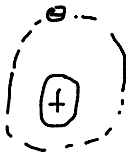
Atoms, Elements and Compounds

Sunday, September 20, 2020 1:55 PM

3.1:
Atomic Structure of the Periodic Table



Atoms consist of protons, electrons and neutrons.

Particle	Mass(amu)	Charge
Proton	1	(+1)
Neutron	1	0
Electron	1/1836	(-)

3 significant figures:
5.02
1.00
28.1

The mass is in terms of **atomic mass unit which means it is relative**. Since the actual mass of protons, neutrons and electrons are incredibly small, they are presented in the manner above

As seen, electrons have very small mass and hence, when calculating the mass of an element, its number of electrons are **negligible.**

Key definitions:

1) Proton number(atomic number): The total number of protons in an atom (if the atom is neutral, the number of protons is equal to the number of electrons)
2) Nucleon number(mass number): (The sum of the number of protons and neutrons)

Key features of the periodic table:

1) Left side is dominantly made up of metals and the right side is made up of non-metals
2) As you go to the right, proton number increases and metallic properties decreases
3) The middle of the periodic table are made up of elements called transition metals
4) Groups are columns and rows are called periods

Isotopes: Isotopes are atoms of the same element which has the same number of protons but different number of neutrons. An example would by C-12 and C-14.

Isotopes have the same chemical properties as since they have same number of protons, they have same number of electrons and hence, **same number of valence electrons.** On the other hand, since they have different masses due to different number of neutrons, they possess different physical properties.

Some isotopes are radioactive (radioisotopes) as they have unstable nucleus and hence, decay to emit radiation. Usage of radioisotopes include:

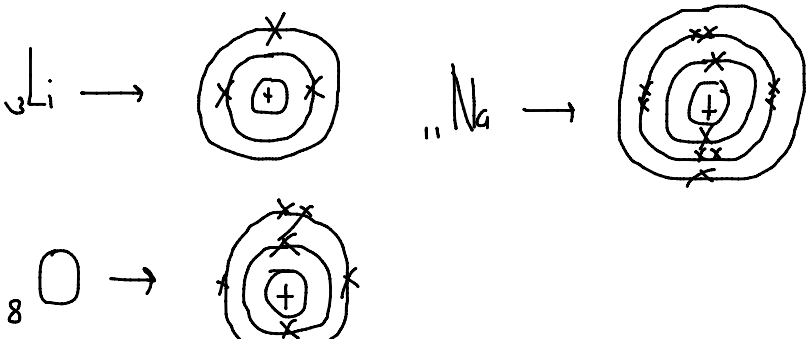
1) Treating cancerous cells as radiation kills cancerous cells (Cobalt-60)
2) Detecting radiation leaks

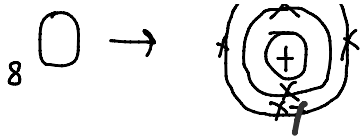
Electrons: Negatively charged particle which orbits the nucleus. In order to remove these electrons, energy is required as there lies attraction between the positively charged nucleus and negatively charged electron.

Electrons fill up shells. Shells are able to accommodate number of electrons in accordance to the formula: $2n^2$

Hence, they fill up in the sequence of: 2,8,18,32

A thing to take note of is that the next shell can only start occupying electrons once the current shell is full. Examples are as of below:





In order to electrons to be stable(to not be able to react), they need to completely fill their last shell. In order to achieve this stability, elements react with each other or react with other elements by either gaining or losing electrons.

Valence electrons are electrons in the outermost shell. This dictates how an element reacts and its chemical properties.

Elements in group 8 or group 0 are called **noble gasses or inert**. This means they are unreactive as their valence shell is completely occupied.

More reactive elements have a greater urge to lose or gain electrons to achieve that full electron shell for stability and hence, form more stable compounds/molecules.

3.2:
Structure and Bonding

3.2.1:
Bonding and the Structure of Matter

Key definitions:
Elements: Made up of one type of atom only.
Extra info: Cannot be separated or divided into smaller components with chemical means. (Physically, it can be divided into proton, neutron and electron)

Molecules: Made up of one or more element/s covalently bonded

Compounds: Made up of more than 1 type of element chemically bonded.

