Chemistry 11

Unit 9 : Bonding

Name:_____ Block:_____

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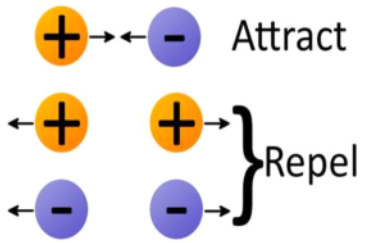
A. THE ELECTRONIC NATURE OF CHEMICAL BONDS

a) Electrostatic Forces Between Charged Particles

Definition: An **electrostatic force** = a force of attraction or repulsion between 2 charged particles.

All bonding is based on the experimentally-derived relationships of electrostatics:

- Opposite charges attract (e^- and proton $\oplus$)
- Like charges repel. (e^- and e^-)
- The greater the charges, the stronger the attractive or repulsive force. (Electronegativity (value))
- The greater the distance between the particles, the weaker the force. (related to atomic radii)



***NOTE**: When considering two examples, if the charge magnitudes are the same, then distance guides you; if relative distances are the same, then decide according to charge magnitude.*

b) Electron Shells Reconsidered

Each period (row) in the periodic table represents an electron shell. (energy level = n) (principal quantum number)

From the 4th shell onwards, the *d*-orbitals lag by 1 *n*-value. ($n-1$)

From the 6th shell onwards, the f-orbitals lag by 2 *n*-values. ($n-2$)

When you reach the end of a given period, that electron shell is FULL and therefore "closed". (→ all e^-)

Definitions:

An open **shell** is one that contains less than the maximum number of electrons

A closed **shell** is one that contains the maximum number of electrons.

The periodic table shows that patterns in the properties of elements are linked to Atomic number

What links atomic number and the properties of elements? electrons (→ # protons => e^-)

As atomic number increases by one, the number of electrons also increases by one. (neutral atoms only)

This means that the elements in the periodic table are also arranged in order of the number of electrons.

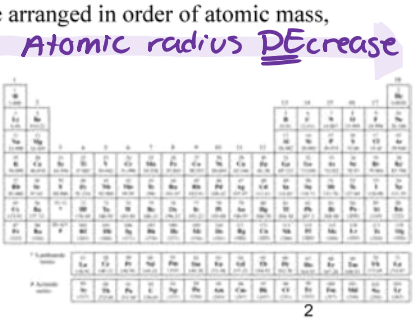
c) What is periodicity?

The Russian chemist ***Dimitry Mendeleev*** observed that when the elements are arranged in order of atomic mass, there are there are ***recurring patterns in certain properties***.

The modern periodic table can be used to analyse trends in properties such as Atomic radius across periods and down groups. (draw this on your diagram)

Atomic radius DEcrease

Atomic radius INcreases.



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d) The Size of Atoms

The radius of an atom (or ion) = the distance from the centre of the nucleus to the valence

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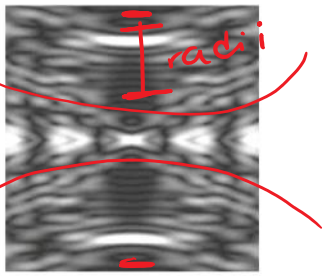
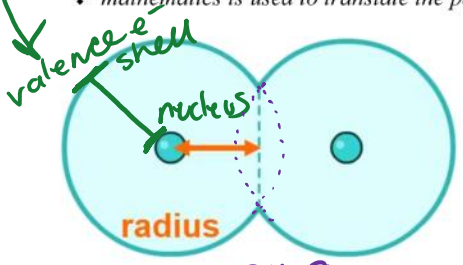
The radius of an atom (or ion) = the distance from the centre of the nucleus to the valence electron shell. The radii of atoms and ions is determined by ***X-Ray diffraction:***



