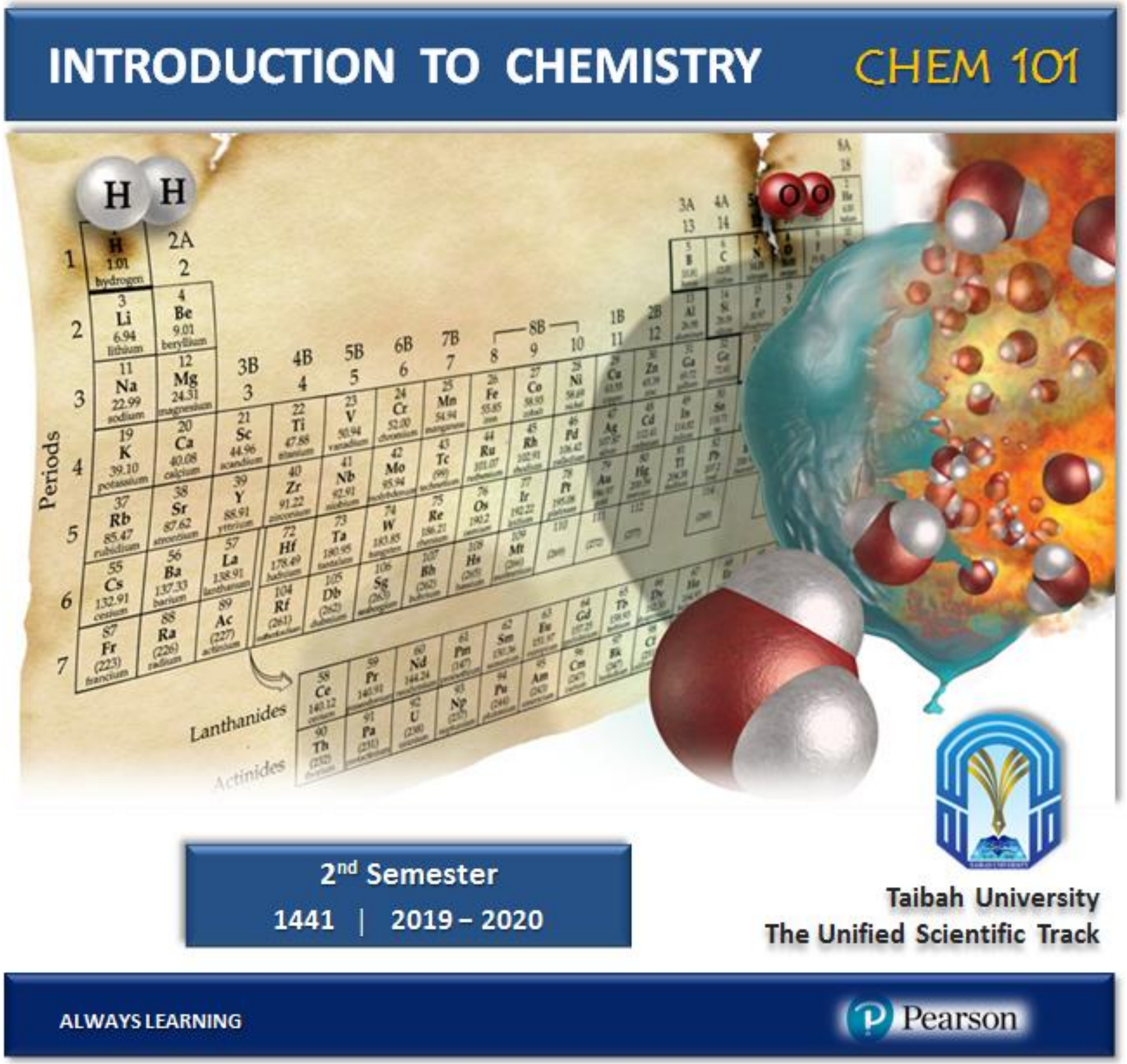


Lecture Presentation

Chapter 2

Atoms, Molecules, Ions, and Periodicity

Topic 04

- Atomic Theory
- Atomic Structure

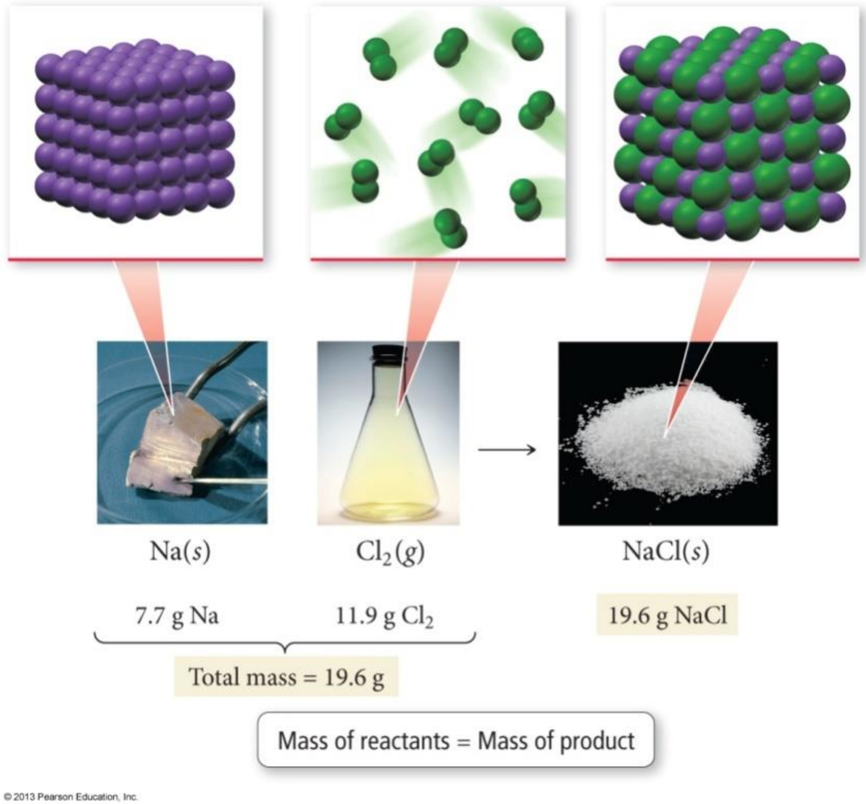
2.2 Modern Atomic Theory and Laws that Led to It

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

Law of 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 produced products
  - Total number of reactants' atoms = Total number of products' atoms



Law of Definite Proportions

Law of Definite Proportions (J. Proust):

➢ A given chemical compound always contains the same elements in the exact same proportions (by mass), regardless to its source or how it was prepared

- **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**

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

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

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

Law of Multiple Proportions

Law of Multiple Proportions (J. Dalton):

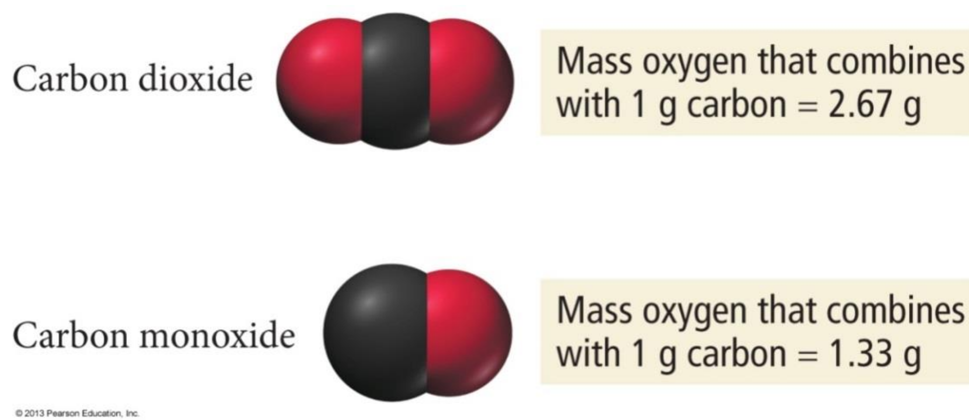
- When two elements, "A" and "B", combine with each other to form two or more compounds, the ratios of the masses of those elements in the formed compounds are simple whole numbers.

✓ **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:**

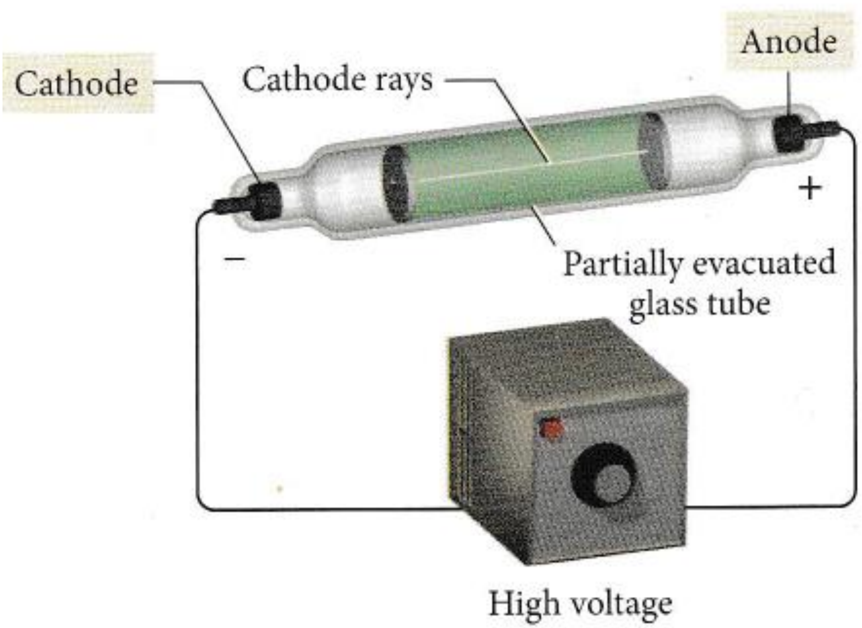
"Fe" to "O" in FeO = (1:1), while in Fe_2O_3 = (2:3)

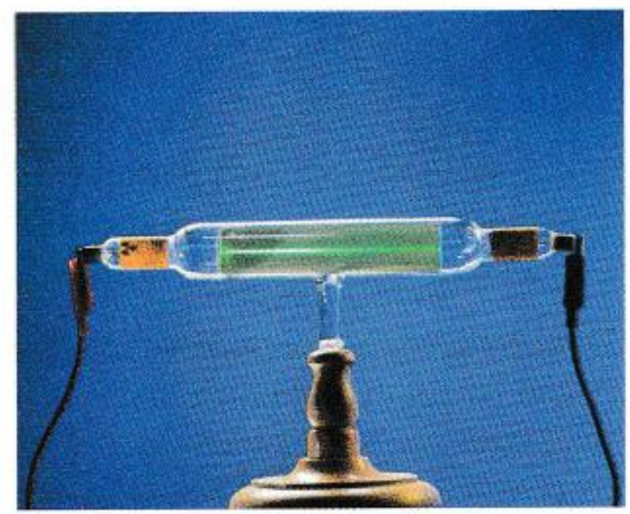


Postulates of Atomic Theory of Matter (J. Dalton):

- 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.

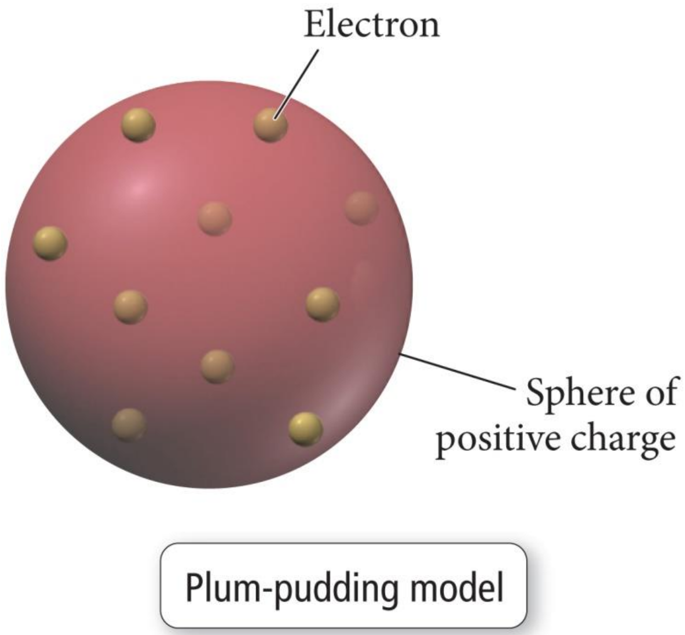
- **Cathode Ray Tube Experiment (J. J. Thomson):**
  - Discovered the **electron** and determined the **electron's charge-to-mass ratio**.





