Intensive notes (Topic 2: Microscopic world I)

Atom

Element is a substance that is made up of only one kind of atoms.			
Atom is the smallest part of an element which has the chemical properties of that element.			
Subatomic particle	**Proton**	**Neutron**	**Electron**
Relative mass	1	1	Negligible
Charge	+1	0	-1
Position	Nucleus	Nucleus	Moving around nucleus

Note 1: An atom is mostly empty space. **Mass is concentrated in the nucleus** of an atom.

Note 2: Atom is electrically neutral because the **number of protons** is equal to the **number of electrons**

Atomic symbol

Atomic number (No. of protons)	Atom	Atomic symbol	Electronic arrangement	Atomic number (No. of protons)	Atom	Atomic symbol	Electronic arrangement
1	Hydrogen*	H	1	11	Sodium#	Na	2,8,1
2	Helium*	He	2	12	Magnesium#	Mg	2,8,2
3	Lithium#	Li	2,1	13	Aluminium#	Al	2,8,3
4	Beryllium#	Be	2,2	14	Silicon^	Si	2,8,4
5	Boron^	B	2,3	15	Phosphorus*	P	2,8,5
6	Carbon*	C	2,4	16	Sulphur*	S	2,8,6
7	Nitrogen*	N	2,5	17	Chlorine*	Cl	2,8,7
8	Oxygen*	O	2,6	18	Argon*	Ar	2,8,8
9	Fluorine*	F	2,7	19	Potassium#	K	***2,8,8,1***
10	Neon*	Ne	2,8	20	Calcium#	Ca	***2,8,8,2***

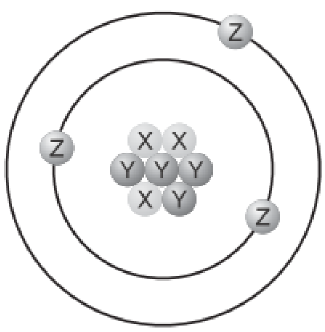
#: Metal ^: Semi-metal *: Non-metal

Note 1: Top 5 most abundant elements on Earth are O > Si > Al > Fe > Ca.

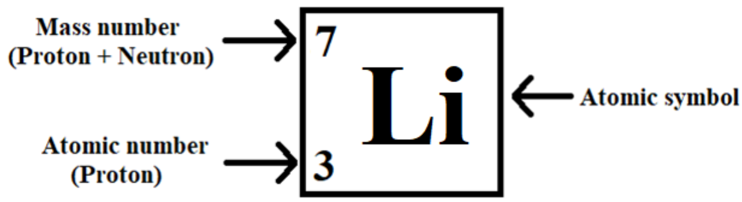
Note 2: Maximum number of electrons a shell can hold can be calculated by $2n^2$, where n is the shell number. Therefore the outermost shell of Ar (2,8,8) is NOT FULLY filled (3^{rd} shell can hold 18 electrons)

Note 3: Each element has its own unique atomic number and atomic symbol.

Full atomic symbol



- Z represents electron (Z moves around nucleus)
- X represents proton (X is inside the nucleus, and the numbers of X and Z are equal)
- Y represents neutron (Y is inside the nucleus)



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Intensive notes (Topic 2: Microscopic world I)

Isotopes

Isotopes are different atoms of the element with **same number of protons** but **different number of neutrons**
Relative atomic mass is the **weighted average** of the relative isotopic masses of all of its naturally occurring isotopes of the ^{12}C = 12.00 scale

Note: Protons and electrons exist in ALL atoms, but not neutrons (e.g. $^{1}_{1}H$ atom does not have neutrons)

Example 1

If the abundance of ^{35}Cl and ^{37}Cl are 75% and 25% respectively, calculate the relative atomic mass of Cl.

Solution: Relative atomic mass = (35) (75%) + (37) (25%)

= 35.5 (No unit!)

Example 2

Magnesium has 3 isotopes, ^{24}Mg, ^{25}Mg and ^{26}Mg. If the abundance of ^{24}Mg is 78.6% and the relative atomic mass of Mg is 24.327, calculate the abundance of ^{25}Mg.

Solution: Let x% be the abundance of ^{25}Mg.