All molecules are compounds but not all compounds are molecules.

Mixtures: substances that are physically combined (methods to separate these mixtures are seen in the last chapter)

Metals vs Non-Metals:

Metals	Non-Metals
strong	brittle
High melting point	Low melting point
Good conductor of heat and electricity(presence of free moving electrons)	Poor conductor of heat and electricity
High density	Low density
Form basic oxides (sodium oxide)	Form acidic oxides (carbon dioxide)
Mostly malleable, ductile and sonorous	
Can be magnetic(Iron, nickel, steel and cobalt)	
Lose electrons in chemical reactions	Gain electrons in chemical reactions
Normally shiny	Normally dull in solid state

Strong: The ability to deal with deformation when load is applied
Malleable: Can be deformed to a different shaped without being broken
Ductile: Can form wires
Sonorous: A resonant sound that is produced when metals are struck

Alloys are mixtures of metals with other metals or other elements. This is done in order to make the metal either have a better appearance or be less malleable(not deformed easily).

This works as when metals are mixed with other substances, the empty spaces in between its atoms are filled by these other substances, making it less malleable.

Common examples:
1) Bronze : Copper and Tin
2) Brass : Copper and Zinc
3) Steel: Iron and carbon

3.2.2:
Ions and Ionic Bonding

Ions are charged particles. These particles are charged due to the gain or lose of electrons. Note that protons and neutrons cannot be lost. Simply said, energy to remove electrons is a lot less.

This is way beyond the syllabus but if you are curious:
The reason why lies in the fact that in order to remove electrons, the energy required is way, way, way less than the energy required to 'remove' a proton or neutron. Electrons are free to move around the nucleus although it is bonded by electrostatic force (when different charges attract) to the nucleus. However, inside the nucleus, there lies a force called strong nuclear force which bonds the neutrons and protons very tightly together. Now, the question arises as to why the electrons are not affected. This is due to the fact that strong nuclear force acts at a very small distance only. Not to mention, protons are electromagnetically shielded by electrons and hence to remove a proton, one must remove all electrons first.

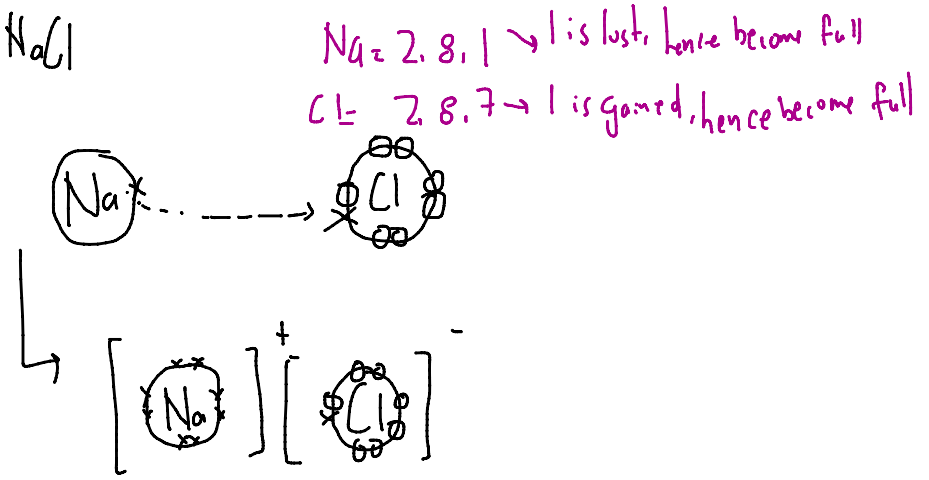
When atoms lose an electron, this means that there are now 1 more proton compared to electron and hence, it is a positive ion. Vice versa is true.

If you notice, elements at the left side of the periodic table have valence electrons that are less than or equal to 4. This means that metals have a tendency to lose electrons instead of gain electrons in order to achieve a full electron shell.