- ❖ *X-Rays are shone through a material producing a diffraction pattern.*
- ❖ *mathematics is used to translate the pattern into the distances between the atoms or ions.*



ex. O_2

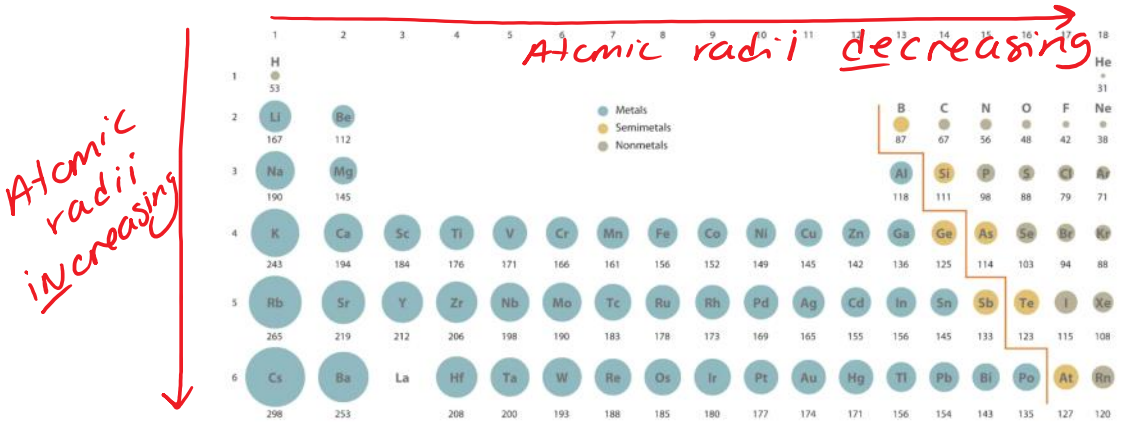
The atomic radius of an element is difficult to precisely define because of the uncertainty over the size of the electron cloud.

Several definitions are used, depending on the type of bonding.

Another definition is half the ***shortest internuclear distance found in the structure of the element.***

Trends in the Sizes of Atoms

You are expected to know - *and be able to explain* - the following observed trends in the size of atoms:

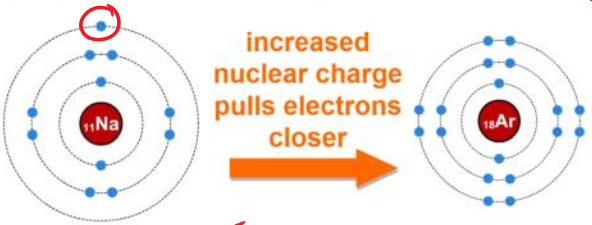


The radii of the atoms (atomic radii) DECREASE as atomic number increases along each period.

→ adding e⁻ ("to the same shell" or energy level)

Reason: Moving with increasing atomic number (from L → R) along a period:

- The number of protons and valence shell electrons are INCREASING
- The electrostatic attractive force between protons and electrons increases
- This increase in the number of protons increases the positive nuclear charge of the atoms, meaning t he nucleus has stronger attraction for the electrons, pulling them in closer to the nucleus. (makes atom smaller)



1 valence e⁻ 8 e⁻ in outer shell.

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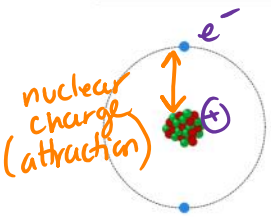
The atomic radii INCREASE as the atomic number increases moving down each group.

Top → Bottom

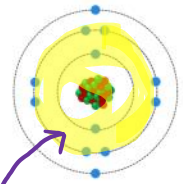
Reason: Moving with increasing atomic number down a group:

- The number of protons and electrons increase dramatically with each new period because a new valence shell/energy level is added every time a new period begins.
- However, each new valence shell is further away from the nucleus than the one before it *(and the distance between charges is more important than the charge magnitudes).*
- Additionally, as the number of core electron layers increase, there is increasing interference or "shielding" between the nuclear charge (p^+) and the valence shell charge (e^-). This is known as the ***Shielding Effect.*** It explains why some atoms are larger than expected.

What is the Shielding Effect?

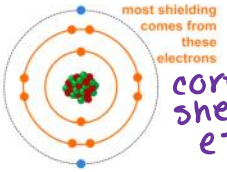


- outer e⁻ are attracted to the positive nuclear charge.
- more p⁺, more attraction between the nucleus and the valence e⁻



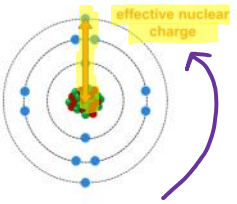
inner (CORE) shells of e⁻ repel the outer valence e⁻. Because of ⊖⊖ charge repulsion

The repulsion starts to cancel out attraction to ⊕ nucleus



core shell e⁻

The reduction in the attraction of valence e⁻ to the nucleus is called "shielding"



The [smaller] amount of attraction valence e⁻ have for the nucleus ... after we take shielding into effect.