24.327 = (24)(78.6%) + (25)(x%) + (26) (100% – 78.6% – x%)

24.327 = 18.864 + (25)(x%) + 5.564 – (26)(x%)

x = 10.1

The abundance of ^{25}Mg is 10.1%

Example 3

Bromine has 2 isotopes, ^{79}Br and ^{x}Br. If the abundance of ^{79}Br is 55% and the relative atomic mass of Br is 79.9, calculate the isotopic mass of ^{x}Br.

Solution: 79.9 = (79) (55%) + (x) (100% – 55%)

x = 81

Note: Isotopes have **SAME chemical properties** because they are **atoms with same electronic arrangement**. They CANNOT be separated by chemical methods. In addition, isotopes have **slightly different physical properties** (e.g. mass). Therefore they CAN be separated by physical method.

Classification of elements

Property	Metal	Non-metal
State	Solid (Except mercury)	Solid or gas (Except bromine)
Appearance	Shiny	Dull
m.p. / b.p.	Usually high	Usually low (Except graphite / diamond)
Hardness	Hard and strong	Brittle
Malleable / Ductile	Yes	No
Density	Usually high	Usually low
Electrical Conductivity	Good conductor	Poor conductor (Except graphite)

Note: Boron and silicon are classified as **metalloid / semi-metal.** They have properties of both metals and non-metals. They conduct electricity under high temperature / with impurities. (Silicon is used in making computer chips)

Intensive notes (Topic 2: Microscopic world I)

Periodic Table

Atomic number → $_{1}$H ← Atomic symbol
1 ← Electronic arrangement

I	II											III	IV	V	VI	VII	0
																	$_{2}$He 2
$_{3}$Li 2,1	$_{4}$Be 2,2											$_{5}$B 2,3	$_{6}$C 2,4	$_{7}$N 2,5	$_{8}$O 2,6	$_{9}$F 2,7	$_{10}$Ne 2,8
$_{11}$Na 2,8,1	$_{12}$Mg 2,8,2											$_{13}$Al 2,8,3	$_{14}$Si 2,8,4	$_{15}$P 2,8,5	$_{16}$S 2,8,6	$_{17}$Cl 2,8,7	$_{18}$Ar 2,8,8
$_{19}$K 2,8,8,1	$_{20}$Ca 2,8,8,2	$_{21}$Sc 2,8,9,2	$_{22}$Ti 2,8,10,2	$_{23}$V 2,8,11,2	$_{24}$Cr 2,8,13,1	$_{25}$Mn 2,8,13,2	$_{26}$Fe 2,8,14,2	$_{27}$Co 2,8,15,2	$_{28}$Ni 2,8,16,2	$_{29}$Cu 2,8,18,1	$_{30}$Zn 2,8,18,2	$_{31}$Ga 2,8,18,3	$_{32}$Ge 2,8,18,4	$_{33}$As 2,8,18,5	$_{34}$Se 2,8,18,6	$_{35}$Br 2,8,18,7	$_{36}$Kr 2,8,18,8

- In modern Periodic Table, elements are arranged in **increasing atomic number. (NOT mass number)**

  Period = Number of occupied electron shells Group = Number of electrons in outermost shell

- Across a period, elements change from metals through semi-metal to non-metals.

Note 1: Hydrogen does NOT belong to any group

Note 2: In Group 0, **Helium has 2** outermost shell electrons while **other elements have 8**

Note 3: Elements between Group II and Group III are **transition metals**

Group I: Alkali metals	Group II: Alkaline earth metals
- Reactivity **increases down the group.** - Melting point **decreases down the group** - Soft metals (can be cut by a knife) and low densities (float on water) - Very reactive (store in paraffin oil to prevent reaction with oxygen) - React with water to give metal hydroxide (alkaline solution) and hydrogen gas	- Reactivity **increases down the group.** - Melting point **decreases down the group** (except Mg) - Denser than Group I elements - Less reactive than Group I elements - React with water less vigorously than Group I elements to give metal hydroxide and hydrogen (except Be) - React with hydrochloric acid to give metal chloride and hydrogen
Group VII: Halogens	**Group 0: Noble gases**
- Reactivity **decreases down the group.** - Melting point **increases down the group** - All react with metals to give ionic compounds - All react with non-metals to give covalent compounds - All are toxic and colored (elements turn darker in color down the group) - All can be decolorized by adding sodium sulphite solution.	- All are colorless gas at room conditions - All are **very unreactive**, no reaction with other elements. (Noble gases are **unreactive** since they have an **octet structure (duplet in helium)** in their outermost shell.)