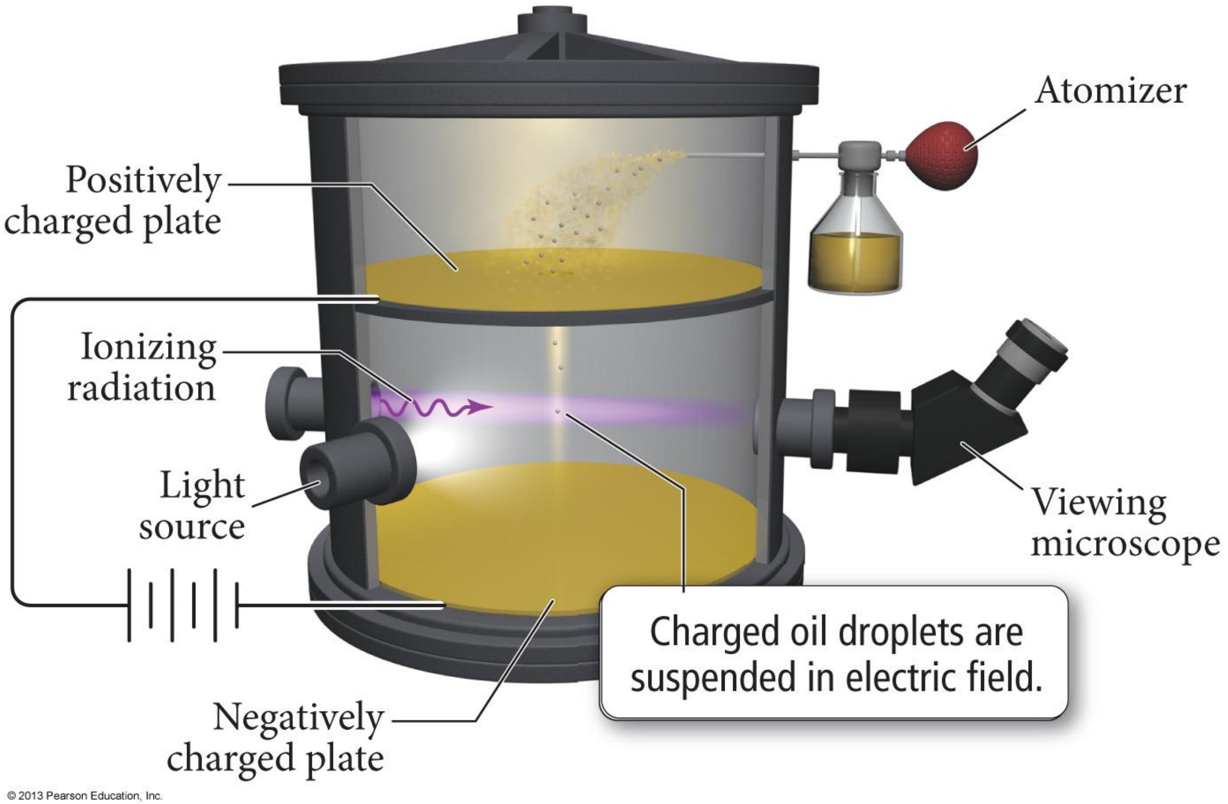
Plum-Pudding Model

- **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

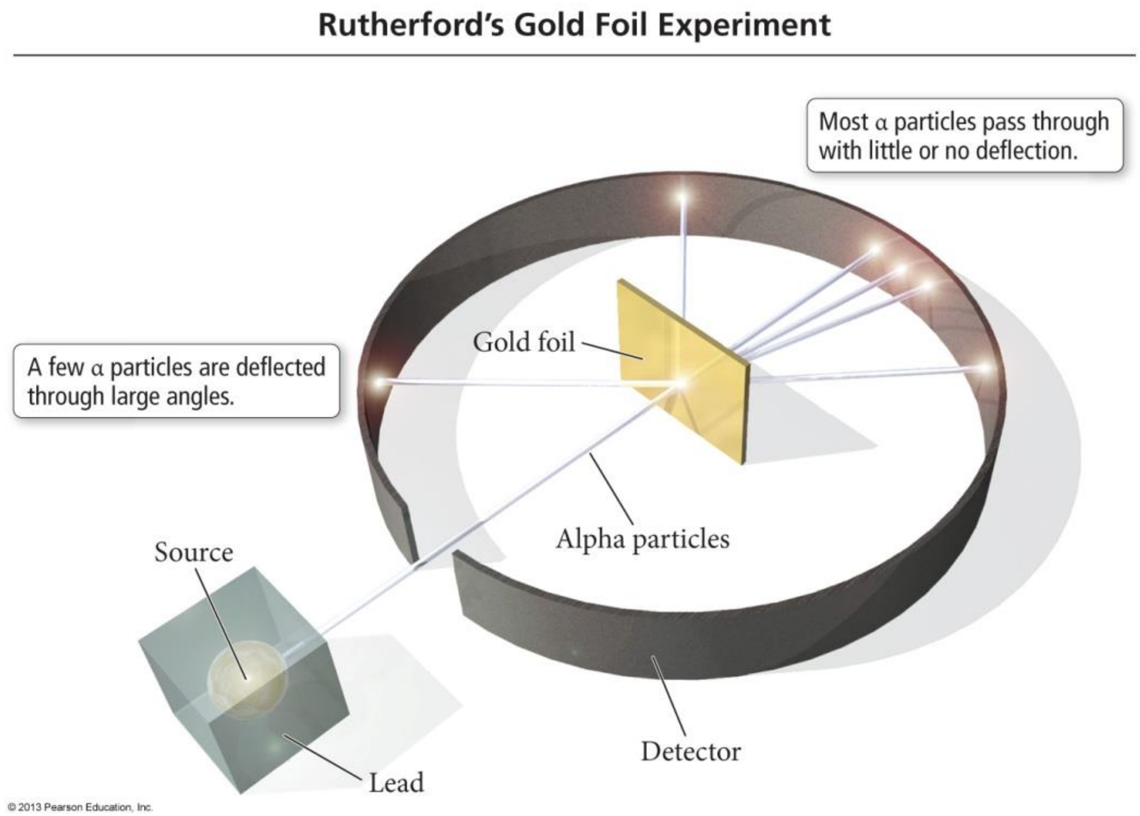


Millikan's Oil Drop Experiment

- **Oil Drop Experiment (R. Millikan):**
  - Led to determining the **electrical charge of the electron.**



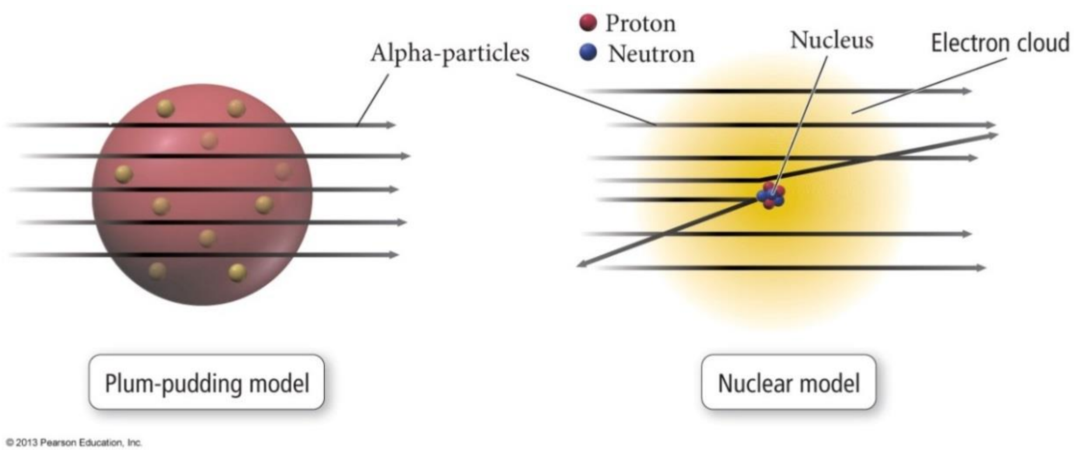
- **Gold Foil Experiment (E. Rutherford):**
- Discovered the atom's **nucleus (protons)** & disapproved the plum-pudding model.



Rutherford's Model (The Nuclear Theory)

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

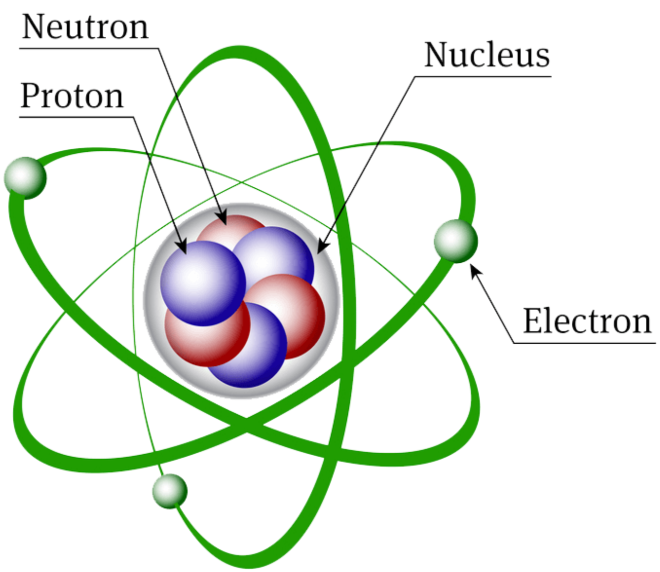
- The atom contains a tiny, dense center called the nucleus.
- The nucleus has essentially the entire mass of the atom.
  - The electrons weigh so little they give practically no mass to the atom.
- The nucleus is positively charged.
  - The amount of positive charge (named: protons) balances the negative 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).



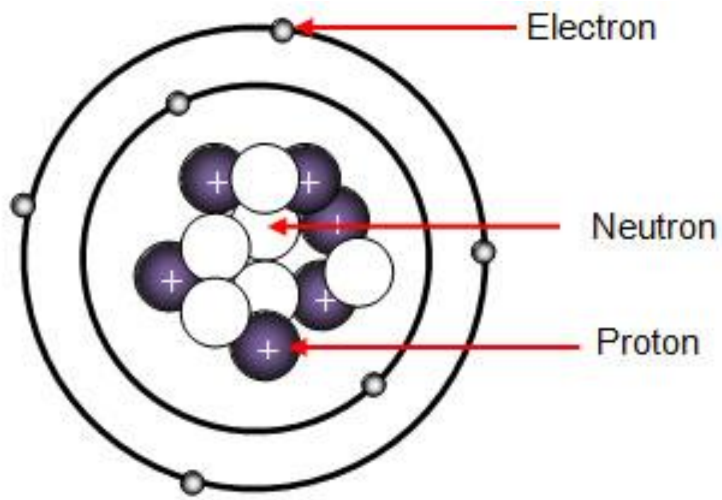
The Discovery of The Neutrons

- **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!



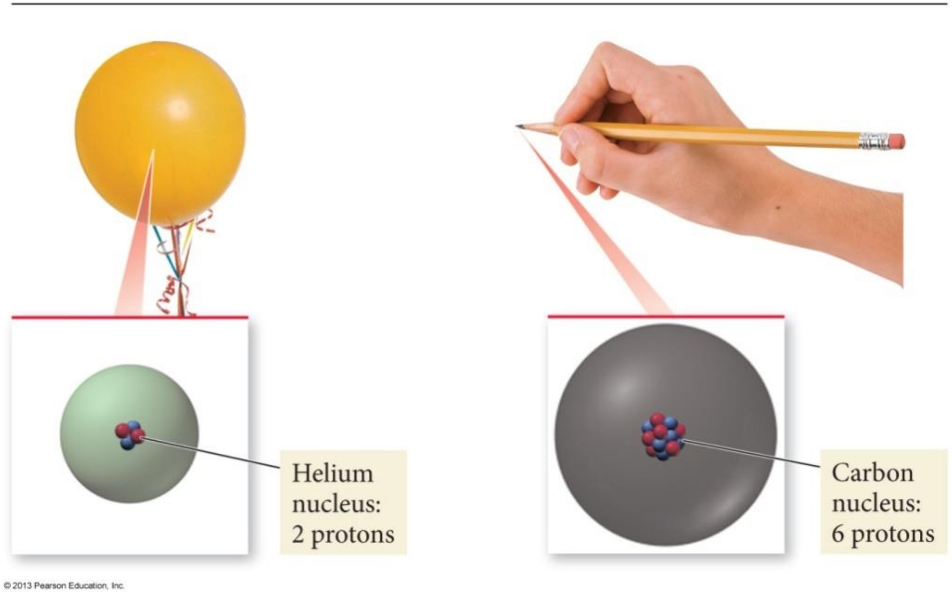
Properties of Subatomic Particles						
Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)	Symbol
Electron	Outside nucleus	-1.602×10^{-19}	1-	0.00055	0.00091×10^{-24}	e , e^-
Proton	Nucleus	1.602×10^{-19}	1+	1.00727	1.67262×10^{-24}	P , P^+, H^+
Neutron	Nucleus	0	0	1.00866	1.67493×10^{-24}	n ,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

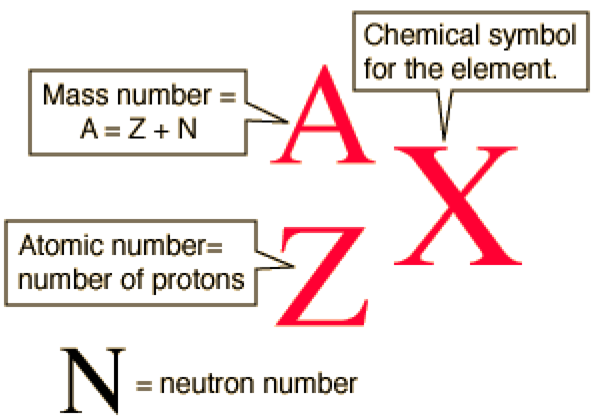


Elements: Defined by their Number of Protons

- **The number of protons** located in an atom's nucleus determines the element's identity.
  - The number of protons in the nucleus of an atom is called the "atomic number" and is referred to as **"Z"**, it's considered as the "fingerprint" of any element.

The Number of Protons Defines the Element





Elements: Defined by their Number of Protons

- Each element has a unique name, symbol, and atomic number.
  - Symbol has either one or two letters: O (oxygen) or Fe (iron)
  - The elements are arranged on the periodic table in order of increasing their atomic numbers.

