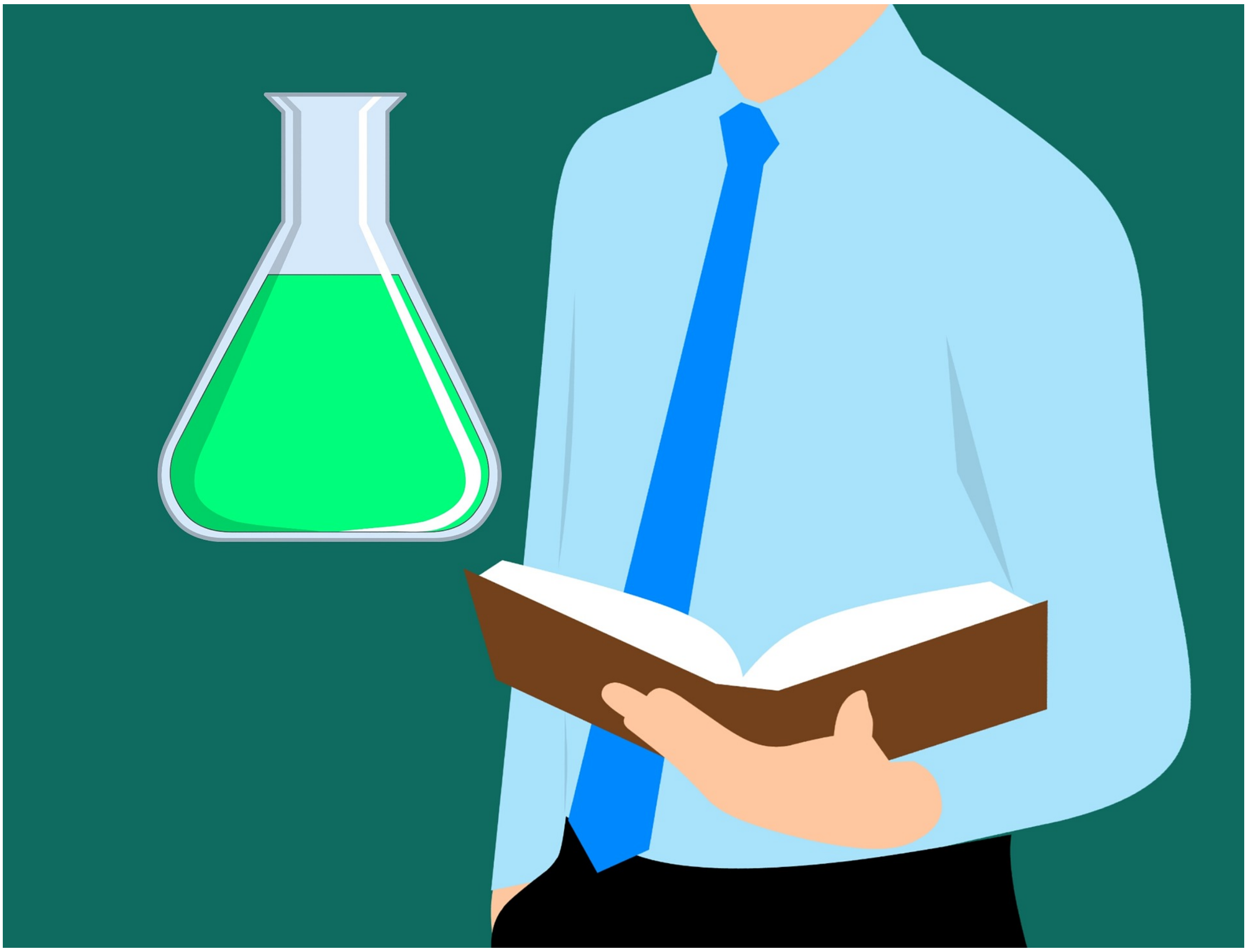

O level Chemistry Notes (Page 2)

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O level Chemistry Notes covering Bonding and Periodic Trends. These notes only show the key points in O'level chemistry curriculum.

Chemical Bonding

There are three main types of chemical bonding:

- Ionic bonding – usually when a metal bonds with a non-metal.
- Covalent bonding – usually when a non-metal bonds with a non-metal.
- Metallic bonding – usually when a metal bonds with a metal.

Ionic bonding (electrovalent)

Ionic bonding occurs when a metal bonds with a non-metal and outer shell electrons of a metal are completely transfered from the metal to the non-metal, giving charged ions that electrostatically attract each other to form a bond.

An ionic bond is defined as the electrostatic attraction between oppositely charged ions.

For example when sodium combines with chlorine, an electron is transferred completely from Na to Cl. If magnesium (2.8.2) combines with chlorine, the magnesium has to lose both its outer shell electrons to two chlorine atoms because each chlorine atom can only accept one electron.

