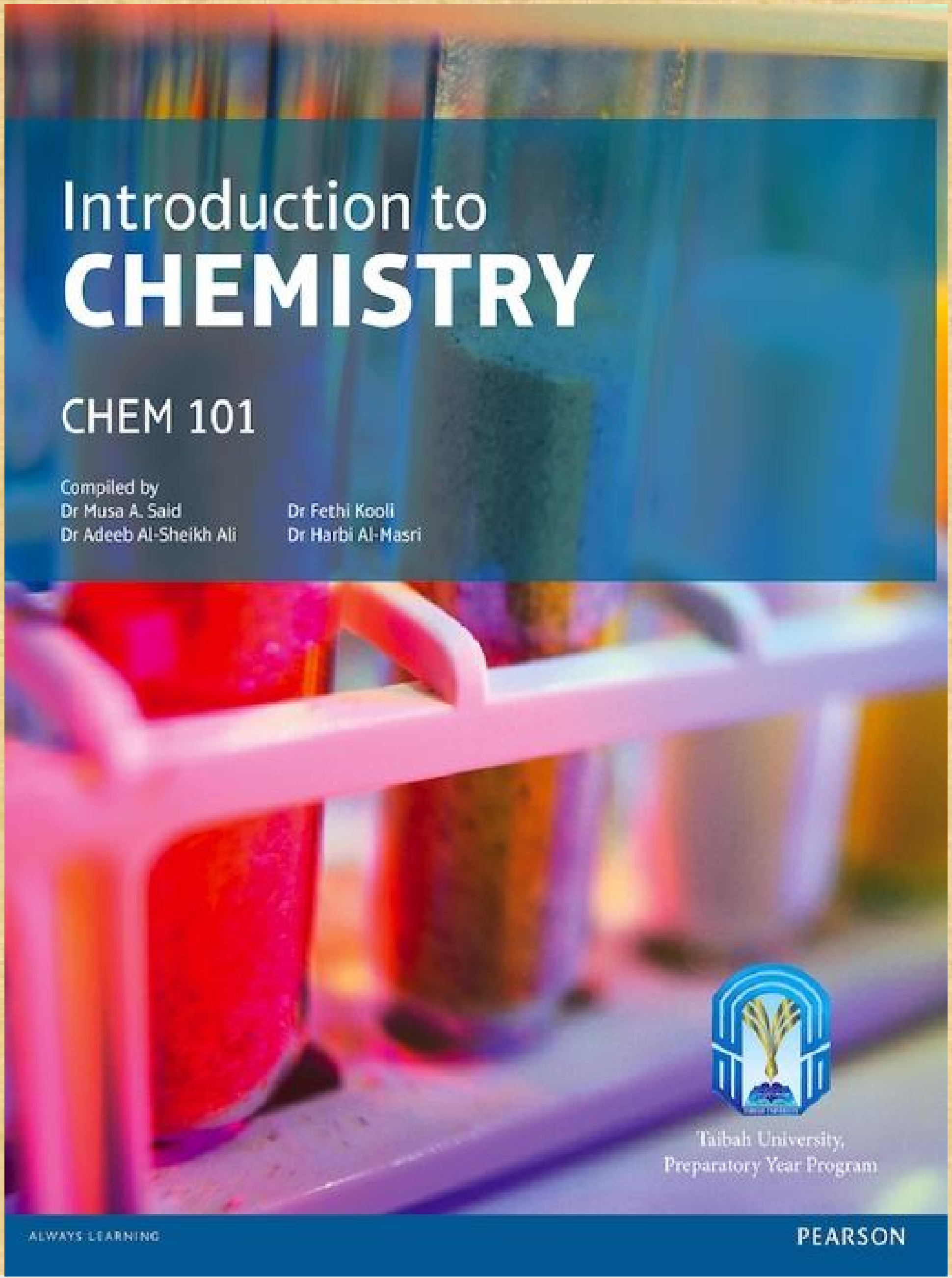

Introduction to
CHEMISTRY

CHEM 101

Compiled by
Dr Musa A. Said
Dr Adeeb Al-Sheikh Ali
Dr Fethi Kooli
Dr Harbi Al-Masri