▲ **The periodic table of the elements**

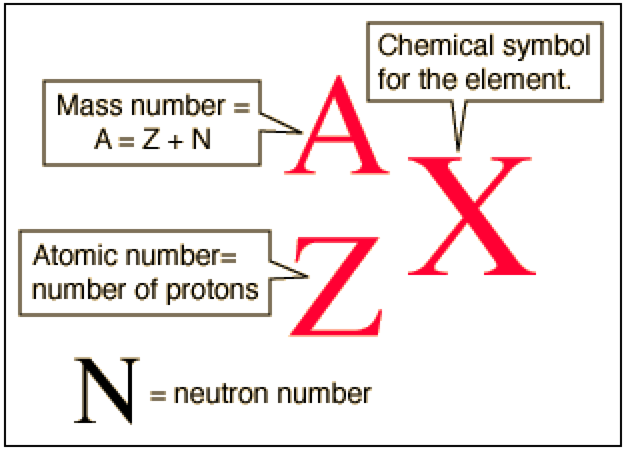
Key: 4 = Atomic number (Z); Be = Chemical symbol; 9.01 = Atomic mass; beryllium = Name

Periods	1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.01 hydrogen																	2 He 4.00 helium
2	3 Li 6.94 lithium	4 Be 9.01 beryllium											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	10 Ne 20.18 neon
3	11 Na 22.99 sodium	12 Mg 24.31 magnesium											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	18 Ar 39.95 argon
4	19 K 39.10 potassium	20 Ca 40.08 calcium	21 Sc 44.96 scandium	22 Ti 47.88 titanium	23 V 50.94 vanadium	24 Cr 52.00 chromium	25 Mn 54.94 manganese	26 Fe 55.85 iron	27 Co 58.93 cobalt	28 Ni 58.69 nickel	29 Cu 63.55 copper	30 Zn 65.39 zinc	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	36 Kr 83.80 krypton
5	37 Rb 85.47 rubidium	38 Sr 87.62 strontium	39 Y 88.91 yttrium	40 Zr 91.22 zirconium	41 Nb 92.91 niobium	42 Mo 95.94 molybdenum	43 Tc (99) technetium	44 Ru 101.07 ruthenium	45 Rh 102.91 rhodium	46 Pd 106.42 palladium	47 Ag 107.87 silver	48 Cd 112.41 cadmium	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	54 Xe 131.29 xenon
6	55 Cs 132.91 cesium	56 Ba 137.33 barium	57 La 138.91 lanthanum	72 Hf 178.49 hafnium	73 Ta 180.95 tantalum	74 W 183.85 tungsten	75 Re 186.21 rhenium	76 Os 190.2 osmium	77 Ir 192.22 iridium	78 Pt 195.08 platinum	79 Au 196.97 gold	80 Hg 200.59 mercury	81 Tl 204.38 thallium	82 Pb 207.2 lead	83 Bi 208.98 bismuth	84 Po (209) polonium	85 At (210) astatine	86 Rn (222) radon
7	87 Fr (223) francium	88 Ra (226) radium	89 Ac (227) actinium	104 Rf (261) rutherfordium	105 Db (262) dubnium	106 Sg (263) seaborgium	107 Bh (262) bohrium	108 Hs (265) hassium	109 Mt (266) meitnerium	110 Ds (281) darmstadtium	111 Rg (280) roentgenium	112 Cn (285)	113 — (284)	114 — (289)	115 — (288)	116 — (292)	117 ** (292)	118 — (294)

Lanthanides	58 Ce 140.12 cerium	59 Pr 140.91 praseodymium	60 Nd 144.24 neodymium	61 Pm (147) promethium	62 Sm 150.36 samarium	63 Eu 151.97 europium	64 Gd 157.25 gadolinium	65 Tb 158.93 terbium	66 Dy 162.50 dysprosium	67 Ho 164.93 holmium	68 Er 167.26 erbium	69 Tm 168.93 thulium	70 Yb 173.04 ytterbium	71 Lu 174.97 lutetium
Actinides	90 Th (232) thorium	91 Pa (231) protactinium	92 U (238) uranium	93 Np (237) neptunium	94 Pu (244) plutonium	95 Am (243) americium	96 Cm (247) curium	97 Bk (247) berkelium	98 Cf (251) californium	99 Es (252) einsteinium	100 Fm (257) fermium	101 Md (258) mendelevium	102 No (259) nobelium	103 Lr (260) lawrencium

Isotopes: When The Number of Neutrons Varies

- **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, but may be physically different.

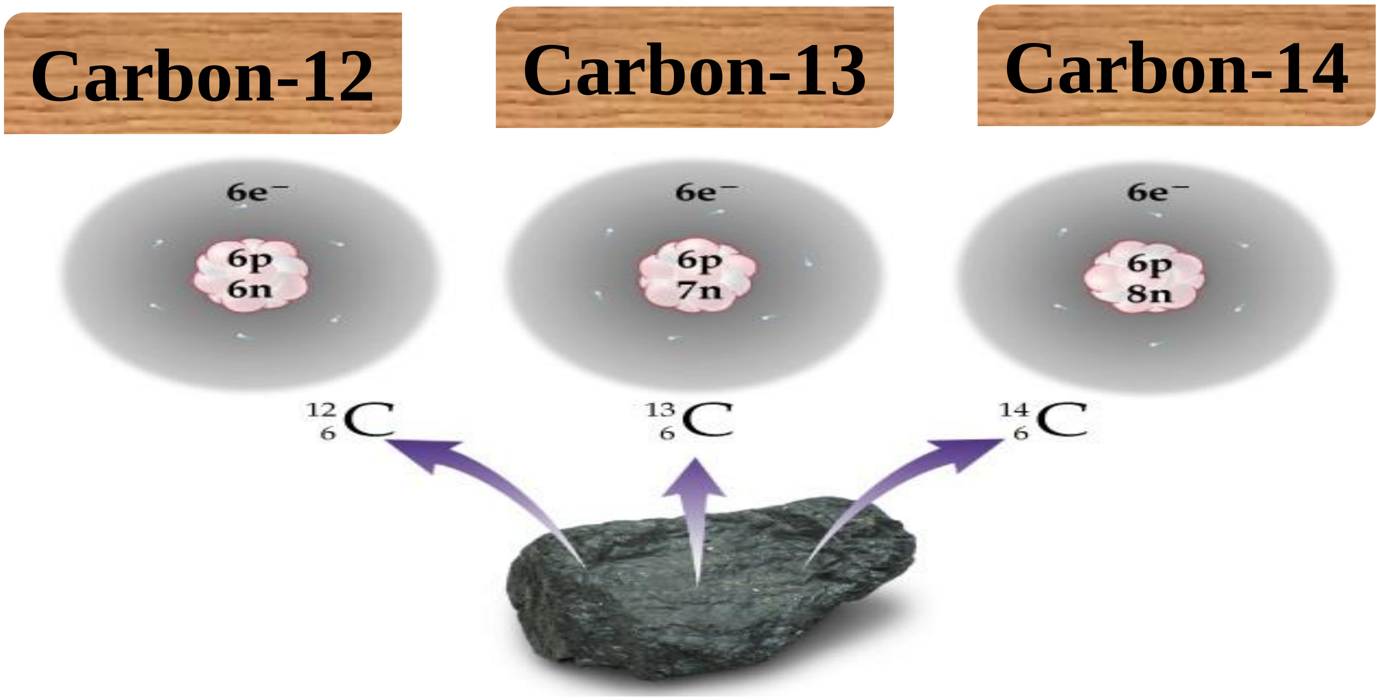


Mass Number (A) = Protons + Neutrons

Note: Isotopes are identified by their "mass numbers"

(example: C–12 , C–13 , C–14)

Isotopes: 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^-

Exercise

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

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

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

Exercise

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

	protons	electrons	neutrons
$^{32}_{16}S$	16	16	16
$^{65}_{29}Cu$	29	29	36
U–240	92	92	148

Ions: Charged Atoms (Cations vs. Anions)

Cations

- **A <u>CATION</u> forms when an atom <u>loses</u> one or more electrons from its outer shell (valence shell, the highest energy level).**
- **Cations are positively charged because the atom has <u>more</u> protons (+) than electrons (–).**
  - Mg atom has 12 protons & 12 electrons.
  - Mg^{2+} ion has 12 protons & 10 electrons.
- **Metal elements tend to form cations.**
- **<u>Example</u>:** $Mg \rightarrow Mg^{2+} + 2\,e^{-}$

Anions

- **An <u>ANION</u> forms when an atom <u>gains</u> one or more electrons into its outer shell (valence shell, the highest energy level).**
- **Anions are negatively charged because the atom has <u>fewer</u> protons (+) than electrons (–).**
  - F atom has 9 protons & 9 electrons.
  - F^{-} ion has 9 protons & 10 electrons.
- **Nonmetal elements tend to form anions.**
- **<u>Example</u>:** $F + 1\,e^{-} \rightarrow F^{-}$

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	________	________

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+}

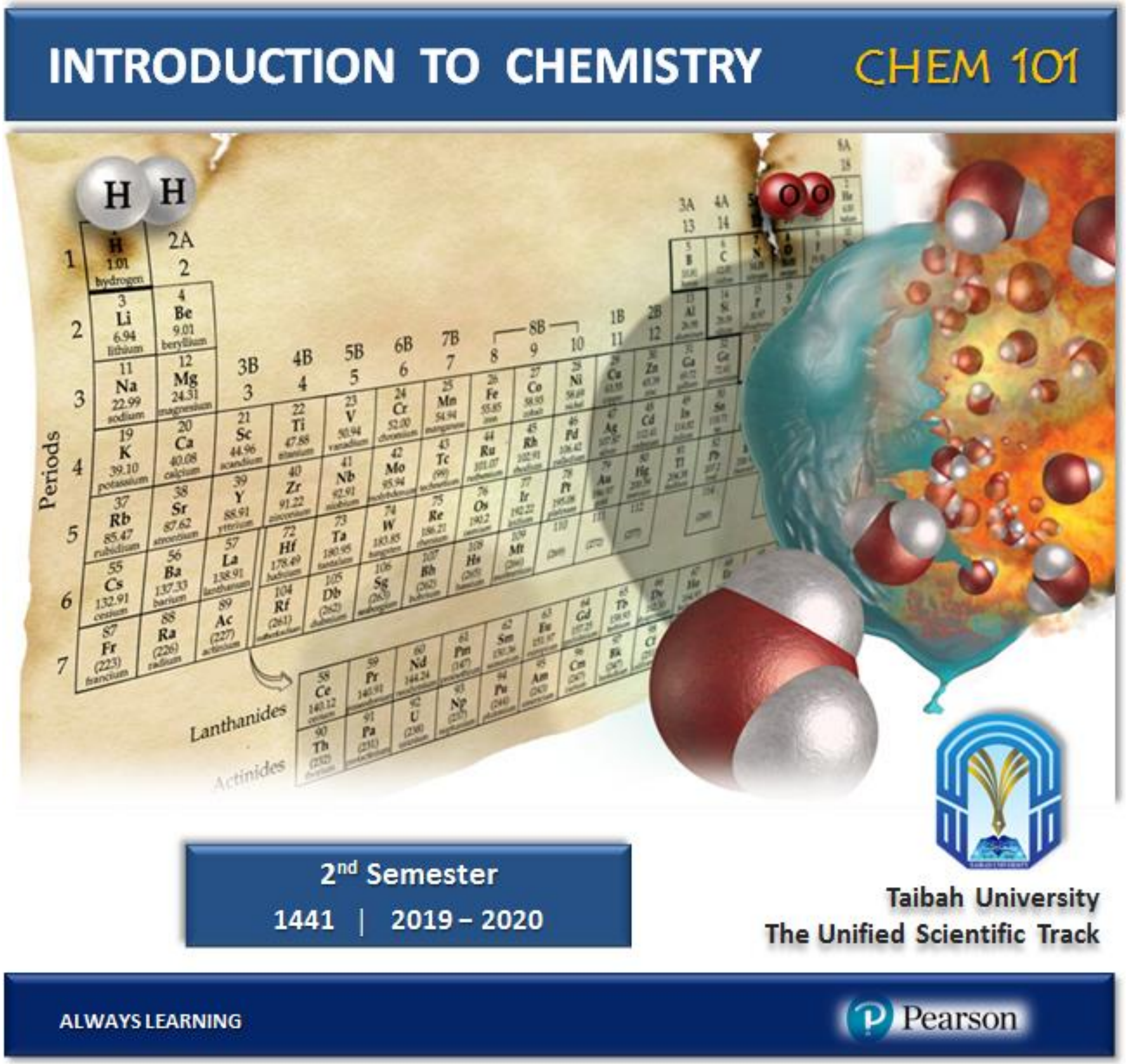
4- Write isotopic symbols of the form $^{A}_{Z}X$ for each isotope:

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



Lecture Presentation

Chapter 2

Atoms, Molecules, Ions, and Periodicity

Topic 05

The Periodic Table of The Elements

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

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**

Major Divisions of the Periodic Table

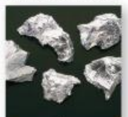
Metals | Metalloids | Nonmetals

Silicon

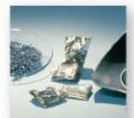
Arsenic

Carbon

Strontium

Chromium

Gold

Copper

Lead

Sulfur

Bromine

Iodine


	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

Lanthanides	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
Actinides	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

Classification of Elements: Metals

- **Metals** lie on the lower left side and middle of the periodic table.

