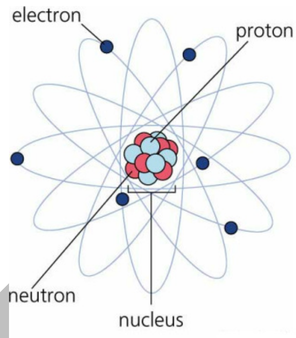
atomic structure

1. What is an atom?

   An atom is the smallest unit of an element.



2. What is inside the atom?

   Subatomic particles - Protons, neutrons and electrons.

3. Where can you find these subatomic particles?

   Protons and neutrons (together known as Nucleons)can be found in the **nucleus**.

   Electrons can be found orbiting around the nucleus in the **electron shells**.

4. What are properties of the subatomic particles?

<u>Particle</u>	<u>Relative charge</u>	<u>Relative mass</u>
Proton	+1	1
Electron	-1	1/1840
Neutron	0	1

5. What do the 2 numbers assigned to each atom in the periodic table represent?

   Atomic (proton) number: Number of protons in an atom.

   This number is element specific - different elements have different proton numbers.

   Since atoms are electrically neutral, this is also equal to the number of electrons in an atom.

   Mass (Nucleon) Number: Total number of protons and neutrons in the atom.

   Example:

   $^{19}_{9}F$ An atom of fluorine will have 9 protons, 9 electrons and 10 (19 – 9) neutrons.

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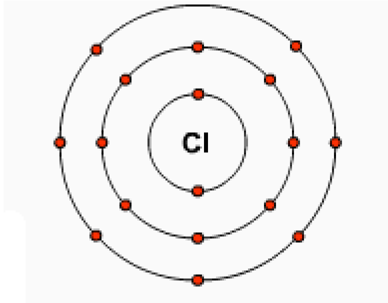
6. What is the electronic structure of an atom?

It shows us how electrons in an atom are arranged in their electron shells. Electrons are arranged according to the following rules.



Chlorine atom

The **1st shell**, closest to nucleus, can contain a max. of **2 electrons**

The **2nd shell** can contain a max. of **8 electrons**

The **3rd shell** can contain a max. of **8 electrons**.

(O-level syllabus)

E.g. Notice how the chlorine atom (it has 17 protons & 17 electrons) has an electronic configuration of **2. 8. 7**?

Note: As no two elements in the periodic table have the same number of protons, no elements will have the same electronic structure!

7. Electrons and their electron shells…

Valence Shell: The electron shell that is furthest from the nucleus – outermost shell.

Valence electrons: Electrons that are found in the valence shell. *(Determines the group an element is placed in the periodic table.)*

Elements with **1** valence electron are in **Group I**.

Elements with **2** valence electrons are in **Group II** etc, etc.

8. How does knowledge of the electronic structure relate to the Periodic Table?

All elements of the **same Group** have the **same number of valence electrons**, and hence undergo similar chemical properties!

All elements of the **same Period** have the **same number of electron shells**!

9. Why do some elements have mass numbers that are not whole numbers?

Isotopes of the element, present in different amounts, exist. The mass number is an average of the atomic masses of all the different isotopes present.

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10. What are isotopes?

Isotopes are **atoms of the same element** with the **same number of protons** but **different number of neutrons**.

Example:

Isotopes	Mass Number	Atomic Number	Neutrons	Protons	Electrons	Elec. Structure	Valence Electrons
Chlorine 35	35	17	**18**	17	17	2, 8, 7	7
Chlorine 37	37	17	**20**	17	17	2, 8, 7	7

From the above, we can deduce the following:-

1. Same number of i. protons, ii. electrons, iii. valence electrons
2. Different number of neutrons.

Hence, both atoms will have **SAME chemical** properties but **DIFFERENT physical** properties.

11. How do we calculate the mass number of isotopes?

Example:

For $^{35.5}_{17}Cl$

Chlorine 35: 75%

Chlorine 37: 25%

Mass number = (0.75 * 35) + (0.25 * 37) = <u>**35.5**</u>

Example

There are 20 atoms of Boron-11 and 80 atoms of Boron-10. Calculate the relative atomic mass of Boron.

Relative atomic mass = [(20 x 11) + (80 x 10)] ÷ 100

= <u>**10.8**</u>

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ionic bonding, covalent bonding and metallic bonding

1. Why do atoms undergo chemical bonding?

   Other than Noble Gases, the atoms of other elements are chemically unstable.

   In order to achieve stability, a Noble gas configuration, atoms can either gain/lose or share valence electrons.

2. What is a Noble gas configuration/Stable electronic configuration?

   An atom is considered stable if it has an octet (8 valence e-) configuration or a duplet (2 valence e-) configuration.