F:	Pale yellow gas	**Cl:**	Greenish yellow gas	**He:**	Fill balloons and airships	**Ne:**	Fill advertising signs	**Ar:**	Fill light bulbs
Br:	Brown liquid	**I:**	Black solid						

- Elements in the same group have **SIMILAR chemical properties** because they have **same number of outermost shell electrons.**

Intensive notes (Topic 2: Microscopic world I)

Ionic bond

Ionic bond is a strong, non-directional electrostatic attraction between oppositely charged ions. It is formed by the transfer of one or more electrons from one atom to another

Formation of ion: Atoms tend to attain the electronic arrangement of the nearest noble gas (stable octet/duplet structure) by gaining or losing electrons (metal) to form an ion.

Metals: Having 1/2/3 outermost shell electrons, tend to lose electrons to form positive ion (cation)

Non-metals: Having 5/6/7 outermost shell electrons, tend to gain electrons to form negative ion (anion)

Formulae and names of ions

Cation (Positively charged ion)

Charge	Formula	Name	(Color)
1+	Na^{+}	Sodium ion	
	K^{+}	Potassium ion	
	H^{+}	**Hydrogen ion**	
	NH_4^{+}	**Ammonium ion**	
	Ag^{+}	**Silver ion**	
	Cu^{+}	Copper(I) ion	
	Hg^{+}	Mercury(I) ion	
2+	Mg^{2+}	Magnesium ion	
	Ca^{2+}	Calcium ion	
	Ba^{2+}	Barium ion	
	Zn^{2+}	**Zinc ion**	
	Pb^{2+}	**Lead(II) ion**	
	Mn^{2+}	Manganese(II) ion	***(Pale pink)***
	Fe^{2+}	Iron(II) ion	***(Green)***
	Co^{2+}	Cobalt(II) ion	***(Pink)***
	Ni^{2+}	Nickel(II) ion	***(Green)***
	Cu^{2+}	Copper(II) ion	***(Blue)***
	Hg^{2+}	Mercury(II) ion	
3+	Al^{3+}	Aluminium ion	
	Fe^{3+}	Iron(III) ion	***(Yellow)***
	Cr^{3+}	Chromium(III) ion	***(Green)***

Anion (Negatively charged ion)

Charge	Formula	Name	(Color)
1-	H^{-}	**Hydride ion**	
	OH^{-}	Hydroxide ion	
	Cl^{-}	Chloride ion	
	Br^{-}	Bromide ion	
	I^{-}	Iodide ion	
	NO_3^{-}	Nitrate ion	
	NO_2^{-}	Nitrite ion	
	HCO_3^{-}	Hydrogencarbonate ion	
	HSO_4^{-}	Hydrogensulphate ion	
	CN^{-}	Cyanide ion	
	ClO_3^{-}	Chlorate ion	
	ClO^{-}	Hypochlorite ion	
	MnO_4^{-}	Permanganate ion	***(Purple)***
2-	O^{2-}	Oxide ion	
	S^{2-}	Sulphide ion	
	SO_4^{2-}	Sulphate ion	
	SO_3^{2-}	Sulphite ion	
	$S_2O_3^{2-}$	Thiosulphate ion	
	CO_3^{2-}	Carbonate ion	
	SiO_3^{2-}	Silicate ion	
	CrO_4^{2-}	Chromate ion	***(Yellow)***
	$Cr_2O_7^{2-}$	Dichromate ion	***(Orange)***
3-	N^{3-}	Nitride ion	
	P^{3-}	Phosphide ion	
	PO_4^{3-}	Phosphate ion	