- **Properties of Metals:**
  - ✓ 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 & not malleable).
  - ✓ They tend to gain electrons when they undergo chemical changes (forming anions).

Classification of Elements: Metalloids

- **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 (groups 1A–2A) | Transition elements (groups 3B–2B) | Main-group elements (groups 3A–8A)

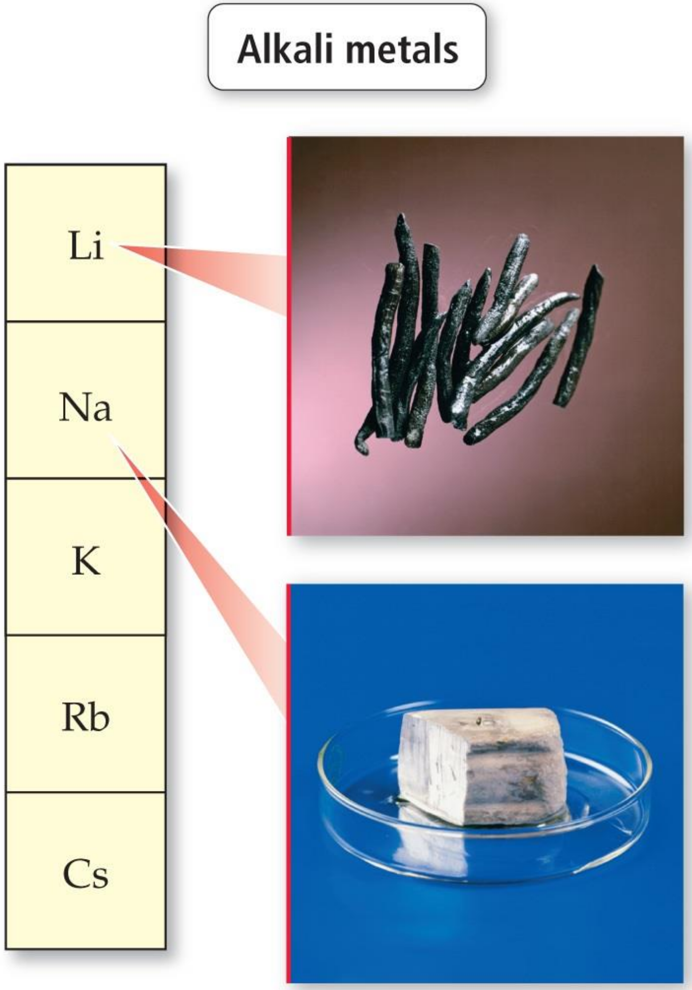
Group number

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**.

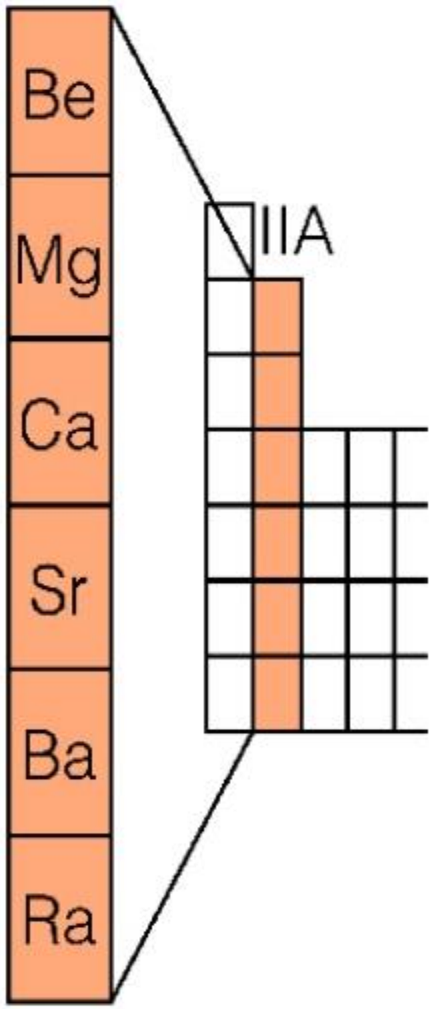
Major Families: Alkali Metals (Group 1A)

- **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.



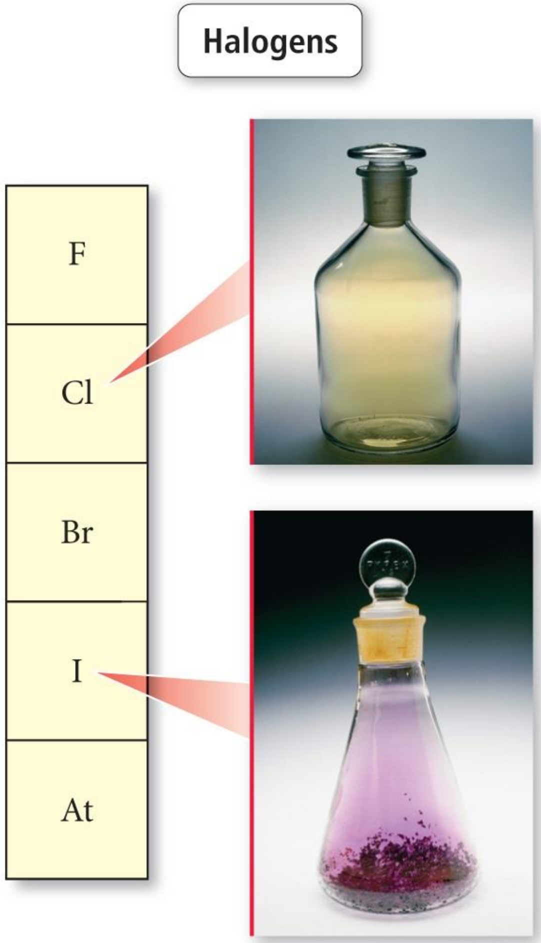
Major Families: Alkaline Earth Metals (Group 2A)

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



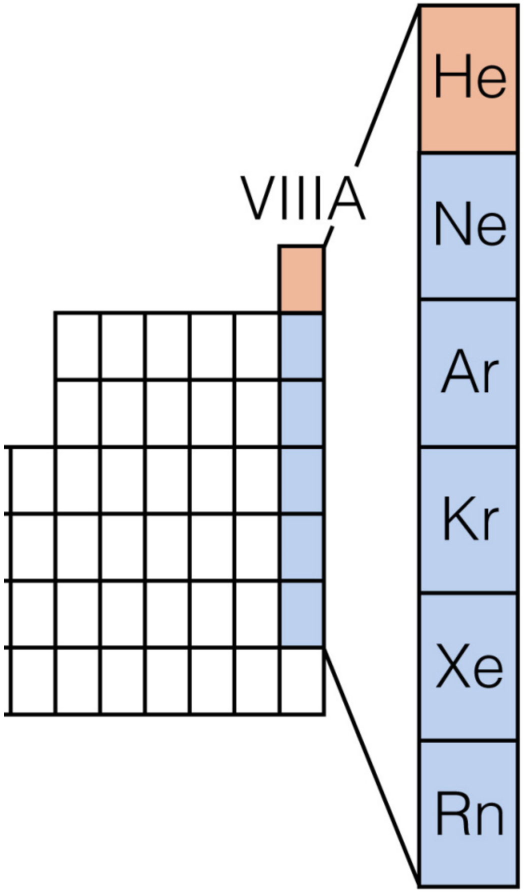
Major Families: Halogens (Group 7A)

- **The group 7A** elements, the **halogens**, are very reactive nonmetals.
- They are always found in nature as salts.
  - 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.

Ions and the Periodic Table

- 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 anion 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**:

- Alkali metals (group 1A): tend to lose one electron to form +1 ions.

- Alkaline earth metals (group 2A): tend to lose two electrons to form +2 ions.

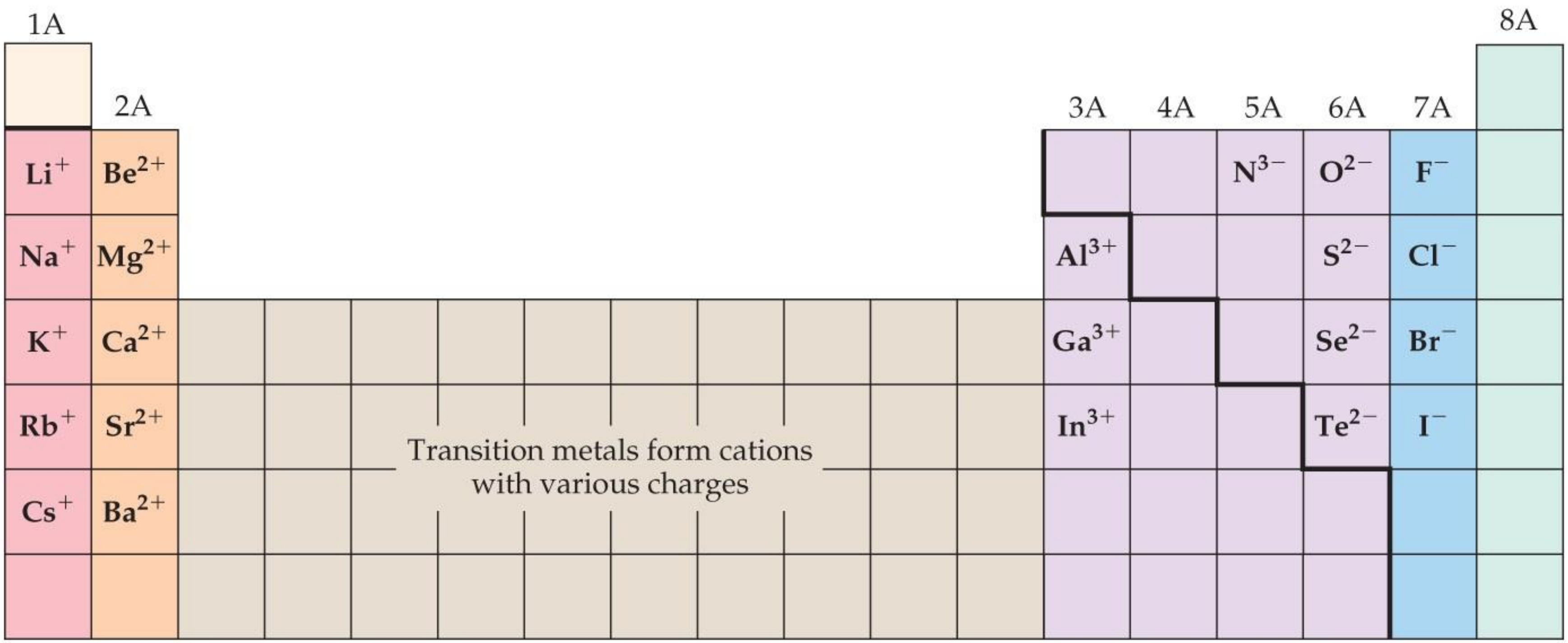
- Group 3A elements: tend to lose three electrons to form +3 ions.

- Oxygen family nonmetals (group 6A): tend to gain two electrons to form −2 ions.

- Halogens (group 7A): tend to gain one electron to form −1 ions.

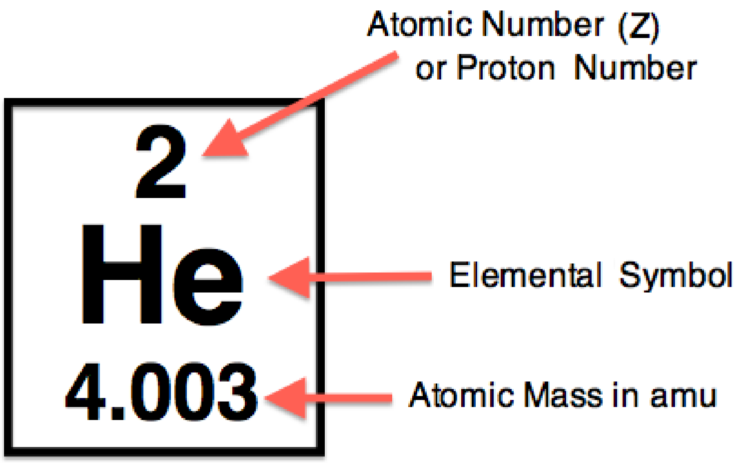
Ions and the Periodic Table

Elements that form ions with <u>predictable charges</u>:



2.7 Atomic Mass: The Average Mass of an Element's Atoms

- **Atomic mass** is sometimes called "atomic weight".



- It represents the average mass of all the isotopes that compose that element, weighted according to the natural abundance (fraction) of each isotope.

2.7 Atomic Mass: How to calculate?

- In general, the atomic mass can be calculated using the equation:

$$\text{Atomic mass} = \sum_n (\text{fraction of isotope } n) \times (\text{mass of isotope } n)$$
$$= (\text{fraction of isotope 1} \times \text{mass of isotope 1})$$
$$+ (\text{fraction of isotope 2} \times \text{mass of isotope 2})$$
$$+ (\text{fraction of isotope 3} \times \text{mass of isotope 3}) + \ldots$$

Note: the fraction of each isotope = its natural abundance (%) / 100

2.7 Atomic Mass: Example

example **Atomic Mass**

Copper has two naturally occurring isotopes: **Cu-63** with mass 62.9396 amu and a natural abundance of 69.17%, and **Cu-65** with mass 64.9278 amu and a natural abundance of 30.83%. Calculate the atomic mass of copper.

