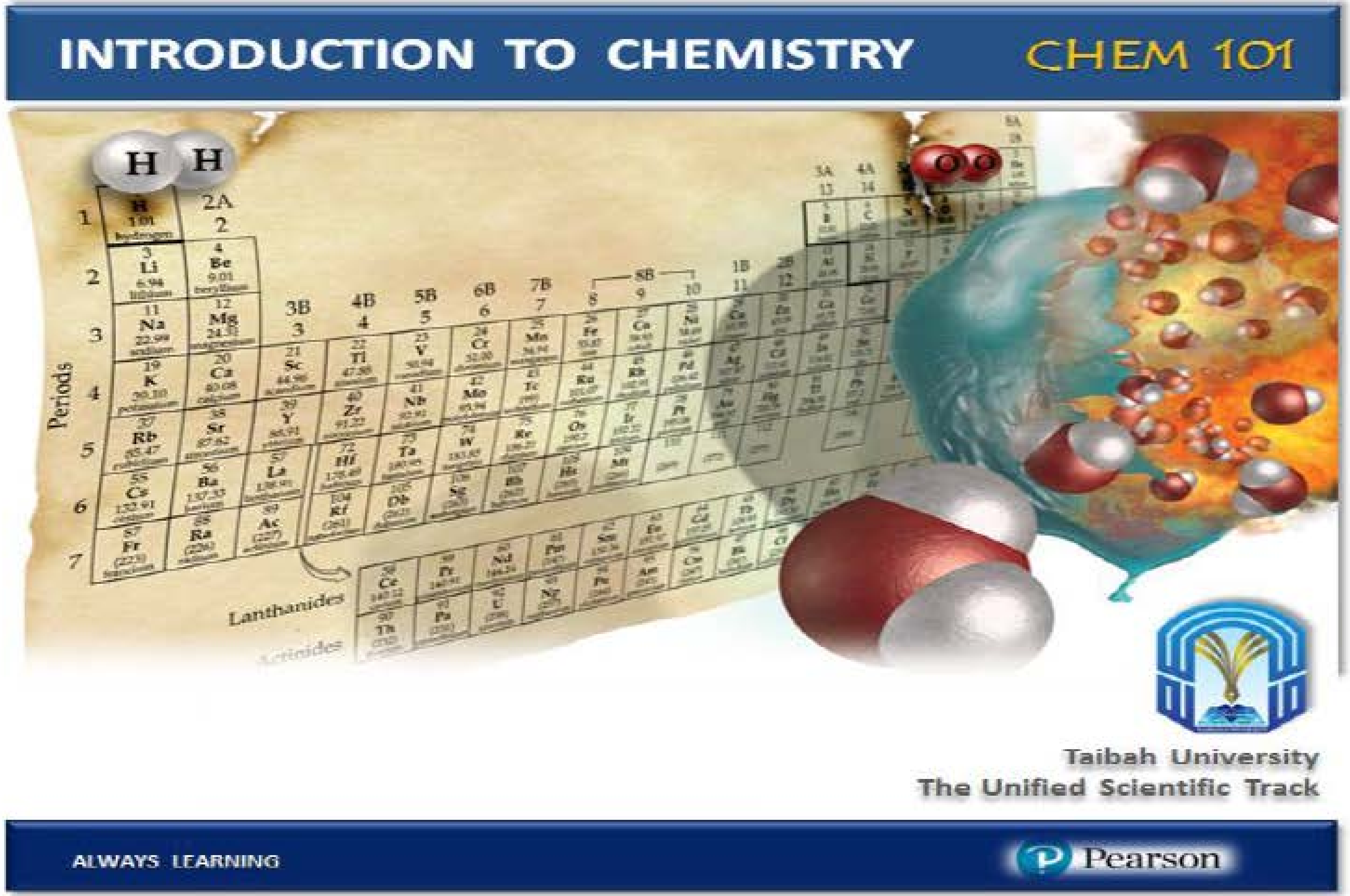


Lecture Presentation

Chapter 2

Atoms, Molecules, Ions, and Periodicity

Topic 04

Atomic Theory and Atomic Structure

2.2 - Modern Atomic Theory and Laws that Led to It

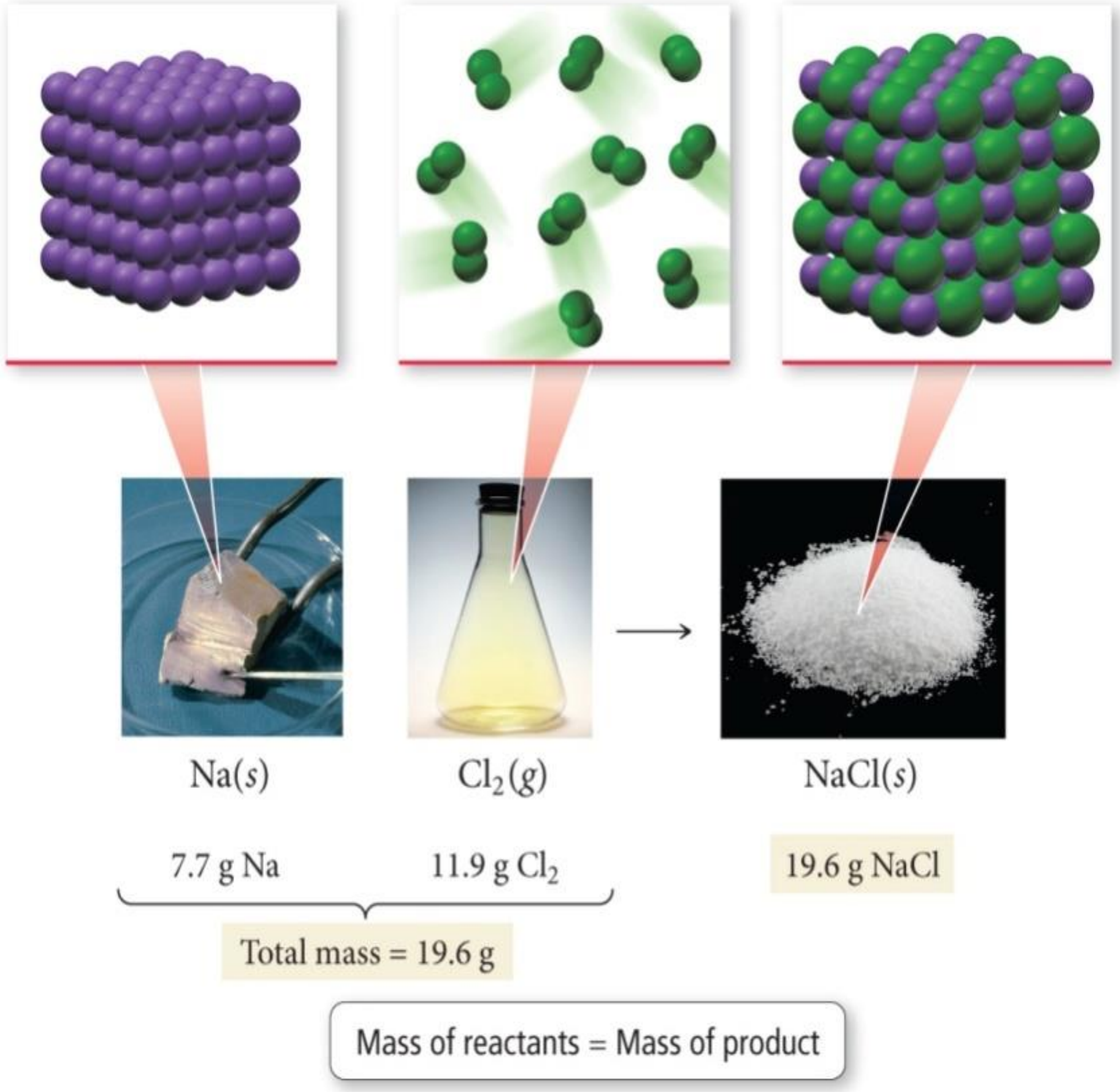
- The theory that all matter is composed of atoms grew out of many observations and laws.
- The three most important laws that led to the development and acceptance of the atomic theory are:

  - Law of the conservation of mass
  - Law of definite proportions
  - Law of multiple proportions

Law of the Conservation of Mass

Law of Conservation of Mass (A. Lavoisier):

- Matter is neither created nor destroyed
- **in a chemical reaction.**
  - Total mass of used reactants
  - = Total mass of products produced
  - Total number of reactant atoms
  - = Total number of product atoms



Law of Definite Proportions (J. Proust):

➢ All samples of a given compound, regardless of their source or how they were prepared, have the same proportions of their constituent elements.

- **For example:** Sodium chloride (NaCl) always has a definite mass-to-mass ratio of chlorine and sodium. This ratio is always the same for any sample of pure NaCl, regardless of its origin:
  - A 100 g sample of NaCl contains 39.3 g Na & 60.7 g Cl

$$\frac{\text{Mass Cl}}{\text{Mass Na}} = \frac{60.7\text{ g}}{39.3\text{ g}} = 1.54$$

  - A 58.44 g sample of NaCl contains 22.99 g Na & 35.44 g Cl

$$\frac{\text{Mass Cl}}{\text{Mass Na}} = \frac{35.44\text{ g}}{22.99\text{ g}} = 1.54$$

Law of Multiple Proportions

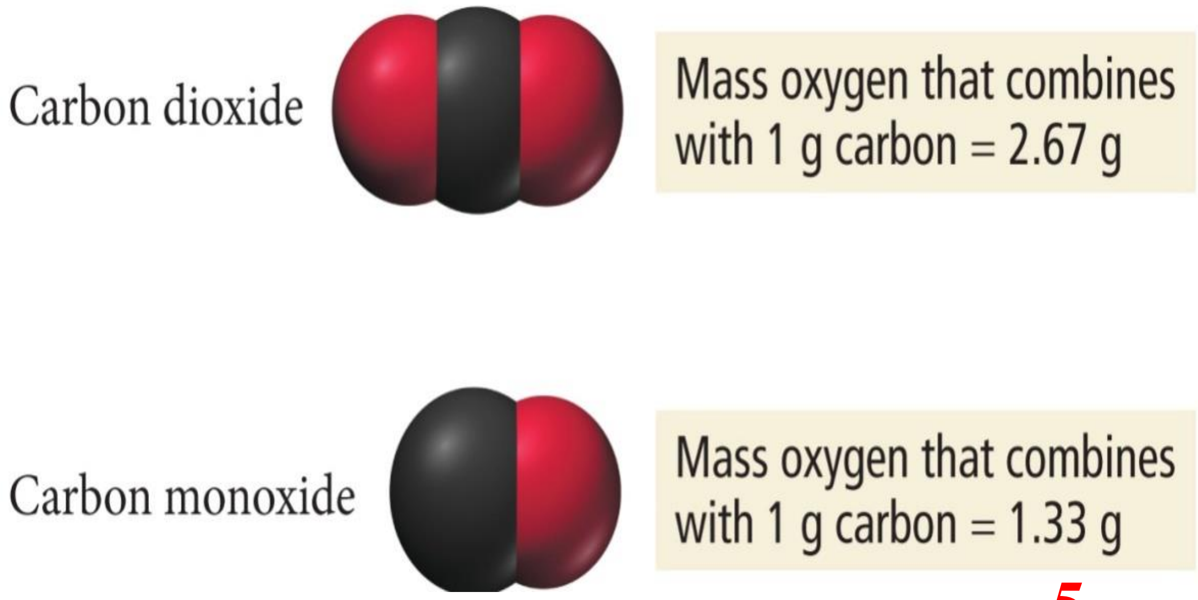
Law of Multiple Proportions (J. Dalton):

- When two elements A and B combine in different proportions; different compounds will be formed.
- **These combinations can be represented as a ratio of small whole numbers.**

✓ **For example:**

- A molecule of carbon dioxide (CO_2) has a ratio of 1 C atom to every 2 atoms of oxygen, or **1:2**.
- A molecule of carbon monoxide (CO) has a ratio of 1 C atom to 1 atom of oxygen, or **1:1**.

✓Another Example:
"Fe" to "O" in
FeO = (1:1),
and in
Fe_2O_3 = (2:3)



Postulates of Dalton's Atomic Theory of Matter:

- Each element is composed of tiny, indestructible particles called atoms.
- An element's atoms are identical in size, mass, and all other properties.
- Molecules are simple whole-number ratios of the combined elements.
- Atoms of one element cannot change into atoms of another element.

7. A molecule of water contains hydrogen and oxygen in a 1:8 ratio by mass. This is a statement of__________.

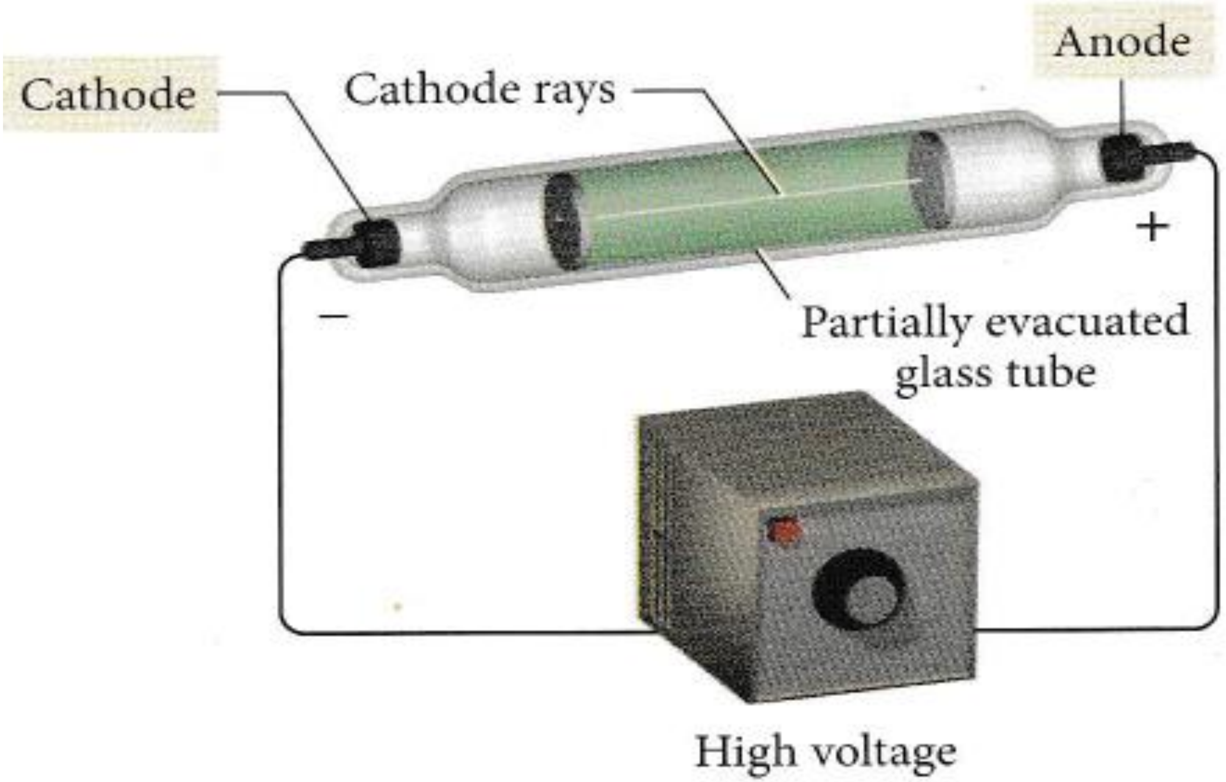
a- the law of multiple proportions
b- the law of definite proportions
c- the law of conservation of mass
d- the law of conservation of energy

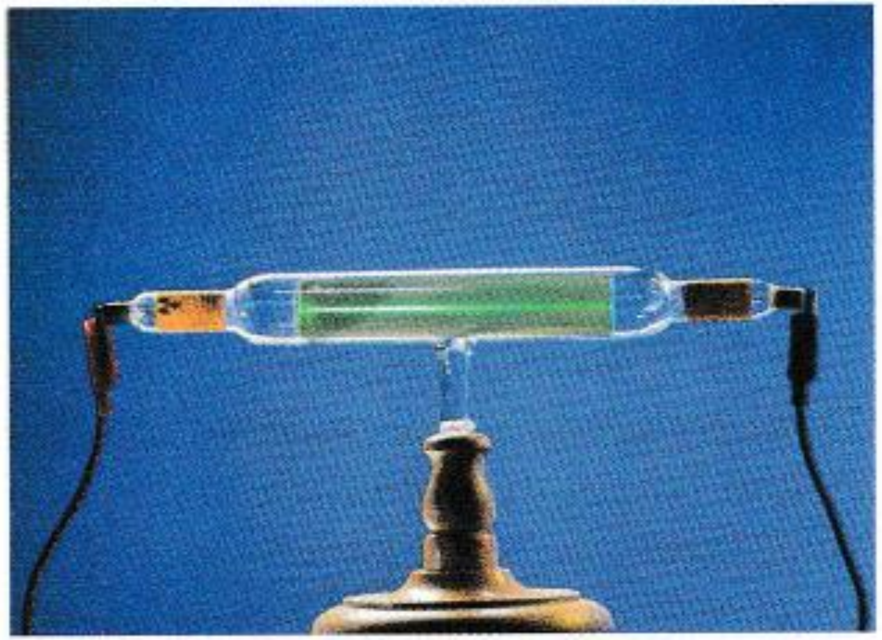
21. Which one of the following is not one of the postulates of Dalton's atomic theory?

a- Atoms are composed of protons, neutrons, and electrons.
b- All atoms of a given element are identical; the atoms of different elements are different and have different properties.
c- Atoms of an element are not changed into different types of atoms by chemical reactions: atoms are neither created nor destroyed in chemical reactions.
d- Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kind of atoms.

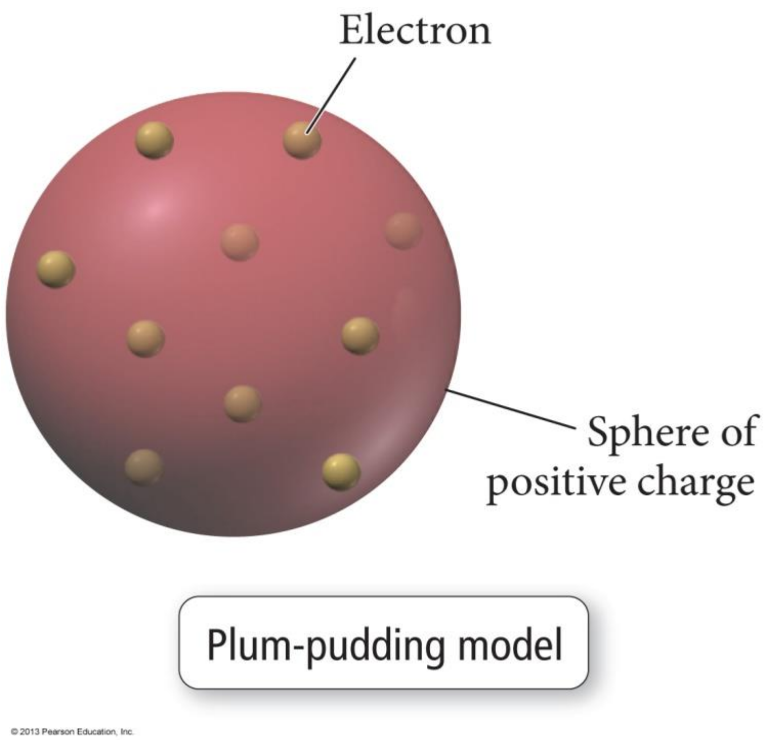
➢J. J. Thomson's Cathode Ray Tube Experiment:

- ➢Discovered the electron and determined the electron's charge-to-mass ratio.

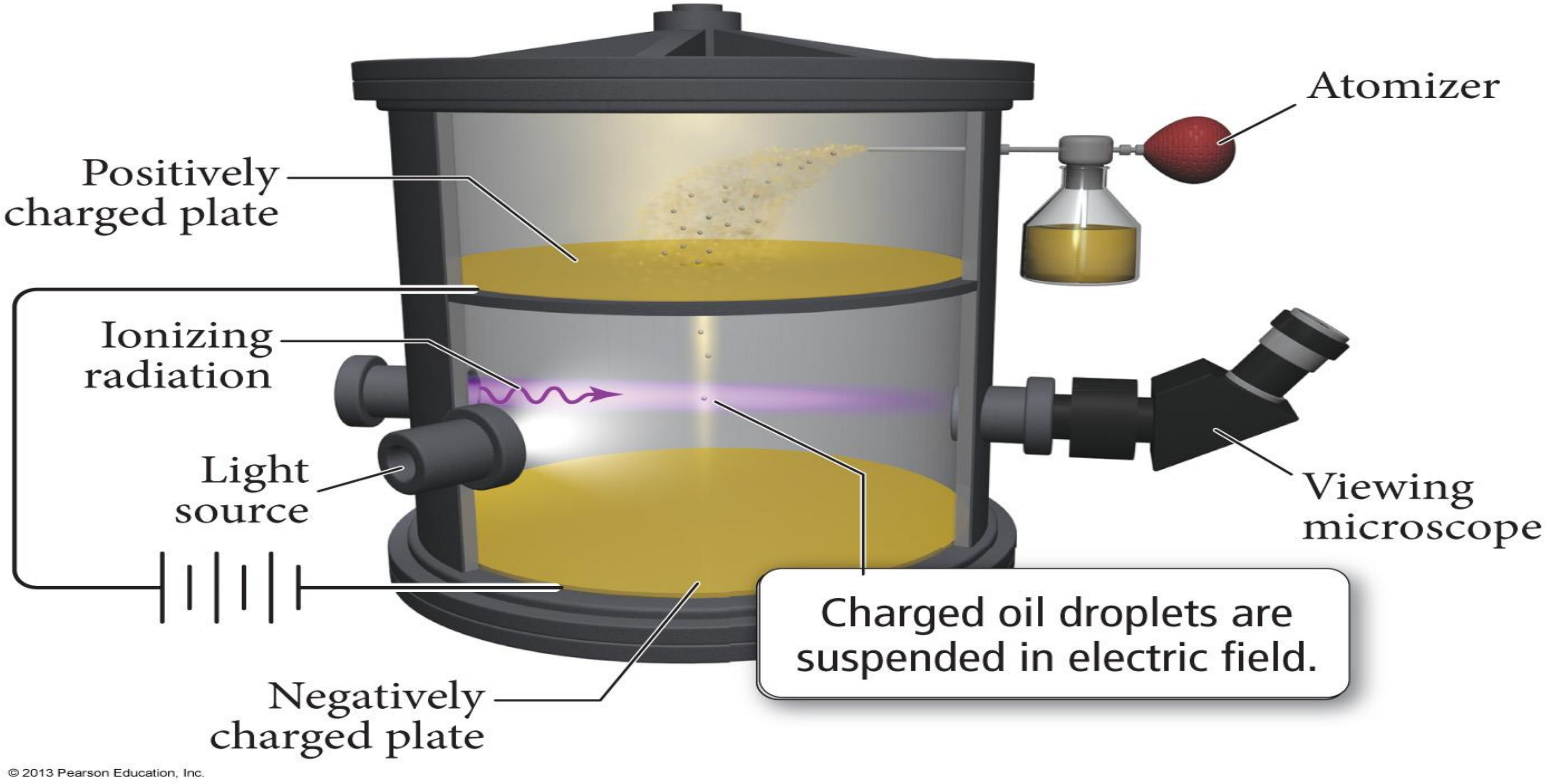




- **Plum-Pudding Model of The Atom (J. J. Thomson)**
  - The atom is composed of a positive cloud of matter in which electrons are embedded.
  - Explains the positive (+), negative (-) charged behavior of matter



- **Millikan's Oil Drop Experiment:**
  - Led to determining the **charge** of the electron.