* Extra e⁻ in the same energy level (going L → K) have very little repulsion ... most shielding is due to FULL core shells.

Explaining the radius ACROSS period 3: what is the role of shielding core electrons? Period 3 L → K

Element	Proton number	Atomic radius (nm)
Na	11	0.190
Mg	12	0.145
Al	13	0.118
Si	14	0.111
P	15	0.098
S	16	0.088
Cl	17	0.079
Ar	18	0.071

Proton number increasing → across period 3, but ***shielding*** remains approximately constant. (+ e⁻ in valence shell. not adding any energy levels)

This causes an increase in effective nuclear charge (p^+) leading to a greater attraction between the nucleus and the outermost electrons.

This pulls these electrons closer to the nucleus and results in a smaller radius.

L → R atomic radii decreasing

* shielding exists, but it isn't increasing

Ions & Atomic Radius

Negative Ions: ⊖ Anions

As extra electrons are added too a neutral atom *(eg O to make O^{-2})* the ion has the SAME positive nuclear charge *(due to protons)*, and an increased number of electrons surrounding the nucleus.

- The electrostatic repulsion INCREASE (e⁻ repelling e⁻)
- The volume occupied by the electrons is larger

NEGATIVE IONS are LARGER
than the corresponding neutral atom

(eg. O atom, and O^{2-} (larger))

Positive Ions: ⊕ Cations

As electrons are **removed** from a neutral atom *(eg Mg to make Mg^{+2})* the ion has the same positive nuclear charge *(due to protons)*, and a decreased number of negative electrons surrounding the nucleus.

- The electrostatic repulsion decreases (fewer e⁻ repelling each other)
- The volume occupied by the electrons decrease

POSITIVE IONS are SMALLER
than the corresponding neutral atom

eg. Mg and Mg^{2+} (smaller)

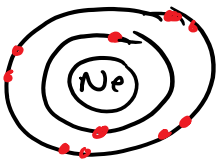
An **open shell** is a shell containing less than it's maximum number of electrons

e) Valence Electrons Reconsidered

Definition: Valence electrons are now thought of as those electrons found in ***open shells.***
They are the electrons that are available for, and ***involved in, chemical bonding.***

(Remember, we are only considering the representative elements for the time being)

Li	Be	B	C	N	O	F	Ne
1	2	3	4	5	6	7	0



no "open" shells.

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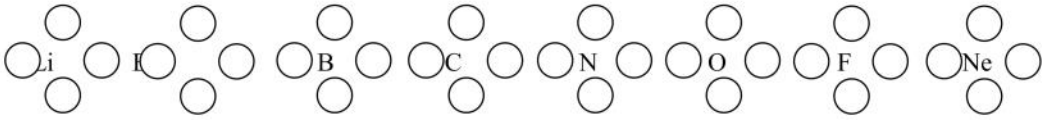
f) The Valence of an Atom

When atoms bond their atomic orbitals are blend together into ***molecular orbitals***. For the representative elements, the ***s*** and ***p*** orbitals blend to form ***sp***-type* bonding orbitals (***d*** and ***f*** orbitals will not be considered). For our purposes, there will always four ***sp***-type molecular orbitals:



(**** The actual details of hybridized molecular orbitals are beyond the scope of Chemistry 11***)

Definition: The **valence** of an atom is defined as _____

- is also the number of e^-s that are normally _____
- is also sometimes referred to as combining capacity

g) Ionization Energy

Definition: The **ionization energy** is the amount of energy needed to _____ a valence electron from one mole of an element in the gas state.

In order to form a _____ ion, an electron must be removed from a neutral atom

EXAMPLE: $Li + energy \longrightarrow Li^+ + e^-$

Comment: Ionization energy is usually expressed in units of kJ/mol. The element must be in the gas state prior to determining the ionization energy so that we know that the energy input is entirely being used to ionize the atom (and not contributing to, for example, a change of state).

- Ionization energy _____ going down a family in the periodic table
- ionization energy _____ going from left to right across a period