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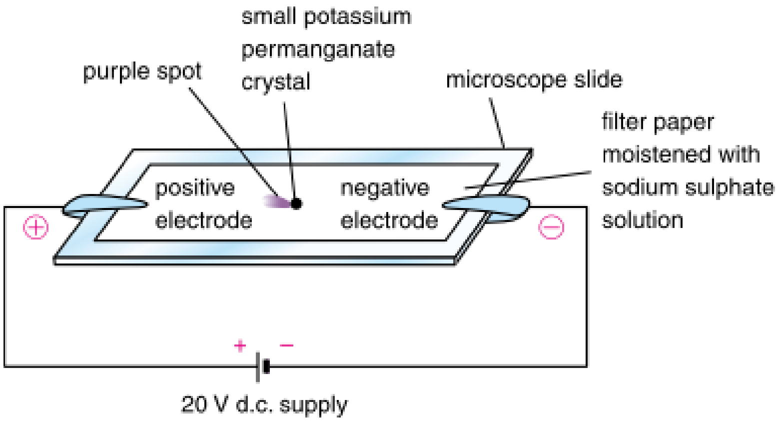
Experiments showing migration of ion

Experiment 1: Using filter paper



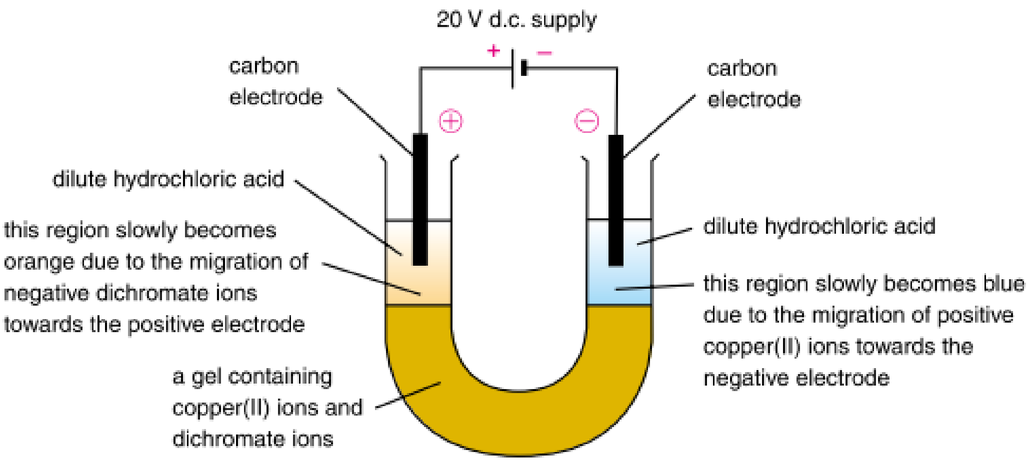
Experimental procedure:

1. Moisten the filter paper with dilute sodium sulphate solution and place it on top of a microscopic slide.
   (Sodium sulphate solution can increase the electrical conductivity of the filter paper by providing mobile ions)
2. Put a small crystal of potassium permanganate crystal at the center on the filter paper
3. Connect each side of the filter paper using electrical wires to the positive and negative pole of the d.c. power supply respectively by crocodile clip.
4. Observe the migration of colored ions after a few minutes. Purple color should move towards positive pole because MnO_4^-, a purple anion, will be attracted by positive pole

Note 1: If the paper is not connected to d.c. supply, the purple color patch will diffuse to all directions.

Note 2: To speed up the experiment, higher voltage can be applied.

Experiment 2: Using U-tube



- Dilute hydrochloric acid can increase the electrical conductivity of the filter paper by providing mobile ions
- Larger amount of compound can be used in this method ➔ Enable easier observation
- The gel can slow down the mixing of the bottom layer in the U-tube with the top layer

Intensive notes (Topic 2: Microscopic world I)

Covalent bond

Covalent bond is the strong directional electrostatic attraction between the shared electrons and the two nuclei of the bonded atoms. It is formed by sharing of outermost shell electrons between two atoms
Molecule is the smallest part of an element or a compound which can exist on its own under room conditions