62. Who had been postulated that electrons are held within a positive charge sphere?

a- Rutherford
b- Dalton
c- Thomson
d- Millikan

1. The discovery of the electron is attained by ....

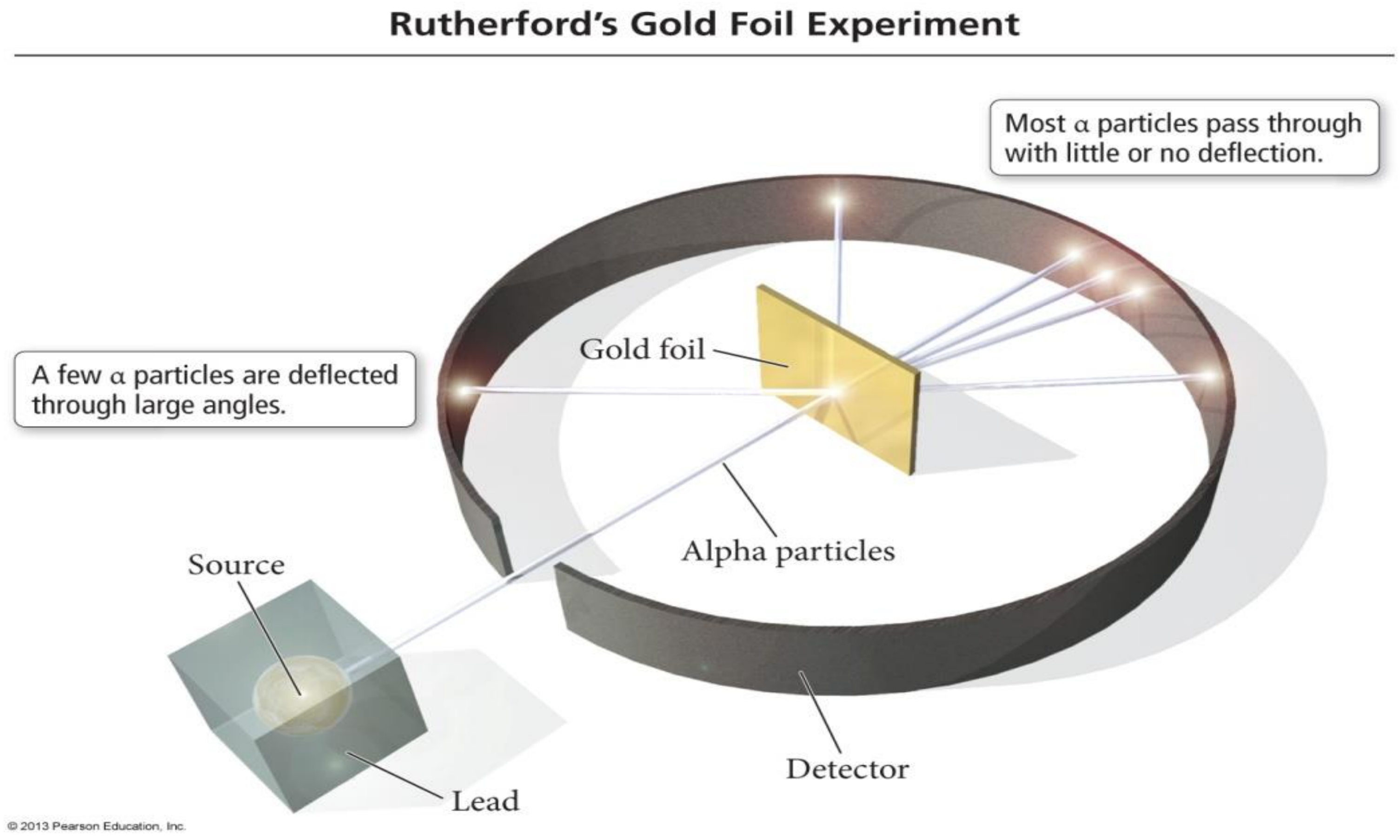
a- Rutherford gold foil experiment
b- Millikan's oil drop experiment
c- Cathode ray tube experiment
d- Dalton's experiment

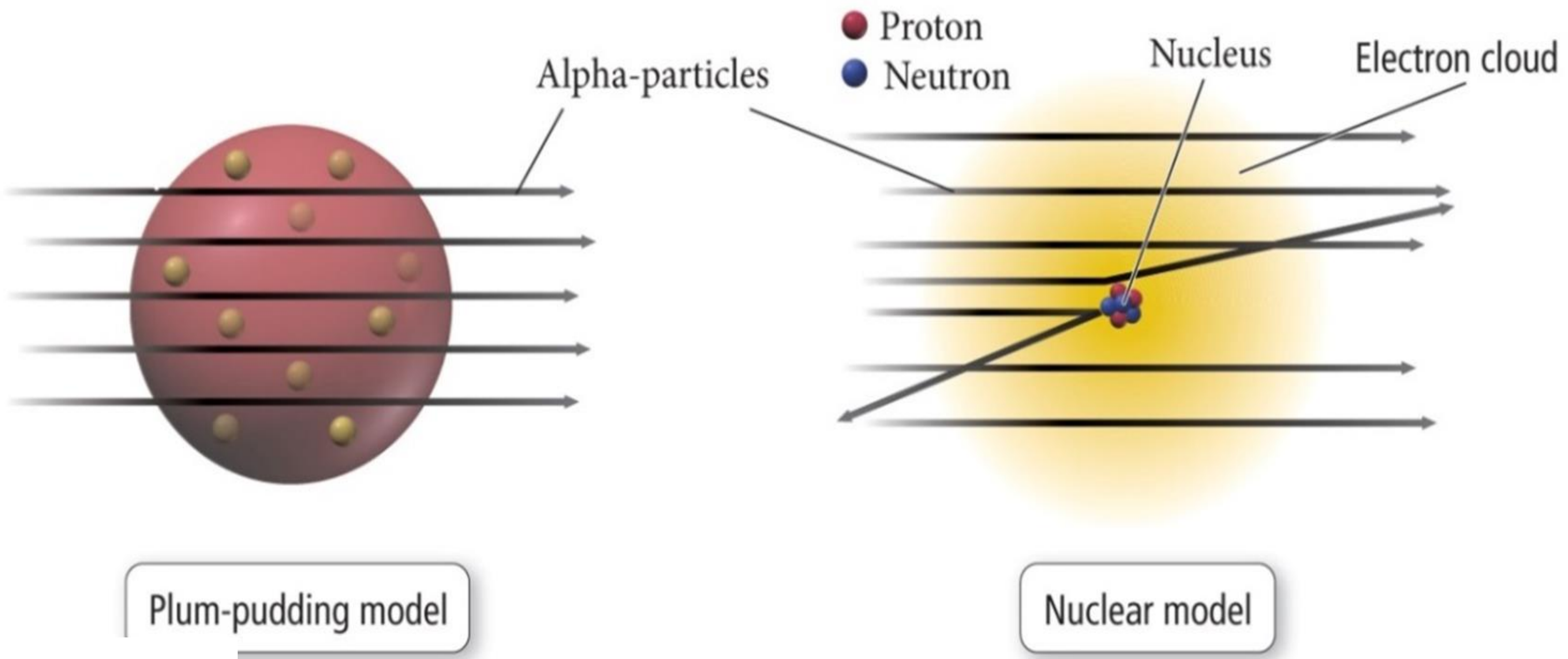
5. The charge of the electron is determined by ....

a- Rutherford gold foil experiment
b- Millikan's oil drop experiment
c- Cathode ray tube experiment
d- Dalton's experiment

2.4 The Structure of The Atom

- **Rutherford's Gold Foil Experiment:**
- Discovered the atom's nucleus (protons) & disapproved the plum-pudding model.




Proton
Neutron
Alpha-particles
Nucleus
Electron cloud
Plum-pudding model
Nuclear model

Rutherford's Model (The Nuclear Theory)

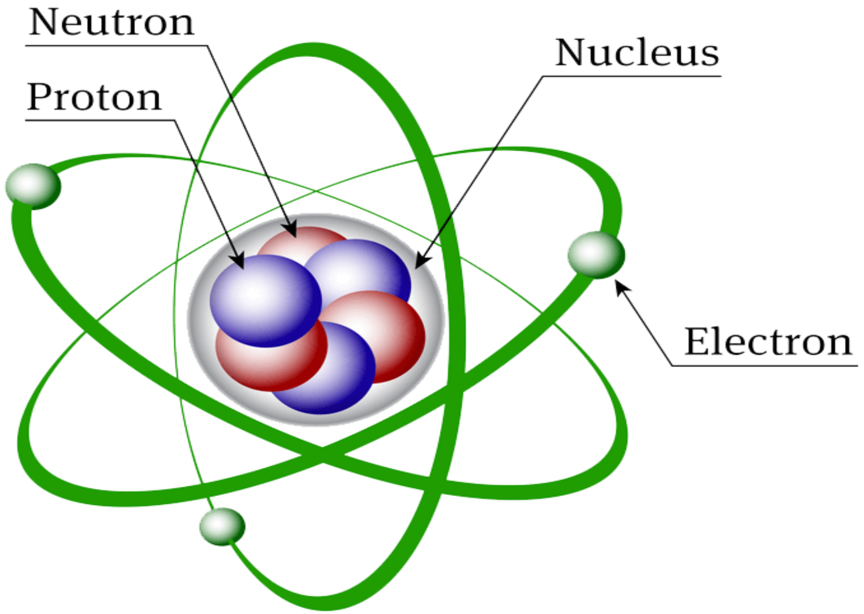
From the gold foil experiment, the following conclusions were proposed:

- The atom contains a tiny, dense center called the nucleus.
- The nucleus has essentially the entire mass of the atom.
  - The electrons weigh so little they give practically no mass to the atom.
- The nucleus is positively charged.
  - The amount of positive charge (named: protons) balances the negative charge of the electrons, so that the atom is electrically neutral.
- The electrons are dispersed in the empty space of the atom surrounding the nucleus (most of the volume of the atom is empty space).

The Discovery of The Neutrons

- **J. Chadwick** was an English physicist who was awarded the 1935 Nobel Prize in Physics for his discovery of the neutrons, neutral particles within the nucleus of the atom.

> ✓ This discovery has explained why the dense nucleus of the atom (protons + neutrons) contains over 99.99% of the mass of the atom. However, it occupies very little of the atom's volume!



8. The gold foil experiment performed in Rutherford's lab________.

a- confirmed the plum-pudding model of the atom
b- led to the discovery of the atomic nucleus
c-was the basis for Thomson's model of the atom
d-utilized the deflection of beta particles by gold foil

9. Cathode rays are________.

a- neutrons | b-X-rays | c-electrons | d-protons

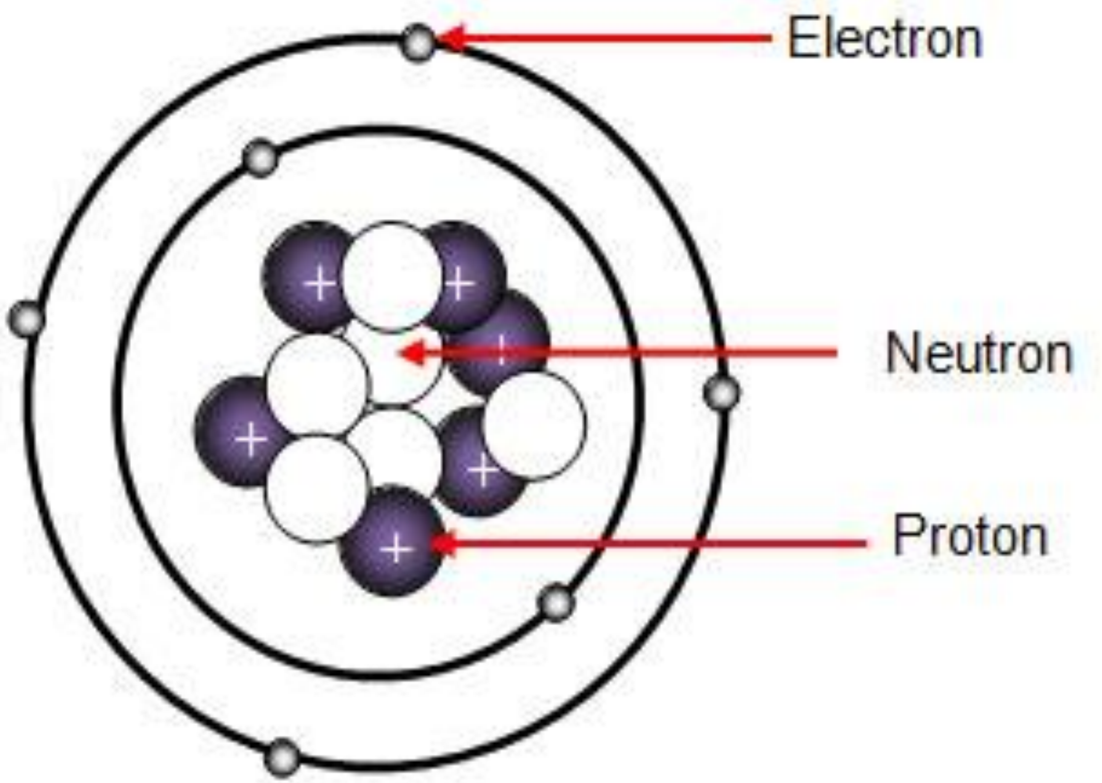
2.5 The Subatomic Particles: Protons, Neutrons and Electrons in Atoms

Elements: are composed of extremely small particles called **atoms**.

Atom: is the basic unit of an element that can enter into chemical combination.

Atom consist of:

1. Electron (e) (-ve charge)
2. Proton (p) (+ve charge)
3. Neutron (n) (neutral)



Properties of Subatomic Particles						
Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)	Symbol
Electron	Outside nucleus	-1.602×10^{-19}	1-	0.00055	0.00091×10^{-24}	e , e^{-}
Proton	Nucleus	1.602×10^{-19}	1+	1.00727	1.67262×10^{-24}	P , P^{+}, H^{+}
Neutron	Nucleus	0	0	1.00866	1.67493×10^{-24}	n , n^{0}

The number of protons

located in an atom's nucleus determines the element's identity

Atomic number: is the number of protons in the nucleus of each atom of an element.

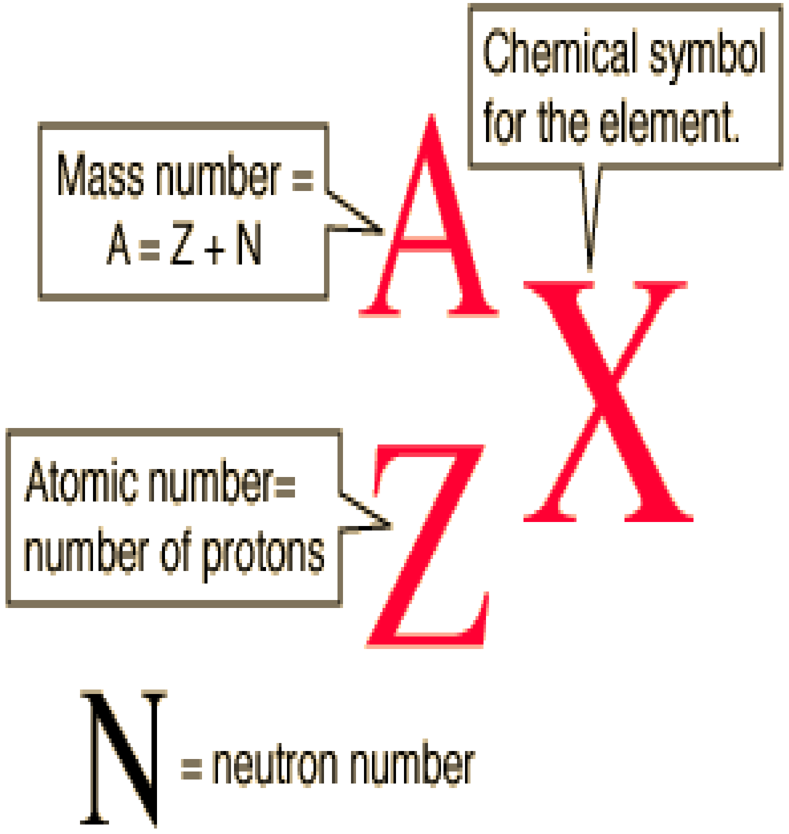
Atomic number (Z) = number of protons in nucleus.

For a neutral atom: Number of protons = number of electrons

Mass number: is the total number of neutrons and protons present in the nucleus of an atom of an element

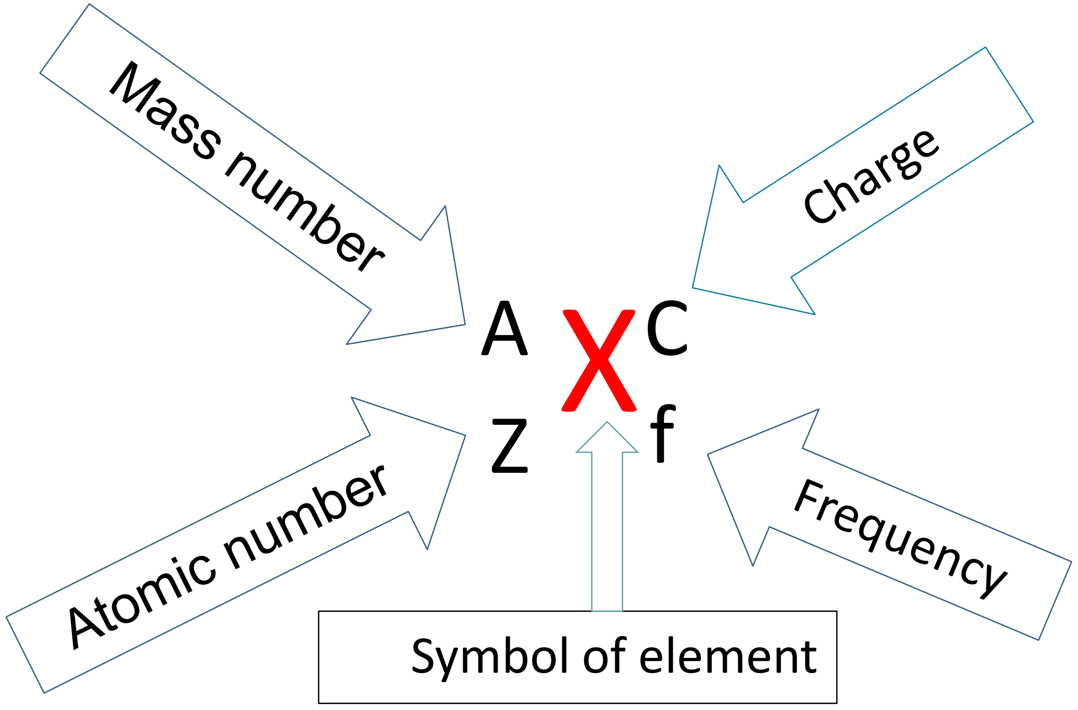
Mass number (A) = atomic number (Z) + number of neutrons

Thus: **the number of neutrons = A - Z**



- The letter c in the chemical symbol ($^{A}_{Z}X^{c}_{f}$) represents:

a- Atomic number b- mass number c- charge d- frequency



Exercise

How many protons, electrons, and neutrons are in the following atoms:

	protons	electrons	neutrons
$^{32}_{16}S$			
$^{65}_{29}Cu$			
$^{240}_{92}U$			

Note: Neutral atoms are having the same number of electrons as protons!

Give the number of protons, neutrons, and electrons in each of the following species:

Elements	$^{20}_{11}Na$	$^{22}_{11}Na$	$^{17}_{8}O$	$^{14}_{6}C$	$^{200}_{80}Hg$
Atomic Number (Z)					
Mass Number (A)					
No. of electrons (e)					
No. of protons (p)					
No. of neutrons (n)					

Examples

1-How many protons, neutrons, and electrons are in $^{14}_{6}C$?

6 protons, 8 =(14 - 6) neutrons, 6 electrons

2-How many protons, neutrons, and electrons are in $^{11}_{6}C$?

6 protons, 5 =(11 - 6) neutrons, 6 electrons

Ions: Losing and Gaining Electrons

- **The Ion:**
- is an atom or group of atoms carrying a positive (+) or negative (-) charge.
- Taking away
- an electron from an atom gives a Cation with a positive charge.
  - More protons in nucleus vs. electrons surrounding nucleus
    - Metals tend to form cations.
- Adding
- an electron to an atom gives an Anion with a negative charge.
  - Fewer protons in the nucleus vs. electrons surrounding nucleus
    - Nonmetals tend to form anions.

Ion

Cation:

an ion with a +ve charge

(lose electron/s)

Na^+, NH_4^+, Mg^{2+}

$\mathbf{Na \rightarrow Na^+ + 1e^-}$

Anion:

an ion with a -ve charge

(Gain electron/s)

Cl^-, OH^-, S^{2-}

$\mathbf{Cl + 1e^- \rightarrow Cl^-}$

Elements: Defined by their Number of Protons

1-How many protons and electrons are in $^{27}_{13}Al^{3+}$?

13 protons, 10 (13 – 3) electrons

2-How many protons and electrons are in $^{78}_{34}Se^{2-}$?

34 protons, 36 (34 + 2) electrons

Use the following table and choose which of the species are positively charged?

