Chemistry Handout

A handout for IBDP Chemistry

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Wir müssen wissen,wir werden wissen

We must know, we would know

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Chapter 1 Periodicity

1.1 Periodic Table

Introduction

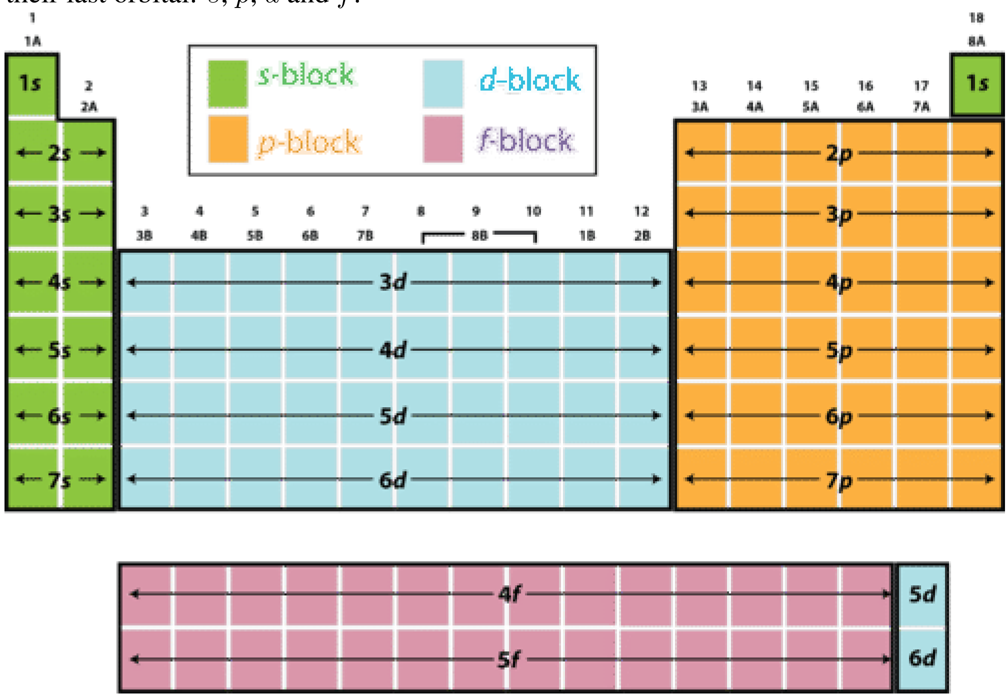
- ❏ *Understanding arrangement of periodic table*
- ❏ *Know the construction of periodic table*
- ❏ *Meaning of period number and group number*
- ❏ *Valence electron and main energy level*
- ❏ *Position of metals, non-metals and metalloids*
- ❏ *Deduce electron figuration from periodic table*

1.1.1 Arrangement of Periodic Table

Definition 1.1 (Periodic table)

The periodic table is a tabular of elements which arrange by increasing atomic number. ♣

Hence we know the basic rule to arrange the periodic table, but there are also other rules. Please review the electron configuration in last chapter(Atomic structure). The periodic table also arrange into 4 blocks associated with their last orbital: s, p, d and f.



Therefore we have a basic idea of how does a periodic table looks like. Then I will give the definition of the columns and rows in the periodic table.

Definition 1.2 (Group)

Group is a vertical column of elements ♣

Of cause the properties of a group is much more than that.

Property

- *Group number is the same as the number of valence electrons.*
- *The chemical property in a group is similar.*

Remark Groups have their own names. Group 1 is known as **Alkali metals**. Group 2 is known as **alkaline earth metals**. Group 17 is known as **halogens**. Group 18 is known as **noble gases(rare gases)**

Definition 1.3 (period)

