PERIODIC TABLE
4.2 Periodicity

(Lecture 3)

LEARNING OUTCOMES

At the end of the lesson the students should be able to :

- Define the first and second ionisation energies.
  Explain the increase in the successive ionisation energies of an element.
- Explain the variations in the first ionisation energy
  - i. across periods 2 and 3.
  - ii. down groups 1 and 2.

4.2.3 Ionization Energies (IE)

The first ionization energy

- the energy required to remove one mole of electron from the outermost orbital in one mole of neutral gaseous atom:

$X_{(g)} \longrightarrow X^{+}_{(g)} + e^{-}$ ΔH = first ionization energy

The second ionization energy

- The energy required to remove one mole of electron from one mole of positive ion in gaseous state:

$X^{+}_{(g)} \longrightarrow X^{2+}_{(g)} + e^{-}$ ΔH = second ionization energy

4.2.3 Ionization Energies (IE)

- The magnitude of IE correlate with the strength of the attractive forces between the nucleus and the outermost electron.
- The lower of IE, the easier to form cation.

Factors Affecting the Ionization Energy

- Atomic radius

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

- Effective nuclear charge

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

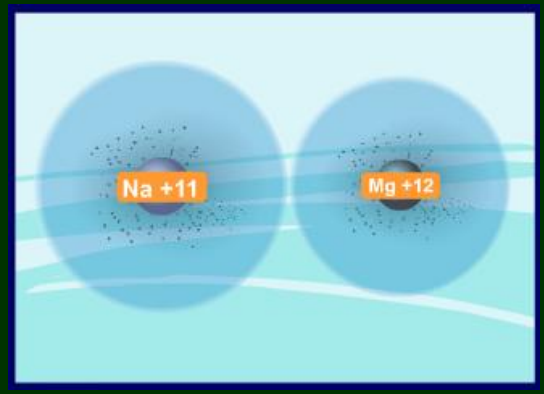
- Shielding effect (screening effect)

  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.

4.2.4 Variation in the First Ionization Energy

a) Ionization Energy Across a Period

- There is a gradual overall increase in IE as we move across a period, although the variation is irregular.
- The reason for the overall trend is the increase in effective nuclear charge felt by the electrons.
- This draws the electrons closer to the nucleus and causes the valence electrons to be held more tightly, which making it more difficult to remove them.



Variation of first IE (kJ/mol) across period 2 :

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

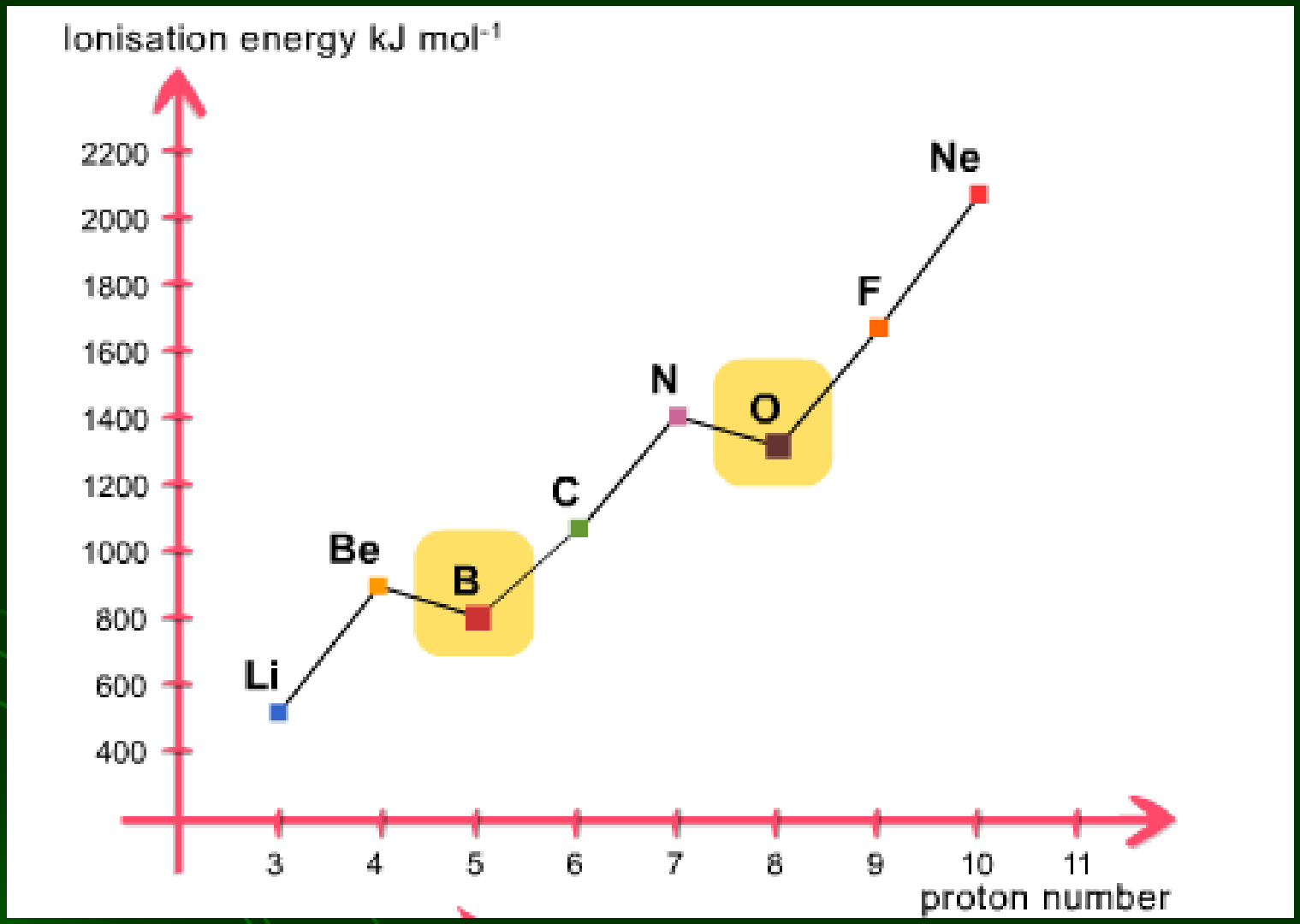
- In general, IE increases from left to right.