Atom or ion element	I	II	III	IV	V
Atom or ion electrons (e)	8	13	8	8	11
Atom or ion protons (p)	6	10	8	10	12
Atom or ion neutrons (n)	6	11	9	7	10

A. III and V   C. II and III   B. **IV and V**   D. I and VI

	A	Z	n	p	e
${}^{24}_{12}Mg^{+2}$					
${}^{31}_{15}P^{-3}$					
${}^{1}_{1}H$					

1. Which of the following expressions represents **two molecules** of water?

A. H_2O
B. H_2O_2
C. $2\ H_2O$
D. $2\ HO_2$

2. The species S^{2-}, F^-, and Cl^- are all…

A. cations
B. anions
C. isotopes
D. Halogens

3. Atoms with the same number of electrons and number of protons are called…

A. ions
B. isotopes
C. neutral atoms
D. different atoms

Isotopes

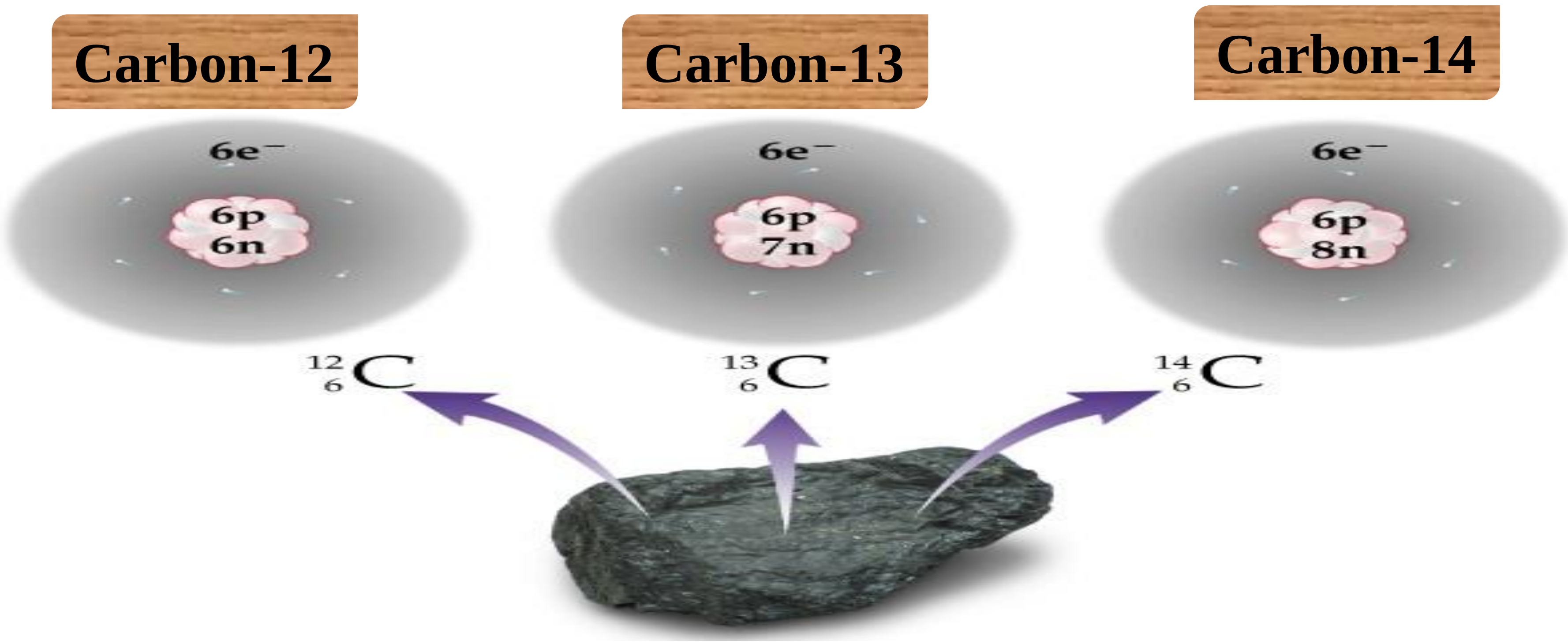
are atoms of one element that have the same number of protons (atomic number) and different number of neutrons. Isotopes differ in mass number because they have different number of neutrons. Isotopes are chemically identical.

Mass Number (A) = Protons + Neutrons

Note: Isotopes are identified by their "mass numbers"

(e.g. C–12 , C–13 , C–14)

c: An Example



	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
protons:	6 p^+	6 p^+	6 p^+
neutrons:	6 n	7 n	8 n
electrons:	6 e^-	6 e^-	6 e^-

Isotopes :

Are different forms of atoms of the same element have the same number of protons (atomic number) but differ in the number of neutrons.

	chlorine Isotopes			Sodium isotopes			Hydrogen isotopes		
	Cl_{17}^{37}	Cl_{17}^{36}	Cl_{17}^{35}	Na_{11}^{23}	Na_{11}^{24}	Na_{11}^{22}	H_{1}^{1}	H_{1}^{2}	H_{1}^{3}
A	37	36	35	23	24	22	1	2	**3**
Z	17	17	17	11	11	11	1	1	**1**
n	20	19	18	12	13	11	0	1	**2**
p	17	17	17	11	11	11	1	1	**1**
e	17	17	17	11	11	11	1	1	**1**

Isobars :

Are the different elements with same mass number but different atomic number (number of protons)

	Isobars :الأيزوبارات				
	K_{19}^{40}	Ca_{20}^{40}		C_{6}^{14}	N_{7}^{14}
A	40	40		14	**14**
Z	19	20		6	**7**
n	21	20		8	**7**
p					
e					

Isotones :

Different elements have the same number of neutrons.

	Isotones :الأيزوتونات				
	Mg_{12}^{24}	Na_{11}^{23}		Zr_{40}^{85}	Sr_{38}^{83}
A					
Z					
n	12	12		45	45
p					
e					

Which of the following pairs are isobars , isotones , isotopes

Mg_{12}^{24}	Na_{11}^{23}

K_{19}^{40}	Ca_{20}^{40}

Cl_{17}^{37}	Cl_{17}^{36}

Assessment

Answer the following questions:

1- Fill in the blanks to complete the table:

Symbol	Z	A	Number of p	Number of e$^-$	Number of n	Charge
________	8	________	________	________	8	2–
Ca^{2+}	20	________	________	________	20	________
Mg^{2+}	________	25	________	________	13	2+
N^{3-}	________	14	________	10	________	________

Answer the following questions:

2- Determine the number of p^+, n^0, and e^- in each atom:

a. $^{14}_{7}N$ b. $^{23}_{11}Na$ c. $^{222}_{86}Rn$ d. $^{208}_{82}Pb$

3- Determine the number of protons and the number of electrons in each ion:

a. Ni^{2+} b. S^{2-} c. Br^{-} d. Cr^{3+}

4- Write isotopic symbols of the form $^{A}_{Z}X$ for each isotope:

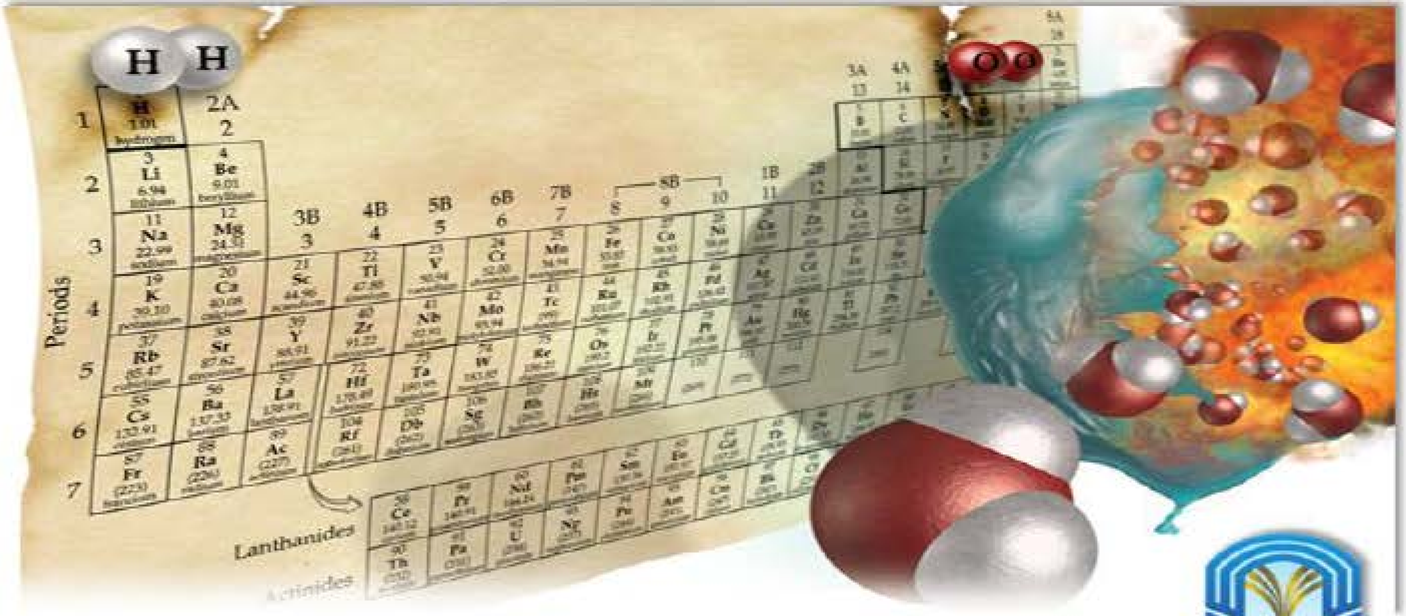
a. the copper isotope with **36** neutrons

b. the oxygen isotope with **8** neutrons

c. the aluminum isotope with **14** neutrons

d. the iodine isotope with **74** neutrons

INTRODUCTION TO CHEMISTRY CHEM 101



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Lecture Presentation

Chapter 2

Atoms, Molecules, Ions, and Periodicity

Topic 05

The Periodic Table: An Overview

2.6 Finding Patterns: The Periodic Law and The Periodic Table

- In 1869, **Dmitri Mendeleev** arranged
- the elements on his table in order

  of increasing **atomic mass**.
- He found that some properties of

  those elements recurred

  in a "**periodic pattern**".

Mendeleev's Periodic Table (1869)

I	II	III	IV	V	VI	VII	VIII		
H 1.01									
Li 6.94	Be 9.01	B 10.8	C 12.0	N 14.0	O 16.0	F 19.0			
Na 23.0	Mg 24.3	Al 27.0	Si 28.1	P 31.0	S 32.1	Cl 35.5			
K 39.1	Ca 40.1		Ti 47.9	V 50.9	Cr 52.0	Mn 54.9	Fe 55.9	Co 58.9	Ni 58.7
Cu 63.5	Zn 65.4			As 74.9	Se 79.0	Br 79.9			
Rb 85.5	Sr 87.6	Y 88.9	Zr 91.2	Nb 92.9	Mo 95.9		Ru 101	Rh 103	Pd 106
Ag 108	Cd 112	In 115	Sn 119	Sb 122	Te 128	I 127			
Ce 133	Ba 137	La 139		Ta 181	W 184		Os 194	Ir 192	Pt 195
Au 197	Hg 201	Ti 204	Pb 207	Bi 209					
			Th 232		U 238				

To be **periodic** means to Exhibit a **repeating pattern.**

Looking for Patterns: Recurring Properties

1 H	2 He	3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne	11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar	19 K	20 Ca

- The color of each element represents its properties.
- We arrange them in rows so that similar properties align in the same vertical columns.

1 H							2 He
3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca						

The Periodic Law

- **Mendeleev** : summarized these observations in the **periodic law**:

The Periodic Law: When the elements are arranged in order of increasing mass; certain sets of properties recur periodically.

- **Mendeleev** : arranged the rows so that elements with similar properties fall in the same vertical columns.

Mendeleev arranged elements in the periodic table according to:

A. number of protons

B. number of electrons

C. mass

D. volume

The Modern Periodic Table

- In 1913, **Henry Moseley** proposed the modern periodic table using
- atomic number instead of atomic mass, as the organizing principle for all the identified elements.

- **The Modern Periodic Table Consists of:**

➢ **7 Rows:** are referred to as **Periods**, the periods are numbered 1–7.

➢ **18 Columns:** are sometimes referred to as **Groups** or **Families**,

➢ they are numbered 1–18 (or the A and B grouping).

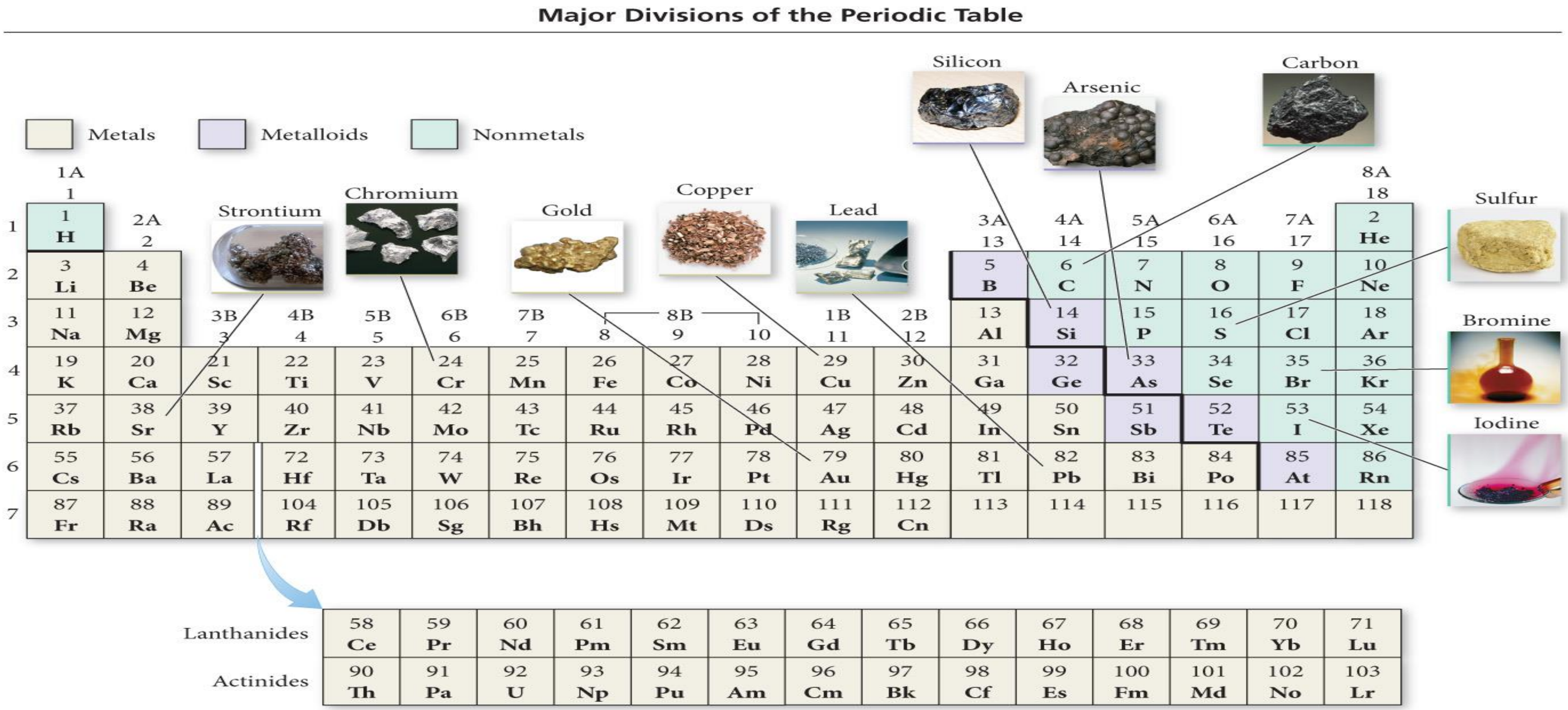
  - They are commonly called "**Families**" because the elements within the column have similar physical and chemical properties.

Classification of Elements

- **Elements in the periodic table are classified**
- **into the following <u>three major divisions</u>:**

  - **Metals**
  - **Nonmetals**
  - **Metalloids**

The Modern Periodic Table: Metals, Nonmetals & Metalloids



Classification of Elements: Metals

- **Properties of Metals:**

- **Metals** lie on the lower left side and middle of the periodic table.
  - ✓They are good conductors of heat and electricity.
  - ✓All metals are solids at room temperature, except mercury (Hg) is a liquid.
  - ✓They can be pounded into flat sheets (malleability).
  - ✓They can be drawn into wires (ductility).
  - ✓They are often shiny.
  - ✓They tend to lose electrons when they undergo chemical changes (forming cations).

- About 75% of the elements in the period table are metals.

Classification of Elements: Nonmetals

- **Nonmetals** lie on the upper right side of the periodic table.

- **Properties of Nonmetals**:

  - ✓ Poor conductors of heat and electricity.
  - ✓ Can be found in all three states of matter (gases, liquids & solids).
  - ✓ Nonmetals with Solid state are brittle (not ductile and not malleable).
  - ✓ They tend to gain electrons when they undergo chemical changes (forming anions).

- **Metalloids** are elements that lie along the zigzag line that divides metals and nonmetals in the periodic table.
- **Properties of Metalloids:**
  - ✓Can exhibit mixed properties of both metals and nonmetals.
  - ✓Solids at room temperature.
  - ✓Known as **semiconductors** for electricity.
  - ✓Poor conductors of heat.

The Modern Periodic Table: Main-group Elements & Transition Elements

Main-group elements | Transition elements | Main-group elements

Group number

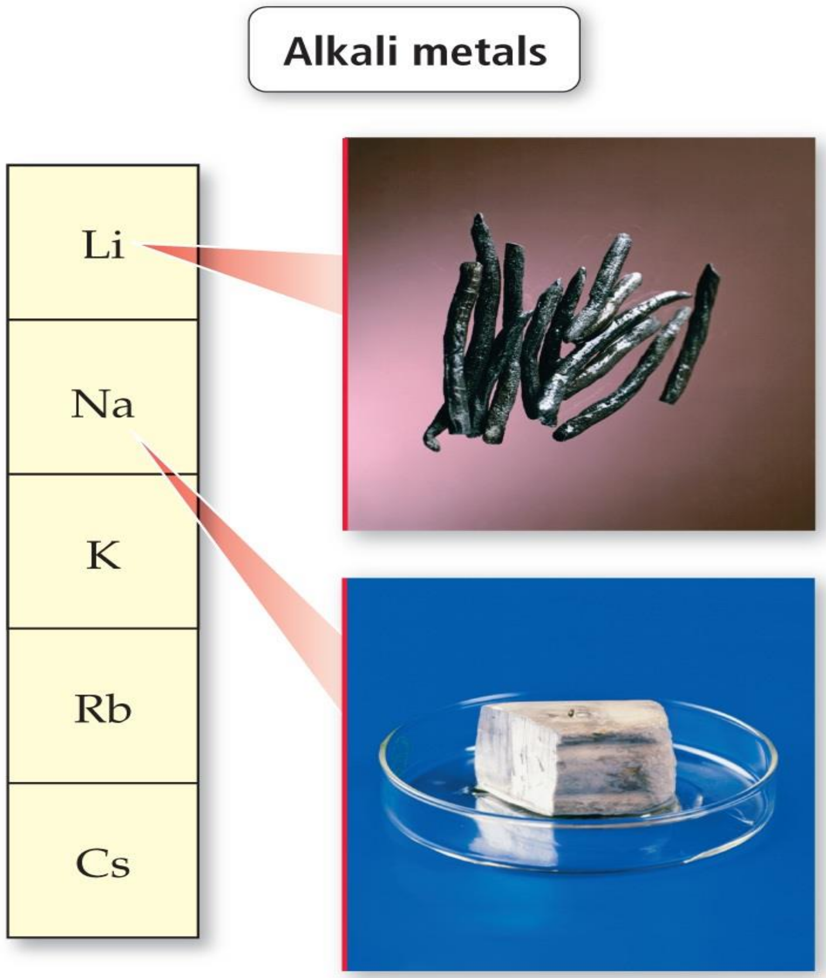
Periods	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113	114	115	116	117	118

- **Main-group elements** (groups with letter **A**): their properties are **largely predictable.**
- **Transition elements** or transition metals (groups with letter **B**): their properties are **less predictable**.

Major Families: Alkali Metals (Group 1A)

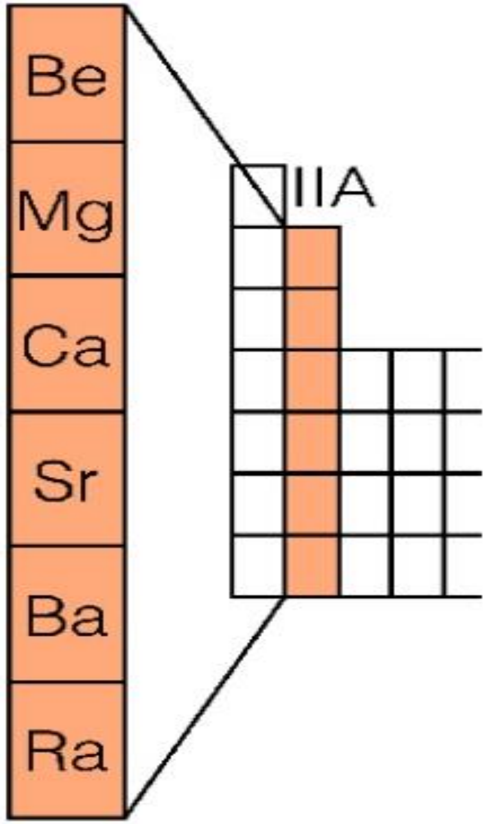
- **The group 1A** elements, called the **alkali metals**, are all highly reactive metals.
  - A marble-sized piece of sodium explodes violently when dropped into water.

Lithium,potassium,and rubidium are also alkali metals.



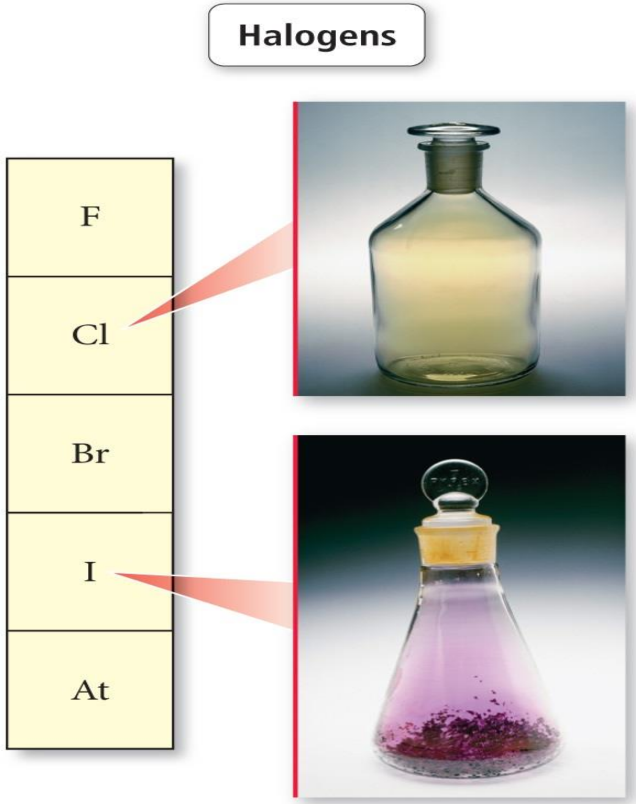
Major Families: Alkaline Earth Metals (Group 2A)

- **The group 2A** elements are called the
- **alkaline earth metals.**
- They are fairly reactive,
- but not quite as reactive as the alkali metals (group 1A).
  - Calcium, for example, reacts fairly vigorously with water.
  - Other alkaline earth metals include magnesium
  - (a common low-density structural metal), strontium, and barium.