Take an example you are in a gym and you carrying dumbbells. You suddenly see the rule that you are only allowed to carry 10,20,30 pieces. Now, if you are a non metal and you are carrying, lets say, 18 pieces, it is easier to carry 2 more than to drop 1 by one 8 dumbbells. If you are a metal and you are carrying 11 dumbbells, it is a lot easier to drop one then to carry 9 dumbbells one by one.

With that being said, when a metal and non metal reacts with each other, electrons are being shared in order to achieve a full outermost electron shell.

Hence, **ionic bonding is the transfer of electrons between non-metals and metals where both are attracted to each other by strong electrostatic force.**
Dot and Cross Diagram (show valence electrons only)



This type of compounds are called ionic compounds. They have unique properties.

Property	Reason
High melting point and boiling point	Strong electrostatic force between charged ions
Can conduct electricity in molten or aqueous state only	Ions are free to move
Giant lattice (ordered pattern/arrangement of ions)	Attraction between positive and negatively charged ions
Usually soluble in water (NaCl)	Beyond syllabus

3.2.3:
Molecule and Covalent Bonding

Molecules are covalently bonded and may comprise of 1 or more type of elements.

In a covalent bond, electrons is being shared instead of lost/gained in order to achieve noble gas configuration. This type of bonding can only take place between non-metals.

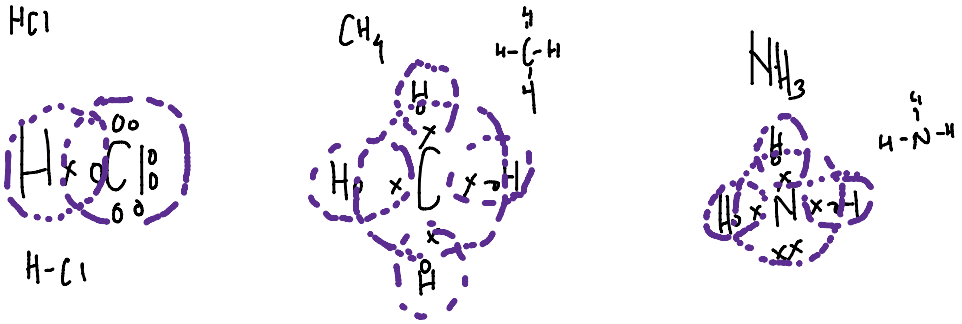
A distinct feature of some elements in the periodic table is that, in nature, they are present as a molecule instead of element. What is unique is that these molecules are made up of the same type of element. They are called diatomic molecules. They are:

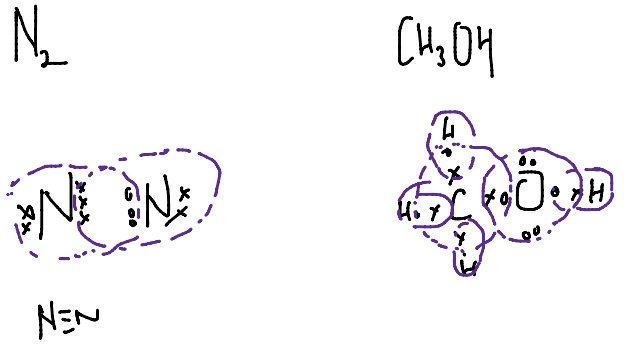
1) Oxygen
2) Nitrogen
3) Hydrogen
4) Fluorine
5) Chlorine
6) Iodine
7) Bromine

O_2, N_2, I_2, Cl_2

Simple molecules have weak inter-molecular forces but strong intra-molecular force(covalent bond itself). Take an example of water. The bond between 1 molecule of water to another is weak but the bond within that one molecule itself(bond between hydrogen and oxygen).