Dative covalent bond

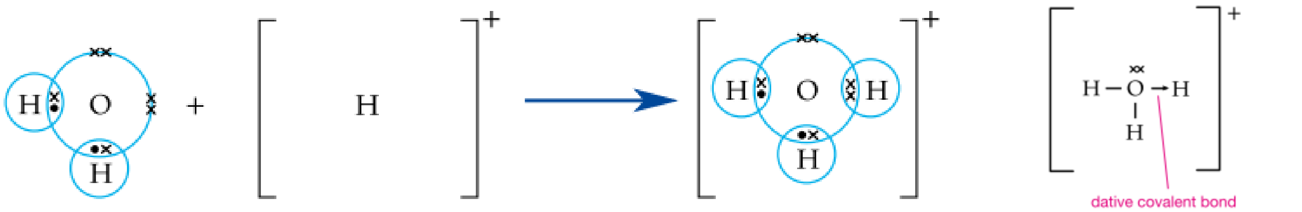
Dative covalent bond is a covalent bond formed between two atoms where both electrons of the shared pair are contributed by the same atom

Q: ***Describe the formation of dative covalent bond in hydronium ion.***

A: H_2O has 2 lone pairs of electrons on O atom.

H^+ does not have any electron in its outermost shell. It can accept 2 more electrons to attain a stable duplet structure.

Dative covalent bond is formed when H_2O donates its lone pair electrons on O to H^+.



Other examples:

Structural formula of H_3O^+

$NH_3 + H^+ \rightarrow NH_4^+$	$NH_3 + BF_3 \rightarrow H_3NBF_3$	$F^- + BF_3 \rightarrow BF_4^-$	CO
$O_2 + O \rightarrow O_3$	$N_2 + O \rightarrow N_2O$	$AlCl_3 + AlCl_3 \rightarrow Al_2Cl_6$	$n\ BeCl_2 \rightarrow (BeCl_2)_n$

Name of covalent compounds

1. Element occurs first in the following series:

   B, Si, C, P, N, H, S, I, Br, Cl, O, F

   (Formula of water is H_2O, NOT OH_2)
2. **Mono (1), di (2), tri (3), tetra (4), penta (5), hexa (6)** can indicate the number of atoms of that element in the compound
3. Name of the second element ends up with **-ide** (e.g. carbon dioxide, not carbon dioxygen)
4. Some hydrogen-containing compounds have special names

   CH_4: Methane **NH_3: Ammonia** **H_2O: Water** **H_2O_2: Hydrogen peroxide**

Summary of chemical bonding

	Ionic bond	Metallic bond	Covalent bond
Nature	**Non-directional** electrostatic attraction between: 1. Cations 2. Anions	**Non-directional** electrostatic attraction between 1. Sea of delocalized electrons 2. Positive metal ions	**Directional** electrostatic attraction between: 1. Shared electrons 2. Two nuclei of the bonded atoms
Formation	Transfer of electron(s) from one atom (group of atoms) to another	Escape of outermost shell electrons from metal atoms	Sharing of outermost shell electrons between two atoms

Electron diagram practice

(a) Sodium chloride ($NaCl$)

(b) Aluminium oxide (Al_2O_3)

(c) Calcium sulphide (CaS)

(d) Magnesium hydride (MgH_2)

(e) Chlorine (Cl_2)

(f) Nitrogen (N_2)

(g) Methane (CH_4)

(h) Water (H_2O)

(i) Methanal (CH_2O)

(j) Nitrosyl chloride ($NOCl$)

(k) Hydrogen peroxide (H_2O_2)

(l) Dinitrogen difluoride (N_2F_2)

(m) Ethene (C_2H_4)

(n) Hydrazine (N_2H_4)

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Intensive notes (Topic 2: Microscopic world I)

