline Solid

Step 1: Lattice point

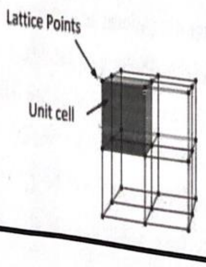
Step 2:
Lattice points then pack together to form two dimensional layers which can either be in the form of a square or closed packed array.



Source:
https://eee.uci.edu/programs/gchem/RDGcrystalstruct.pdf

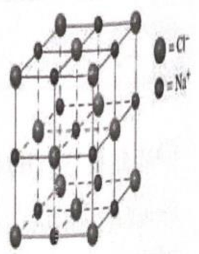
Step 3:
The two dimensional layers of lattice points come together to form a three dimensional structure which we usually call a **unit cell.** The **unit cell** is the smallest repetition of lattice points to form a crystal.

Step 4:
A collection of a large number of unit cells forms the structure of a crystal. Each unit cell stacked together forms a simple cubic lattice. Each of the lattices are stacked in an organised pattern to give the crystalline structure.



- In a crystal lattice, each atom (or ion) is surrounded by a certain number of atoms (or ions). We call this number, the **coordination number (CN)** of a particle within a particular crystal. It gives a measure of how tightly atoms (spheres) are packed together and a higher coordination number signify that the spheres are closer to each other.
- CN is always the same when only one kind of atom exists in a structure (especially in pure metals). But, when two or more different kinds of particles (atoms or ions) exist in a structure (especially in ionic and covalent compounds) the CN for a particular particle is the number of particles of the other kind, which are found next to the given particle.

Example:



Note: 6 Cl^- ions surround one Na^+ ion and vice versa. The ratio of the coordination between Cl^- and Na^+ ions is 6:6.

There is only one type of amorphous solid. However, there are several different types of crystalline solids, depending on the identity of the units that compose the crystal. Crystalline solids can be categorised into four types and these are as follows:

- Ionic solids or crystals;
- Metallic solids or crystals;
- Molecular solids or crystals; and
- Covalent network solids or crystals.

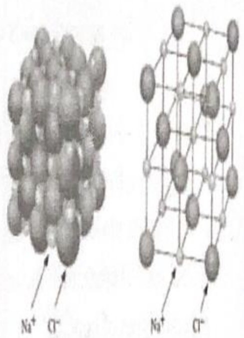
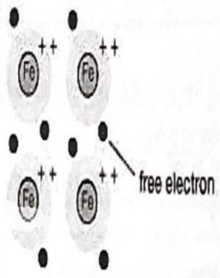
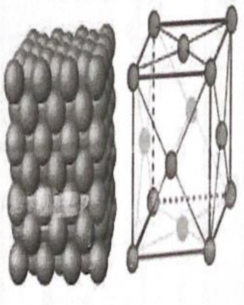
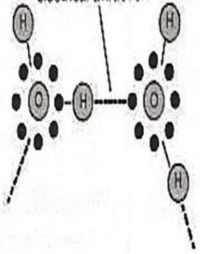
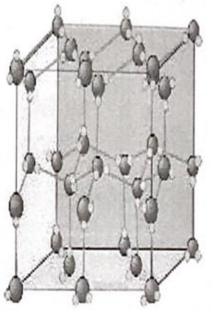
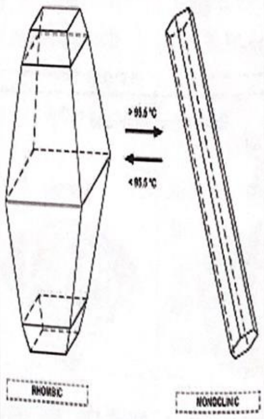
Figure 9.10 outlines the general characteristic of each type enabling one type to be distinguished from the other. As illustrated ionic crystals consist of ions (cations and anions) while molecular crystals are either comprised of polar or non-polar molecules. Whereas, metallic and covalent crystals are both composed of atoms with different types of bonding.

Figure 9.10: General Characteristic of four Type of Crystals

Type of Crystals

Type	Characteristic
Ionic crystals	- Solids that are comprised of **positve and negative ions** which **occupy the crystal lattice points.** - The size of the ions may differ.
Metallic crystals	- A **free atom** of the same element **occupies every lattice point.**
Molecular crystals	- **Atoms or molecules occupy the crystal lattice points.** - These crystals are held and packed together as closely as their shapes and sizes would allow.
Covalent network crystals	- **Atoms** that are **covalently bonded** to other atoms at neighbouring lattice sites **occupy lattice points.** - That is, atoms are held together in an extensive 3-dimensional network entirely by covalent bonds.

Table 2: Bonding Type and Examples of Crystalline Solids.