Major Families: Halogens (Group 7A)

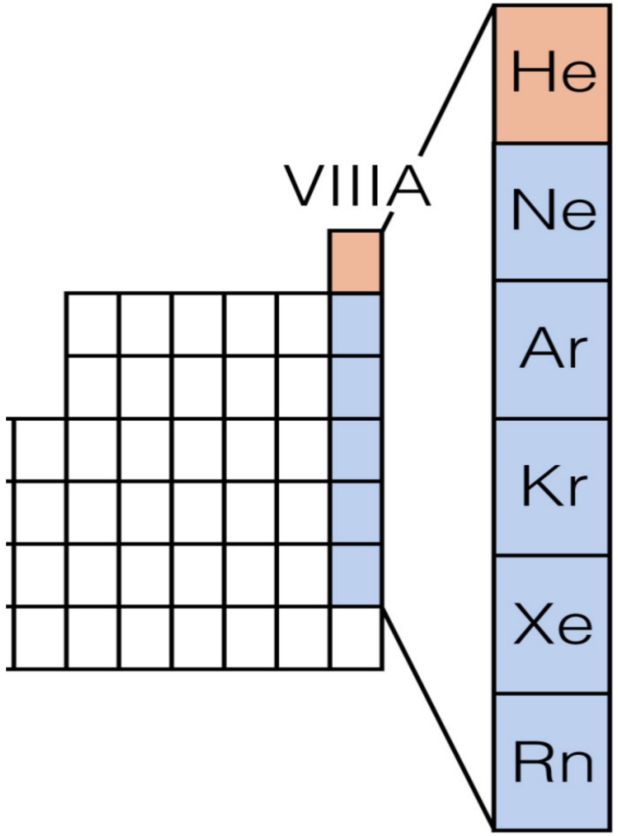
- **The group 7A** elements, the **halogens**, are very reactive nonmetals.



- They are always found in nature as a salt.
- **Chlorine, a greenish-yellow gas with a pungent odor**
- **Bromine, a red-brown liquid that easily evaporates into a gas**
- **Iodine, a purple solid**
- **Fluorine, a pale-yellow gas**

Major Families: Noble Gases (Group 8A)

- **The group 8A** elements, called the **noble gases**,
- are mostly unreactive (inert).
  - The most familiar noble gas is helium, used to fill buoyant balloons.
  - Other noble gases are neon (often used in electronic signs),
  - argon (a small component of our atmosphere), krypton, and xenon.



Ions and the Periodic Table

- **A main-group metal**
- tends to lose electrons, forming a cation
- with the same number of electrons as the nearest noble gas.
- **A main-group nonmetal**
- tends to gain electrons, forming an anion
- with the same number of electrons as the nearest noble gas.

- For the **main-group elements**

that form cations with predictable charge,

the charge is equal to the group number:

(for example: sodium, Na, of group 1A, forms the cation Na^{1+}).

- For the **main-group elements**

that form anions with predictable charge,

the charge is equal to the group number minus eight:

(for example: nitrogen, N, of group 5A, forms the cation N^{3-}).

- **Transition elements:** may form different ions with variable charges:
  - ✓ **(for example: iron (Fe) can form the cations: Fe^{2+} or Fe^{3+}**
  - ✓ **, also copper (Cu) can form the cations: Cu^{1+} or Cu^{2+}).**

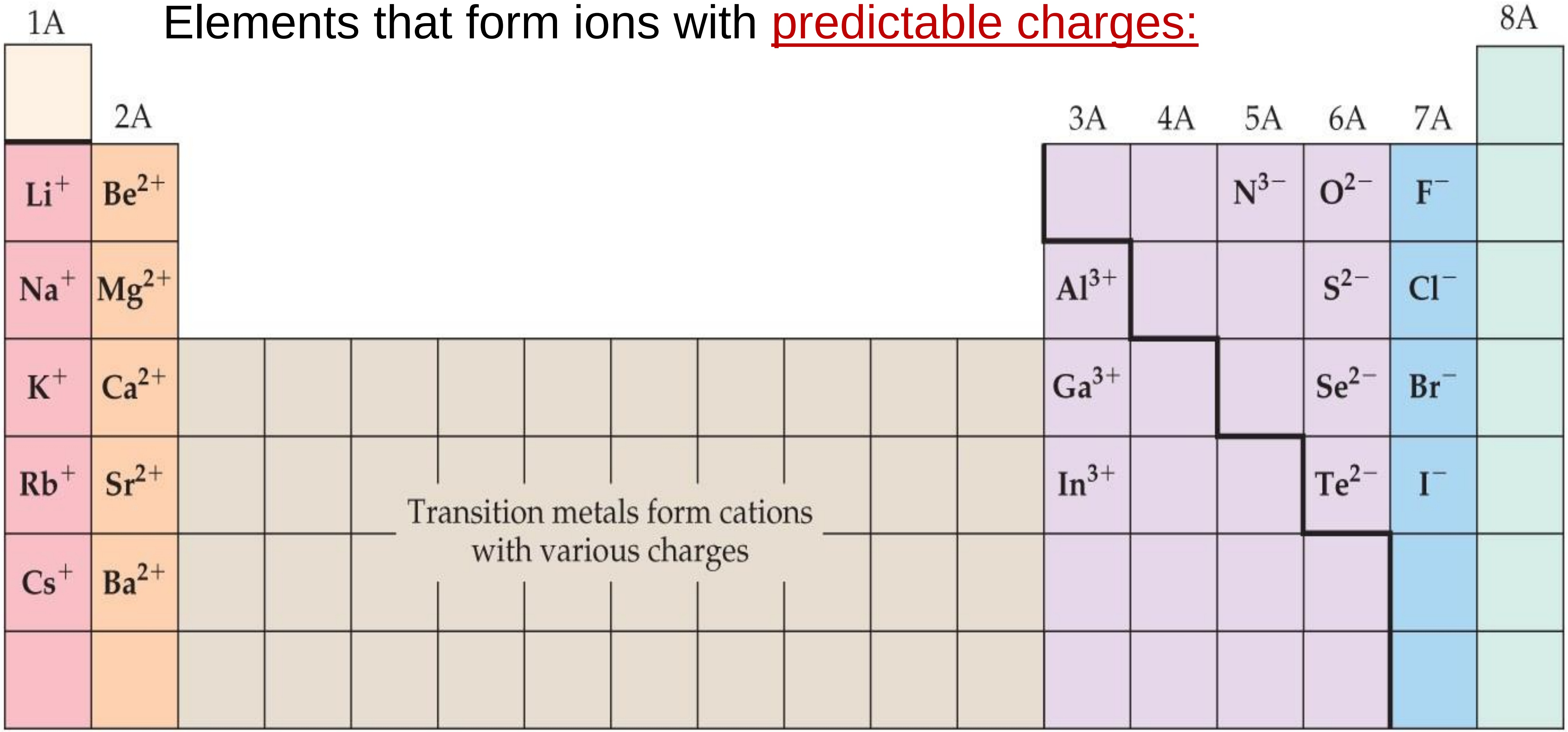
Ions and the Periodic Table

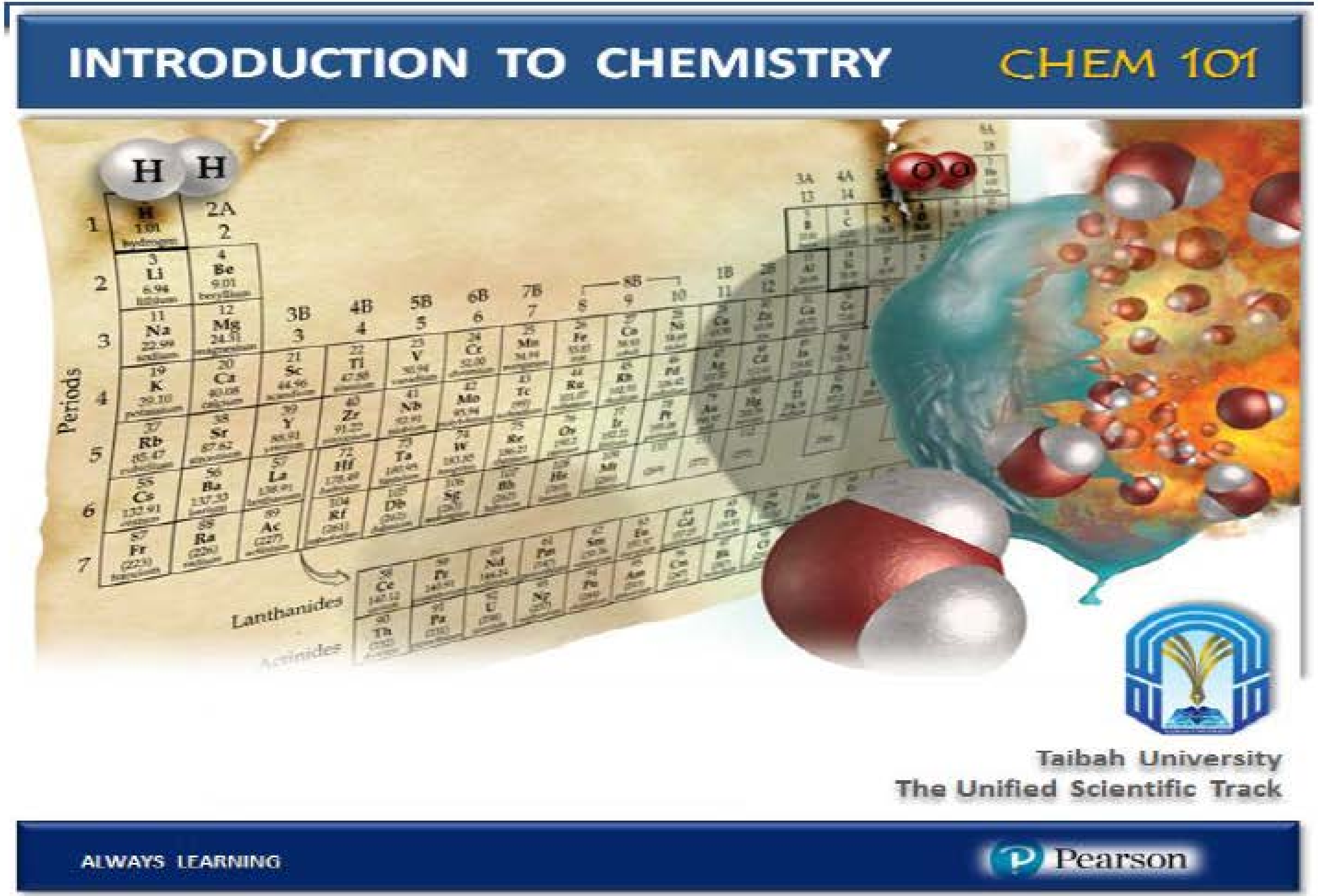
- **In general, the charge of ions of main-group elements can be predicted from their group number:**

- The alkali metals (group 1A): tend to lose one electron to form +1 ions.
- The alkaline earth metals (group 2A): tend to lose two electrons to form +2 ions.
- The halogens (group 7A): tend to gain one electron to form −1 ions.
- The oxygen family nonmetals (group 6A): tend to gain two electrons to form −2 ions.

Ions and the Periodic Table

Elements that form ions with predictable charges:





Lecture Presentation

Chapter 2

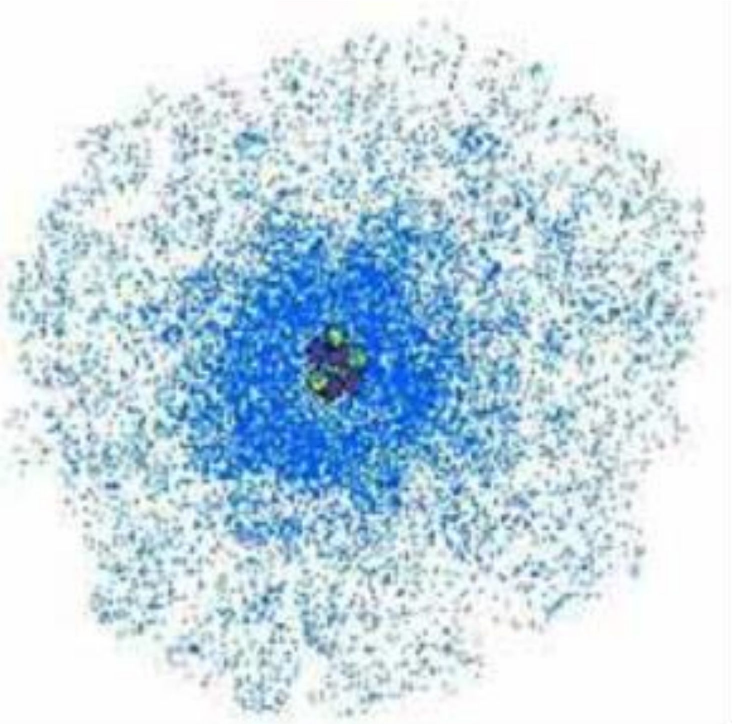
Atoms, Molecules, Ions, and Periodicity

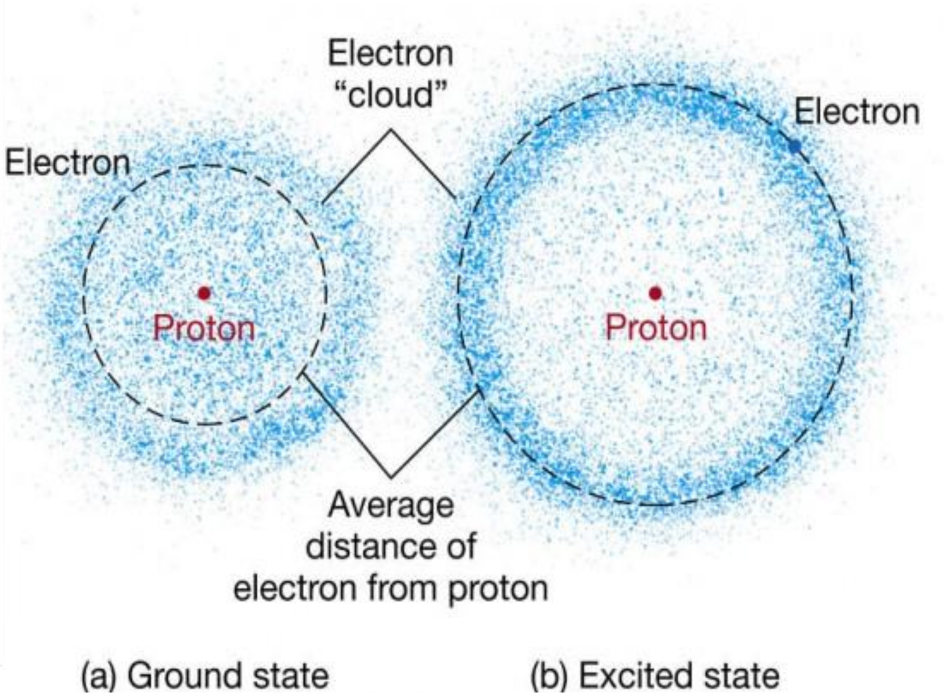
Topic 06

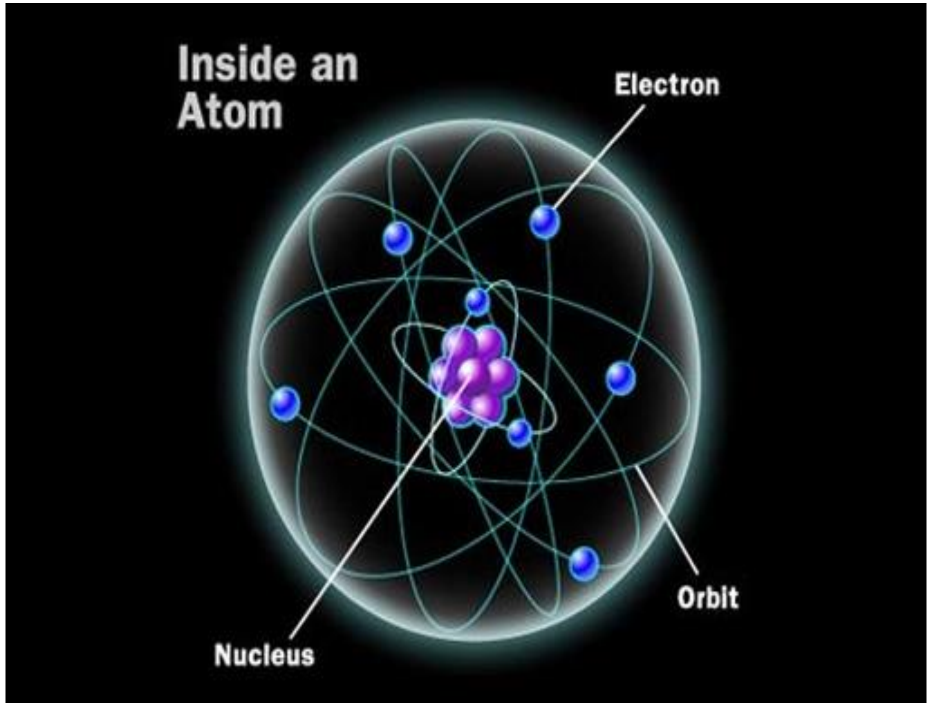
Electron Configurations

2.9- Orbitals and Quantum Numbers

- **Niels Bohr's Model:** the electrons move in spherical orbits at fixed distances from the nucleus (similar to structure of the solar system).
- **Erwin Schrödinger** develops mathematical equations to describe the motion of electrons in atoms. His work leads to the electron cloud model.







➢ According to Quantum Mechanics:

Electron location around the atom's nucleus is described by the four quantum numbers:

- n (principle energy level)
- l (orbital type: s, p, d, f…)
- m_l (orientation of orbital)
- m_s (spin of electron in orbital)

➢ Principal Quantum Number, *n*

- The principal quantum number,

n, <u>describes the energy level</u> on which the orbital resides.

- The values of *n* are integers > 0

$$n = 1, 2, 3, 4, 5, 6, 7$$

➢ Angular momentum Quantum Number, l

- This quantum number defines the shape of the orbital.
- Allowed values of l are integers ranging from 0 to $n-1$.
- We use letter designations to communicate the different values of l and, therefore, the shapes and types of orbitals.

Value of l	0	1	2	3
Type of orbital	*s*	*p*	*d*	*f*

- Magnetic Quantum Number, m_l

  - Describes the three-dimensional orientation of the orbital.

  - Values are integers ranging from $-l \leq m_l \leq l$

$$m_l = -l\ ,\ (-l+1)\ ,\ (-l+2)\ ,\ldots,\ -2\ ,\ -1\ ,\ 0\ ,\ 1\ ,\ 2\ ,\ldots (l-1)\ ,\ (l-2)\ ,\ +l$$

➢ Spin Quantum Number, m_s

- It designates the direction of the electron spin
- and may have a spin of +1/2, represented by ↑, or –1/2, represented by ↓.
- The significance of the electron spin quantum number is its determination of an atom's ability to generate a magnetic field or not.

Ex-)-Which one of the following sets of quantum numbers is not possible?

	n	l	m_l	m_s
Row 1	4	3	−2	+1/2
Row 2	3	0	1	−1/2
Row 3	3	0	0	+1/2
Row 4	2	1	1	−1/2
Row 5	2	0	0	+1/2

A. Row 1   B. Row 2   C. Row 3   D. Row 4   E. Row 5

Electrons in an orbital with $l = 3$ are in a

A. *d* orbital.
B. *f* orbital.
C. *g* orbital.
D. *p* orbital.
E. *s* orbital

- Which of the following sets of quantum numbers refers to a 3p orbital?

a- **n** = 3, l = 0, **m**$_l$ = 0, **m**$_s$ = + 1/2

[redacted]- **n = 3, l = 1, m$_l$ = -1, m$_s$ = + 1/2**

c- **n** = 3, l = 2, **m**$_l$ = 1, **ms** = + 1/2

d- **n** = 3, l = 3, **m**$_l$ = -2, **m**$_s$ = + 1/2

Which of the following sets of quantum numbers refers to a 2s orbital?

a- **n** = 1, l = 2, **m**$_l$ = 2, **m**$_s$ = + 1/2

b- **n** = 1, l = 2, **m**$_l$ = 1, **m**$_s$ = + 1/2

c- **n** = 2, l = 2, **m**$_l$ = 0, **m**$_s$ = + 1/2

[redacted] **n = 2, l = 0, m$_l$ = 0, m$_s$ = + 1/2**

- How many orbitals have the following quantum numbers: **n** = 6, l = 2 , **m**$_l$ = -2?

a- 0 [redacted] 1 c- 3 d- 6

The lowest energy state of an atom is referred to as its

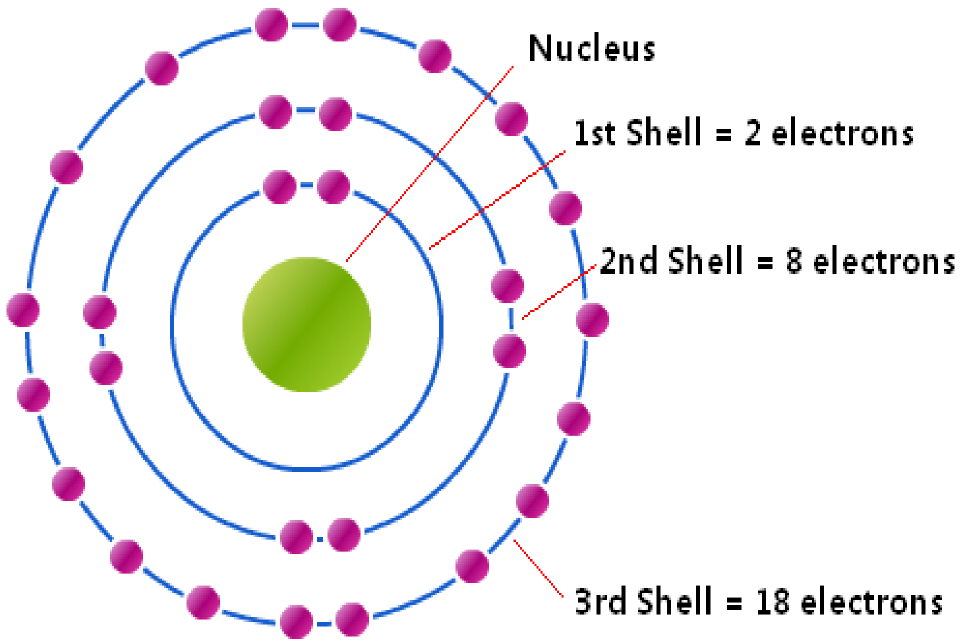
a) bottom state.
b) **ground state.**
c) fundamental state.
d) original state.

.All s orbitals are

a) shaped like four-leaf clovers.
b) dumbbell-shaped.
c) spherical.
d) triangular.

Electron configuration:

is how the electrons are distributed among the various atomic orbitals in an atom.