Structure and properties of different substances

Structure	**Giant ionic**	**Simple molecular**	**Giant covalent**	**Giant metallic**
Particles	Cations, Anions	Molecules	Atoms	Metal atoms
Bonding	**Strong ionic bond (electrostatic attractions)** between cations and anions	**Strong covalent bond** between atoms **weak van der Waals' force** between molecules	**Strong covalent bond** between atoms	**Strong metallic bond (electrostatic attraction)** between metal cations and delocalized electrons
Examples	- Compounds made of metal and non-metal - Ammonium compounds	- Non-metals - Compounds made of non-metals - Compounds made of non-metals and semi-metals	- Silicon dioxide (quartz) - Graphite - Diamond - Silicon - Boron	- Metals
State	Hard solid	Gas, liquid, soft solid	Hard solid	Hard solid (except **group I** and **Hg**)
Melting point and boiling point	High due to **strong ionic bond** between cations and anions	Low due to **weak van der Waals' force** between molecules	Very high due to **strong covalent bonds** between atoms	High due to **strong metallic bond** between metal ions and delocalized electrons (except group I metals and Hg)
Solubility	- Soluble in **polar solvents** (water) - Insoluble in non-polar solvents	- Usually soluble in **non-polar solvent**	Usually insoluble in both polar and non-polar solvents	Metals do not dissolve in polar and non-polar solvents, except those can **react with water**
Electrical conductivity	Conducts electricity only in **aqueous** state or **liquid** state (Conduct electricity by mobile ions)	Non-conductor since they are neutral molecules (except those **acids/alkalis** that can **ionize** in water to give mobile ions)	Non-conductor (except **graphite** as it can conduct electricity by delocalized electrons)	Conducts electricity only in **solid** state or **liquid** state (Conduct electricity by delocalized electrons)

Substances that can conduct electricity:

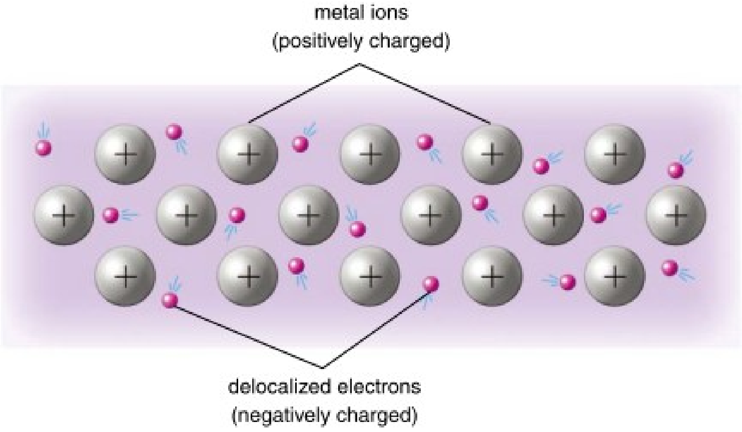
1. Metals (A conductor, conduct electricity by **delocalized electrons without** decomposition)
2. Graphite (A conductor, conduct electricity by **delocalized electrons without** decomposition)
3. Aqueous / molten salts (An electrolyte, conduct electricity by **mobile ions with** decomposition)
4. Aqueous acids / alkalis (An electrolyte, conduct electricity by **mobile ions with** decomposition)

Giant metallic structure

All metals have giant metallic structure. Outermost shell electrons of metal atoms are delocalized

Giant lattice of positive metal ions are surrounded by a sea of delocalized electrons

- Metallic bond is the strong, non-directional electrostatic attraction between the delocalized electrons and metal ions

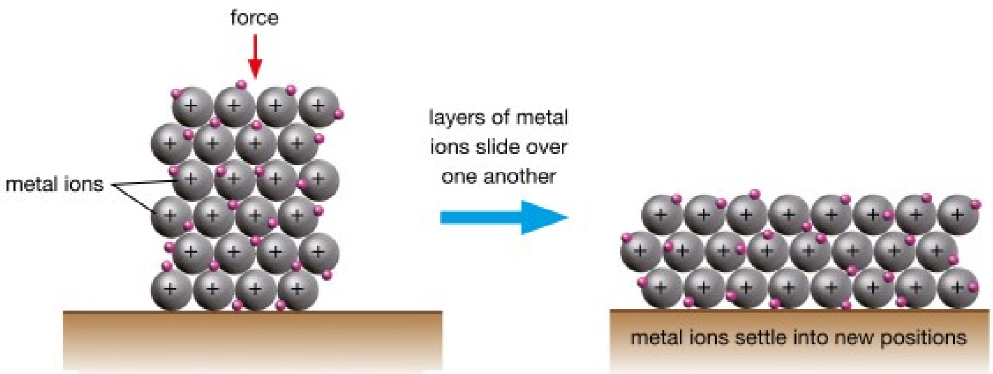


Properties of substances with giant metallic structure:

- Good conductor of heat and electricity ➔ Due to the presence of <u>delocalized electrons</u>
- High melting point ➔ <u>Metal ions and delocalized electrons</u> are held by <u>strong metallic bond</u>
- High density ➔ Metal ions are packed closely

Q: Why are metals malleable and ductile?

