

Super lab

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 configuration from periodic table

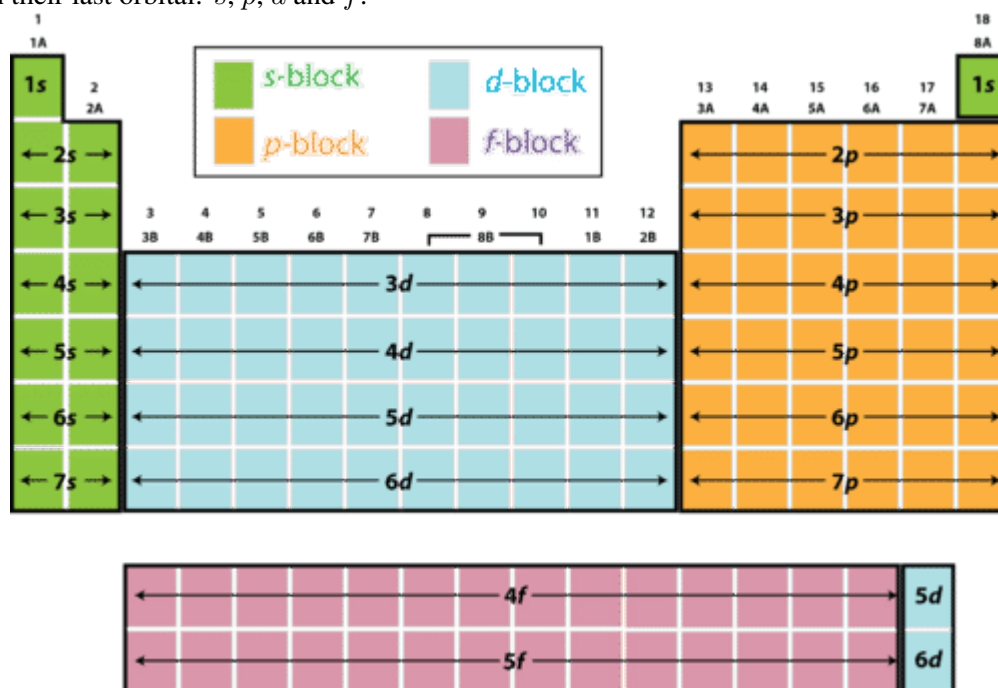
1.1.1 Arrangement of Periodic Table

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 course 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-metals in a period but we will learn it in the after next few concept.

Group 1 1A	2 2A	3B	4B	5B	6B	7B	8	9	10	11B	12B	13A	14A	15A	16A	17A	18 8A
1 H Hydrogen 1.0078	2 He Helium 4.0026																
3 Li Lithium 6.938	4 Be Beryllium 9.0122											5 B Boron 10.806	6 C Carbon 12.009	7 N Nitrogen 14.006	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.084	15 P Phosphorus 30.974	16 S Sulfur 32.059	17 Cl Chlorine 35.446	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.63	33 As Arsenic 74.922	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.798
37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.96	43 Tc Technetium 98.9062	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33		72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)		104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (269)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (268)	111 Rg Roentgenium (268)	112 Cn Copernicium (268)	113 Uut Ununtrium (268)	114 Fl Flerovium (268)	115 Uup Ununpentium (268)	116 Lv Livermorium (268)	117 Uus Ununseptium (268)	118 Uuo Ununoctium (268)
		Lanthanides															
		Actinides															
		57 La Lanthanum 138.91	58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97	
		89 Ac Actinium (227)	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (262)	

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 depends on it.

Example 1.1 Group 1 has 1 valence electron, and Group 2 has 2 valence electrons. Group 17 has 7 valence electrons.



Note Please note that you should remember all the definitions. In the exam paper all the questions are related with the axiom or the definition.

1.2 Periodicity

Definition 1.5 (Periodicity)

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.

Axiom 1.1

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**.

Axiom 1.2

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.

Axiom 1.3

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.

Axiom 1.4

Ionic radius of negative ions is bigger than atomic radius.



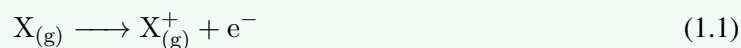
This is because the nuclear charge remain but the number of electrons increase, therefore greater repulsion between electrons.