2.10- Electron Configuration: *s, p, d and f* Sublevels

- **The number of orbitals and maximum number of electrons in each sublevel:**

✓ Each **orbital** in any sublevel is able to hold a maximum of 2 electrons:

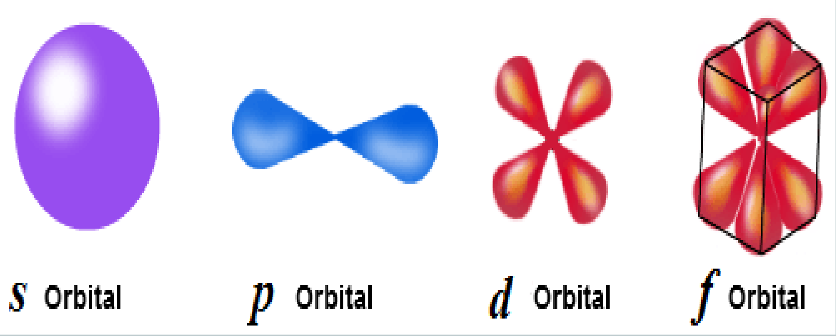
- The ***s*** sublevel has only one orbital and can therefore hold only 2 electrons.
- The ***p*** sublevel has three orbitals and can therefore hold 6 electrons.
- The ***d*** sublevel has five orbitals and can therefore hold 10 electrons.
- The ***f*** sublevel has seven orbitals and can therefore hold 14 electrons.

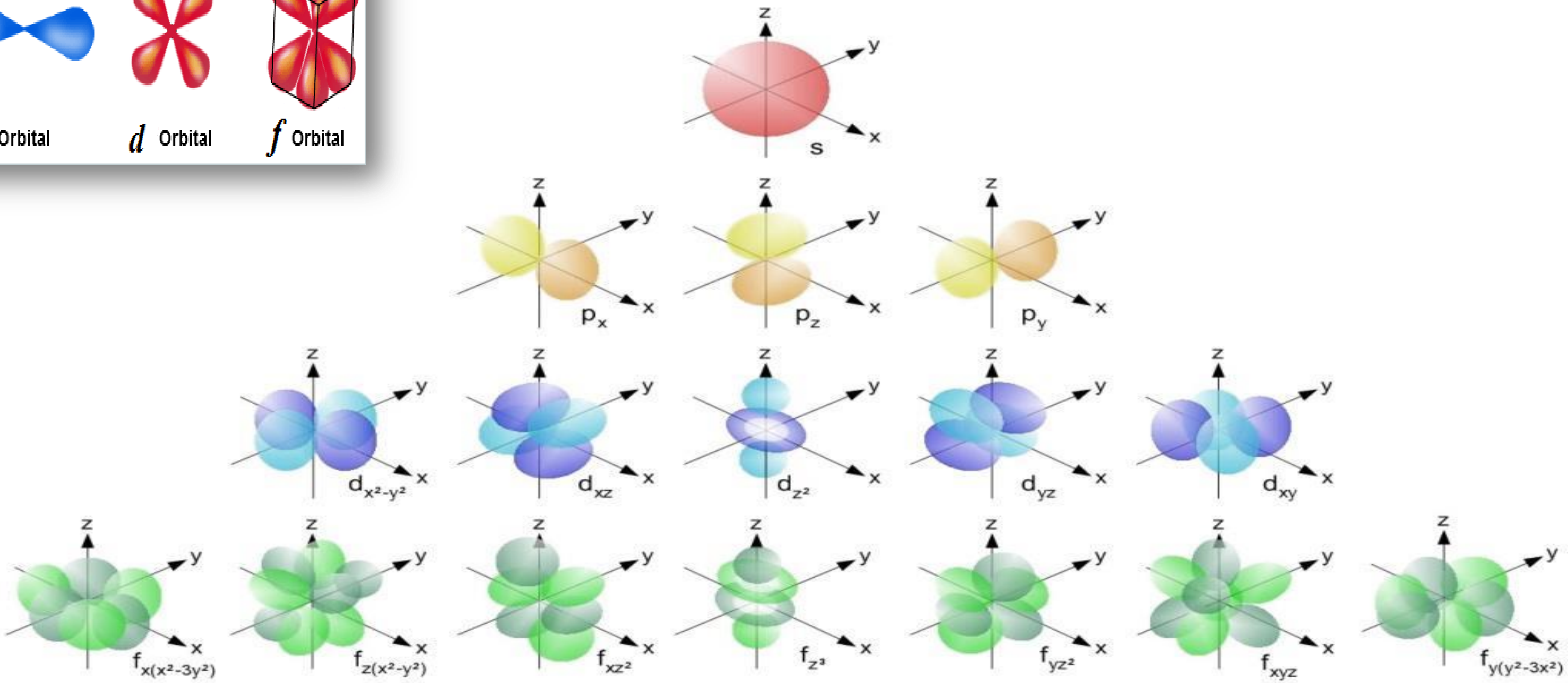
✓ The maximum number of electrons that can occupy a specific energy level can be calculated using the following formula:

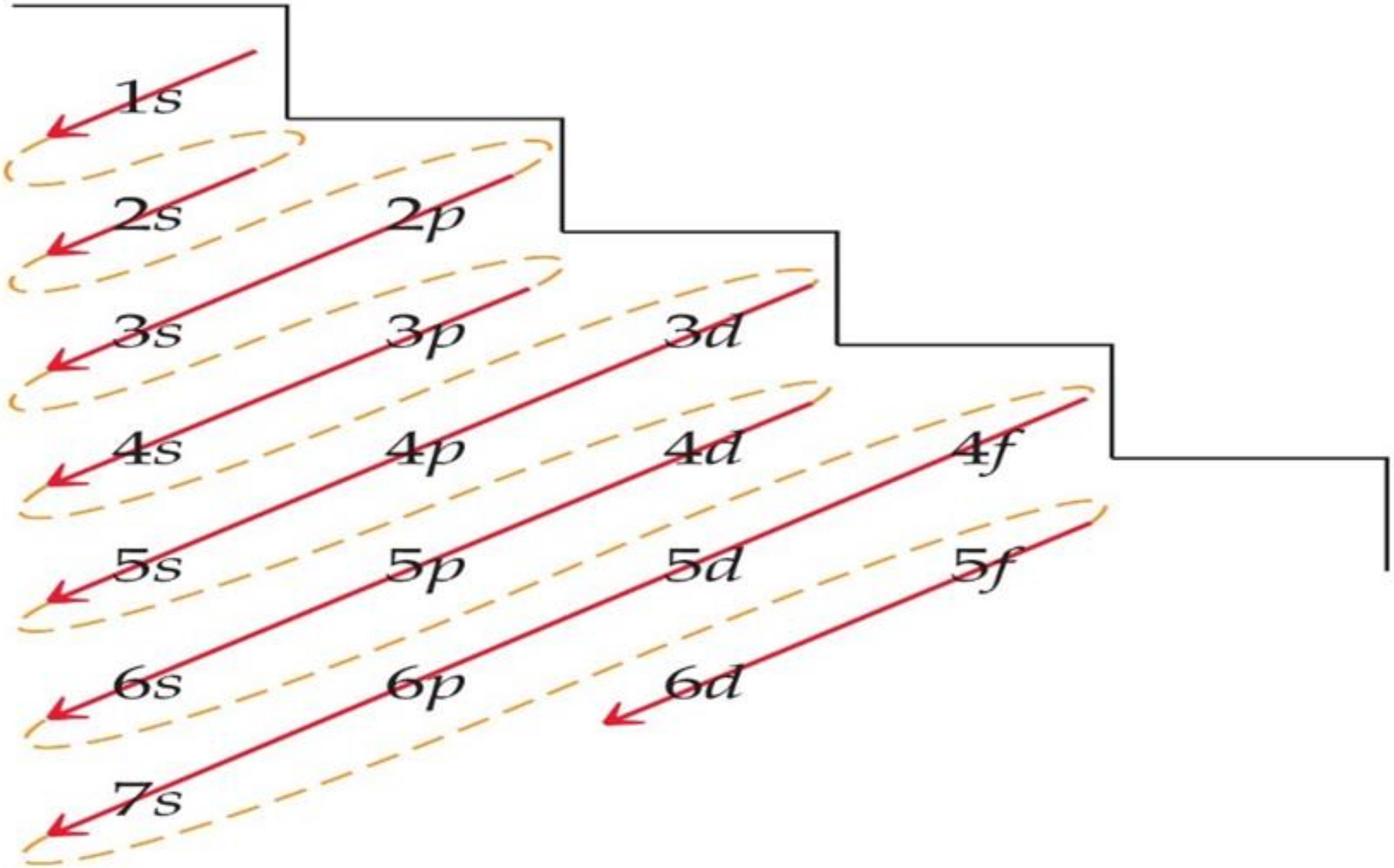
where n = the principal quantum number (the number of the energy level).

Electron Capacity = $2n^2$

2.10- Electron Configuration: Shapes of *s, p, d & f* orbitals






1s
2s
2p
3s
3p
3d
4s
4p
4d
4f
5s
5p
5d
5f
6s
6p
6d
7s

. What is the maximum number of orbitals described by the quantum numbers: $n = 3 \quad l = 2$

a) 1
b) 3
c) 5
d) 9

). What is the maximum number of orbitals described by the quantum numbers: $n = 4$

a) 7
b) 14
c) 32
d) 48

..The maximum number of electrons that can occupy an energy level described by the principal quantum number, n, is

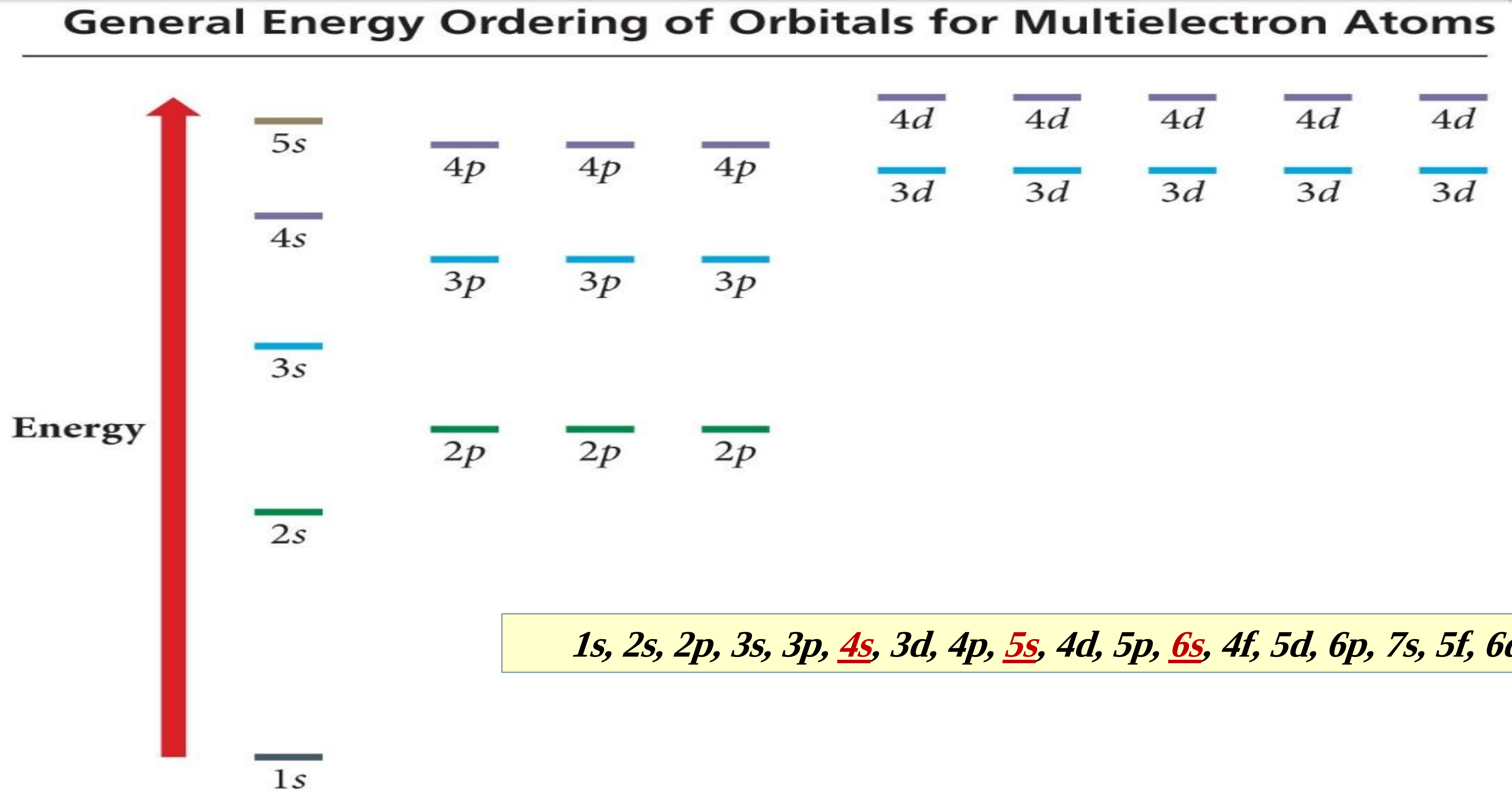
a) $n + 1$
b) $2n$
c) $2n^2$
d) n^2

1-Aufbau Principle ("Fill up" electrons):
the electrons are added one by one to the atomic orbitals in lowest energy orbitals.

2-The Pauli Exclusion Principal:
NO two electrons in an atom can have the same set of four quantum numbers.

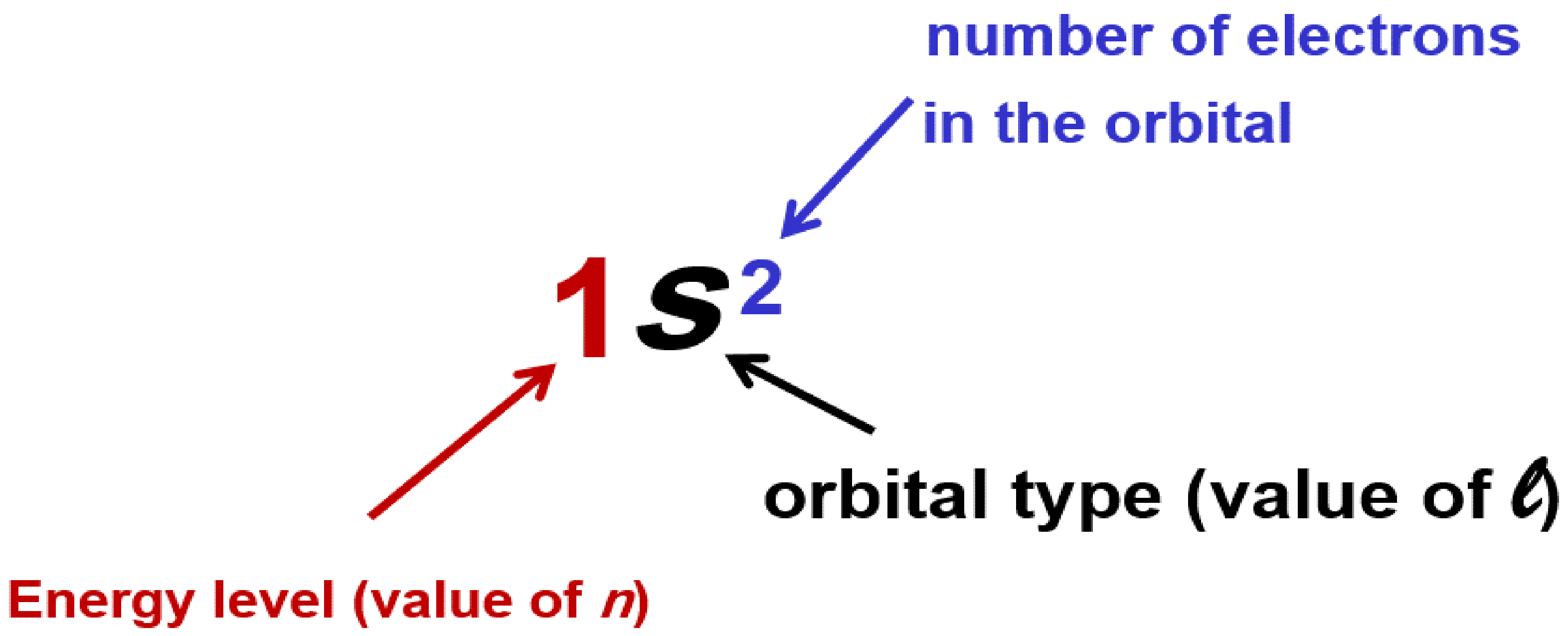
3-Hund's Rule:
The most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins.

Electron Configurations: Ordering of Orbital Filling



1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p

Example: the electron configuration for **He**, atomic number = 2



2.10 Electron Configurations: Representing The Electron Configurations of Atoms

Example: The four quantum numbers for each of the two electrons in helium atom:

	Electron Configuration	Orbital diagram
He	$1s^2$	↑↓ $1s$

n	l	m_l	m_s
1	0	0	$+\frac{1}{2}$
1	0	0	$-\frac{1}{2}$

- **Rules of the aufbau principle** (aufbau: is a German word meaning “building”):

1. Lower-energy orbitals fill before higher-energy orbitals.
2. An orbital can hold only two electrons,

   which must have opposite spins (**Pauli exclusion principle**).

1. If two or more degenerate orbitals are available, follow **Hund’s rule**.

Hund’s Rule:

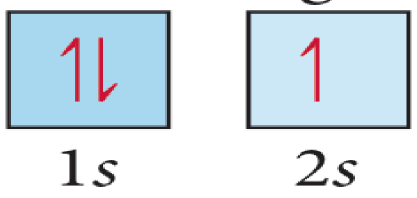
when filling degenerate orbitals, the electrons fill them singly first, with parallel spins.

Electron Configurations: Examples

- **Lithium (Li)** has an atomic number of 3, so to be neutral it must have 3 electrons:

Electron configuration	
Li	$1s^22s^1$

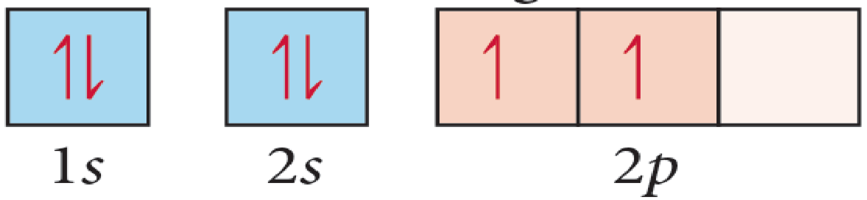
Orbital diagram



- **Carbon (C)** has an atomic number of 6, so to be neutral it must have 6 electrons:

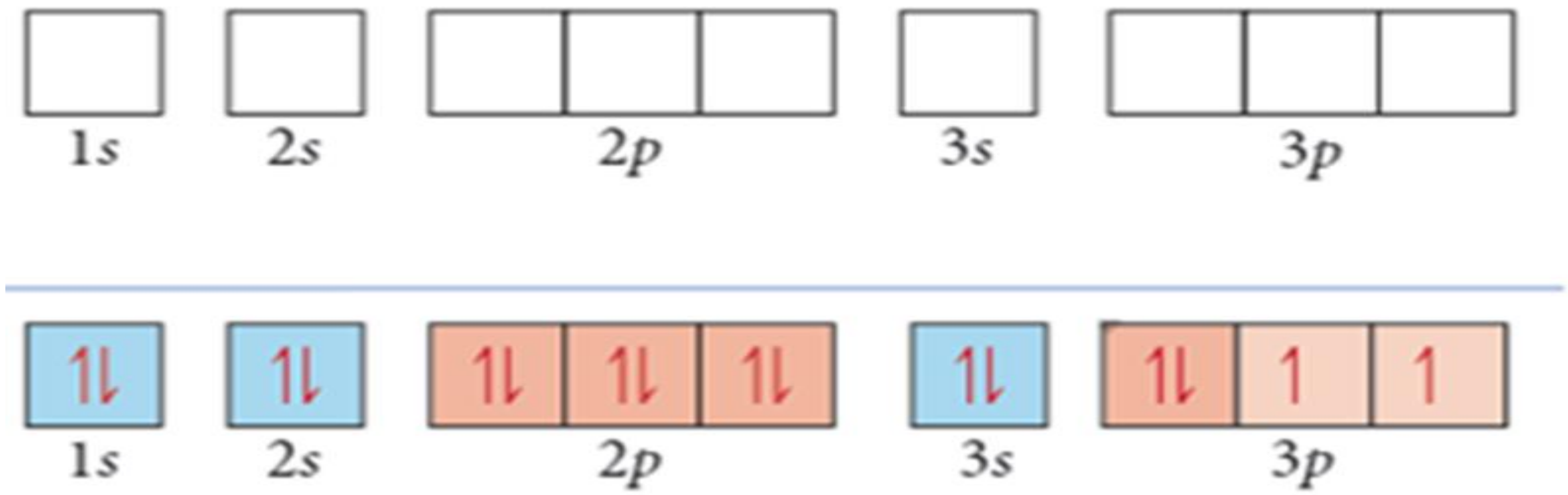
Electron configuration	
C	$1s^22s^22p^2$

Orbital diagram



Writing Orbital Diagrams

Write an orbital diagram for sulfur and determine the number of unpaired electrons.