Similarly, when K (2.8.8.1) combines with S (2.8.6), two K atoms each lose one electron each, and one S atom gains two electrons.

Giant Ionic Substances

Once the ions are formed, they attract one another. A sodium ion attracts negative chloride ions from all directions to form a regular giant ionic lattice. Ionic solids have lots of strong ionic bonds that require a lot of energy to break. Therefore they have the following properties:

- high melting point: NaCl melts at 801
- they don't conduct electricity when solid, but they do conduct when molten or in solution, since the ions become free to move and can carry charge and undergo electrolysis.
- usually soluble in water. Water is a polar molecule (with one negative end and one positive end) and can cluster around the ions, allowing them to separate, and so overcome the strong attractive electrostatic forces which hold the lattice together.

Covalent Bonding

When two non-metal atoms combine they both need to gain electrons, and they can do this by sharing two electrons in a covalent bond. A covalent bond is defined as the electrostatic attraction between the positively charged protons in the nucleus and the negatively charged shared electrons.

In carbon dioxide, carbon (2.4) needs to form four bonds, and oxygen (2.6) needs to form two, so two double-bonds result (O=C=O). The covalent bond is strong, but it binds two specific atoms together (unlike the ionic attractions, which occur in all directions).

Simple Molecular Structures

A molecular structure consists of small molecules, with weak forces of attraction (intermolecular forces) between molecules.

When a molecular substance is melted or boiled, it is only necessary to provide a small amount of energy to break these weak attractions, so they have low melting points and boiling points. Molecular substances are gases, liquids, or low-melting solids at room temperature.

They usually share the following properties:

- low melting points (melting only involves breaking the weak attraction between molecules).
- low boiling points (like melting)
- do not conduct electricity in solid, or when melted, or in solution, as they have no charged particles.
- often dissolve in non-polar solvents, like hexane; usually insoluble in water.

Giant Covalent structures

If a non-metal atom can form three or four bonds, it is possible for it to form giant structures linked by covalent bonds. There are two forms of carbon which have giant structures.

In diamond each atom is covalently bonded to four neighbours, and each of those to three others, and so on throughout the whole crystal.

Graphite consists of layers of hexagons with strong covalent bonds holding each C atom to its three neighbours.

Both diamond and graphite have very high melting points. When elements are found to exist in more than one crystalline form they are referred to as allotropes. Diamond and Graphite are therefore allotropes of Carbon.

Metallic bonding

In metallic bonding metals give up their outer electrons to be shared with all their neighbours. These "delocalised" electrons form a mobile "sea" of electrons which flows between the positively charged ions (nuclei).

The positive ions themselves pack as tightly as possible in a giant structure bound by oppositely charged electrons.

A metallic bond is therefore defined as the electrostatic attraction between the positively charged metal ions and the negatively charged delocalized electrons.

Giant metallic substances have the following properties:

- They conduct electricity because the electrons are free to flow between the ions and carry the electrical charge.
- They are malleable and ductile because the ions can slide over each other, but continue to attract each other strongly in their new positions, so that the metallic bonds do not break but distorts instead.
- They have high melting points because the metallic bonds are strong and require a lot of energy to break.

The Periodic Table

Elements in the Periodic Table are put in order of increasing atomic number and arranged according to electronic structure.

Group

The chemical properties of elements depend on the number of electrons in the outer shell, so we place them in vertical groups which all have the same number of electrons in the outer shell. After element 20 the electron arrangement becomes more complicated, but still the number of electrons in the outer shell of any element is equal to the group number.

Hydrogen is not normally placed in any of the main groups due to the fact that is has one shell and one electron in that shell which is supposed to be full with 2 electrons.This makes hydrogen behave:

1. as a group 1 element because it has 1 outer electron.
2. as a group 7 element because it has a shortage of 1 electron in the outer shell like group 7 elements.

Period

A horizontal row in the table is called a period. Elements in the same period as the same number of shells.

Periodic Trends

If the elements are listed in order of atomic number, similar elements appear at regular intervals. This is a periodic property. There is a regular pattern of properties across one period, and a similar pattern across the next period.

The most obvious pattern is the change from reactive metals on the left (Group 1), through less reactive elements in the middle, to increasingly reactive non-metals in Group 7, followed by the very unreactive gases in Group 0.

Soluble oxides of metals are alkaline, and soluble oxides of non-metals are acidic, there is also a pattern of alkaline oxides on the left giving way to increasingly acidic oxides across the period.

Within the same group elements are generally very similar, though they may show a regular trend in their properties. The similarity occurs because the atoms have the same number of electrons in their outer shell which determines their chemical nature.

Group 1 (Alkali Metals)

This group contains Li (lithium), Na (sodium), K (potassium), Rb (rubidium), Cs (caesium).

These are called the alkali metals because they form alkaline oxides. They all have one electron in their outer shell. They are all reactive metals, with a valency (combining power) of 1.