   Stable electronic configurations:

   Helium: *2*

   Neon: *2, 8*

   Argon: *2, 8, 8*

Ionic Bonding

3. What happens when an atom gains/loses valence electrons?

   Atoms that have gained/lost electrons form charged particles called **ions**.

   In general, atoms of **Metals lose valence electrons** to form positively charged ions known as Cations.

   Atoms of **Non-metals gain valence electrons** to form negatively charged ions called Anions.

Example:

$^{24}_{12}Mg \rightarrow Mg^{2+} + 2e^-$

Magnesium has 12 electrons. Its electronic configuration is 2, 8, 2.

In order for it to obtain a stable electronic configuration, it will lose 2 valence electrons to form 2, 8.

$^{19}_{9}F + e^- \rightarrow F^-$

Fluorine has 9 electrons. Its electronic configuration is 2, 7.

In order for it to obtain a stable electronic configuration, it will gain 1 valence electron to form 2, 8.

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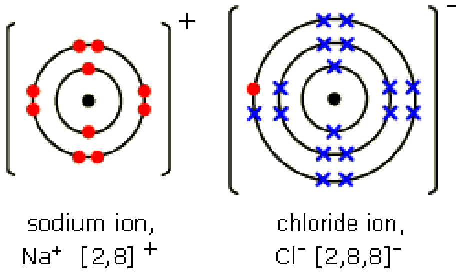
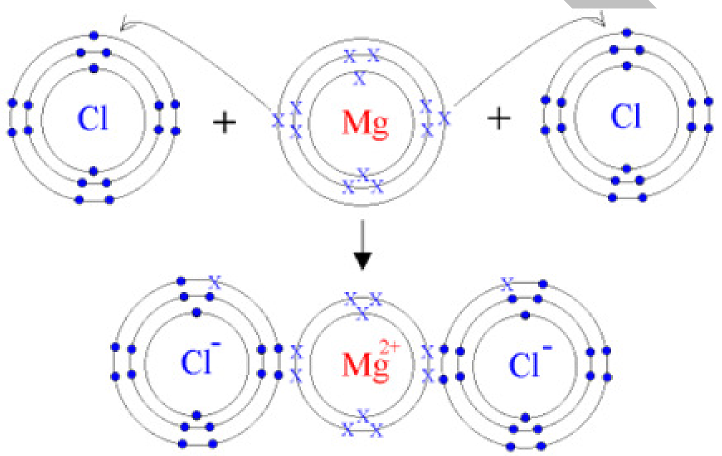
4. How is an **<u>ionic compound</u>** formed?

When a positively charged ion is placed close to a negatively charged ion, something happens. Oppositely charged ions attract. This attraction force is known as **Electrostatic Forces of Attraction**, which make the **ionic bond**!

In general, atoms of Metals and Non-metals will combine chemically to form ionic compounds.





Notice how Magnesium loses 2 valence electrons. And how these valence electrons are gained by 2 chlorine atoms?

5. **Some important things to take note of

a. Common ions and their charges

<u>Cations</u>		**<u>Anions</u>**	
Name	**Formula**	**Name**	**Formula**
Sodium	Na^{+}	Chlorides	Cl^{-}
Ammonium	NH_4^{+}	Hydroxides	OH^{-}
Calcium	Ca^{2+}	Nitrates	NO_3^{-}
Magnesium	Mg^{2+}	Oxides	O^{2-}
Zinc	Zn^{2+}	Carbonates	CO_3^{2-}
Lead	Pb^{2+}	Sulfates	SO_4^{2-}
Copper	Cu^{2+}	Phosphates	PO_4^{3-}
Silver	Ag^{+}		
Iron	Fe^{2+} / Fe^{3+}		

b. Writing chemical formula of ionic compounds

The net charge of an ionic compound is equal to zero. This means that the number of positive charges and negative charges must add to zero.

Example:

Zn^{2+} & Cl^{-}

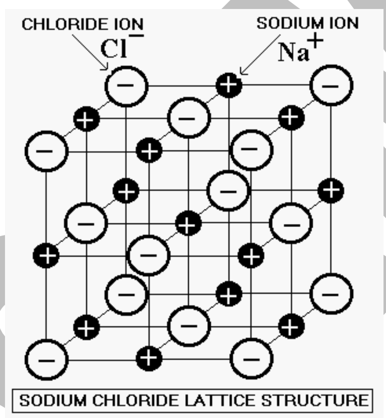
In order for the compound to have a net 0 charge, 2 chloride ions are required to bond with 1 zinc ion. Hence the formula of the compound is $ZnCl_2$.