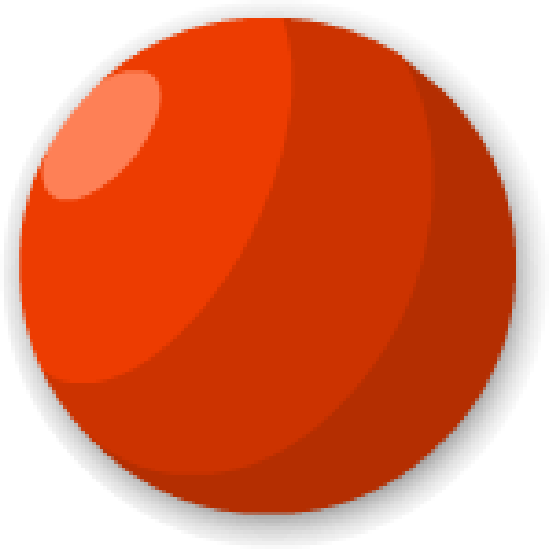
4.2.4 Variation in the First Ionization Energy

Anomalous cases

- **Between group 2 and 13. Ionization energy of Be > B**
- **Electronic configuration of**
- **Be : $1s^2\ 2s^2$ (completely filled orbital)**
- **B : $1s^2\ 2s^2\ 2p^1$ (partially filled orbital)**
- **Completely filled orbital is more stable than partially filled orbital**
- **More energy is needed to remove the electron.**
- **Therefore ionization energy of Be is higher.**

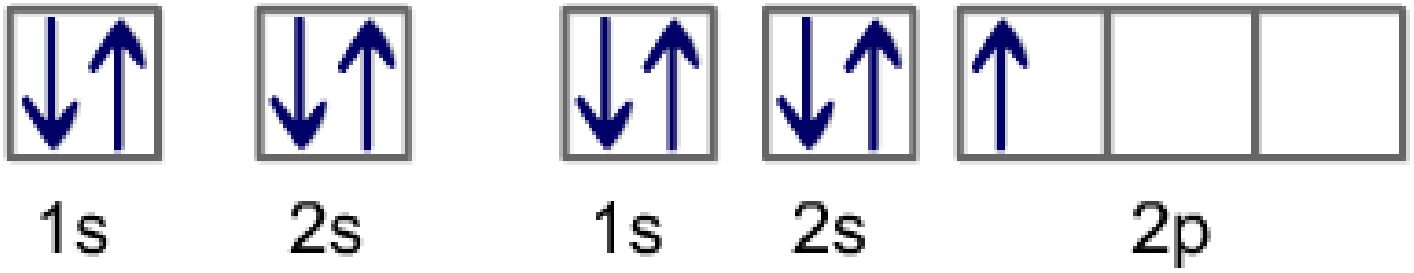
- Between group 2 and 13.
- Ionization energy of N > O
- Electronic configuration of
- N : $1s^2\ 2s^2\ 2p^3$ (half filled orbital)
- O : $1s^2\ 2s^2\ 2p^4$ (partially filled orbital)
- Half filled orbital is more stable than partially filled orbital
- More energy is needed to remove the electron.
- Therefore ionization energy of N is higher.