Solution

Convert the percent natural abundances into decimal form by dividing by 100.	Fraction Cu-63 $= \frac{69.17}{100} = 0.6917$ Fraction Cu-65 $= \frac{30.83}{100} = 0.3083$
calculate the atomic mass using the equation given in the text.	Atomic mass $= 0.6917(62.9396\ \text{amu}) + 0.3083(64.9278\ \text{amu})$ $= 43.5\underline{3}53\ \text{amu} + 20.0\underline{1}72\ \text{amu} = 63.5525 = 63.55\ \text{amu}$

- **Exercise**: Naturally occurring chlorine consists of 75.77% chlorine-35 atoms (mass 34.97 amu) and 24.23% chlorine-37 atoms (mass 36.97 amu). Calculate the atomic mass of chlorine.

- **Answer**: The Atomic Mass of Cl = 35.45 amu

Assessment

Answer the following questions:

1- Element X has three isotopes (see the table), the atomic mass of this element is _____ amu.

Isotope	Abundance	Mass
^{53}X	25.00 %	52.62
^{56}X	37.00 %	56.29
^{58}X	38.00 %	58.31

2- Which pairs of elements do you expect to be similar? Why?

a. N and Ne b. Mo and Sr c. Ar and Kr d. Cl and I e. P and Pd

3- Determine whether or not each element is a main-group element:

a. tellurium b. potassium c. vanadium d. manganese

4- Predict the charge of the monoatomic ion formed by each element:

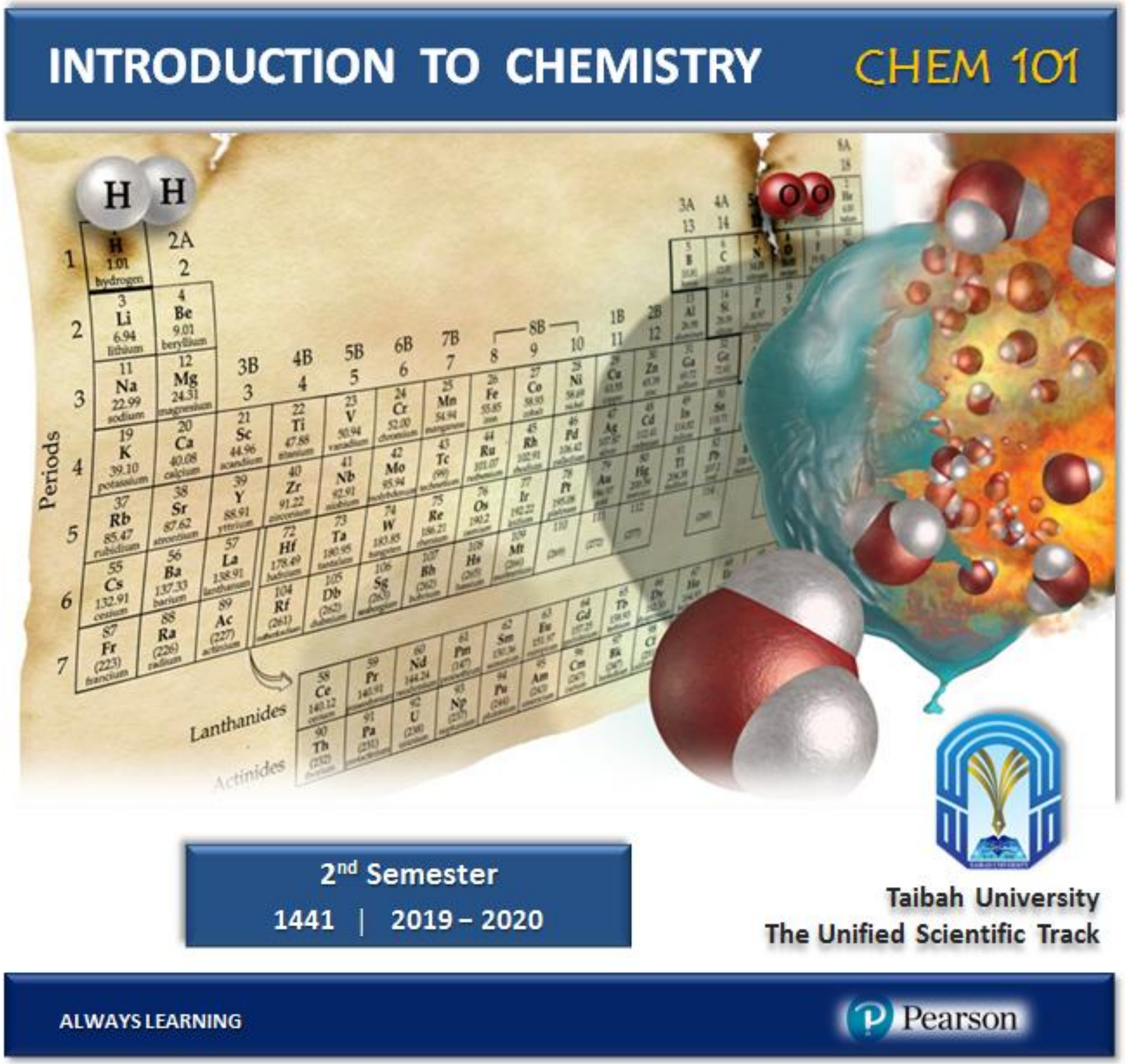
a. O b. K c. Al d. Rb e. N

5- Using a copy of the periodic table, write the name of each element and classify it as a metal, nonmetal, or metalloid:

a. Na b. Mg c. Br d. N e. As

6- Using a copy of the periodic table, classify each element as an alkali metal, alkaline earth metal, halogen, or noble gas:

a. sodium b. iodine c. calcium d. barium e. krypton



Lecture Presentation

Chapter 2

Atoms, Molecules, Ions, and Periodicity

Topic 06

Electron Configurations of The Atoms

Orbitals and Quantum Numbers

- **N. Bohr's Model:** the electrons move in spherical orbits at fixed distances from the nucleus (similar to structure of the solar system).
- **E. Schrödinger** developed 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**

✓ Quantum numbers designate specific **shells**, **subshells, orbitals, and spins** of electrons. This means that they describe the **characteristics** of an electron in an atom:

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

Orbitals and Quantum Numbers

- **Principal Quantum Number, *n***

  - Indicates the energy level (also known as: *shell*) in which the electron's orbital resides.

  - The values of ***n*** are integers > 0

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

Orbitals and Quantum Numbers

- **Angular Momentum Quantum Number, ℓ**

  - Indicates the sublevel of the electron and 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*

Orbitals and Quantum Numbers

- **Magnetic Quantum Number, m_ℓ**
  - Describes the three-dimensional orientation of the orbital.
  - Values are integers ranging from $-\ell \le m_\ell \le \ell$

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

Orbitals and Quantum Numbers

- **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 ↓.
- Due to the spinning of the electron, it generates a magnetic field.
- No two paired electrons can have the same spin value.

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

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

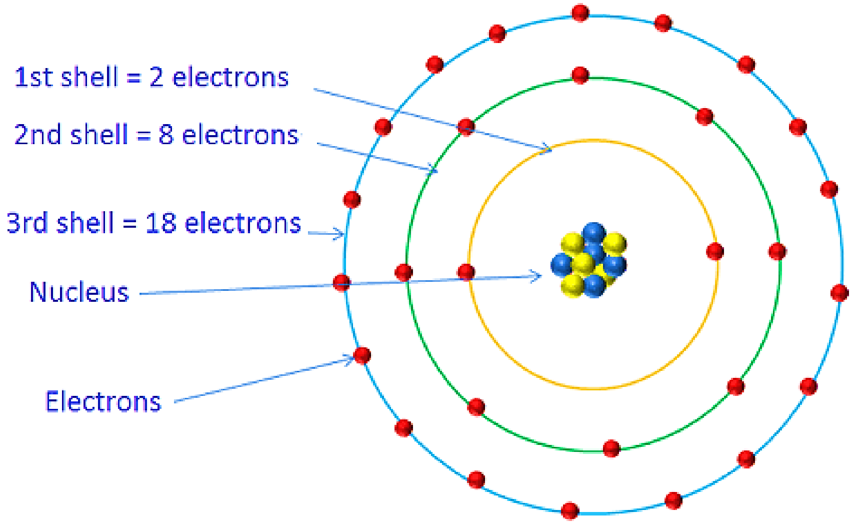
✓ **Each orbital in any sublevel is able to hold a maximum of 2 electrons:**

- The ***"s"*** sublevel has 1 orbital and can therefore hold only 2 electrons.
- The ***"p"*** sublevel has 3 orbitals and can therefore hold 6 electrons.
- The ***"d"*** sublevel has 5 orbitals and can therefore hold 10 electrons.
- The ***"f"*** sublevel has 7 orbitals and can therefore hold 14 electrons.

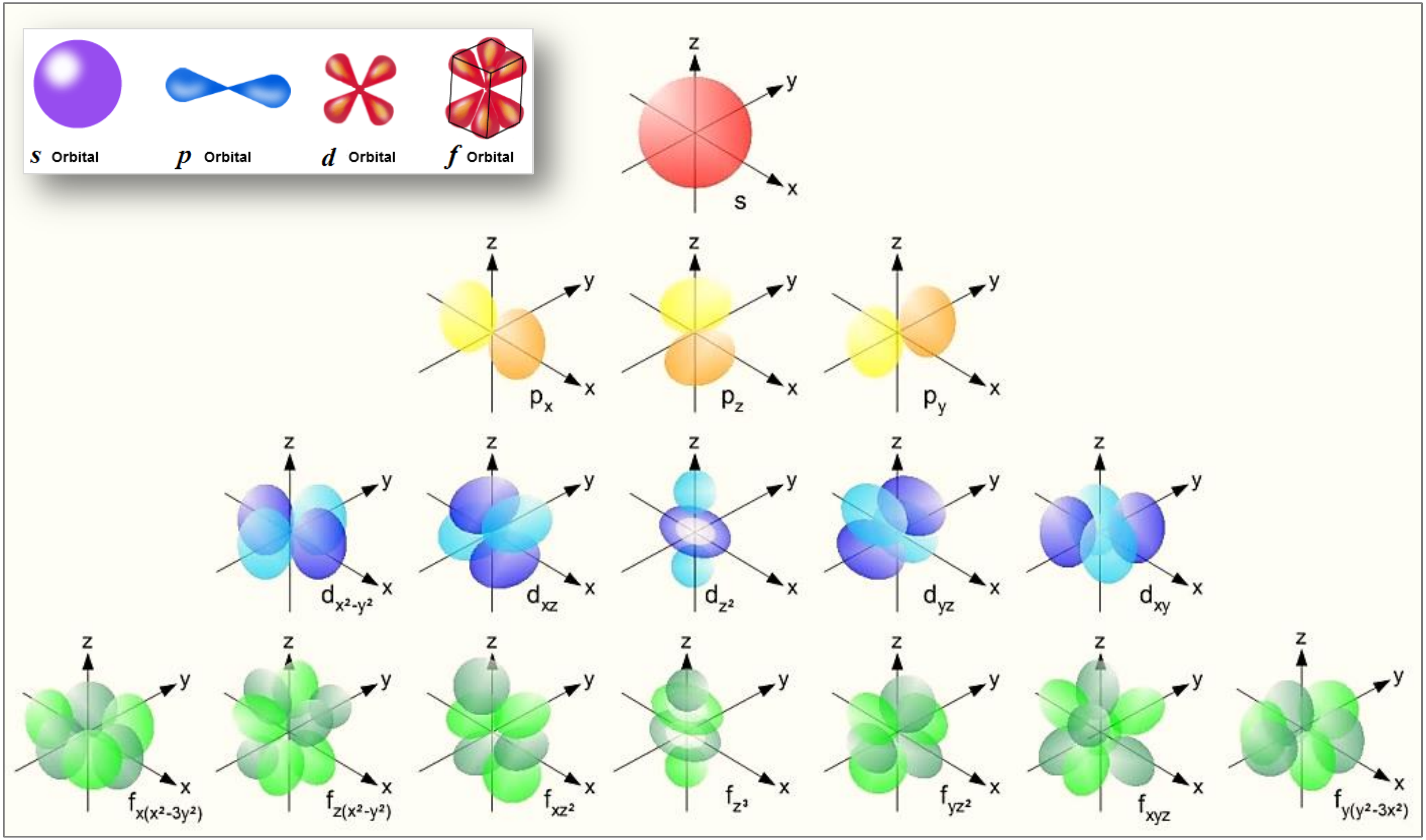
The maximum number of electrons that can occupy a specific energy level (shell) can be calculated using:

Electron Capacity = $2n^2$

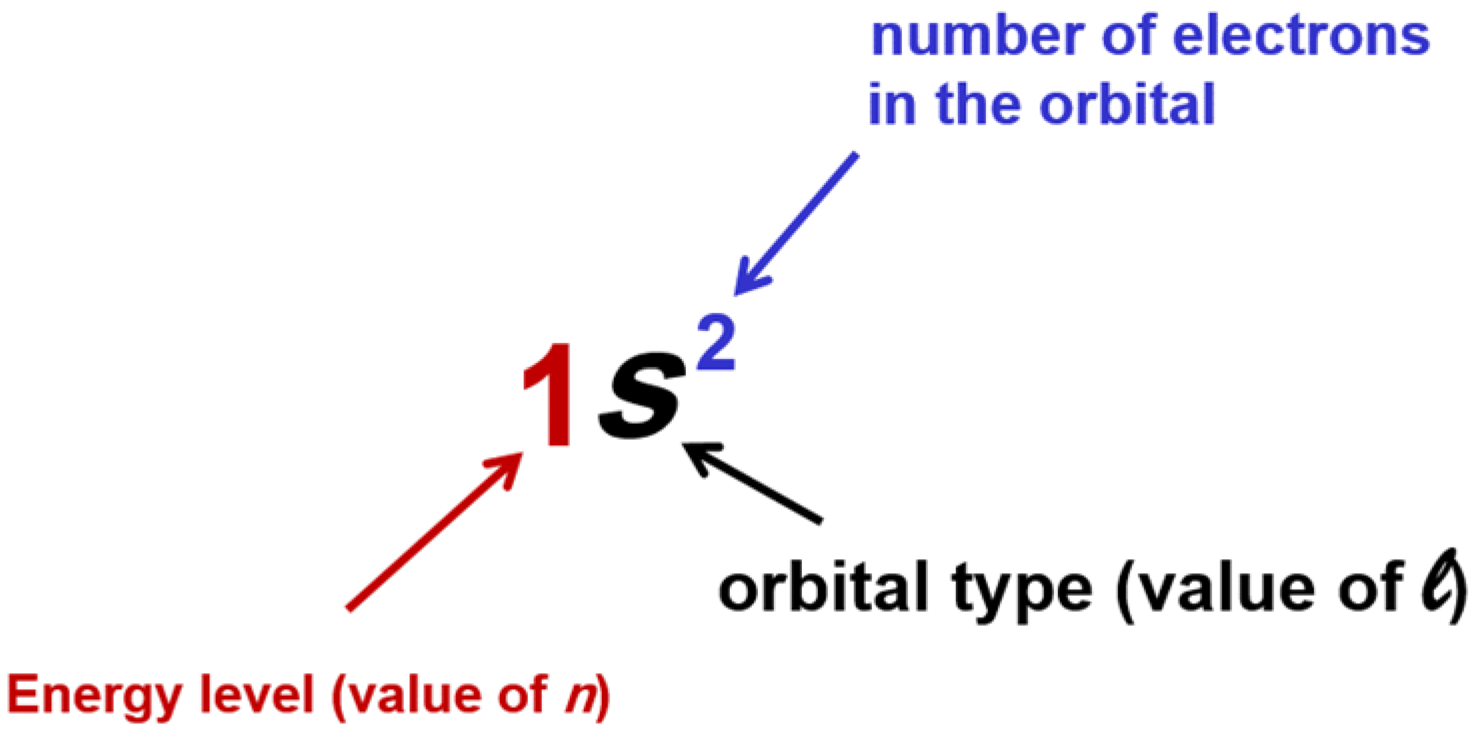
where ***n*** = the principal quantum number (number of the energy level).



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



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

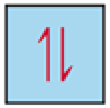


Electron Spin and The Pauli Exclusion Principle

- **Pauli Exclusion Principle:**

  In the same atom; **No** two electrons can have the same four quantum numbers.

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	$-\frac{1}{2}$

Electron Configurations of Atoms

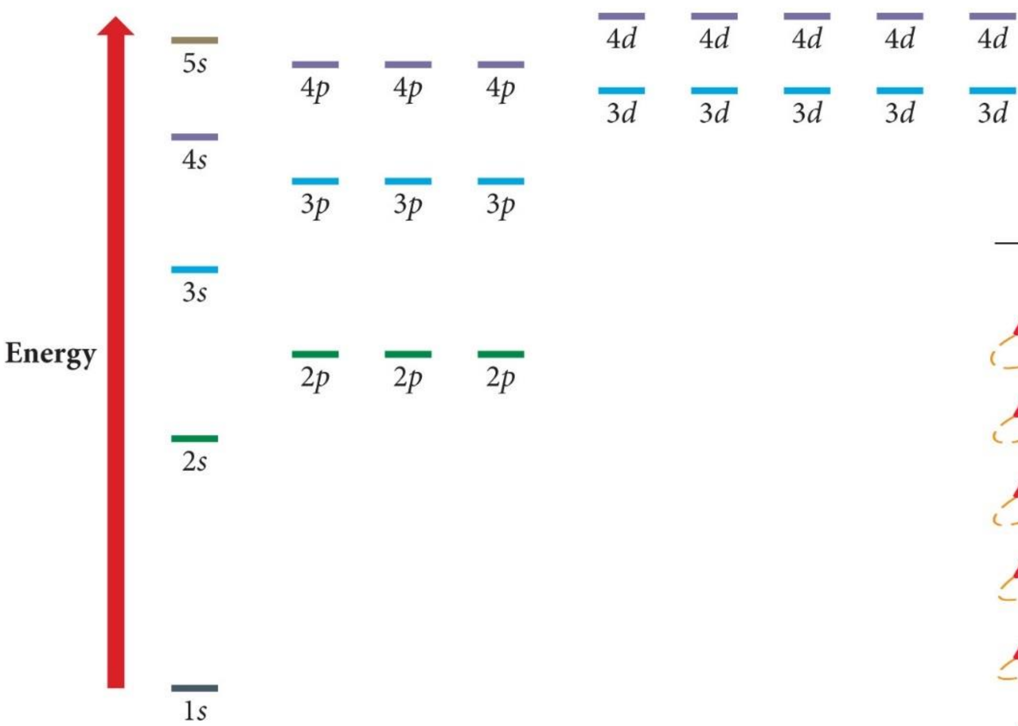
- **Rules of the aufbau principle** (*aufbau*: is a German noun meaning "building-up"):

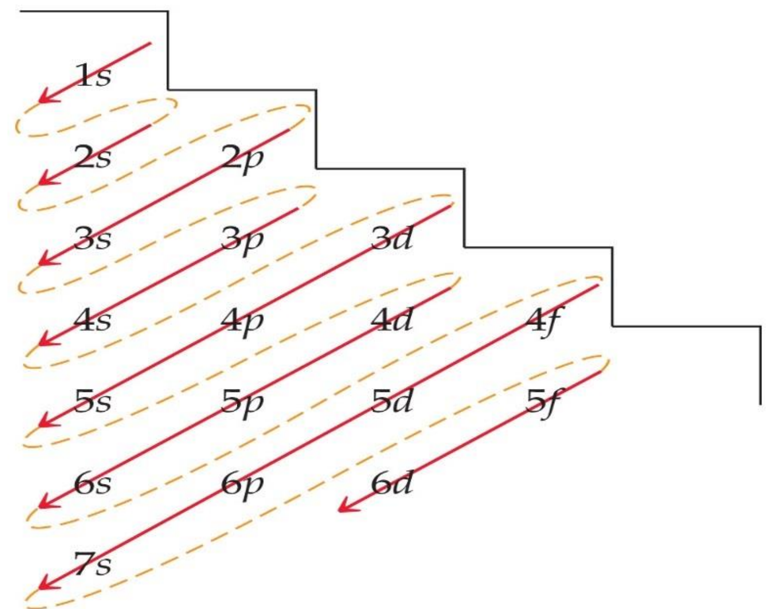
1. Lower-energy orbitals fill before higher-energy orbitals (**aufbau principle)** (*aufbau*: is a German noun meaning "building-up").
2. An orbital can hold only two electrons, which must have opposite spins (**Pauli exclusion principle**).
3. If two or more degenerate orbitals are available, follow **Hund's rule**.

Hund's Rule: when filling degenerate orbitals, the electrons fill them singly first, with parallel spins.

Electron Configurations: Ordering of Orbital Filling

General Energy Ordering of Orbitals for Multielectron Atoms





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

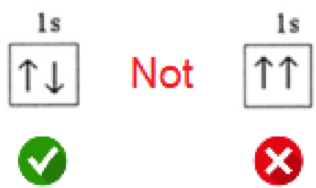
Electron Configurations: Summary

Summary of Orbital Filling Rules:

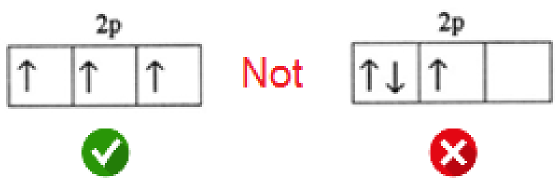
- **The Aufbau Principle:** Lower energy orbitals fill before higher energy orbitals. Orbitals fill in the following order:

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

- **Pauli Exclusion Principle:** Each Orbital (in each sublevel) can hold no more than two electrons. When two electrons occupy the same orbital, their **spins must be opposite**:

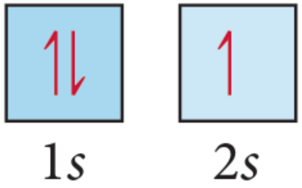


- **Hund's Rule:** When orbitals of identical energy are available, electrons first occupy these orbitals **singly** with parallel spins **before pairing**:

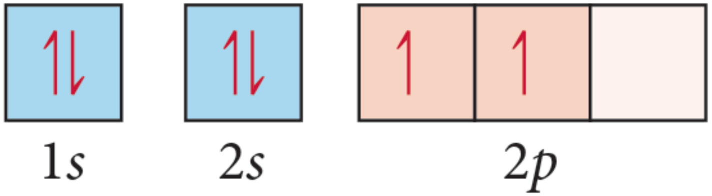


Electron Configurations: Examples

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

Electron configuration		Orbital diagram
Li	$1s^2 2s^1$	

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

Electron configuration		Orbital diagram
C	$1s^2 2s^2 2p^2$	

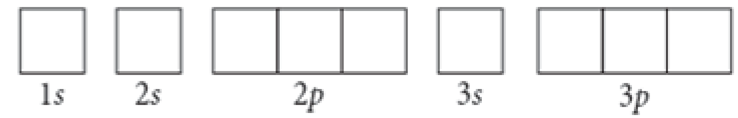
Electron Configurations: Examples

EXAMPLE Writing Orbital Diagrams

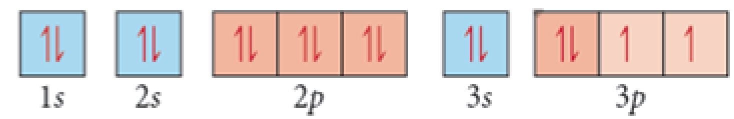
Write an orbital diagram for sulfur and determine the number of unpaired electrons.

SOLUTION

Since sulfur is atomic number 16 it has 16 electrons and the electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^4$. Draw a box for each orbital putting the lowest energy orbital ($1s$) on the far left and proceeding to orbitals of higher energy to the right.



Distribute the 16 electrons into the boxes representing the orbitals allowing a maximum of two electrons per orbital and remembering Hund's rule. You can see from the diagram that sulfur has two 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	$1s^2 2s^2 2p^6 3s^2 3p^1$

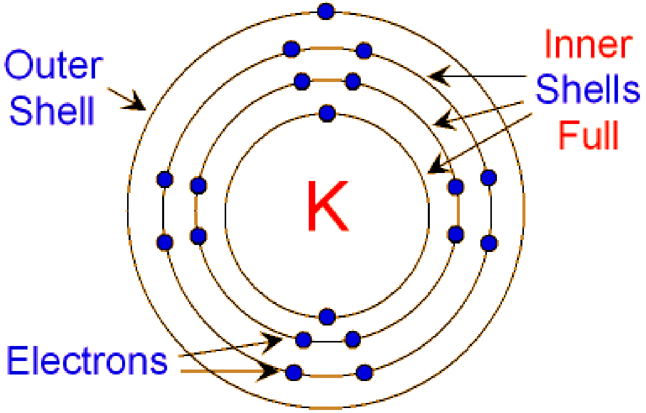
Electron Configurations: Examples

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	$1s^2 2s^2 2p^6$	$1s^2 2s^2 2p^2 2p^2 2p^2$	↑↓	↑↓	↑↓	↑↓	↑↓

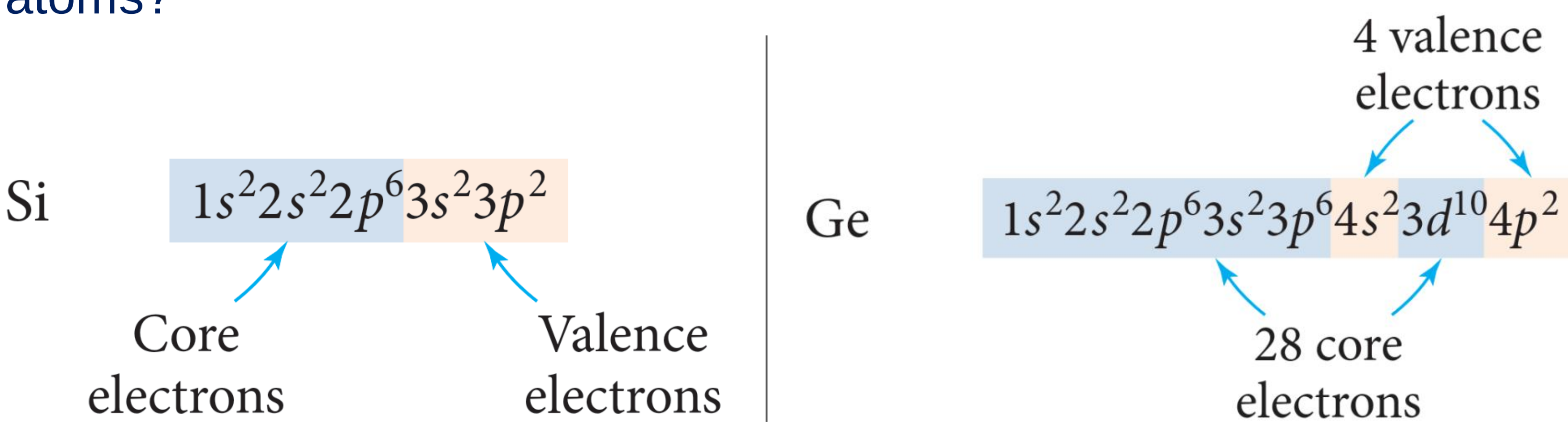
- **Valence Electrons:** electrons in all the sublevels within the highest principal energy level (***n***) in the atom. (*also known as "the valence shell" or "the outer shell")*

  - The number of "valence electrons" is very important in determining both chemical and physical behavior of the atom.
  - Valence electrons in atoms participate in:
    - ✓ **Bonding (ionic and covalent)**
    - ✓ **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, also known as the inner shells).