Example: Write the electron configuration for the following elements:

Mg, P, Br, and Al

Mg	$1s^2\,2s^2\,2p^6\,3s^2$
P	$1s^2\,2s^2\,2p^6\,3s^2\,3p^3$
Br	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^5$
Al	$1s^2\,2s^2\,2p^6\,3s^2\,3p^1$

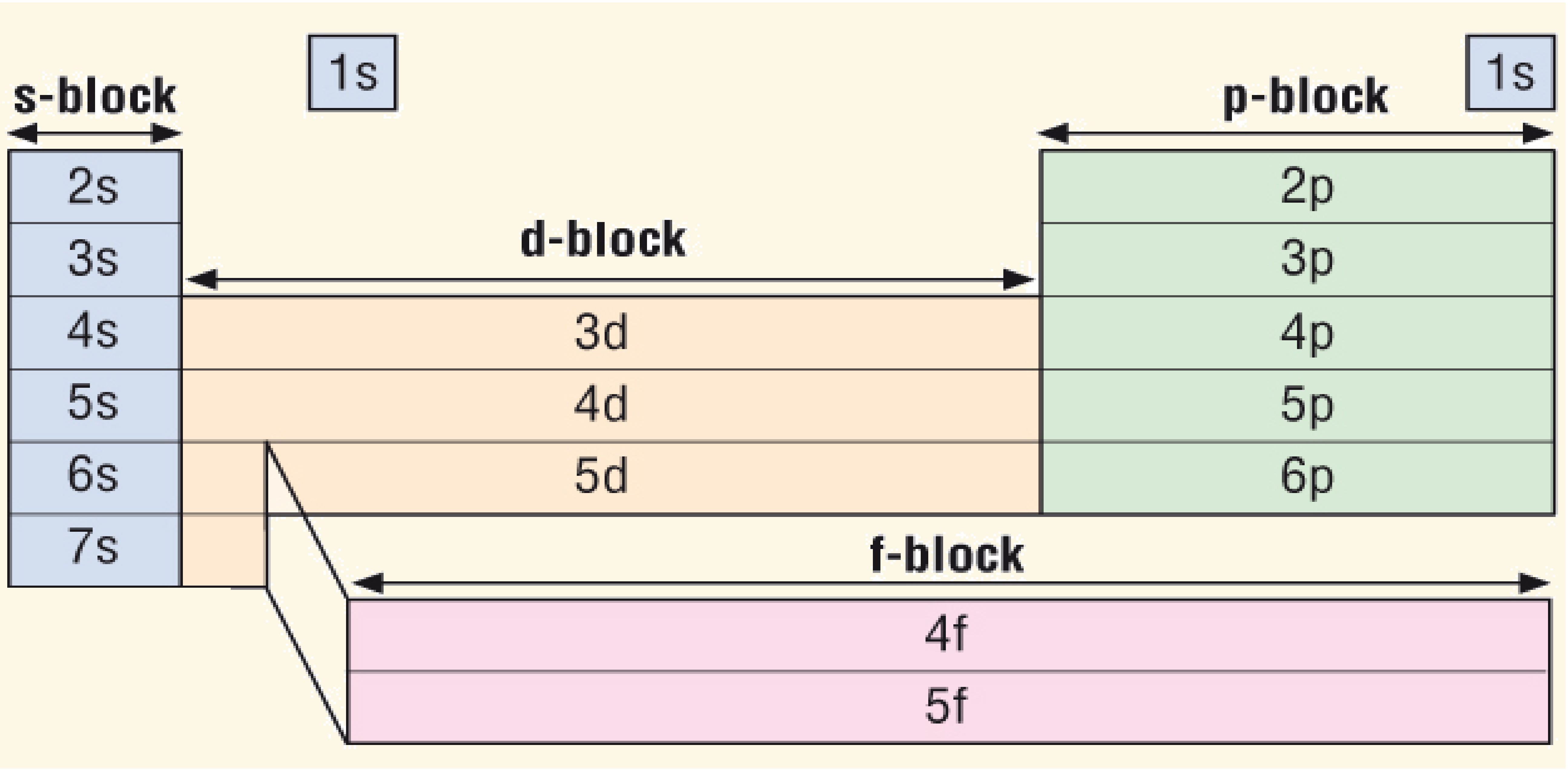
Electron Configurations: Examples

Electron Configurations of the First Ten Elements

	Electron Configurations		Orbital Box Diagrams				
	Condensed	*Expanded*	1*s*	2*s*	2*p*		
H	$1s^1$		↑				
He	$1s^2$		↑↓				
Li	$1s^2 2s^1$		↑↓	↑			
Be	$1s^2 2s^2$		↑↓	↑↓			
B	$1s^2 2s^2 2p^1$		↑↓	↑↓	↑		
C	$1s^2 2s^2 2p^2$	$1s^2 2s^2 2p^1 2p^1$	↑↓	↑↓	↑	↑	
N	$1s^2 2s^2 2p^3$	$1s^2 2s^2 2p^1 2p^1 2p^1$	↑↓	↑↓	↑	↑	↑
O	$1s^2 2s^2 2p^4$	$1s^2 2s^2 2p_x^2 2p^1 2p^1$	↑↓	↑↓	↑↓	↑	↑
F	$1s^2 2s^2 2p^5$	$1s^2 2s^2 2p^2 2p^2 2p^1$	↑↓	↑↓	↑↓	↑↓	↑
Ne	$1s^2 2s^2 2p^6$	$1s^2 2s^2 2p^2 2p^2 2p^2$	↑↓	↑↓	↑↓	↑↓	↑↓

. A possible set of quantum numbers for the last electron added to complete an atom of sodium Na in its ground state is

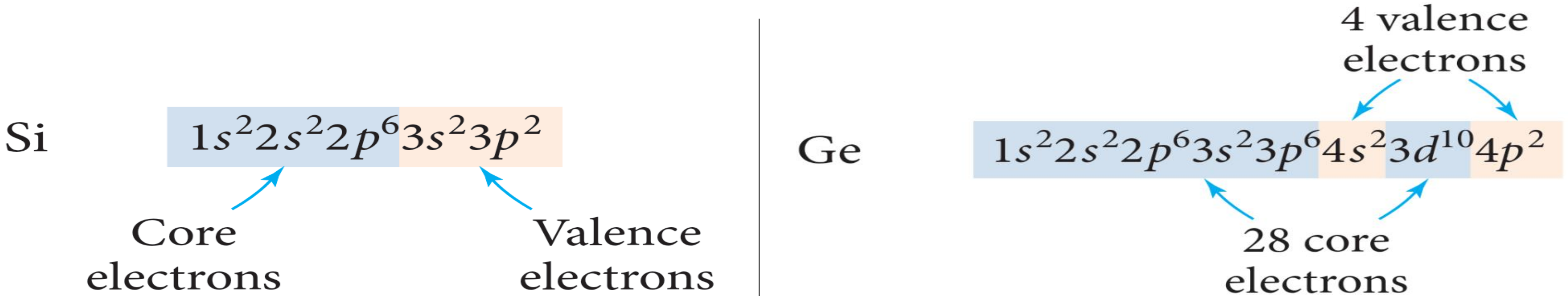
a) $n = 3$, $l = 1$, $m_l = 0$, $m_s = ½$
b) $n = 3$, $l = 0$, $m_l = 0$, $m_s = ½$
c) $n = 2$, $l = 1$, $m_l = -1$, $m_s = ½$
d) $n = 2$, $l = 0$, $m_l = -1$, $m_s = ½$


1s
s-block
2s
3s
4s
5s
6s
7s
d-block
3d
4d
5d
1s
p-block
2p
3p
4p
5p
6p
f-block
4f
5f

2.11 Electron Configurations: Valence Electrons & Core Electrons

- **Valence Electrons:** electrons in all the sublevels within the highest principal energy level (***n***).
  - ✓ One of the most important factors in the way an atom behaves, both chemically and physically, is the number of its "valence electrons".
  - ✓ The highest principal energy level is also known as "the valence shell"
  - ✓ Valence electrons in atoms participate in:
    - ✓ Bonding
    - ✓ Making cations (by losing e^-)
    - ✓ Making anions (by gaining e^-)
- **Core Electrons:** electrons in all lower energy levels (i.e. all shells except the valence shell).

Example: How many valence and core electrons are in **Si** and **Ge** atoms?



Exercise: Draw the orbital diagram and indicate how many valence and core electrons are in: Ne, Kr, Al, Cl, O, F, S and Be neutral atoms (atoms in their ground states, i.e. not ions)?

Exercise: Draw the orbital diagram and indicate how many valence and core electrons are in:

Ne, Kr, Al, Cl, O, F, S and Be neutral atoms (atoms in their ground states, i.e. not ions)?

Orbital Blocks of the Periodic Table

s-block elements
d-block elements
p-block elements
f-block elements

Groups

Periods	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1	1 **H** $1s^1$																	2 **He** $1s^2$
2	3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
3	11 **Na** $3s^1$	12 **Mg** $3s^2$											13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^23d^1$	22 **Ti** $4s^23d^2$	23 **V** $4s^23d^3$	24 **Cr** $4s^13d^5$	25 **Mn** $4s^23d^5$	26 **Fe** $4s^23d^6$	27 **Co** $4s^23d^7$	28 **Ni** $4s^23d^8$	29 **Cu** $4s^13d^{10}$	30 **Zn** $4s^23d^{10}$	31 **Ga** $4s^24p^1$	32 **Ge** $4s^24p^2$	33 **As** $4s^24p^3$	34 **Se** $4s^24p^4$	35 **Br** $4s^24p^5$	36 **Kr** $4s^24p^6$
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^14d^4$	42 **Mo** $5s^14d^5$	43 **Tc** $5s^24d^5$	44 **Ru** $5s^14d^7$	45 **Rh** $5s^14d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^14d^{10}$	48 **Cd** $5s^24d^{10}$	49 **In** $5s^25p^1$	50 **Sn** $5s^25p^2$	51 **Sb** $5s^25p^3$	52 **Te** $5s^25p^4$	53 **I** $5s^25p^5$	54 **Xe** $5s^25p^6$
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	57 **La** $6s^25d^1$	72 **Hf** $6s^25d^2$	73 **Ta** $6s^25d^3$	74 **W** $6s^25d^4$	75 **Re** $6s^25d^5$	76 **Os** $6s^25d^6$	77 **Ir** $6s^25d^7$	78 **Pt** $6s^15d^9$	79 **Au** $6s^15d^{10}$	80 **Hg** $6s^25d^{10}$	81 **Tl** $6s^26p^1$	82 **Pb** $6s^26p^2$	83 **Bi** $6s^26p^3$	84 **Po** $6s^26p^4$	85 **At** $6s^26p^5$	86 **Rn** $6s^26p^6$
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	89 **Ac** $7s^26d^1$	104 **Rf** $7s^26d^2$	105 **Db** $7s^26d^3$	106 **Sg** $7s^26d^4$	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112 **Cn**	113 **Uut**	114 **Uuq**	115 **Uup**	116 **Uuh**	117 **Uus**	118 **Uuo**

Lanthanides	58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}6d^1$
Actinides	90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$	103 **Lr** $7s^25f^{14}6d^1$

Electron configurations of Ions

- The **sulfur atom** has 6 valence electrons:

  S atom = $1s^2 2s^2 2p^6 3s^2 3p^4$

– To have 8 valence electrons, sulfur must gain 2 more e^- forming anion:

  S^{2-} anion = $1s^2 2s^2 2p^6 3s^2 3p^6$

- The **magnesium atom** has 2 valence electrons:

  Mg atom = $1s^2 2s^2 2p^6 3s^2$

– When magnesium forms a cation, it loses its 2 valence electrons:

  Mg^{2+} cation = $1s^2 2s^2 2p^6$

Assessment

Answer the following questions:

1- Name an element in the fourth period of the periodic table with:

a. five valence electrons   b. a complete outer shell

2- Write full orbital diagrams for each element:

a. N   b. F   c. Mg   d. Al   e. K

3- Determine the number of valence electrons in each element.

a. Ba   b. Cs   c. Ne   d. S   e. C

4- The complete electron configuration of sulfur is ________.

A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

B) $1s^2\ 2s^2\ 2p^{10}\ 3s^2$

C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^2$

E) $1s^4\ 2s^4 2\ p^6\ 3s^2$

5- Which one of the following is the correct electron configuration for a ground-state nitrogen atom?

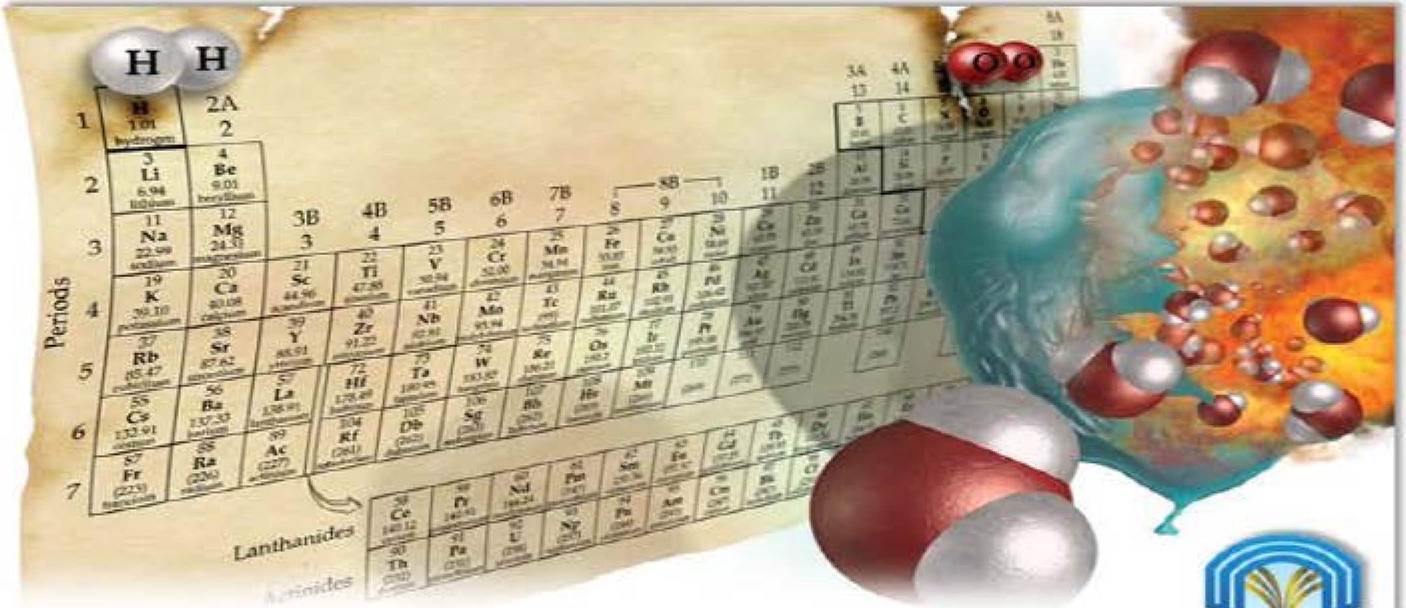
A) 1s [↑↓] 2s [↑↓] 2p [↑↓][↑][ ]

B) 1s [↑↓] 2s [↑↑] 2p [↑][↑][↑]

C) 1s [↑↑] 2s [↑↓] 2p [↑][↑][↑]

D) 1s [↑↓] 2s [↑↓] 2p [↑][↑][↑]

INTRODUCTION TO CHEMISTRY CHEM 101



Taibah University
The Unified Scientific Track

ALWAYS LEARNING Pearson

Lecture Presentation

Chapter 2

Atoms, Molecules, Ions, and Periodicity

Topic 07

The Periodic Trends

2.13 Periodic Trends: Moving Across The Periodic Table

- **Periodic Trends**: are the properties that show patterns when examined across the periodic table (i.e. when moving across periods or down the groups).

- The periodic trends of the following properties will be discussed:
    - ✓The Effective Nuclear Charge
    - ✓Atomic Radii (the sizes of atoms)
    - ✓Ionic Radii (the sizes of ions)
    - ✓Ionization Energy
    - ✓Electron Affinities
    - ✓Metallic Character
    - ✓Electronegativity

2.13 Periodic Trends: The Effective Nuclear Charge

➢ **Effective Nuclear Charge** (Z_{eff}): It is the pull force an electron "feels" from the nucleus (protons).

- The closer the electrons are to the nucleus, the greater the "pull" on the electrons.
- The greater the Z_{eff}, the more tightly the electrons are held and the more energy needed to remove the electrons.
  - Electrons located farthest from the nucleus experience less Z_{eff}.

➢ **General trend in Z_{eff} :**

✓ **Z_{eff} increases going across periods.**

✓ **Z_{eff} decreases going down groups.**

$$Z_{effective} = Z - S$$

Where, **Z** is the nuclear charge, and **S** is the number of electrons in lower energy levels.

2.13 Periodic Trends: The Effective Nuclear Charge

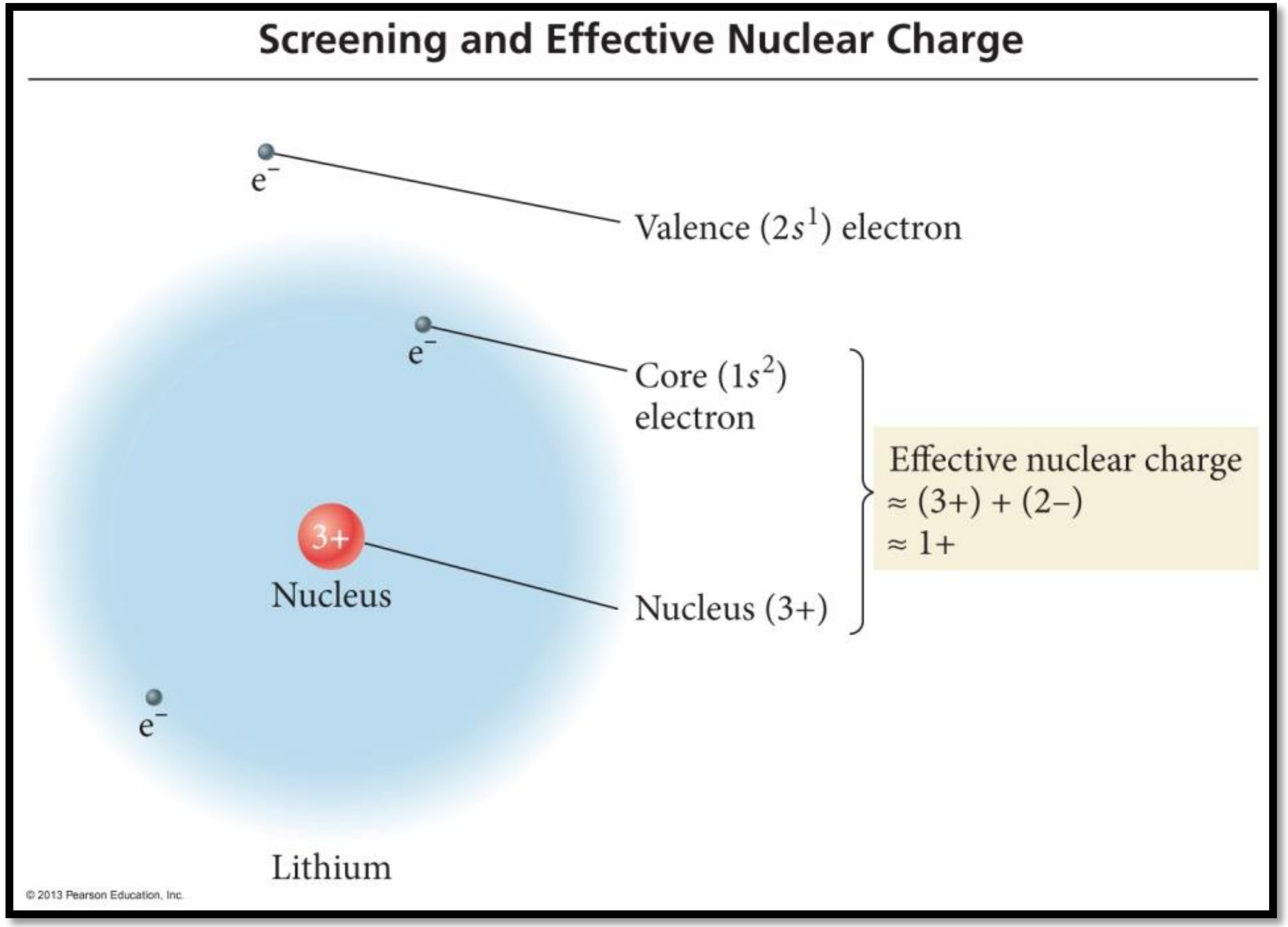
- Z_{eff} increases across a period owing to incomplete shielding by inner electrons in atomic orbitals (subshells).

- Shielding ability of subshells:

  s > p > d > f

- Estimate Z_{eff}

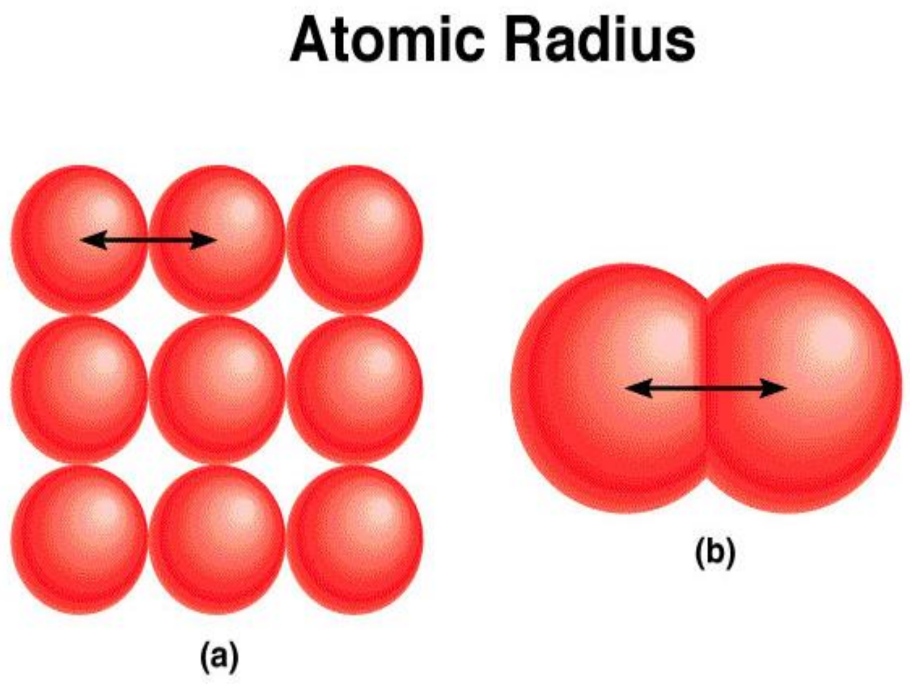
  = [Z (atomic number) - (number of inner electrons)]
  - Pull felt by ***2s*** electron in Li: Z_{eff} = 3 - 2 = 1



Atomic Radius: (نصف القطر الذري)

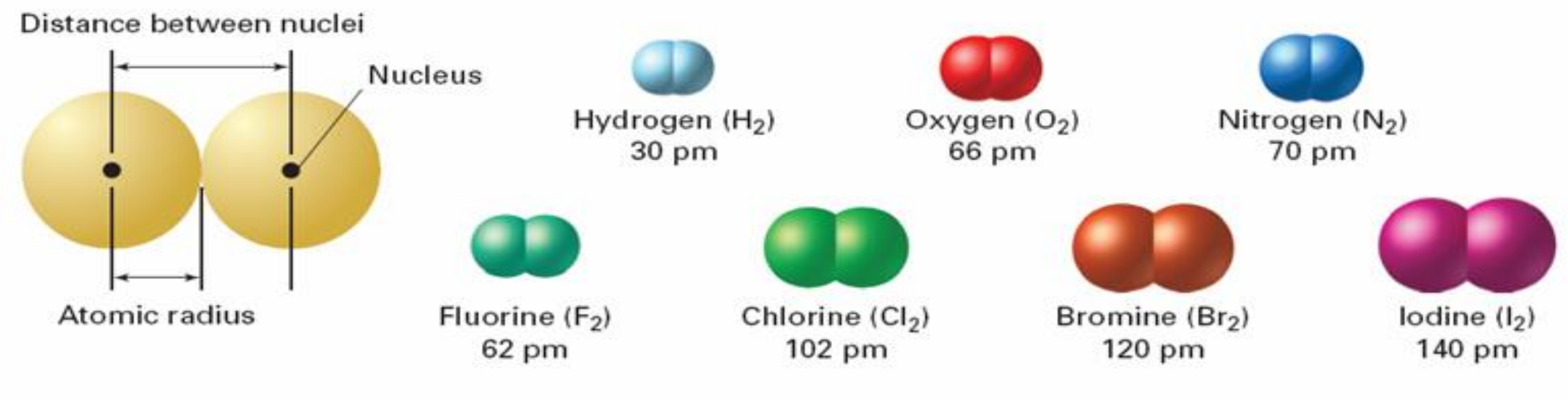
is a term used to describe the size of the atom,

and it is one-half (½) the distance between the two nuclei in two adjacent metal atoms (a) or in a diatomic molecule (b).