Ionisation energy kJ mol-1
2200
2000
1800
1600
1400
1200
1000
800
600
400
Li
Be
B
C
N
O
F
Ne
3
4
5
6
7
8
9
10
11
proton number



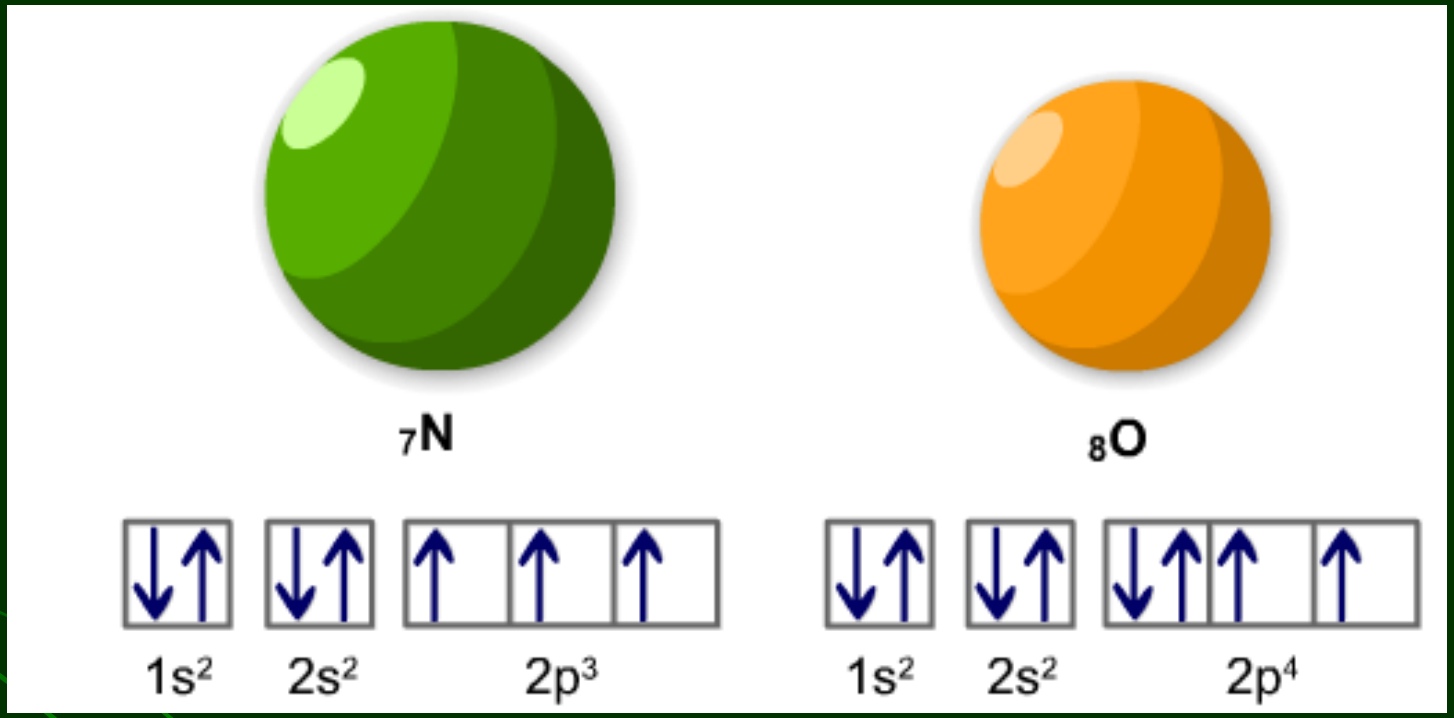
$_4Be$

$_5B$



4.2.4 Variation in the First Ionization Energy

- **b) Ionization Energy Down Within Group**
- Going down the group, the shielding effect and atomic size increases.
  - the outer electrons are farther from the nucleus and held less tightly by the nucleus.
  - So, less energy is needed to remove their first electron.
  - therefore IE decreases.


$_7$N
$_8$O
1s²
2s²
2p³
1s²
2s²
2p⁴