Mg^{2+} & NO_3^{-}

In order for the compound to have a net 0 charge, 2 nitrate ions are required to bond with 1 magnesium ion. Hence the formula of the compound is $Mg\ (NO_3)_2$.

6. What is the structure of an ionic compound?

Ionic compounds are made up of ions that are arranged regularly in a giant ionic lattice structure. It can also be referred to as a giant crystal lattice and is hard.



7. What are some properties of ionic compounds?

These are found in the table below.

SAMPLE

Covalent bonding

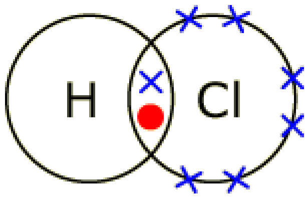
8. What about compounds that are made up of only non-metals? How do they obtain stable electronic configuration?

Valence electrons are shared between atoms of non-metals. This forms the covalent bond.

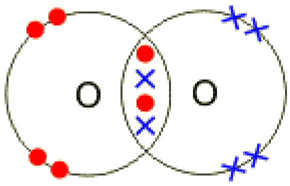
*If only 1 pair of electrons is shared, it is called a **single covalent bond**.*

*If 2 pairs of electrons are shared, it is called a **double covalent bond**.*

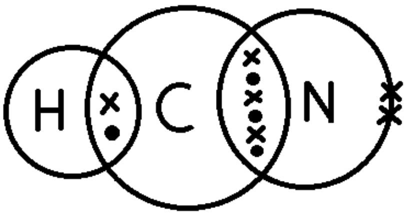
*If 3 pairs of electrons are shared, it is called a **triple covalent bond**.*



Single bond



Double bond



Triple bond

Notice how each atom obtains either a duplet or octet configuration by sharing their valence electrons?

9. How many types of covalent substances are there?

1. Simple covalent
2. Giant macromolecular

10. What are some properties of covalent compounds?

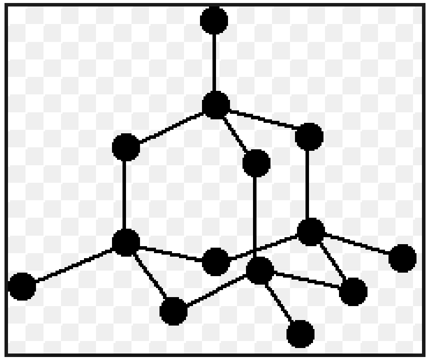
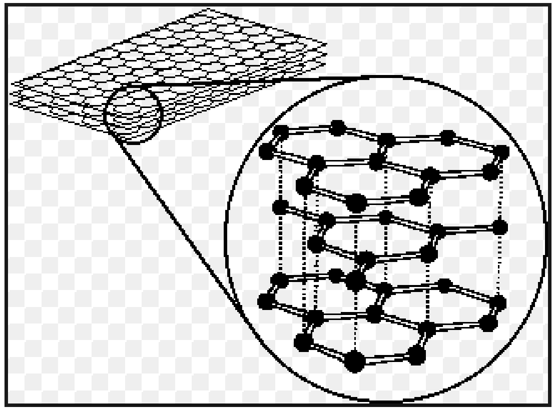
These are found in the table below.

SAMPLE

chemical bonding quick revision guide

Properties of Ionic Compounds	SIMPLE covalent Molecules and their properties
1. **Crystalline solids & Hard** 2. **High m.p and b.p **** Large amount of thermal energy needed to overcome strong electrostatic forces of attraction between oppositely charged ions. !!Note: the **larger the charges** of the positive and negative ions, the **stronger the ionic bonds** → higher m.p! 3. **Low volatilities** Cannot evaporate easily because of the strong electrostatic forces of attraction between ions. Note: Volatility is inversely proportional to b.p. => High b.p ; Low volatility. Vice versa. 5. **Conduct electricity in aq. and molten states **** Free moving ions. Can carry current. Note: in solid state, ions are held in fixed positions, hence cannot conduct electricity. 6. **Mostly soluble in water; insoluble in organic solvents ****	1. **Low mp and bp. **** Weak intermolecular forces of attraction (Van Der Waals') between molecules, hence little heat energy needed to overcome. !!Note: the intramolecular bonds (covalent bonds) are very strong. This is DIFFERENT from the intermolecular forces of attraction. 2. **Volatility** High volatility → Evaporate easily. Weak VDW forces between molecules. Diffuse easily. 3. **Solubility **** Mostly insoluble in water, soluble in organic solvents. Note: there are exceptions. Hydrogen chloride, glucose and ammonia are covalent compounds but are highly soluble in water. 4. **Lack of electrical conductivity. **** Simple covalent molecules DO NOT conduct electricity in any state as they lack free moving ions and mobile electrons. Note: exception is graphite. It conducts electricity.