- **Atomic Radius:** is an average radius of an atom based on measuring large numbers of molecules of elements and compounds.

✓ There are several methods for measuring the radius of an atom, and they give slightly different numbers.
  - ✓ Van der Waals radius = nonbonding
  - ✓ Covalent radius = bonding radius



2.13 Periodic Trends: Atomic Radii (Sizes of Atoms)

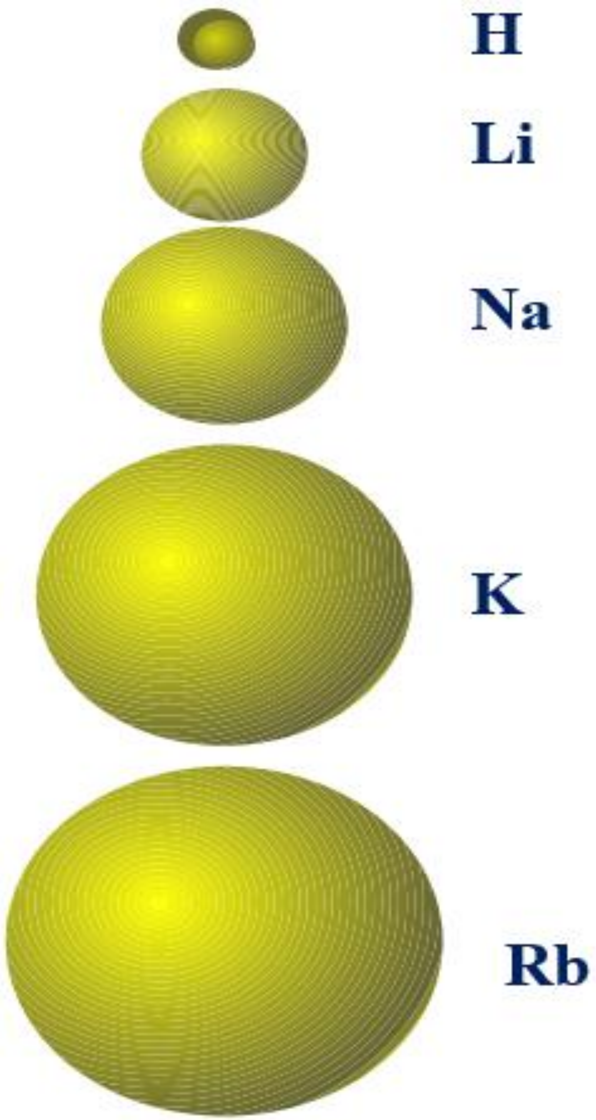
➢ **General trend in atomic radii:**

✓ Atomic radius decreases across period (left to right)

- ✓ Adding electrons to same valence shell
- ✓ Effective nuclear charge increases
- ✓ Valence shell held closer

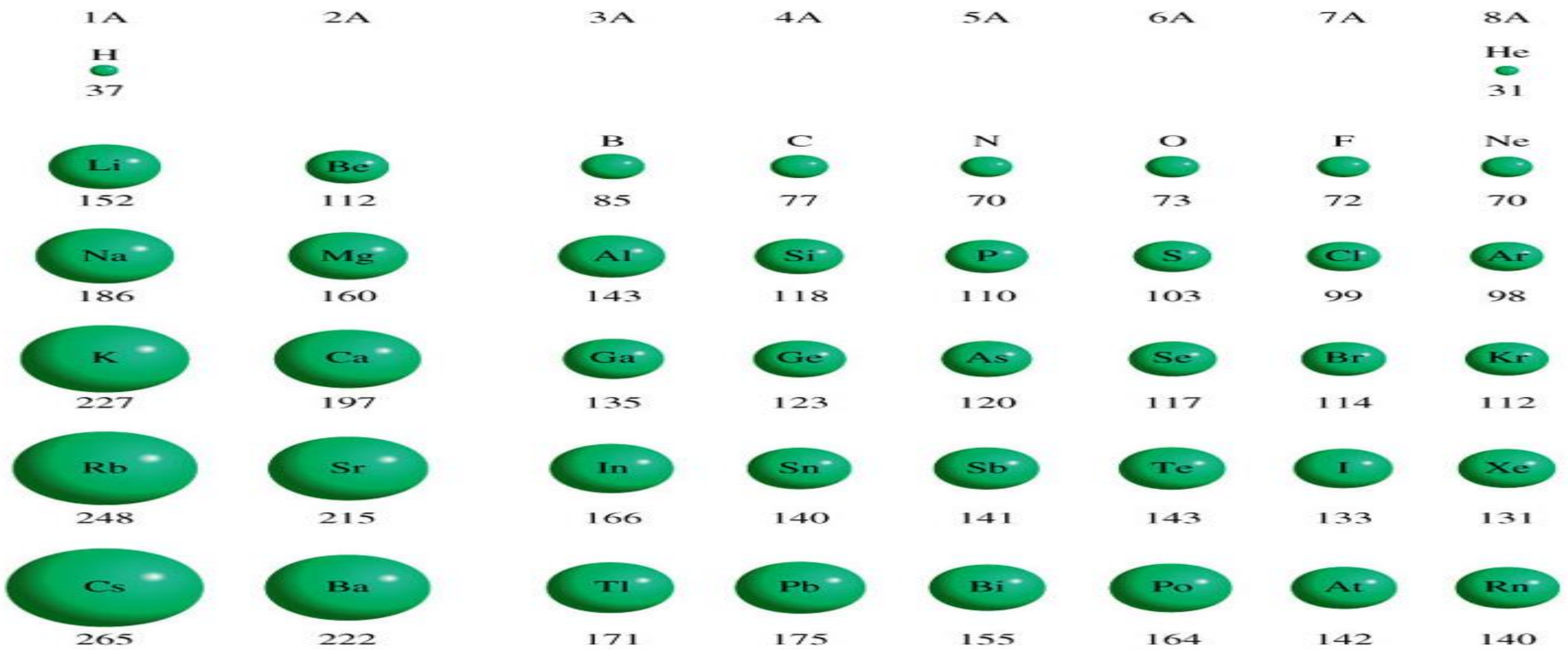
✓ Atomic radius increases down group

- ✓ Valence shell farther from nucleus
- ✓ Effective nuclear charge fairly close



Decreasing atomic radius

Increasing atomic radius

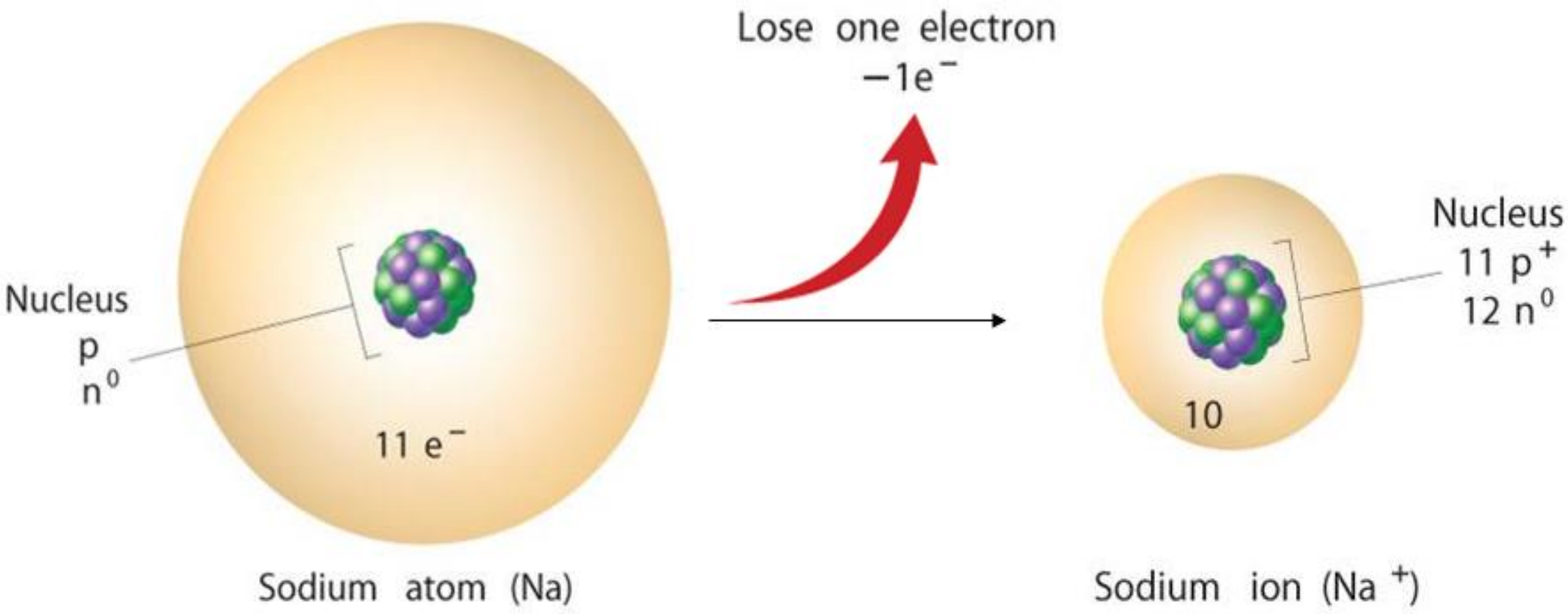


2.13 Periodic Trends: Ionic Radii (Sizes of Ions)

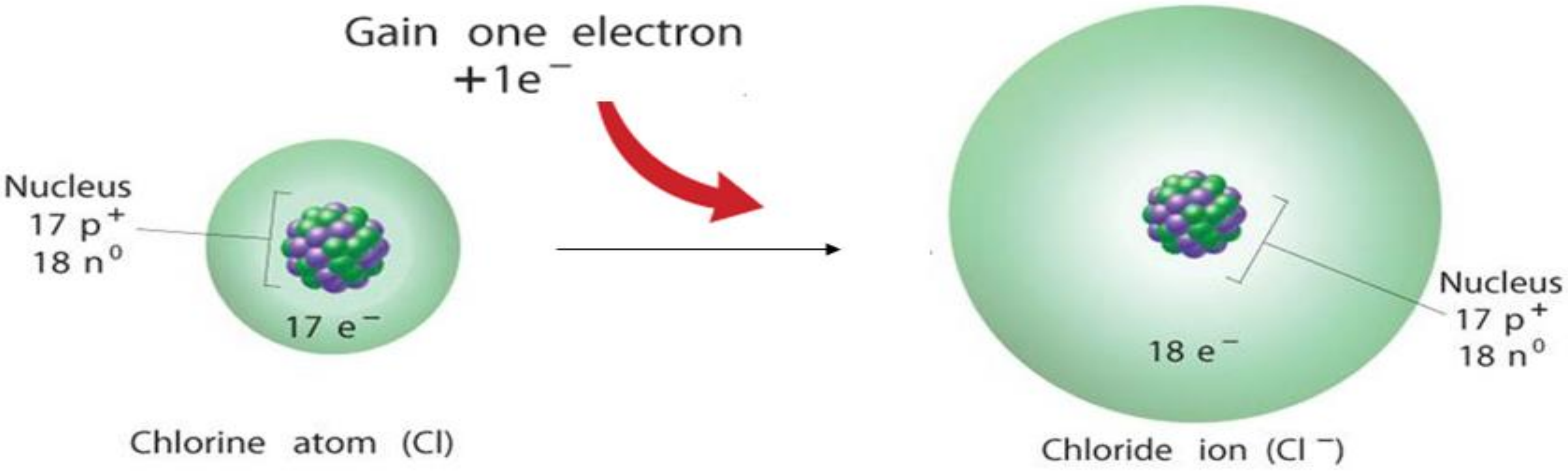
- **Ionic Radius:** is the interatomic distances in ionic compounds.
- ✓ Ions in the same group have the same charge.
- ✓ Ion size increases down the column.
  - ✓ Higher valence shell, larger
- ✓ Cations are smaller than their neutral atoms.
- ✓ Anions are larger than their neutral atoms.
- ✓ Cations are smaller than anions.
  - ✓ Except Rb^{+} and Cs^{+} bigger or same size as F^{-} and O^{2-}.
- Larger positive charge = smaller cation
  - ✓ For isoelectronic species
  - ✓ Isoelectronic = same electron configuration
- Larger negative charge = larger anion
  - ✓ For isoelectronic species

2.13 Periodic Trends: Ionic Radii (Sizes of Ions)

- Metals elements lose valence electrons to form cation ions.
- Cation radii are always smaller than atomic radii.

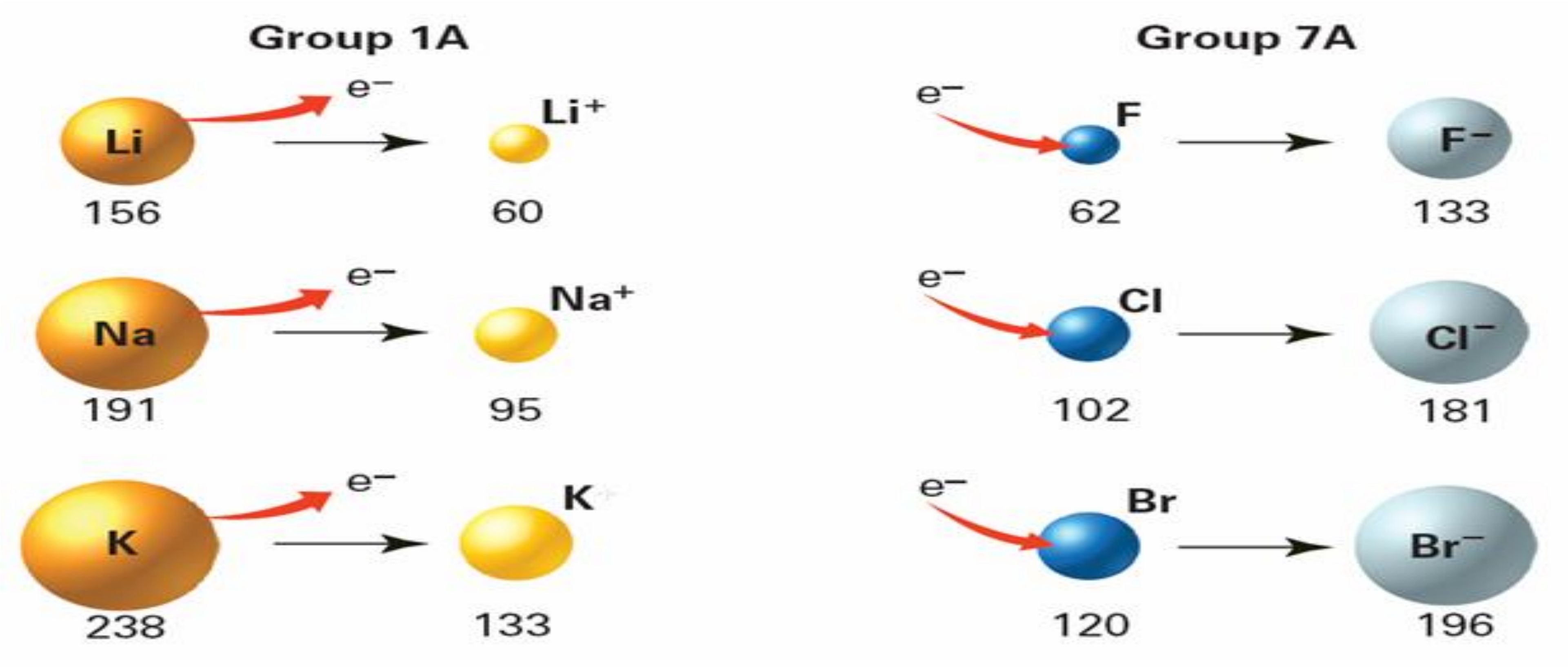


- Non-metal elements gain valence electrons to form anion ions.
- Anion radii are always larger than atomic radii.



Trends in Ionic Size

Relative radius of some atoms vs. their ions (in angstroms A°): ➢



Example:

Referring to a periodic table, arrange the following atoms in order of <u>increasing atomic radius</u>: P, Si, N.

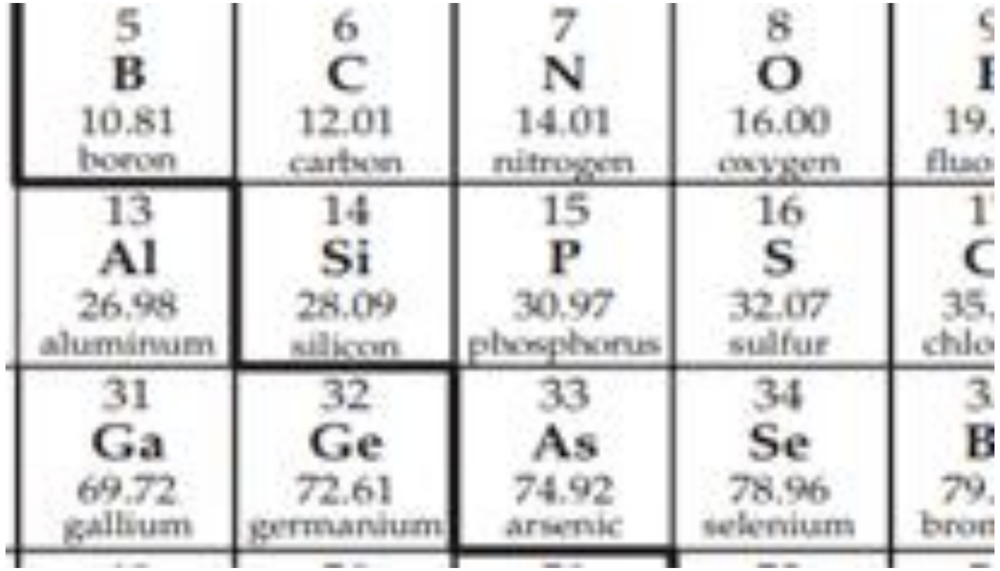
<u>Solution:</u>

1. N & P are in the same group → N is smaller than P

2. P & Si are in the same period → P is smaller than Si

The arrangement of <u>increasing</u> atomic radius is:

N < P < Si



5 B 10.81 boron	6 C 12.01 carbon	7 N 14.01 nitrogen	8 O 16.00 oxygen	9 F 19. fluo
13 Al 26.98 aluminum	14 Si 28.09 silicon	15 P 30.97 phosphorus	16 S 32.07 sulfur	1 C 35. chlo
31 Ga 69.72 gallium	32 Ge 72.61 germanium	33 As 74.92 arsenic	34 Se 78.96 selenium	3 B 79. brom

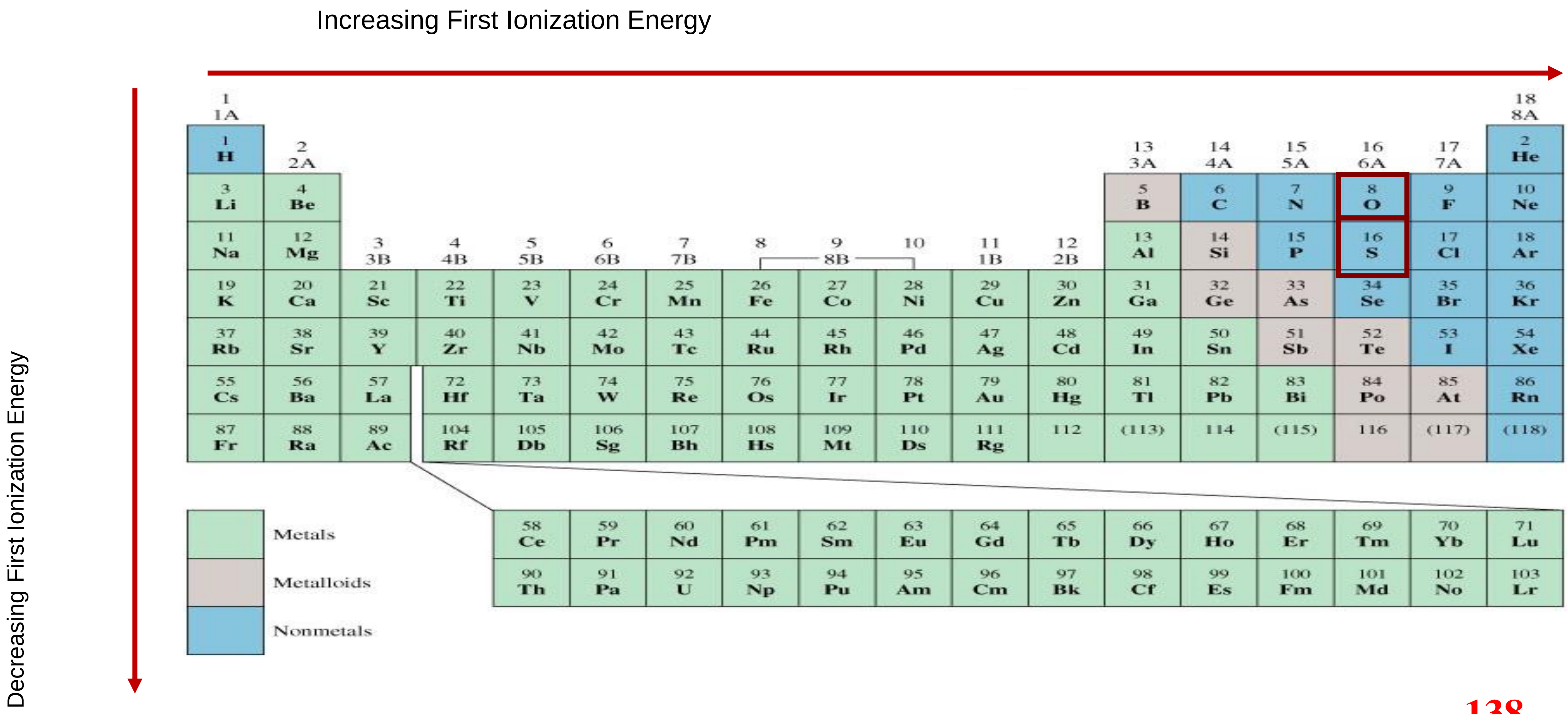
2.13 Periodic Trends: Ionization Energy

- **Ionization Energy (IE): the minimum energy needed to remove an electron from an atom or ion.**
  - ✓ **Measured in gaseous state**
  - ✓ **For endothermic process**
  - ✓ **Valence electron easiest to remove, lowest IE:**

$$M(g) + IE_1 \rightarrow M^{1+}(g) + 1\ e^-$$

$$M^{+1}(g) + IE_2 \rightarrow M^{2+}(g) + 1\ e^-$$

- ✓ **First ionization energy (IE_1) = energy to remove electron from a neutral atom.**
- ✓ **Second ionization energy (IE_2)= energy to remove from +1 ion, etc.**


Increasing First Ionization Energy
Decreasing First Ionization Energy
1 1A
2 2A
13 3A
14 4A
15 5A
16 6A
17 7A
18 8A
3 3B
4 4B
5 5B
6 6B
7 7B
8
9 8B
10
11 1B
12 2B
1 H
2 He
3 Li
4 Be
5 B
6 C
7 N
8 O
9 F
10 Ne
11 Na
12 Mg
13 Al
14 Si
15 P
16 S
17 Cl
18 Ar
19 K
20 Ca
21 Sc
22 Ti
23 V
24 Cr
25 Mn
26 Fe
27 Co
28 Ni
29 Cu
30 Zn
31 Ga
32 Ge
33 As
34 Se
35 Br
36 Kr
37 Rb
38 Sr
39 Y
40 Zr
41 Nb
42 Mo
43 Tc
44 Ru
45 Rh
46 Pd
47 Ag
48 Cd
49 In
50 Sn
51 Sb
52 Te
53 I
54 Xe
55 Cs
56 Ba
57 La
72 Hf
73 Ta
74 W
75 Re
76 Os
77 Ir
78 Pt
79 Au
80 Hg
81 Tl
82 Pb
83 Bi
84 Po
85 At
86 Rn
87 Fr
88 Ra
89 Ac
104 Rf
105 Db
106 Sg
107 Bh
108 Hs
109 Mt
110 Ds
111 Rg
112
(113)
114
(115)
116
(117)
(118)
58 Ce
59 Pr
60 Nd
61 Pm
62 Sm
63 Eu
64 Gd
65 Tb
66 Dy
67 Ho
68 Er
69 Tm
70 Yb
71 Lu
90 Th
91 Pa
92 U
93 Np
94 Pu
95 Am
96 Cm
97 Bk
98 Cf
99 Es
100 Fm
101 Md
102 No
103 Lr
Metals
Metalloids
Nonmetals

2.13 Periodic Trends: Ionization Energy

- **General trend in first ionization energy:**

  - ✓ **First IE generally decreases down the group.**
    - **Valence electron farther from nucleus**
  - ✓ **First IE generally increases across the period.**
    - **Effective nuclear charge increases**

- **Factors Affecting Ionization Energy:**

1- Nuclear charge: the larger the nuclear charge, the greater the ionization energy.

