

General Chemistry

Periodic Table II

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Periodic Table

by Aini Norhidayah

<http://ocw.ump.edu.my/course/view.php?id=479>

Chapter Description

- Expected Outcome:

At the end of the lecture, the students should be able to understand and solve the problems regarding on the periodic table, representative and transition elements.

- Reference:

- Chemistry for matriculation semester 1, Tan Yin Toon, Sheila Shamuganathan. Companion website.



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Ionization energy

Minimum energy required to remove an electron from a gaseous atom or ion in its ground state.

- Unit: kJ/mol



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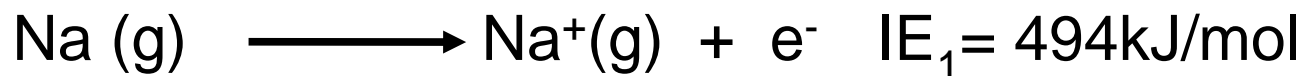
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1st ionization energy

Minimum energy required to remove one electron from the outermost orbital of neutral gaseous atom in its ground state.

eg



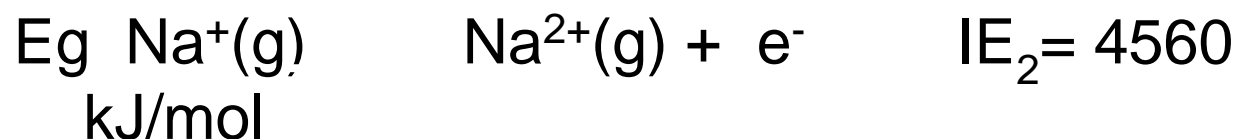
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2nd ionization energy

Energy required to remove one electron from positive ion in gaseous ion in its ground state



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1. Atomic radius

The valence electrons of an atom with a larger radius experience a less attraction towards nucleus, hence possesses a low ionization energy.

2. Effective nuclear charge

The higher the nuclear charge, the stronger the attraction forces between the nucleus and electrons. This causes the ionization energy to increase.



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3. Shielding effect (screening effect)

The shielding effect of the electrons of the inner orbitals causes the outer electrons to be less attracted to the nucleus and thus decrease the magnitude of ionization energy.



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1st ionization energy- crossing a period

- Effective nuclear charge increase causes the size of the atoms to decrease.
- Valence electrons are closer to the nucleus .
- Attractions between the nucleus and the valence electrons are stronger.
- More difficult to remove an electrons from each atom.
- More energy is needed to remove the outermost electron from an atom.
- Thus, ionisation energy (IE) increase.



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1st ionization energy- moving down a period

- Number of shell increase, the shielding effect increase, atomic size also increase.
 - Valence electrons are farther away from the nucleus and held less tightly by the nucleus.
 - Attractions between the nucleus and the valence electrons are weaker.
 - Easier to remove an electron from each atom.
- So, less energy is needed to remove the first electron.- Thus, ionization energy decrease.



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First IE across period 2

Group	1	2	13	14	15	16	17	18
Elements Period 2	Li	Be	B	C	N	O	F	Ne
IE	520	900	801	1086	1402	1314	1681	2081

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First IE across period 3

Group	1	2	13	14	15	16	17	18
Elements Period 3	Na	Mg	Al	Si	P	S	Cl	Ar
IE	496	738	578	786	1012	1000	1251	1521



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Electronegativity

-Tendency of an atom to attract electrons (or electron density) towards itself.



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Electronegativity across a period

- Proton number increase, nucleus charge become more positive. So, atomic radius decreases.
- Atomic size become smaller because of Z_{eff} .
- Attraction between nucleus and outer electron become stronger.
- Hence, the atom has greater relative tendency to attract electron to itself.
- Therefore when across a period, the electronegativity increases.



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Electronegativity down a group

- Although the positive charge on the nucleus increases, this is more than the offset by the increase in the atomic radius and the additional screening effect of the extra electron shells.
- So, going down a group, the atomic size increase because the screening effect increases.
- The attraction of the nucleus and outer electron become weaker.
- Hence, the atom has smaller relative tendency to attract electron to itself.
- Therefore when going down a group, the electronegativity decreases.



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Electronegativity

- Picture source:

http://www.chemhume.co.uk/ASCHEM/Unit%201/Ch3IMF/Images%203/electronegativity_values.jpg

Electronegativity increases 

Period	Electronegativity increases 																	
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.6
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo
Lanthanides	*	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27		
Actinides	**	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.3		



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Metallic character

Ability to conduct electricity or heat.

Depends on:

- i. Atomic radius/size
- ii. Ionization energy



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Metallic character across period 3

- decrease in metallic character
- Crossing a period, Z_{eff} increases; the attraction between nucleus and valence electrons in the valence shell getting stronger.
- Thus, the atomic radius/size decreases, the ionization energy increases and it is difficult to remove the electrons from an atom.
- Therefore, a cation difficult to form and the metallic character decreases.
- The elements change from metals to metalloid and finally, non-metals.

Group	1	2	13	14	15	16	17	18
Element	Na	Mg	Al	Si	P	S	Cl	Ar
Character	Metal			Metalloid	Non-metal			

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Metallic character down group 1 and 17

- The metallic character increases.
- This is because when going down a group, screening effect increases due to the outer electrons enter new energy levels (the number of electron shells increases).
- Thus, the atomic radius/size increases, the ionization energy decreases and it is easier to remove the electrons from an atom.



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Melting and boiling point- group 1

- Decrease when descending group 1.
- Due to the proton number of the element increase which means the atomic radius/size increase too.
- The outermost electrons are further from the nucleus.
- This causes the attraction between the nucleus of the atom and the electrons in the metallic bond to weaken.



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Melting and boiling point- group 17

- The melting point and boiling point increase when descending the group.
- The increase is due to the proton number of the element increase which means the atomic radius/size increase too.
- The number of electrons in the molecules increases and the molecules get larger. Thus, the intermolecular forces (van der Waals forces) between the neighbouring molecules gets stronger.



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Melting and boiling point- across period 3

The variation can be discussed in three parts:

- i. Metallic bond
- ii. Giant covalent structure
- iii. Simple molecular structure



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Na to Al melting point

- Na, Mg and Al are metals. In metals, the atoms are held together by strong metallic bonds.
- Metallic bond is an electrostatic forces between positive metal ions and sea of negative valence electrons.
- Strength of the metallic bond is proportional to the number of valence electrons in the metallic atoms.
- Valence electrons of Na, Mg, Al are 1, 2 and 3 respectively. Thus, melting point and boiling point increase from Na to Al.



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Si

- Si has strong covalent bonds holding the atoms together in three-dimensional array.
- Each Si is tetrahedrally covalent bonded to four other Si atoms infinitely.
- High energy is needed to break the strong covalent bonds.



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P to Ar

P, S, Cl and Ar are non-metals.

- P_4 , S_8 , Cl_2 have simple molecular structure while Ar exist as monatomic.
- The covalent bonds within the molecules are very strong, but the van der Waals forces of attraction between the molecules are very weak.
- The van der Waals forces of attraction increase as the molecular size increases.



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Acid base Characteristic of oxide element

- For Period 3, when reacted with oxygen, the elements will form : basic oxide, amphoteric and acidic oxide
- Metals form basic oxides, whereas non-metals form acidic oxides.
- Metal oxides are ionic compounds, whereas non-metallic oxides are molecular covalent compounds.
- Silicon is a metalloid.



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