They have the following physical properties:

- They are silver, shiny metals when freshly cut, though they tarnish rapidly and are kept under oil to protect them from air and water.
- They have unusually low densities for metals (Li, Na and K will float on water). These densities increase down the group.
- They become softer going down the group: lithium is quite hard, sodium is as soft as cheese, and caesium is like putty.
- The melting points decrease going down the group they are all unusually low for metals.

They react with water as follows:

They react with water with increasing vigour down the group.

- Lithium floats, reacts quite vigorously, and fizzes giving off hydrogen.
- Sodium melts to a ball, fizzes around the surface bubbling vigorously giving off hydrogen.
- Potassium reacts violently, melting to a silver ball, catching fire to burn with a mauve flame, and fizzing around.
- Caesium explodes.

When they react with water, metal hydroxide (not oxide) and hydrogen are formed.

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

Down group the atoms become larger and the outer shell electron is held less strongly since it is further away (and shielded from) from the attraction of the protons in the nucleus. It becomes easier to lose it to form positive ions and so the elements become more reactive.

Compounds of group 1 elements have the following properties:

They are almost all white, crystalline solids (all ionic) which are soluble in water.

Group 7 (Halogens)

Group 7 contains the elements F (fluorine), Cl (chlorine), Br (bromine), I (iodine). They all have seven electrons in their outer shells. Edexel IGCSE Chemistry Revision Notes They are reactive non-metals, with a valency (combining power) of 1 since they have a shortage of 1 electron in the outer shell.

They are all diatomic:

- F_2
- Cl_2
- Br_2
- I_2

Physical Properties of halogens:

Element	F_2	Cl_2	Br_2	I_2
State	gas	gas	liquid	solid
Melting point	–220	–101	–7	114
Boiling point	–188	–35	59	184
Colour	yellow	green/yellow	red/brown	purple/black

With increasing atomic number (down the group):

- melting points and boiling points increase (because the attractive forces between the molecules increase as the molecules get larger).
- they become darker. Solid iodine looks black, but its vapour is purple, and it is purple when dissolved in organic solvents like hexane.

Chemical properties of halogens

Since fluorine is extremely reactive and dangerous to use, we shall consider the other three members.

- All will bleach dyes, like litmus, though chlorine is rapid, bromine slow, and iodine is so slow that it needs warming.
- All will react with most metals, on warming, to form salts:

$2Na + Cl_2 \rightarrow 2NaCl$ (sodium chloride)

$Zn + I_2 \rightarrow ZnI_2$ (zinc iodide)

- Each halogen will displace the ones below it in the group (i.e. the less reactive ones) from their salts. So chlorine will displace bromine from bromides, and iodine from iodides; bromine will displace iodine from iodides.

We can show this with an ionic equation:

$Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(l)$ [solution turn reddish-brown due to the displaced Bromine]

$Cl_2(g) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(l)$ [solution turn purple due to the displaced Iodine]

- Down the group, the atoms become larger and the protons in the nucleus attracts a new electron less strongly, since the outer shell is further away (and more strongly shielded). Therefore halogens become less reactive going down the group (opposite of Group 1).

Properties of halogen compounds (halides):

- Sodium chloride, sodium bromide and sodium iodide are all white, crystalline solids (NaCl, NaBr and NaI) which are soluble in water. Their solutions react with silver nitrate solution to form a precipitate of the corresponding silver salt (which is insoluble):

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3$ (aq) [forms a white precipitate of (AgCl)]

- Hydrogen halides (HCl, HBr and HI) are prepared by direct combination between the halogen and hydrogen:

e.g. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ [Hydrogen chloride gas]

- Hydrogen halides are all colourless gases which are very soluble in water. They have a simple molecular structure (small molecules, e.g. HCl). They dissolve in water to form strong acids (hydrochloric acid, HCl; hydrobromic acid, HBr; and hydriodic acid, HI).

e.g. $HCl(g) \rightarrow HCl$ (aq) [Hydrochloric acid]

In methyl benzene the hydrogen chloride HCl does not split into H^+ ions. So a solution of hydrogen chloride in methyl benzene is not acidic. Hydrochloric acid is only an acid if water is present because in water HCl dissociates into H^+ ions and Cl^- ions.

Group 0 – The Inert (Noble) Gases

Also known as group 8, this group includes He (helium), Ne (neon), Ar (argon), Kr (krypton), Xe (xenon). They are all unreactive gases that have full outer shells of electrons (2 for He, 8 for all the others), which is why they don't react with other elements.

Physical properties of group 8 elements:

They are all colourless gases at room temperature. Their boiling points and densities increase with atomic number:

Uses

- helium: filling airships and balloons
- neon: filling gas discharge tubes (red advertising lights – "neon signs")
- argon: filling light bulbs (filling "pearl" bulbs)
- krypton: making lasers.

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