2- Shielding effect: the greater the shielding effect, the less the ionization energy.

3- Radius: the greater the distance between the nucleus and the outer electrons of an atom, the less the ionization energy.

4- Sublevel: an electron from a full or half-full sublevel requires additional energy to be removed.

<u>Example :</u> Which atom should have a smaller first ionization energy: oxygen (O) or sulphur (S)?

<u>Solution:</u>

O & S are in group 6A

O: [He] $2s^2\ 2p^4$

S: [Ne] $3s^2\ 3p^4$

4A 14	5A 15	6A 16	7A 17	2 He 4.00 heliu
6 C 12.01 carbon	7 N 14.01 nitrogen	8 O 16.00 oxygen	9 F 19.00 fluorine	10 Ne 20.18 neon
14 Si 28.09 silicon	15 P 30.97 phosphorus	16 S 32.07 sulfur	17 Cl 35.45 chlorine	18 Ar 39.95 argon
32 Ge 72.61 germanium	33 As 74.92 arsenic	34 Se 78.96 selenium	35 Br 79.90 bromine	36 Kr 83.80 krypto

The valence electrons in S are farther from the nucleus → removing of them is easier

Thus: I_1 (S) < I_1 (O)

- **Electron Affinity (EA):** is the energy change associated with the gaining of an electron by the atom in the gaseous state.

$$Cl(g) + 1\,e^- \longrightarrow Cl^-(g) \qquad EA = -349\ kJ/mol$$

- EA can be either endothermic or exothermic in nature.
  - Why either energy exchange?
    - It is due to electron-electron repulsion within orbitals and the volume of the atom.
- **General trends in electron affinity:**
  - EA increases across a period.
    - EA becomes more positive due to increase in Z_{eff}.
  - EA decreases down a group.
    - EA becomes less positive due to decrease in Z_{eff}.

Increasing Electron affinity

Decreasing Electron affinity

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112	(113)	114	(115)	116	(117)	118
			58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
			90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

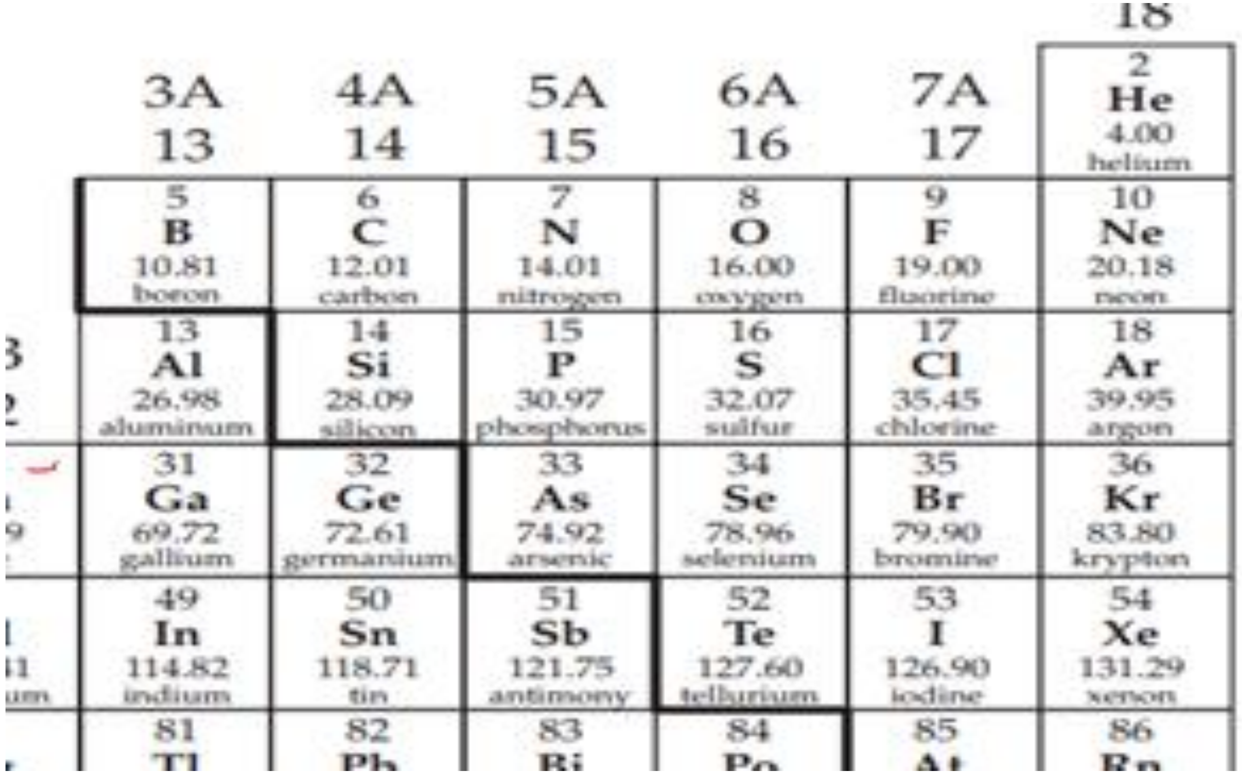
Which choice correctly lists the elements in order of <u>decreasing</u> electron affinity?

a. O, Cl, B, C

b. O, Cl, C, B

c. Cl, O, C, B

d. Cl, O, B, C



2.14 Periodic Trends: Metallic Character

- **Metallic Character:** is how closely an element's properties match the ideal properties of a metals.
  - – More malleable and ductile, better conductors, and easier to ionize
- **General trends in metallic character:**

- Metallic character decreases across a period.
  - ✓ Metals found at the left of the period and nonmetals to the right
- Metallic character increases down the column.
  - ✓ Nonmetals found at the top of the middle main group elements and metals found at the bottom

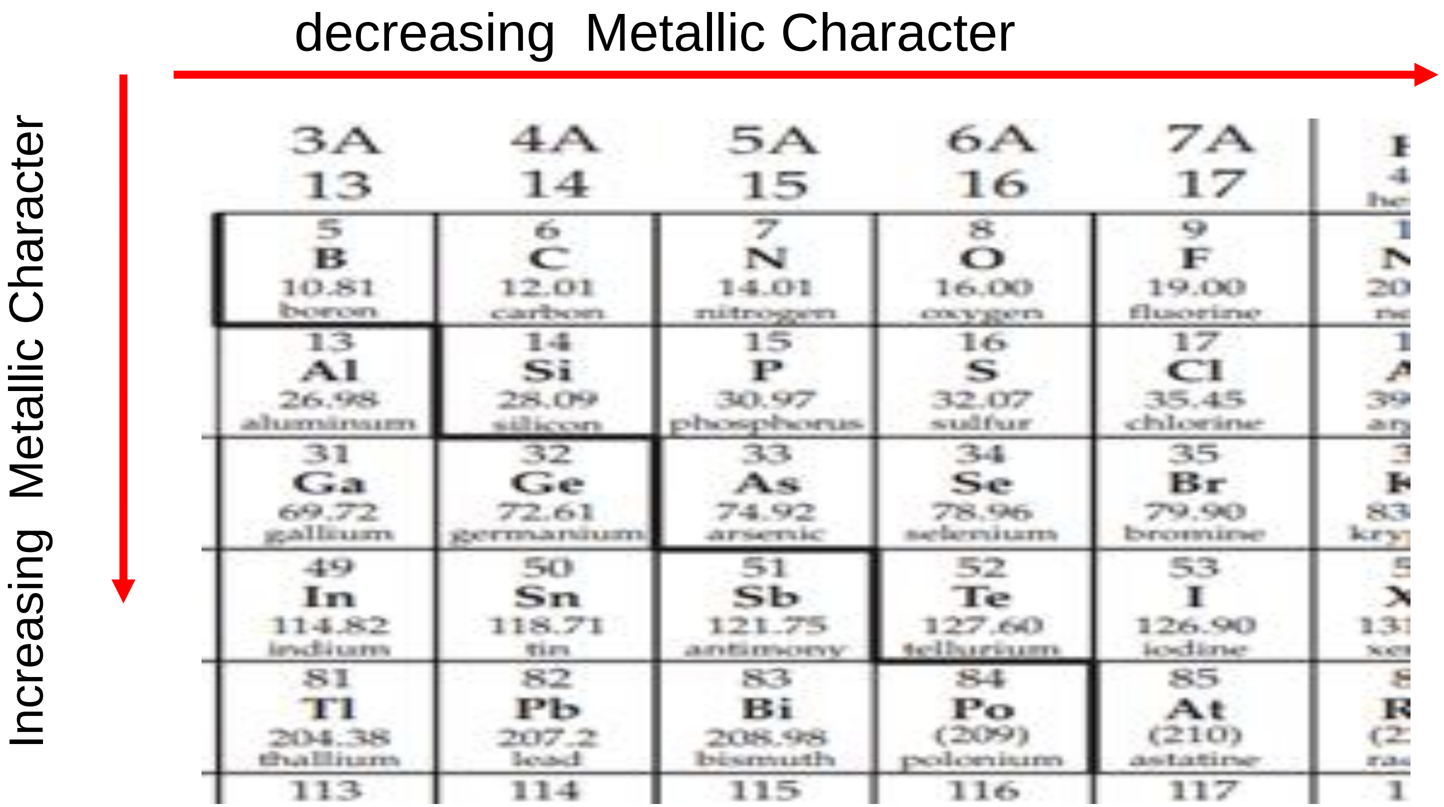
Example :choose the <u>more</u> metallic element from following:

(a)Sn or Te

(b) P or Sb

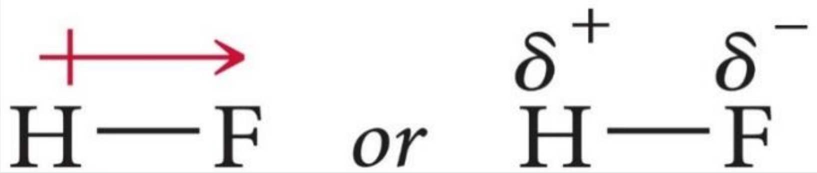
Sn>Te

Sb > P



3A 13	4A 14	5A 15	6A 16	7A 17
5 B 10.81 boron	6 C 12.01 carbon	7 N 14.01 nitrogen	8 O 16.00 oxygen	9 F 19.00 fluorine
13 Al 26.98 aluminum	14 Si 28.09 silicon	15 P 30.97 phosphorus	16 S 32.07 sulfur	17 Cl 35.45 chlorine
31 Ga 69.72 gallium	32 Ge 72.61 germanium	33 As 74.92 arsenic	34 Se 78.96 selenium	35 Br 79.90 bromine
49 In 114.82 indium	50 Sn 118.71 tin	51 Sb 121.75 antimony	52 Te 127.60 tellurium	53 I 126.90 iodine
81 Tl 204.38 thallium	82 Pb 207.2 lead	83 Bi 208.98 bismuth	84 Po (209) polonium	85 At (210) astatine
113	114	115	116	117

- **Electronegativity (EN):** is the ability of an atom in a molecule to attract electrons to itself.
  - ✓ This attraction or pulling of electrons causes a separation of charge within the bond.
  - ✓ Dipole moment is formed.
  - ✓ The greater the difference, the more POLAR the bond.



- **General trends in electronegativity:**
  - ✓ Electronegativity increases across a period.
  - ✓ Electronegativity decreases down a group.

Ex-)- Which of these atom is the *most* electronegative?

A-Li   B-Al   C-P   D- O

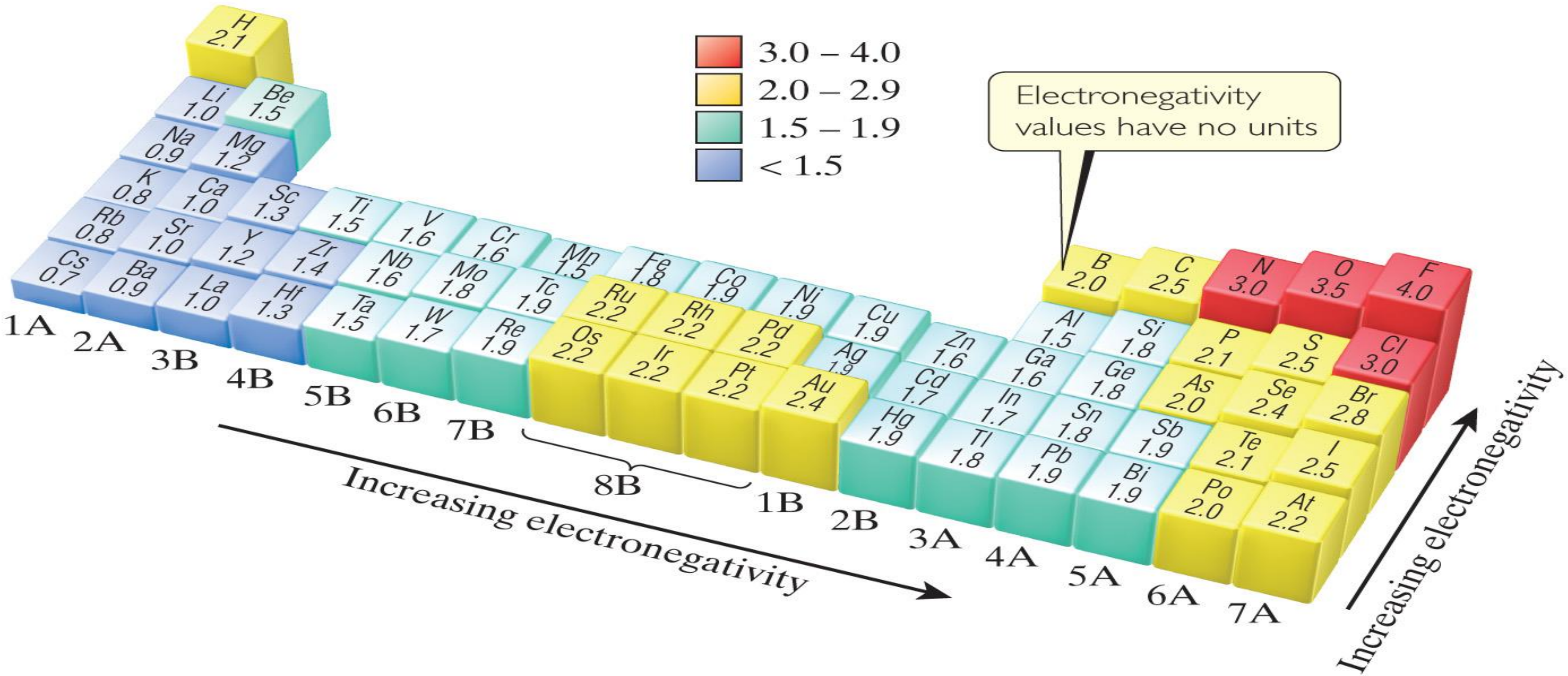
Ex-)- Which of these elements has the *greatest* electronegativity?

A-Na   B-Mg   C - F   D-O

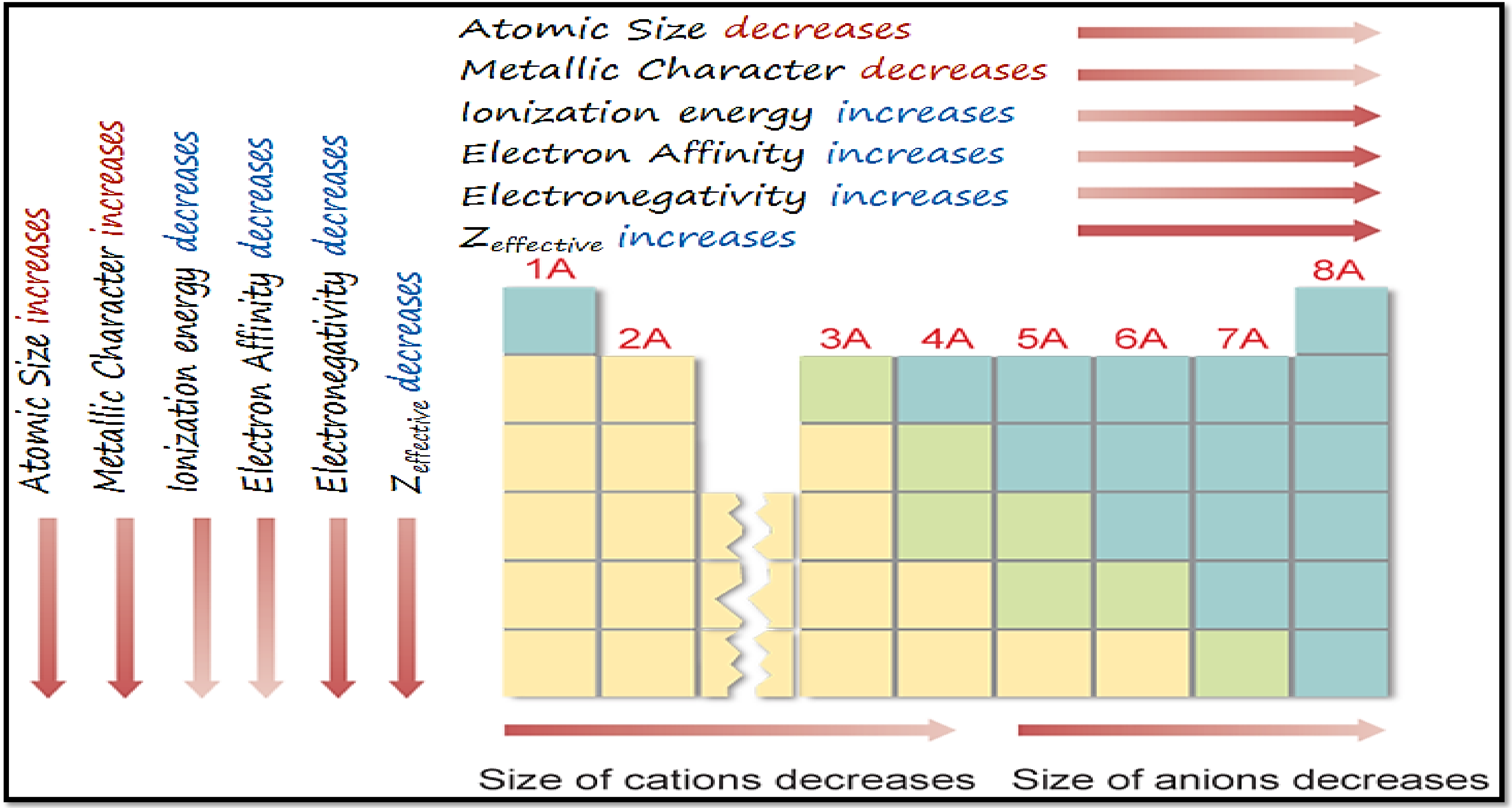
Ex-)- Which of these elements is the *least* electronegative?

A-Li   B-Al   C-P   D-N

2.14 Periodic Trends: Electronegativity



2.14 Periodic Trends: A Summary



Review مراجعة

(1) The element that has the valance electron configuration $3s^2\ 3p^3$ is:

a) Carbon

b) Nitrogen

c) Phosphorus

d) Neon

(2) Titanium (Ti) element is found in the periodic table in

a) s-block

b) P-block

c) d-block

d) f-block

(3) What the electronic configuration for **Co**

a) [Ar] $4s^2$ $3d^5$

b) [Ar] $4s^2$ $3d^7$

c) [Ar] $4s^1$ $3d^6$

d) [Ar] $4s^2$ $3d^4$

(4) Arrange the following in order of **<u>increasing</u> first ionization** energy:
F, K, P, Ca, and Ne.

a) K < Ca < P < F < Ne

b) Ne < F < Ca < K < P

c) P < F < Ne < K < Ca

d) K < F < P < Ne < Ca

(5) Which of these elements is most likely to be a **good** conductor of electricity?

a) N

b) S

c) He

d) Fe

(6) magnesium ion, $_{12}Mg^{2+}$, has

a) 12 protons and 13 electrons.

b) 24 protons and 26 electrons.

c) 12 protons and 10 electrons.

d) 24 protons and 22 electrons.

Answer the following questions:

1. Arrange these elements: Mg, Na, Cl, S, Ar, Si, and P, in order of:

a. decreasing atomic radius.

b. increasing ionization energy.

c. decreasing electronegativity.

d. increasing metallic character

2. Choose the <u>more metallic</u> element from each pair:

a. Sr or Sb

b. Be or Ba

c. Ti or Cu

d. S or Si

3. Choose the largest atom from each pair:

a. Al or Cl   b. Si or C   c. S or Se   d. Ne or Xe

4. Arrange the elements in order of increasing atomic radius: Ca, Rb, S, Si, Ge, F.

5. Arrange these elements in order of increasing electronegativity: C, N, O, Be, B.

6. Define each term and indicate what happens for each of them when moving right to left within a period of the periodic table?

a. Electronegativity

b. Ionization energy

c. Atomic radius

d. Metallic character

e. Electron affinity

a. Electronegativity : is the ability of an atom in a molecule to attract electrons to itself.

b. Ionization energy: is the minimum energy (kJ/mol) required to remove an electron from a gases atom in its ground state.

c. Atomic radius

d. Metallic character: is how closely an element's properties match the ideal properties of a metals.

e. Electron affinity: is the negative of the energy change that occurs when an electron is accepted by an atom in the gases state