giant molecular structures

Properties	**Diamond** (similar to Silicon dioxide, SiO_2)	**Graphite**
Structure	All carbon atoms connected to each other by very strong covalent bonds in a 3 dimensional structure. Each carbon atom is bonded to 4 other carbon atoms by very strong covalent bonds. 	Flat layers of carbon atoms. Each carbon atom is connected to 3 other atoms by strong covalent bonds in a 2 dimensional hexagonal structure. 
Electrical conductivity	Diamond does not conduct electricity Each carbon atom uses up its 4 valence electrons for bonding with 4 other carbon atoms. No free electrons to move and conduct elec.	Graphite can conduct electricity In each layer, a carbon atom is bonded to 3 other carbon atoms, leaving 1 valence electron free. These electrons are delocalized, can travel along the layers, and conduct electricity.

Hardness or softness	Giant molecular structure held together by very strong covalent bonds resulting in a rigid tetrahedral structure. A lot of energy required move each atom hence diamond is hard.	Carbon atoms are held together by very strong covalent bonds in hexagonal layers. However, intermolecular forces of attraction between layers is weak, hence layers can slide past one another, making graphite soft.
Melting point	Very high melting point Structure Giant molecular structure with millions of carbon atoms held together by very strong covalent bonds that extend throughout the network. A large amount of heat energy is needed to break these bonds -> high m.p.	Very high melting point. Carbon atoms held together by strong covalent bonds in hexagonal layers. A lot of heat energy needed to break these bonds -> high m.p.

chemical bonding and atomic structure

Model answers for questions that keep coming out!

Questions 1,2 and 3 are typical questions on ionic compounds. It can be applied to any other type of ionic compound.
Questions 4 and 5 are questions on covalent compounds. Similarly, it can be applied to any other covalent compound.
Questions 6-9 are on atomic structures, electrons etc.

1. Explain why sodium chloride has a very high melting point.

Sodium chloride is an ionic compound.

A lot of heat energy is needed to overcome the strong ionic bonds (electrostatic forces of attraction) between the oppositely – charged **ions** (Na^+ and Cl^-), to separate the ions during melting, so sodium chloride has a high melting point.

2. Explain why sodium chloride will conduct electricity when molten (aqueous).

When molten (aqueous), the **ions** are free to move throughout the liquid (solution). The **mobile ions** carry their **charges** with them as they move throughout the liquid (solution), thus allowing sodium chloride to conduct electricity.

3. Explain why sodium chloride does not conduct electricity when solid.

When solid, the ions (Na^+ and Cl^-) are held in fixed positions by strong ionic bonds.

The ions are not able to move in the solid, so there are no mobile ions to carry the current, so sodium chloride does not conduct electricity when solid.

4. Explain why carbon dioxide has a low boiling point.

Carbon dioxide is a simple covalent compound.

Little heat energy is needed to overcome the weak intermolecular forces of attraction between the carbon dioxide **molecules** during boiling, so carbon dioxide has a low boiling point.

5. Explain why carbon dioxide does not conduct electricity in any state.

Carbon dioxide is a covalent compound.

There are no free moving ions or mobile electrons (charge carriers) to conduct electricity.

6. Explain why the melting/boiling point of chlorine is lower than iodine?

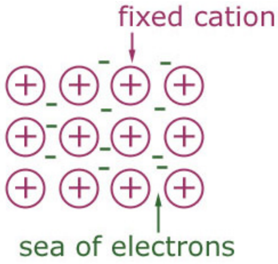
The size of a chlorine molecule is smaller than the size of an iodine molecule. As molecular size increases, the intermolecular forces of attraction (Van Der Waal Forces of attraction) increase. Hence melting point and boiling points increase.

(This can be used to describe why smaller covalent molecules have lower melting and boiling points compared to larger covalent molecules)

metallic bonding

1. What is the metallic bond?

   The strong attractive force between the positively charged metal ions and sea of delocalized electrons.



2. How does this affect properties of metals?

   i. Metals have high melting points.

   The electrostatic forces of attraction between the positively charged metal ions and the sea of delocalized electrons are very strong. A very large amount of thermal energy is needed to break these bonds.

   ii. Metals are good conductors of electricity

   Metals have mobile delocalized valence electrons within the metal lattice structure. These are able to travel from the negative terminal to the positive terminal of the electrical circuit.

   iii. Metals are good conductors of heat.

   The mobile delocalized electrons conduct heat. They are able to travel and collide with the electrons nearby.

   iv. Metals are malleable and ductile.

   The sea of delocalized electrons "glues" the atoms in the metal together, preventing the metal from fracture.