Formation of covalent bonds:





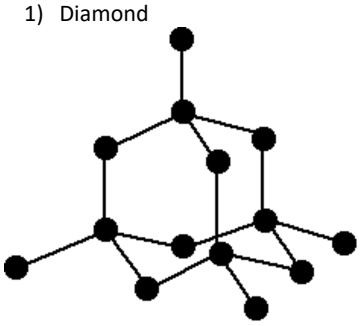
Property	Reason
Low m.p and b.p	Weak intermolecular force between molecules
Poor conductor of electricity and heat	No free flowing electrons or ions
Most cases, insoluble in water	Beyond syllabus

Ionic	Covalent
High m.p and b.p due to strong electrostatic force of attraction	Low m.p and b.p due to weak force of attractions
Good conductor of electricity in aqueous or liquid state due to presence of free flowing ions	Poor conductor of heat and electricity in any state
Usually soluble in water	Usually insoluble in water
Not volatile	Volatile

3.2.4:
Macromolecules

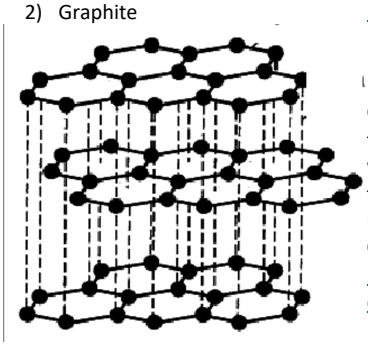
Macromolecules are very large molecules. In this syllabus, 3 are studied. These structures are also called giant covalent.

1) Diamond



(Tetrahedral)
Diamond is an allotrope of carbon. Each carbon atom is covalently bonded to 4 other carbon atoms hence there are not free moving electrons or ions. This explains why it poorly conducts electricity and heat. It has very high melting and boiling points as well as being the hardest natural substance. This is due to the nature of the molecule where there are no weak bonds and each carbon atom is covalently bonded to 3 other atoms. Hence, it is used as a cutting tool as well as in jewelry
Misconception alert: Most students get confuse when covalent bonds are mentioned as being very strong. Do not get confused with intermolecular forces and intra molecular forces. Intermolecular force is the covalent bond itself where an atom bonds with another atom whereas intramolecular force is the bonding with one molecule to another. Since diamond are all covalently bonded to each other, it has very high m.p and b.p.
https://www.chemguide.co.uk/atoms/structures/giantcov.html

2) Graphite



Graphite, just like diamond, is an allotrope of carbon. One carbon atom is bonded to 3 other carbon atoms hence, there lies a free moving electron. This explains why carbon can conduct electricity well. Graphite has high m.p and b.p as well as the covalent bonds within the layers are very strong but between the layers, there lies weak van der wall forces. This explains why graphite can act as a lubricant. Graphite is used to make pencils.
https://www.researchgate.net/figure/The-atomic-structure-of-graphite-The-dashed-lines-indicate-the-weak-connection-between_fig1_3811818

3) Silicon dioxide

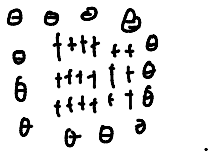
(Tetrahedral)
Makes up sand. Each silicon atom is bonded to 4 oxygen atoms and each oxygen atom is bonded to 2 silicon atoms. Similar to diamond, each atom is covalently bonded, hence, has very high m.p and b.p. There are not free moving electrons and hence, poorly conducts electricity.

http://www.chm.bris.ac.uk/motm/silica/silicah.htm

3.2.5:
Metallic bonding

Metallic bonding is when positive ions are surrounded by a sea of delocalized electrons. This bonding explains a lot of properties that metals possess.

1) High melting and boiling point due to strong electrostatic force between positive ions and sea of delocalized electrons
2) Conduct electricity and heat well due to free flowing electrons
3) Is malleable and ductile as layers of ions can slide over each other



Exams tips:
1) Understand fully properties of metals and non metals
2) Be able to differentiate inter and intramolecular forces
3) Explain why graphite is a lubricant and can conduct electricity

4) Be able to explain why metals are conductors of electricity
5) Be able to draw diagrams for all 3 bondings
6) Ionic compounds can only conduct electricity at molten or aqueous state due to free moving **ions** not electrons