Crystal type	Bonding	Examples
Ionic crystals	- Ions are connected together by electrostatic attractions (ionic bonds).  ionic bonding electron transferred from Na to Cl (Source: https://www.britannica.com/science/crystal/Types-of-bonds#ref506334)	- All chlorides, oxides, carbonates and sulfides of metals, such as NaCl, CaF_2, $MgCl_2$, MgO, and $NaNO_3$ to name a few.  (Source: https://slideplayer.com/slide/7562643/)
Metallic crystals	- Metal atoms are held together by metallic bonds.  metallic bonding ions surrounded by free electrons Source: https://www.britannica.com/science/crystal/Types-of-bonds#ref506334	- All metallic elements such as Na, Mg, Al, Fe, and Cu, to name a few.  Copper metal crystal (Source: https://chem.libretexts.org/Textbook_Maps/Introductory_Chemistry/Book%3A_Introductory_Chemistry_(CK-12)/08%3A_Ionic_and_Metallic_Bonding/8.11%3A_Crystal_Structure_of_Metals)
Molecular crystals	- Polar molecules held together by dispersion forces, dipole-dipole forces or hydrogen bonding.  molecular bonding weak electrical attraction binds molecules Source: https://www.britannica.com/science/crystal/Types-of-bonds#ref506334	- Water (ice), sucrose, and sulfur dioxide.  - Ice crystals are composed of molecules that are packed in tetrahedral manner. The packing is very open with the result that water at temperatures below 4 °C is denser than ice. (Source: https://chem.libretexts.org/Textbook_Maps/General_Chemistry/Book%3A_Chem1_(Lower)/07%3A_Solids_and_Liquids/7.03%3A_Hydrogen-Bonding_and_Water)
	- Non-polar molecules or atoms held together by dispersion forces.  Rhombic and monoclinic sulfur (Source: http://www.microscopy-uk.org.uk/mag/artmay18/rh-sulphur.pdf)	- Carbon dioxide (dry ice), phosphorous, naphthalene, sulfur, and iodine. - Rhombic sulfur or alpha sulfur is formed at room temperature and is the most stable crystal form. The crystals result from the packing together of ring-shaped sulfur molecules, each consisting of **eight** sulfur atoms, which are linked together. - When cooled liquid sulfur forms needle-like crystals known as **monoclinic or beta sulfur**. This is the stable form of solid sulfur found at temperatures above 96 °C.

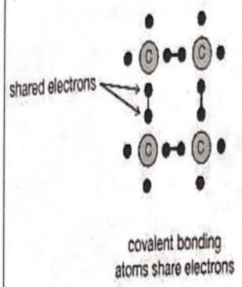
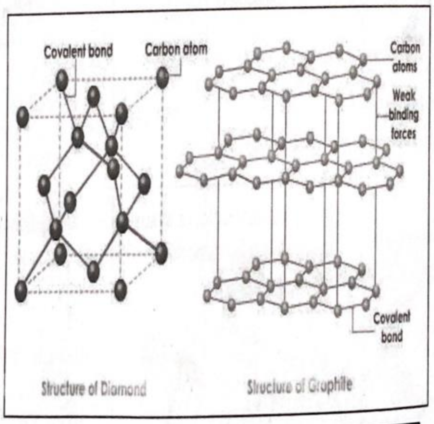
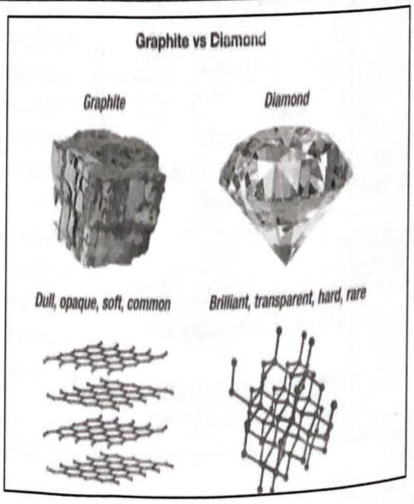
Crystal type	Bonding	Examples
Covalent network crystals	- Atoms are held together by covalent bonds.  Source: https://www.britannica.com/science/crystal/Types-of-bonds#ref506334	- Carbon (diamond and graphite), silicon dioxide or quartz (SiO_2), silicon carbide (SiC), and zinc sulfide (ZnS). - In the diamond crystal, the atoms are packed **tetrahedrally**. The particles are not built in layers but held in a three dimensional network by covalent bonds. - In graphite carbon atoms form layers of 6-member rings. The forces between the layers are very weak (Van der Waals forces), hence the layers can slide very easily over each other.
	  (Source: https://www.lifechanyuan.org/2018/06/25/a-revelation-from-the-structure-of-graphite-and-diamond/)	

Table 3: Physical Properties of Crystalline Solids

Physical Properties	Type			
	Ionic	Metallic	Molecular	Covalent network
Appearance	- Dull surface	- Quite lustrous (shiny)	- Dull surface	- Dull surface
Melting Point	- Ranges from moderate to high or high to very high melting point.	- This would entirely depend on electron configuration. - Ranges from low to very high melting point.	- Atoms or non-polar molecules- very low to moderate melting point. - Polar molecules-low to moderate melting point.	- Very high melting point
Conductivity	- In the solid state it is a non-conductor but in the liquid/molten state it is a good electrical conductor.	- Good conductor of heat and electricity.	- Poor conductor of heat and electricity in both solid and liquid state.	- Poor conductor of heat and electricity*
Ductile and Malleability	- Is hard but brittle as it falls apart when stress is applied.	- Hardness range from soft to very hard. - Is able to be deformed when stress is applied, thus are ductile and malleable.	- Soft	- Very hard and at high temperatures will either sublime or melt.
Density	- Relatively dense	- Usually high density	- Low density	- Low density

*Many exceptions exist. For example, graphite has a relatively high electrical conductivity within the carbon planes, and diamond has the highest thermal conductivity of any known substance.