A: Metals ions and delocalized electrons are held by strong metallic bond. When force is applied, non-directional metallic bond can still hold the metal atoms together in new position.



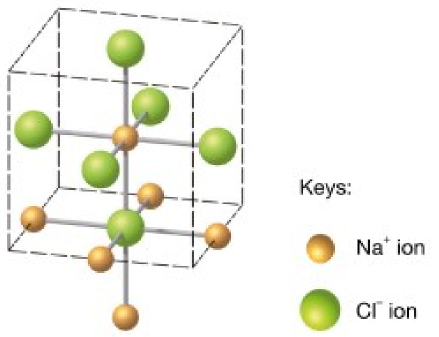
Q: How do metals conduct electricity?

A: In a piece of metal, delocalized electrons move in all directions. When both ends of the metal are connected to a battery, the delocalized electrons move in one direction only.

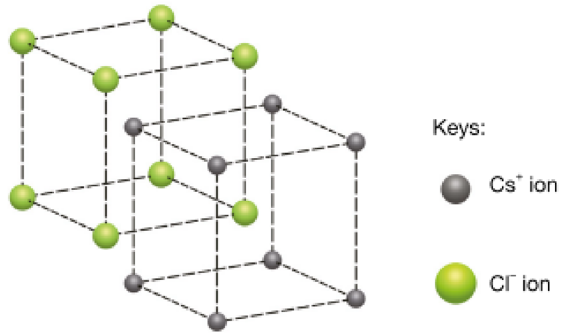
Q: Why is there no chemical change when metal metals conduct electricity?

A: Delocalized electrons move towards the positive pole of the battery, leaving the metal. At the same time, equal number of electrons move into the other end of the metal from the negative pole. Number of electrons in the metal remains unchanged at any moment.

Giant ionic structure



Sodium chloride



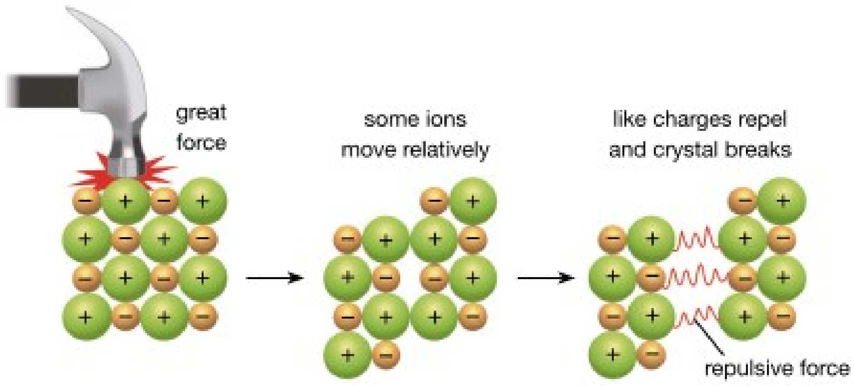
Caesium chloride

Properties of substances with giant ionic structure:

- Conductor of electricity in aqueous state and molten state ➔ Due to the presence of mobile ions
  (Ions in solid state are NOT mobile. Therefore ionic compounds do not conduct electricity in solid state)
- High melting point ➔ Cations and anions are held by strong ionic bond

Q: Why ionic compounds are brittle?

A: The oppositely charged ions are held together by strong ionic bonds. When force is applied, the relative movement of the ions brings ions of the same charge close to each other. This will result in repulsion and the crystal will break. Therefore, ionic compounds are brittle



Q: Using sodium chloride as an example, explain why ionic compounds are soluble in water but insoluble in organic solvent

A: When solid sodium chloride is added to water, attraction exists between ions in sodium chloride and water molecules. This attraction causes the sodium ions and chloride ions to move away from the crystal and go into the water. Then water molecules surround the ions. The ions are said to be hydrated. No such attraction exists between the ions in sodium chloride and molecules of organic solvents. Thus, sodium chloride is insoluble in organic solvents.