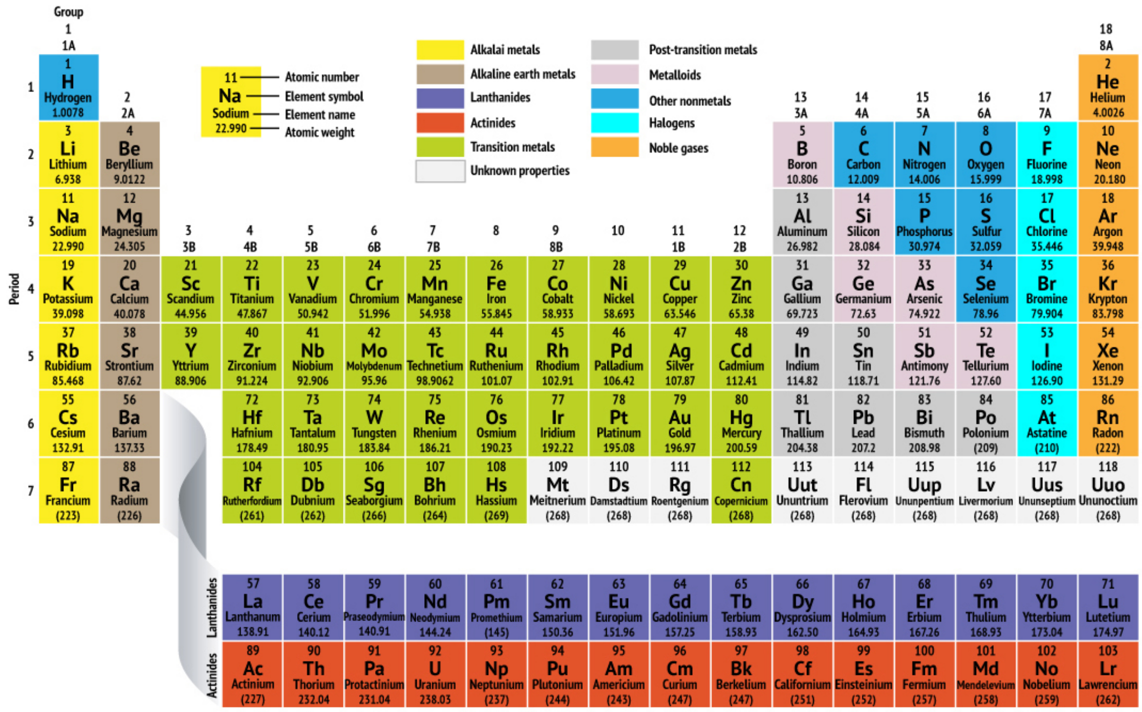
Period is a horizontal row of elements ♣

As same as group, period also have some properties.

Property

- *Period number is the same number of main energy level of the atom*
- *All sub-levels except the outer most sub-level will be full*
- *Metals are on the left and non-metals are on the right*

Note *Although I have give the trend of metals and non-matals in a period but we will learn it in the after next few concept.*



Let's have a look at the periodic table. The different kinds of elements are separated by different colors. There are few kinds of elements I will give a brief idea in this section but mainly talk about it later:

- Metals
  - Conductor
  - Oxidized(Lose electron)
- Non-metals
  - Reduced(gain electron)

Definition 1.4 (valence electrons)

Valence electrons are electrons in the outer shell (the highest main energy level) of an atom. ♣

The group number of an element is related to the number of valence electrons. The chemical property is depends on it.

Theorem 1.1
The Group number is equal to the valence electrons in this group.

Example 1.1 Group 1 has 1 valence electron, and Group 2 has 2 valence electrons

Note *Please note that you should remember all the definitions. In the exam paper2 all the questions are related with the theorem or the definition.*

1.2 Periodicity

Definition 1.5 (Periodicty)
Periodicity is repeating trends or patterns of chemical and physical properties in elements.

Let's talk about the trends of physical properties.

1.2.1 The trends in atomic radii

Definition 1.6 (Atomic radii)
Atomic radii is the distance from the nucleus to the outermost electron

Let's list some factors which will effect the radii of a atom.

- 1. The nuclear charge
- 2. Shielding effect
- 3. The electron shells

By these factors we can easily see that the atomic radii decrease cross a period.

Theorem 1.2
The atomic radii decrease cross a period

The reason that atomic radius decreases across a period is that nuclear charge increases across the period with no significant change in shielding effect. The shielding remains approximately constant because atoms in the same period have the same number of inner shells.

Theorem 1.3
Atomic radius increases down a group.

The radii of elements down a group increase is because the number of electron shells increases. In addition, the shielding effect counteracts with the increase of nuclear charge down a group.

1.2.2 The trends in ionic radii

Definition 1.7 (Ionic radius)
The ionic radius is a measure of the size of an ion.

In general, the ionic radii of positive ions are smaller than their atomic radii, and the ionic radii of negative ions are greater than their atomic radii.

Theorem 1.4

Ionic radius of positive ions is smaller than atomic radius.

This is because the nuclear charge remain but the number of electrons decrease at the same time. Hence it leads to a greater attraction between electrons and nuclei.

Theorem 1.5

Ionic radius of negative ions is bigger than atomic radius.

This is because the nuclear charge remain but the number of electrons increase, therefore greater repulsion between electrons.