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 there 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

Periods	Groups 1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	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.

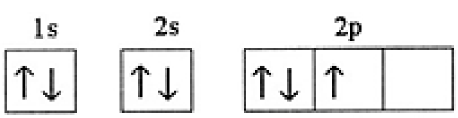
a. Ba   b. Cs   c. Ne   d. S   e. C

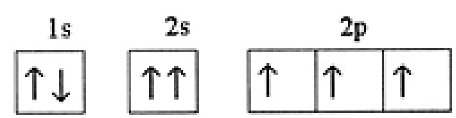
4- The complete electron configuration of sulfur is ________.

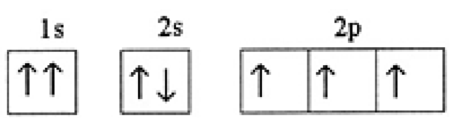
A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ B) $1s^2\ 2s^2\ 2p^{10}\ 3s^2$

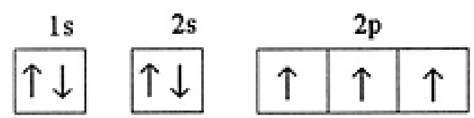
C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^2$ E) $1s^4\ 2s^4 2\ p^6\ 3s^2$

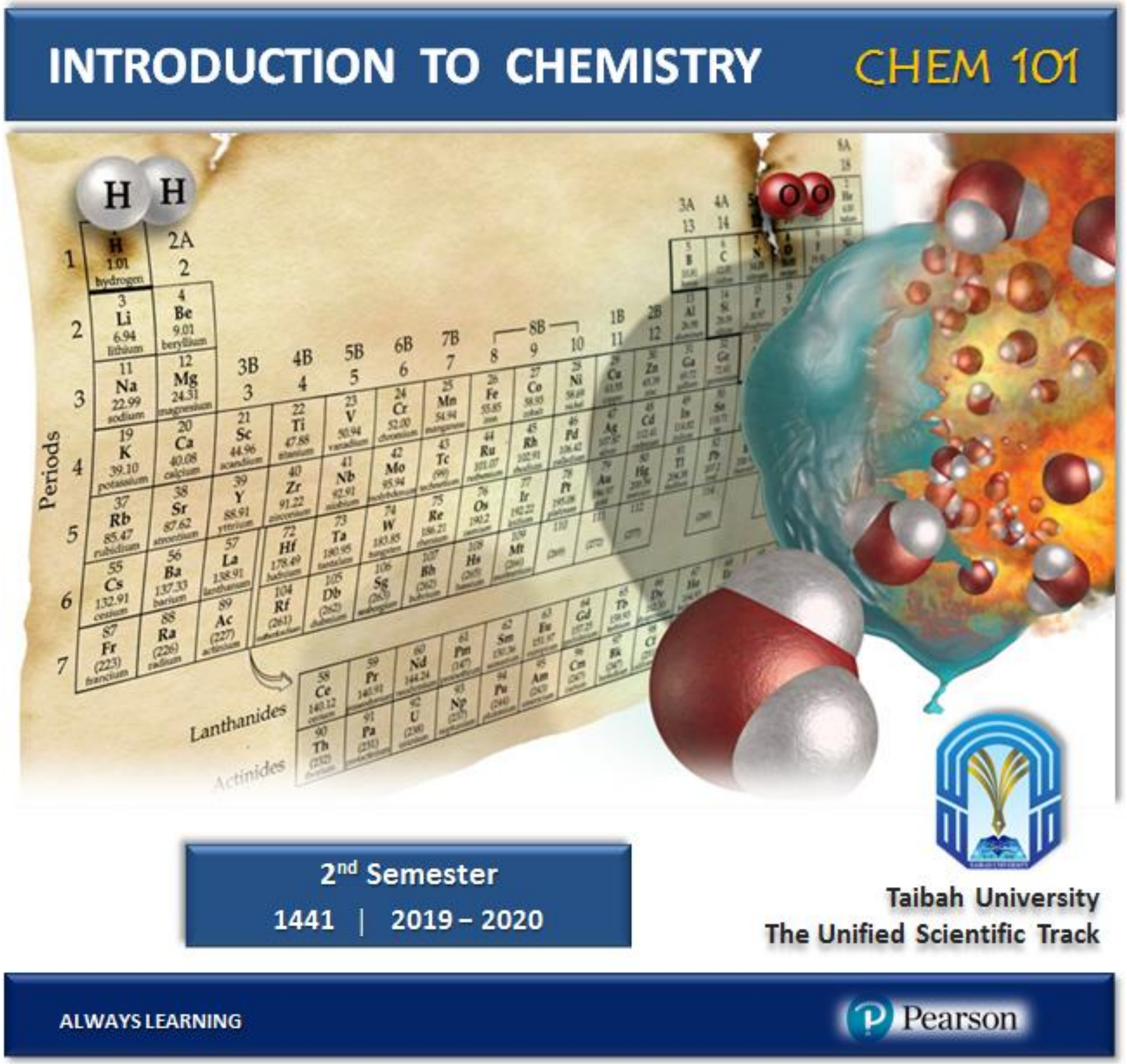
5- Which one of the following is the correct electron configuration for a ground-state nitrogen atom?

A) 

B) 

C) 

D) 



Lecture Presentation

Chapter 2

Atoms, Molecules, Ions, and Periodicity

Topic 07

The Periodic Trends

2.11 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

2.11 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.

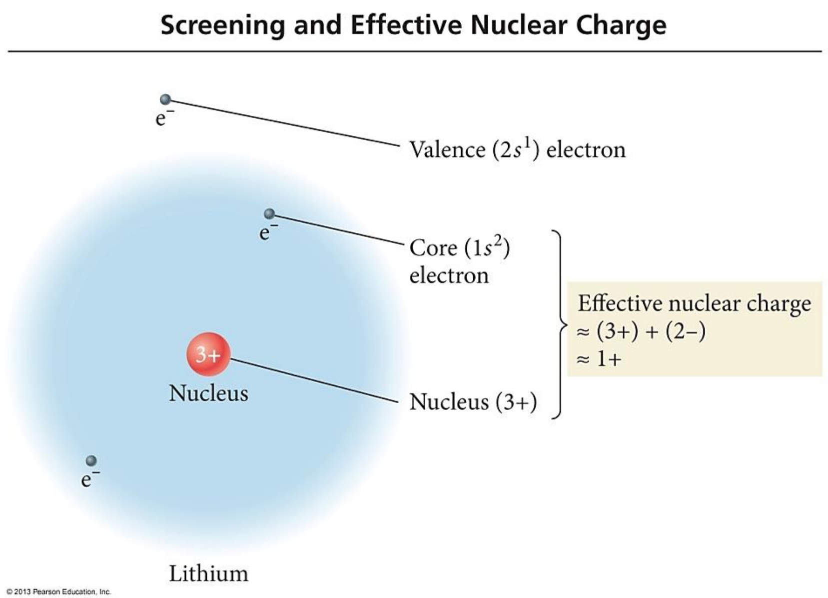
2.11 Periodic Trends: The Effective Nuclear Charge

- 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

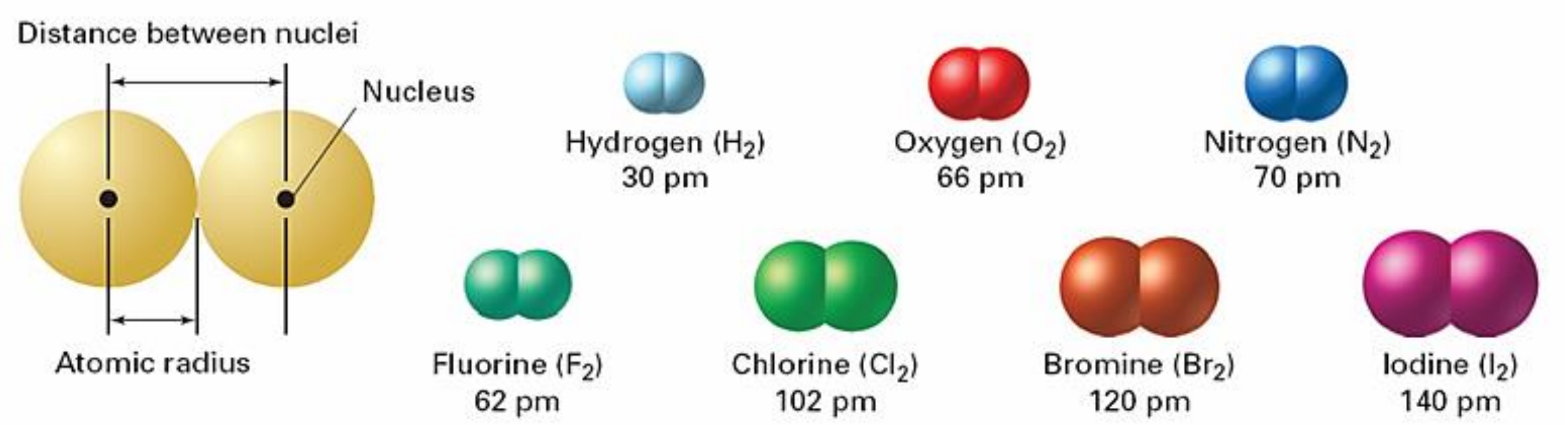
- Estimate Z_{eff} = [Z (atomic number) - (number of inner electrons)]
  - Pull felt by ***2s*** electron in Li:

    $Z_{eff} = 3 - 2 = 1$



2.11 Periodic Trends: Atomic Radii (Sizes of Atoms)

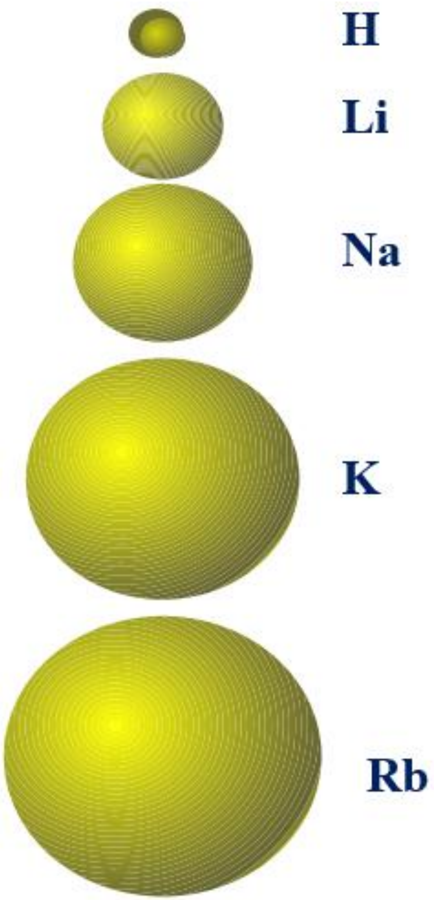
- **Atomic Radius:** is an average radius of an atom based on measuring large numbers of molecules of elements and compounds.
- ✓ There are several methods for measuring the radius of an atom, and they give slightly different numbers.
  - ✓ Van der Waals radius = nonbonding
  - ✓ Covalent radius = bonding radius