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.



 **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

Axiom 1.5

First Ionization energy increases across a period



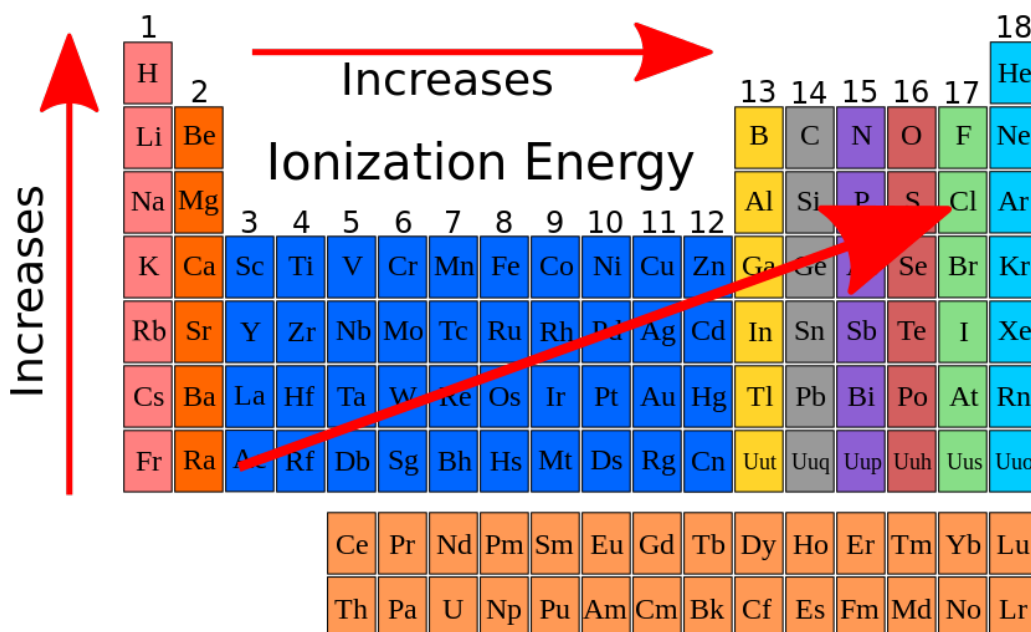
Proof The increase in nuclear charge across a period causes an increase in the attraction between the outer electrons and the nucleus makes the electrons more difficult to remove. At the same time the shielding effect is doesn't matter because in the same main energy level it has no significant change.

Axiom 1.6

First Ionization energy decreases down a group



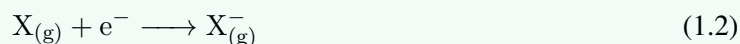
Proof The electron being removed is from the energy level furthest from the nucleus so it gets easier to remove valence electrons as atomic radius increases down a group.



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.



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.

Axiom 1.7

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*.

Axiom 1.8

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.

Axiom 1.9

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.

Axiom 1.10

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.

Axiom 1.11

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.

Axiom 1.12

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.

Axiom 1.13

Metallic character decreases across a period.