1.2.3 Trends in I.E.

The full definition of first ionization energy is: the energy required to remove one electron from each atom in one mole of gaseous atoms under standard conditions.

Definition 1.8 (First ionization energy)

First ionization energy is the energy required to remove one electron from each atom in one mole of gaseous atoms under standard conditions.

$$X_{(g)} \longrightarrow X^{+}_{(g)} + e^{-} \tag{1.1}$$

Note *Note that the few preconditions "gaseous", "standard conditions" are required.*

There are few factors will effect the I.E.

- 1. Shielding effect
- 2. The distance between the electrons and nucleus
- 3. Nuclear charge

Theorem 1.6

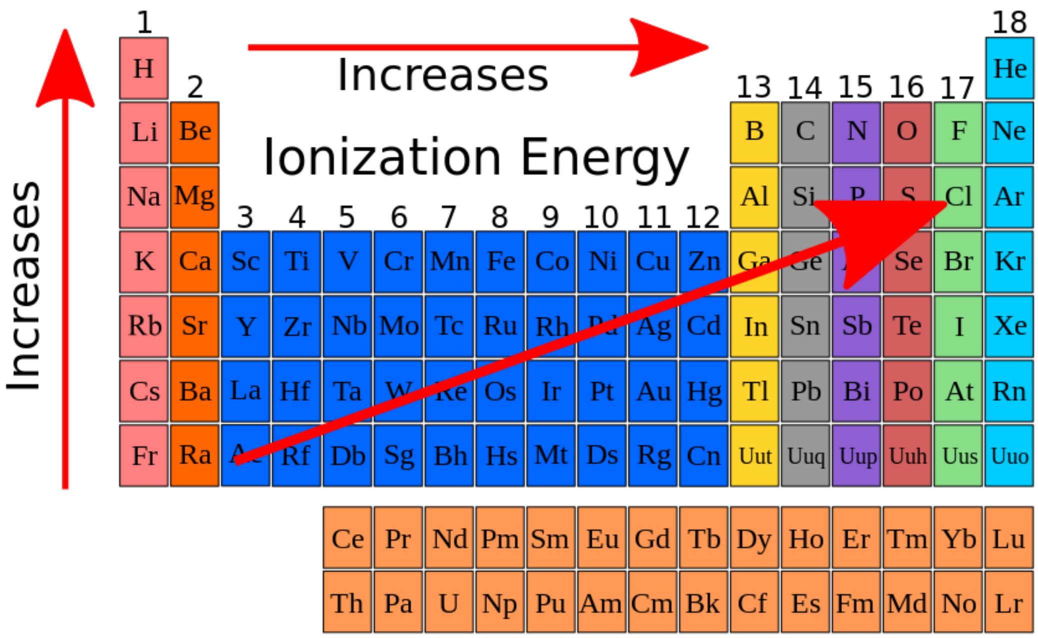
Ionization energy increases across a period

Proof The increase in nuclear charge across a period causes an increase in the attraction between the outer electrons and the nucleus makes the electrons more difficult to remove. At the same time the shielding effect is doesn't matter because in the same main energy level it has no significant change.

Theorem 1.7

Ionization energy decreases down a group

Proof The electron being removed is from the energy level furthest from the nucleus so it gets easier to remove valence electrons as atomic radius increases down a group.



1.2.4 Trends in electron affinity

Definition 1.9 (Electron affinity)

Electron affinity is the energy released when one mole of an electron is added to one mole gaseous atoms.

$$X_{(g)} + e^{-} \longrightarrow X^{-}_{(g)} \tag{1.2}$$

♣

Generally, metals have a low EA and non-metals have a higher EA. The greater the distance between the nucleus and the outer energy level, the weaker the electrostatic attraction and the less energy is released when an electron is added to the atom.

Theorem 1.8

The general trend of electron affinity across a period is that the electron affinity becomes more exothermic. ♡

Proof This is because of an increase in nuclear charge and a decrease in atomic radius from left to right across the period.

1.2.5 Electronegativity

Definition 1.10 (Electronegativity)

Electronegativity is a measure of the attraction of an atom in a molecule for the electron pair in the covalent bond of which it is a part.

♣

Note *Note that while you write the definition you should not forget that the electronegativity is defined in molecule and in covalent bond.*

Theorem 1.9

Electronegativity decreases down a group

Proof It is because of the size of the atoms increases down a group. At the same time the effective nuclear charge felt by the bonding electrons is approximately the same.

Theorem 1.10

Electronegativity increase cross a period.