2.11 Periodic Trends: Atomic Radii (Sizes of Atoms)

- **General trend in atomic radii:**

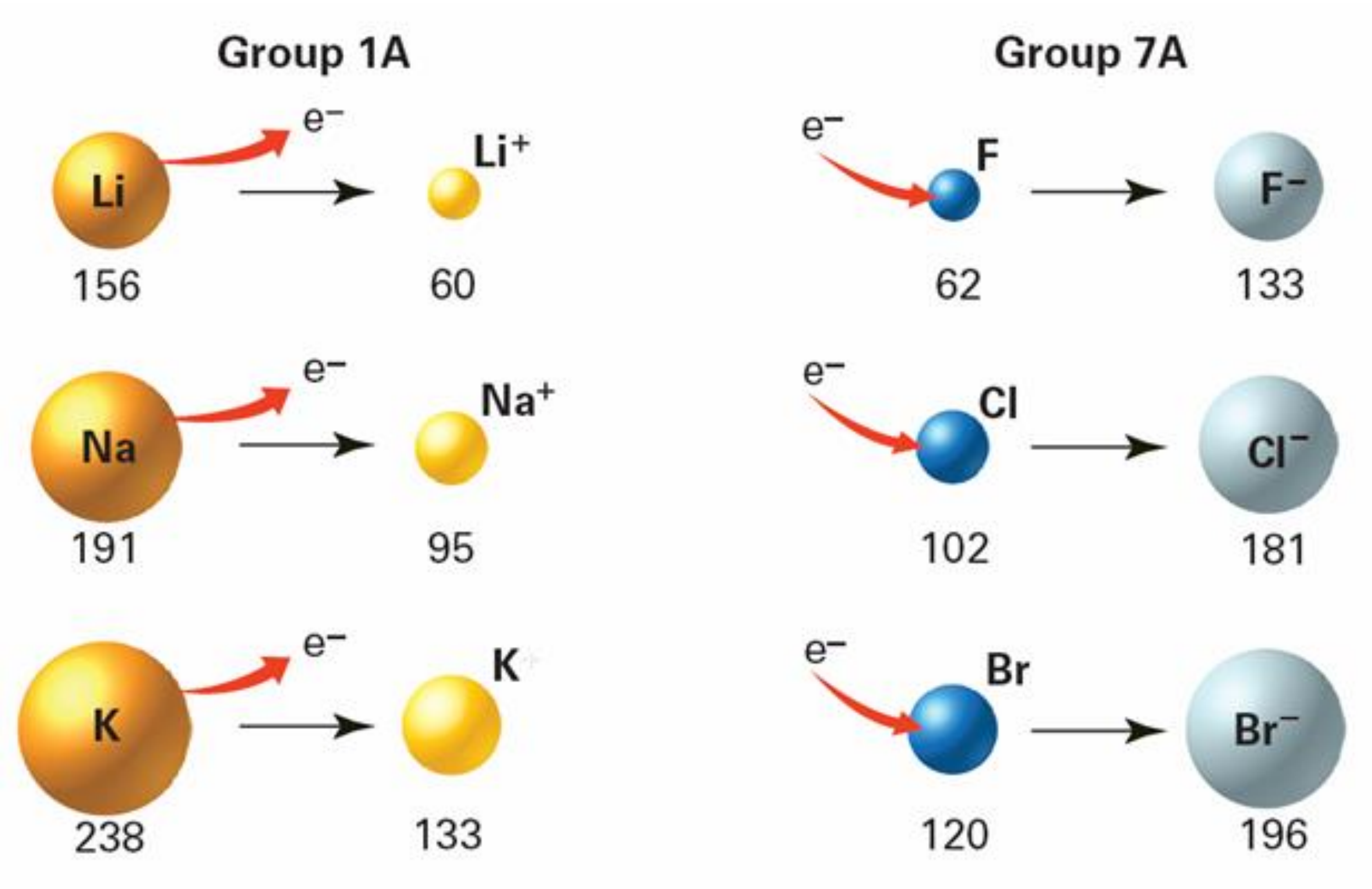
- ✓ Atomic radius decreases across the period (left to right).
  - ✓ Adding electrons to same valence shell
  - ✓ Effective nuclear charge increases
  - ✓ Valence shell held closer
- ✓ Atomic radius increases down the group (up to down).
  - ✓ Valence shell farther from nucleus
  - ✓ Effective nuclear charge fairly close



2.11 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

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



- **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 is the easiest to remove, i.e. has the 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.

2.11 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 lower the ionization energy.

3- **Radius**: the greater the atomic radius, the lower the ionization energy.

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

2.12 Periodic Trends: Electron Affinities

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

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

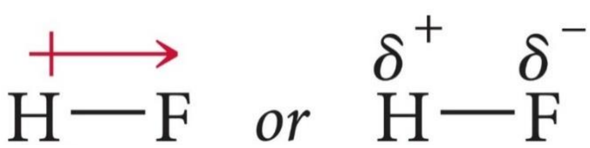
- EA can be either endothermic or exothermic in nature.
  - Why either energy exchange?
    - It is due to electron-electron repulsion within orbitals and the volume of the atom.
- **General trends in electron affinity:**
  - EA increases across a period.
    - EA becomes more positive due to increase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

2.12 Periodic Trends: Metallic Character

- **Metallic Character:** is how close an element's properties to the ideal properties of metals.
  - More malleable and ductile, better conductor, and easier to lose electrons to form cations.
- **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

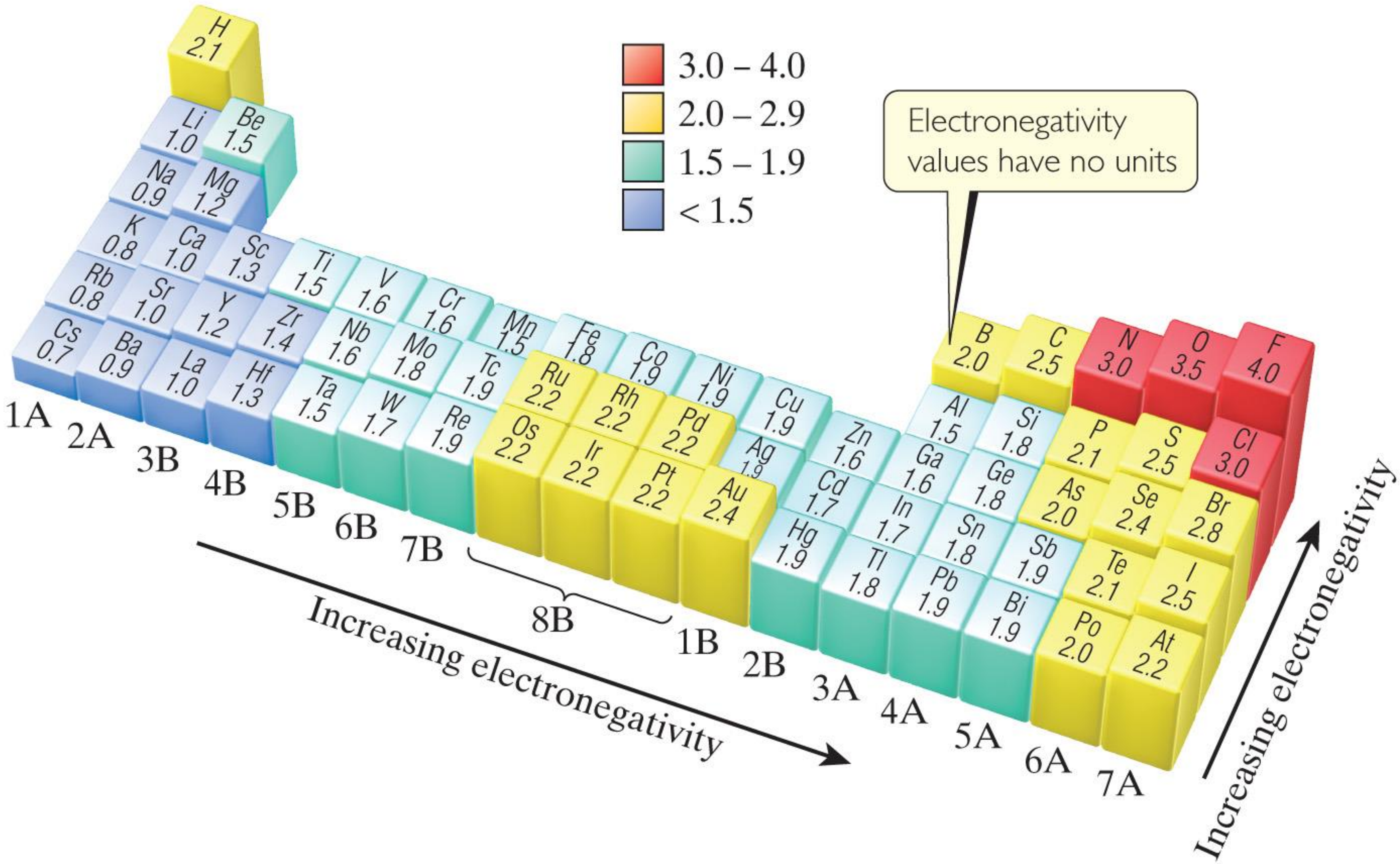
2.12 Periodic Trends: Electronegativity

- **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.

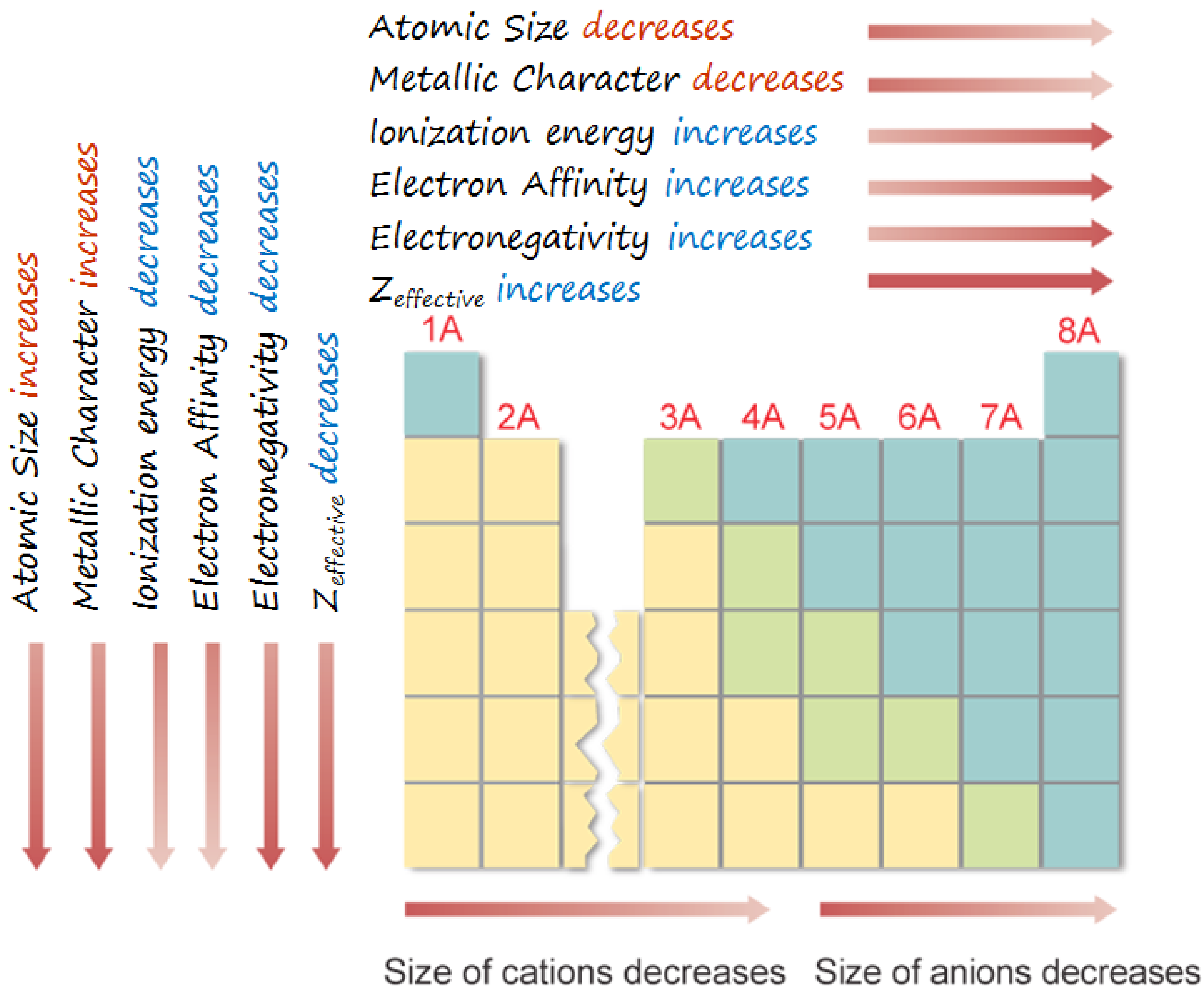


- **General trends in electronegativity:**
  - ✓ Electronegativity increases across a period.
  - ✓ Electronegativity decreases down a group.

2.12 Periodic Trends: Electronegativity



2.12 Periodic Trends: A Summary



Assessment

Answer the following questions:

1. Arrange these elements: Mg, Na, Cl, 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 <u>largest</u> 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 <u>increasing</u> atomic radius: Ca, Sc, As, Co, Fe.

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

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

a. Electronegativity
b. Ionization energy
c. Atomic radius
d. Metallic character
e. Electron affinity

▲ Periodic Table of the Elements

Key

Atomic number → 6; Symbol → C; Name → Carbon; atomic mass → 12.01

An element

Main group (groups 1–2, 13–18); Transition metals (groups 3–12)

Period	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 Hydrogen 1.008																	2 He Helium 4.003
2	3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
3	11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
4	19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.41	31 Ga Gallium 69.72	32 Ge Germanium 72.64	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
5	37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3
6	55 Cs Cesium 132.9	56 Ba Barium 137.3	57 La Lanthanum 138.9	72 Hf Hafnium 178.5	73 Ta Tantalum 180.9	74 W Tungsten 183.8	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
7	87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (267)	105 Db Dubnium (268)	106 Sg Seaborgium (271)	107 Bh Bohrium (272)	108 Hs Hassium (270)	109 Mt Meitnerium (276)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (280)	112 Cn Copernicium (285)	113 Nh Nihonium (284)	114 Fl Flerovium (289)	115 Mc Moscovium (288)	116 Lv Livermorium (293)	117 Ts Tennessine (293)	118 Og Oganesson (294)

Lanthanides 6	58 Ce Cerium 140.1	59 Pr Praseodymium 140.9	60 Nd Neodymium 144.2	61 Pm Promethium (145)	62 Sm Samarium 150.4	63 Eu Europium 152.0	64 Gd Gadolinium 157.3	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0
Actinides 7	90 Th Thorium 232.0	91 Pa Protactinium 231.0	92 U Uranium 238.0	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)