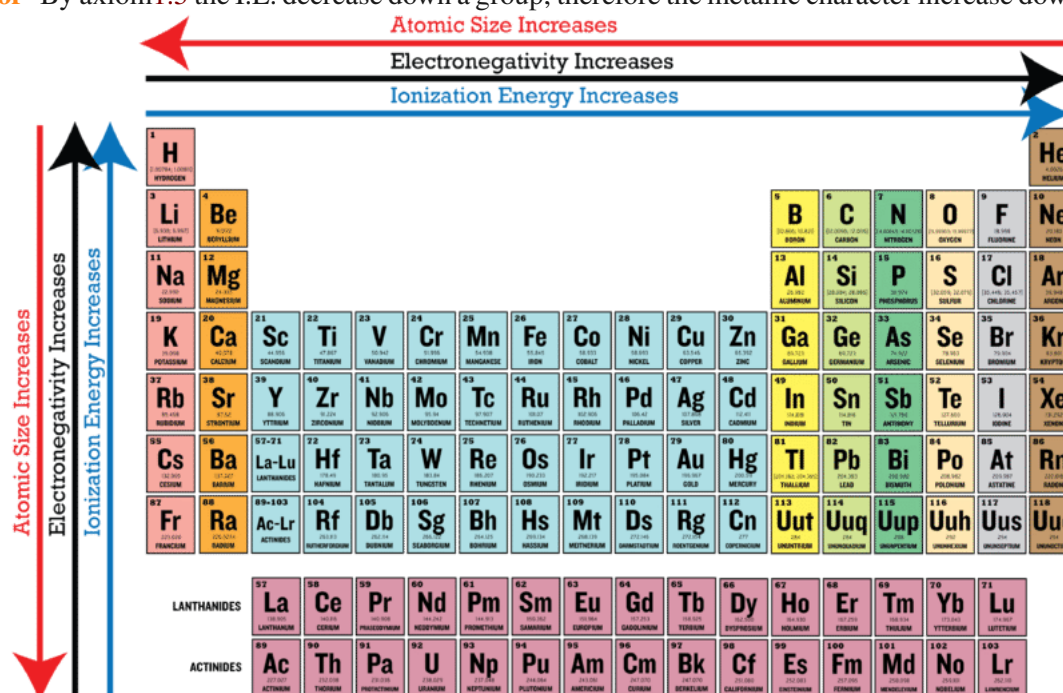
Proof By the axiom 1.5 we know that the I.E. is increase cross a period, hence the metallic character decrease cross a group. 1.12

Axiom 1.14

Metallic character increases down a group.



Proof By axiom 1.5 the I.E. decrease down a group, therefore the metallic character increase down a group. 1.12



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.

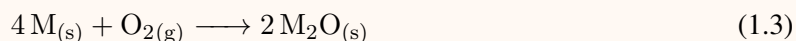
Property

- **Atomic/Ionic radius increases** down the group as there are more electron shells. *1.2*
- **First ionization energy decrease** down the group as the valence electron is further from the nucleus so its easier to remove. *1.6*
- **Electronegativity decreases** because of increased distance and shielding. *1.8*
- **Melting points decrease** as atoms become larger and therefore metallic bonds becomes weaker. *1.10*
- **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

Axiom 1.15

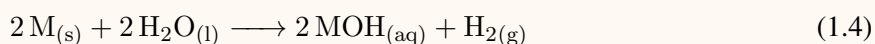
Reaction of a alkali metals M with oxygen.



Property M_2O is a basic oxide that will dissolve in water to form an alkaline solution

Axiom 1.16

Reaction of a alkali metal with water



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 (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. *1.2*
- **First ionization energy decrease** down the group as the valence electron is further from the nucleus so its easier to remove. *1.6*
- **Electronegativity decreases** because of increased distance and shielding. *1.8*
- **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	Giant Metallic			Giant Covalent	Simple Covalent			
Structure of Oxide								
Formula of oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₄ O ₁₀	SO ₃	Cl ₂ O ₇	
					P ₄ O ₆	SO ₂	Cl ₂ O	
Acid-Base character of oxide	Basic		Amphoteric	Acidic				No oxide

1.2.10.1 Reactions to remember

- $\text{Na}_2\text{O}_{(\text{s})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow 2 \text{NaOH}_{(\text{aq})}$
- $\text{MgO}_{(\text{s})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{Mg}(\text{OH})_{2(\text{aq})}$
- $\text{P}_4\text{O}_{10(\text{s})} + 6 \text{H}_2\text{O}_{(\text{l})} \longrightarrow 4 \text{H}_3\text{PO}_{4(\text{aq})}$
- $\text{SO}_{3(\text{g})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{H}_2\text{SO}_{4(\text{aq})}$



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

Chapter 2 Transition metal D-block elements(HL)

Transition metals have variable oxidation states, display catalytic and magnetic properties, form complex ions with ligands and have colored compounds.

Definition 2.1 (Transition metal)

Transition metal is elements which have incomplete d-block in one or more oxidation state.