Proof The reason for this is for these the increase in nuclear charge across the period with no significant change electrons in shielding. And by the nuclear charge increase, the ability to attract electron pair in covalent bond is increase.

Property

Metals have low electronegativity because they lose electrons easily.
Non-metals have high electronegativity as they gain electrons to complete their outer shell.
Hence we can use electronegativity to measure to what extent the elements are non-metals.

1.2.6 Trends in melting point and boiling point

Definition 1.11 (Melting point)

The melting point of a substance is the temperature at which it changes state from solid to liquid.

The melting point depends on the type of bonding (covalent, ionic or metallic), structure (ionic lattice, molecular covalent, giant covalent, or metallic structures), and strength of metallic bond. We will talk about these later.

Theorem 1.11

The melting points of the Alkali metals (group 1) decrease down the group.

Note *Note that the Group 1 is called Alkali metals which means "base elements". All the metals are metallic bond.*

Proof This occurs because the attractive forces between the delocalized electrons(free electrons) and the nucleus decrease owing to the increase in distance. The increase in nuclear charge is counteracted by the increase in shielding.

Theorem 1.12

The melting and boiling points of the halogens(group 17) increase down the group.

Proof This is because as the molecules become larger, the attractive forces between them increase. These shorter-range attractive forces are known as London dispersion forces and increase with the number of electrons in atoms or molecules.

1.2.7 Metallic Character

In a general idea Metals are elements at the left side of periodic table.

- Metals are shiny solids that are excellent thermal and electrical conductors. They are ductile and malleable.

- Metals are reducing agents and form cations. Their oxides and hydroxides behave like bases and neutralize acids.

Definition 1.12 (Metallic character)

Metallic character is How easily an atom can lose electrons ♣

Therefore the metallic character of elements can be compared in terms of first ionization energies.

Theorem 1.13

In general, reactive metals have low ionization energies but reactive non-metals have high ionization energies. ♡

Remark This definition is not precise in some ways.

Metallic character is displayed by metals, which are all on the left-hand side of the periodic table including alkali metals(group 1), alkaline earth metals(group 2), transition metals(d-block) the lanthanide and actinides(f-block), and the basic metals. From left to right across a period there is a decrease in metallic character and an increase in nonmetallic character. Going down a group, the metallic character increases and the first ionization energy decreases. The more reactive the metal, the greater the metallic character of the metal.

Theorem 1.14

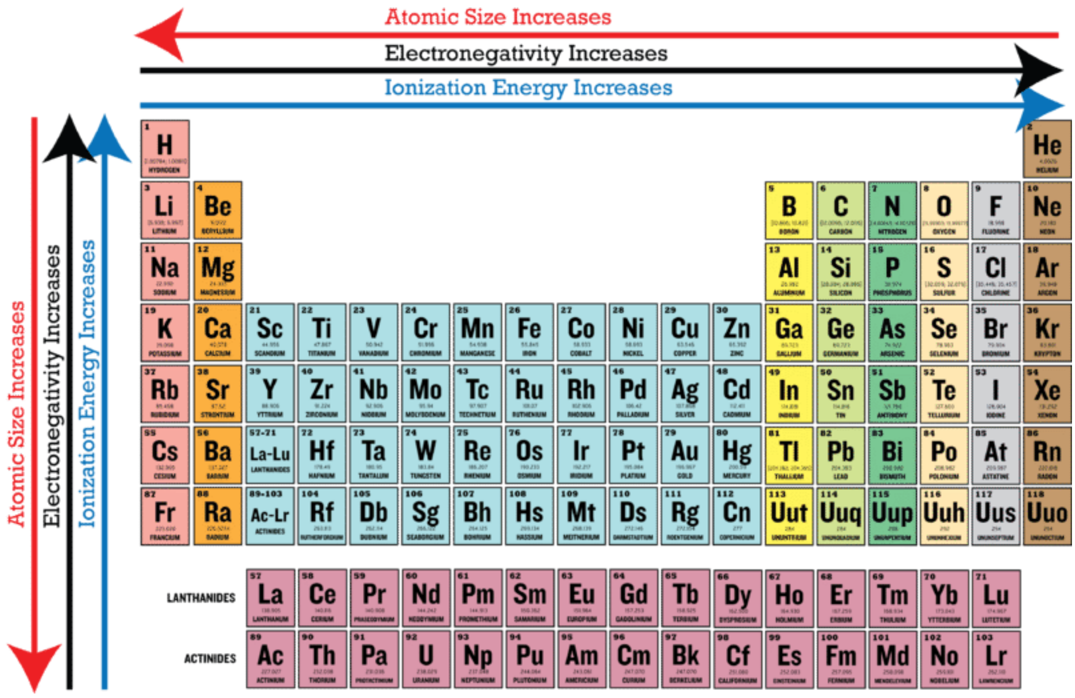
Metallic character decreases across a period. ♡