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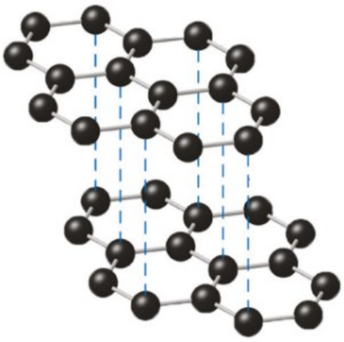
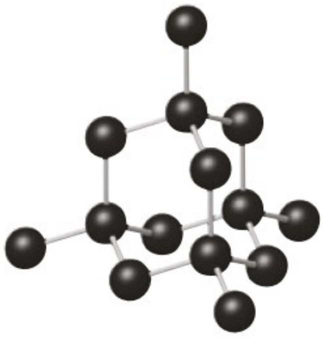
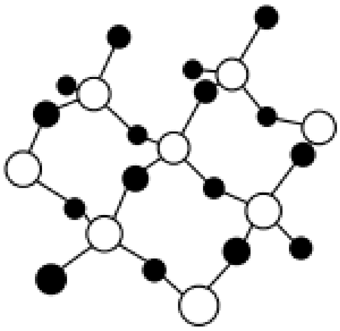
Giant covalent structure



Graphite



Diamond



Silicon dioxide (Quartz)

Properties of substances with giant covalent structure:

- Non-conductor ➔ Absence of delocalized electrons and mobile ions
  (Except graphite which can conduct electricity by delocalized electrons)
- High melting point ➔ Atoms are held by strong covalent bond

Q: Explain why graphite is brittle while diamond is hard

A: In graphite, graphite layers are held together by weak van der Waals' forces. Graphite layers can be separated easily.
In diamond, carbon atoms are held together by strong covalent bonds. The relative movement of atoms is restricted.

Q: Explain why graphite can conduct electricity while diamond cannot.

A: In graphite, each carbon atom is covalently bonded to 3 other carbon atoms in its layer and 1 outermost shell electron is delocalized. These delocalized electrons can move freely within a graphite layer
In diamond, each carbon atom is covalently bonded to 4 other carbon atoms. There are no delocalized electrons in diamond.

Q: Explain why substances with giant covalent structure cannot dissolve in both polar and non-polar solvents.

A: It is difficult to break the covalent bonds and separate the atoms.

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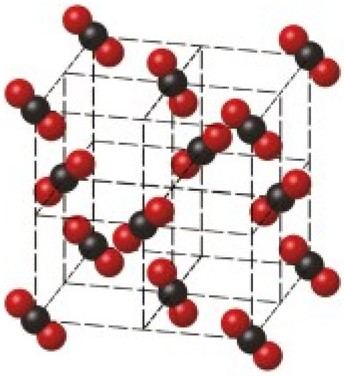
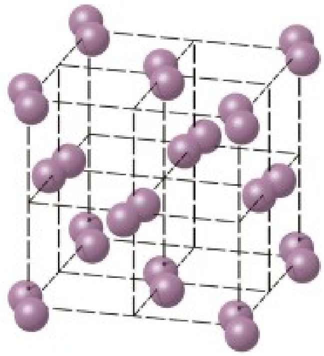
Simple molecular structure



Carbon dioxide



Iodine

Properties of substances with simple molecular structure:

- Non-conductor ➔ Absence of delocalized electrons and mobile ions
  (Except acids and alkalis which can conduct electricity in aqueous state)
- Low melting point ➔ Molecules are held by weak van der Waals' forces
  Note: When melting substances with simple molecular structures (e.g. methane), only weak van der Waals' forces between molecules are going to overcome. The strong covalent bonds between atoms will NOT be broken.

Q: Explain why substance with simple molecular structure are brittle

A: As the molecules are held together by weak van der Waals' forces.

Q: Using iodine as an example, explain why substance with simple molecular structure are generally soluble in organic solvents but not water.

A: The attraction between water molecules is quite strong (due to the presence of strong hydrogen bonds). When iodine is added to water, the weak attraction (van der Waals' forces) between iodine and water molecules is NOT strong enough to overcome the attraction between the water molecules. Iodine molecules and water molecules therefore cannot mix together.
In heptane (an organic solvent), heptane molecules are held together by weak attraction (van der Waals' forces).
When iodine is added to heptane (an organic solvent), iodine and heptane molecules can mix together readily.

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