 **Note** Zinc is not considered a transition element as it does not have an incomplete d orbital.

2.0.0.1 Oxidation state

Axiom 2.1

- All the d-block elements (except scandium and zinc) can exist in more than one oxidation state.
- There is no increase in successive ionization energies in transition metals.
- All transition elements can form ions with an oxidation number of +2 (by losing the 2e from the 4s shell). So, transition elements show an oxidation state of +2 when the electrons are removed.
- In addition, each element can form a number of ions with different oxidation numbers
-



•

2.0.1 Redox reactions and redox

Definition 2.2

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



2.1 Complex ions

Definition 2.3 (Complex ions)

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



Definition 2.4 (Ligands)

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



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

Chapter 3 Bonding and Structure

Introduction

- ☐ Understanding the covalent bond
- ☐ Understanding the ionic bond
- ☐ Understand VSEPR theory
- ☐ Molecular and ionic structure
- ☐ Intermolecular forces
- ☐ Hybridization(HL)

Definition 3.1 (Chemical Bonding)

A chemical bond is a lasting attraction between atoms or ions that enables the formation of molecules and crystals.



Note In general chemical bonding will be either ionic or covalent.

3.1 Covalent bond

Definition 3.2 (Lewis structure)

Lewis structures are diagrams that show the bonding between atoms of a molecule, as well as the lone pairs of electrons that may exist in the molecule.

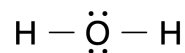


Remark While drawing lewis structure we often use the dot or lines.

For example the lewis structure of sulfur looks like:



Example 3.1 The molecule of H_2O in lewis structure is like:

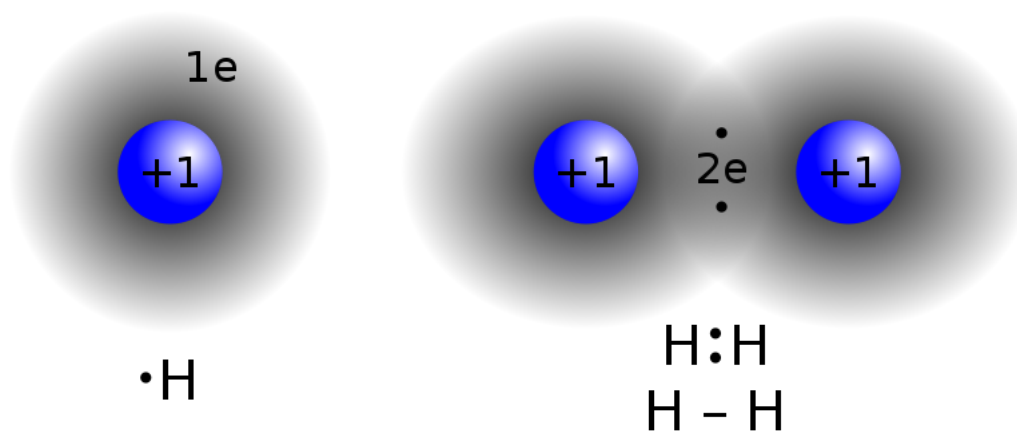


In this case you can see that the 6 dots between S it is it's outer most electrons which also named as "valence electron". In the lewis structure of H_2O you can spot that Oxygen has "lines" between Hydrogen, which is a way to represent the attraction force, we named it as "covalent bond".

Definition 3.3 (Covalent bond)

Covalent bond is a chemical bond formed by the sharing of a pair of electrons between atoms.





[1]

As you can see that this H_2 molecule is formed by two H atom that each of them share an electron, and the $1s$ orbital begin to overlap.

3.2 Ionic bond

- Cation is ion with positive charge.
- Anion is ion with negative charge.

Definition 3.4 (Ionic bond)

Ionic bond is an electrostatic attraction between oppositely charged ions.



Axiom 3.1 (The rule of formation of ions)

Elements in the main groups of the periodic table that electrons are gained or lost to achieve the electron configuration of the nearest noble gas.



It is in need to mention that this rule is not the precise. But in the case that we are high school students it is not indispensable to understand the full process.

3.2.1 Ionic substances

One concept in the ionic substances is that there are no individual ionic bonds. Therefore Ions typically pack into extremely regular crystalline structures (In an arrangement that minimizes the lattice energy maximizing attractions and minimizing repulsions).

3.2.2 Physical properties of ionic bond

Typically, ionic substances are high-melting solids.