Proof By the theorem1.6 we know that the I.E. is increase cross a period, hence the metallic character decrease cross a group.1.13

Theorem 1.15

Metallic character increases down a group. ♡

Proof By theorem1.6 the I.E. decrease down a group, therefore the metallic character increase down a group.1.13



1.2.8 Group 1-The alkali metals

The alkali metals are a group of very reactive metals. By the few sections before we can outline the properties of it.

Property

- ***Atomic/Ionic radius increases*** *down the group as there are more electron shells.1.3*
- ***First ionization energy decrease*** *down the group as the valence electron is further from the nucleus so its easier to remove.1.7*
- ***Electronegativity decreases*** *because of increased distance and shielding.1.9*
- ***Melting points decrease*** *as atoms become larger and therefore metallic bonds becomes weaker.1.11*
- ***Reactivity increases down the group*** *as the valence electron is easier to lose, due to shielding.*

1.2.8.1 Reactions of the elements in group 1

Theorem 1.16

Reaction of a alkali metals M *with oxygen.*

$$4\,\mathrm{M_{(s)}} + \mathrm{O_{2(g)}} \longrightarrow 2\,\mathrm{M_2O_{(s)}} \tag{1.3}$$

Property $\mathrm{M_2O}$ *is a basic oxide that will dissolve in water to form an alkaline solution*

Theorem 1.17

Reaction of a alkali mental with water

$$2\,\mathrm{M_{(s)}} + 2\,\mathrm{H_2O_{(l)}} \longrightarrow 2\,\mathrm{MOH_{(aq)}} + \mathrm{H_{2(g)}} \tag{1.4}$$

Property *An alkaline solution is formed. The alkali metal hydroxides are strong bases and ionize completely in aqueous solution.*

1.2.9 Group 17- Halogens

The elements in group 17 are known as the halogens. They are all nonmetals consisting of diatomic molecules ($\mathrm{X_2}$)

Property *Same trends as alkali metals except melting points increase as Van der Waal forces becomes greater with more electrons.*

- ***Atomic/Ionic radius increases*** *down the group as there are more electron shells.1.3*
- ***First ionization energy decrease*** *down the group as the valence electron is further from the nucleus so its easier to remove.1.7*
- ***Electronegativity decreases*** *because of increased distance and shielding.1.9*
- ***Melting points increase*** *as Van der Waal forces becomes greater with more electrons.*
- ***Reactivity decreases down the group*** *as with each consecutive element the outer shell gets further from the nucleus. So the attraction between the nucleus and electrons gets weaker, so an electron is less easily gained.*

1.2.10 Oxides of period 3 elements

Definition 1.13 (Amphoteric)

Amphoteric oxides react both with acids and with bases.

All period three oxides will react with water to form either an acidic or alkali (basic) solution:

Group	1	2	13	14	15	16	17	18
Element	Na	Mg	Al	Si	P	S	Cl	Ar
Structure of element Structure of Oxide	Giant Metallic			Giant Covalent	Simple Covalent			
Formula of oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7	
					P_4O_6	SO_2	Cl_2O	
Acid-Base character of oxide	Basic		Amphoteric	Acidic				No oxide

1.2.10.1 Reactions to remember

- $Na_2O_{(s)} + H_2O_{(l)} \longrightarrow 2\,NaOH_{(aq)}$
- $MgO_{(s)} + H_2O_{(l)} \longrightarrow Mg(OH)_{2(aq)}$
- $P_4O_{10(s)} + 6\,H_2O_{(l)} \longrightarrow 4\,H_3PO_{4(aq)}$
- $SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(aq)}$

Note *You should remember all the formula before the exam.*

Chapter 2 Transition metal D-block elemnets(HL)

Transition metal and D-block elements are

2.0.1 Redox reactions and redox

Definition 2.1

Disproportionation is the simultaneous increase and decrease in oxidation state of the same element. ♣

2.1 Complex ions

Definition 2.2 (Complex ions)

Complex ion is an ion which contain a d-block elements in the center and surrounded by ligands. ♣

Definition 2.3 (Ligands)

Ligands are molecules or negative ions which have long pair of electrons. ♣

Ligands will share their lone pair with empty orbitals in the central d-block metal ion.