Taibah University,
Preparatory Year Program

ALWAYS LEARNING
PEARSON

Lecture Presentation

Chapter 2

Atoms, Molecules, Ions and Periodicity

الذرات والجزيئات، والأيونات وتواترها

Matter: Its Theories & Laws

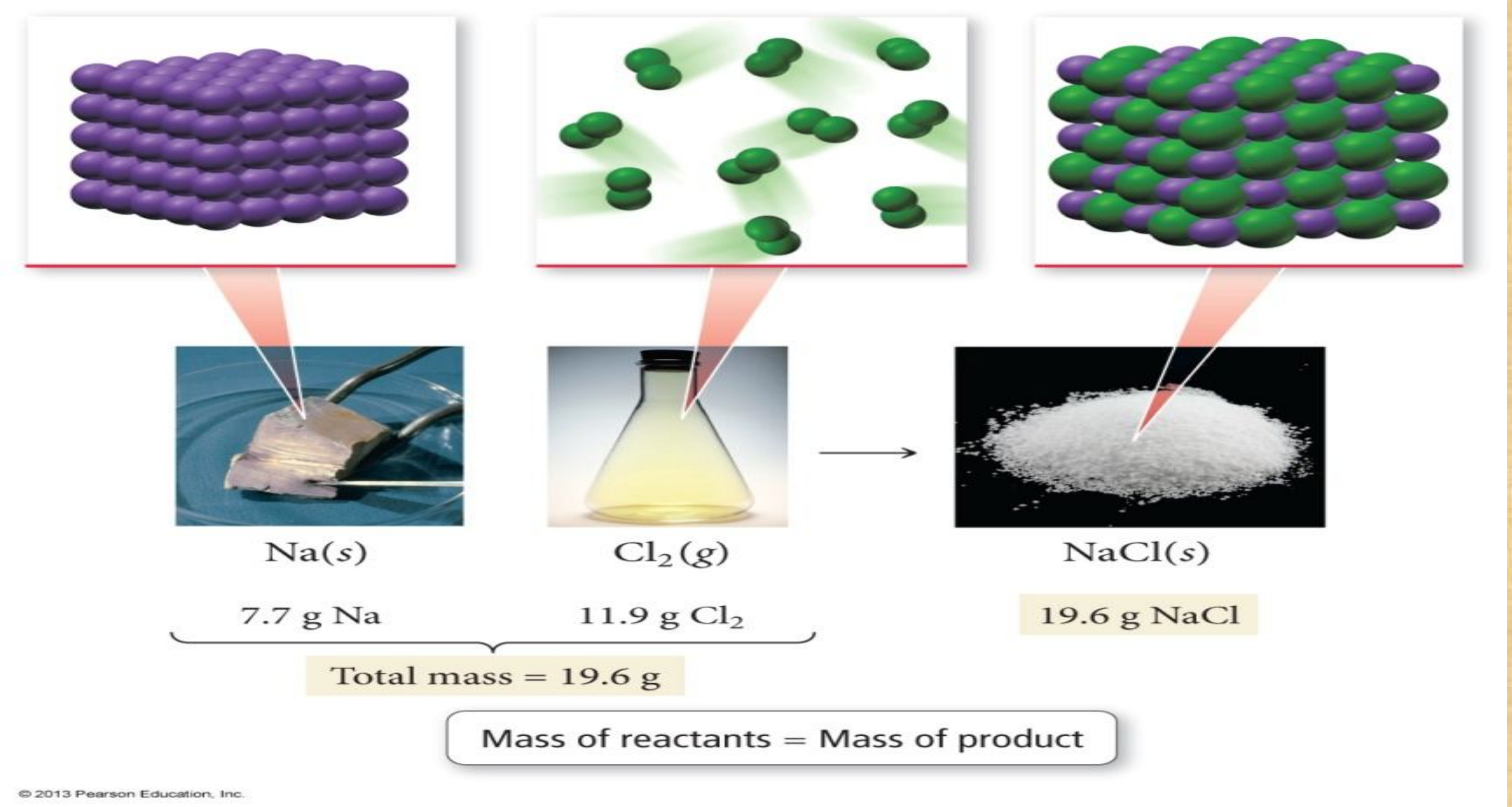
• Law of the Conservation of Matter	قانون حفظ المادة
• Atomic Theory of Matter	النظرية الذرية للمادة
• Law of Definite Proportions	قانون النسب المحددة
• Law of Multiple Proportions	قانون النسب المتعددة

Law of the Conservation of Matter

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



Atomic Theory of Matter

Dalton’s atomic theory of matter proposes that:

- Atoms are small, discrete, indivisible pieces of matter.

- All elements are made up of particles called atoms.
- An element’s atoms are identical in size, mass, & chemical properties.
  - Scientists did not know about isotopes.
    - Isotopes are elemental atoms that differ in their mass due to different number of neutrons.

- Molecules (compounds) are formed when two or more elements combined.
- Molecules are simple whole-number ratios of the combined elements.

Law of Definite Proportions

Law of definite proportions:

- **For a given compound, the elements always combine in the same proportion.**
- **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 molecule (NaCl) is always a 1:1 ratio of one sodium atom to chlorine atom.**
  - **A 100.0 g sample of NaCl contains 39.3 g Na & 60.7 g Cl.**

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

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

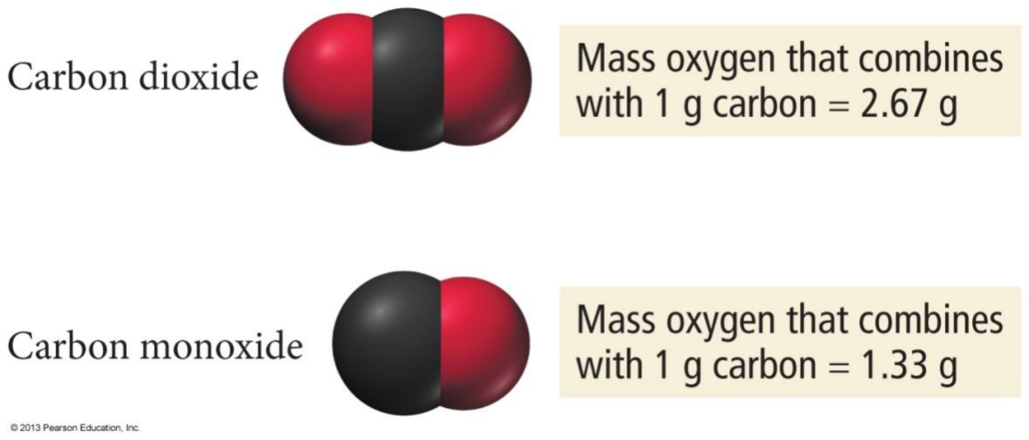
$$\frac{\text{Mass Cl}}{\text{Mass Na}} = \frac{35.44\ \text{g}}{22.99\ \text{g}} = \underline{1.54}$$

Law of Multiple Proportions

Law of multiple proportions:

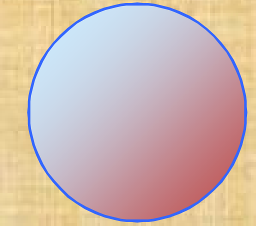
- Two elements A and X can form different compounds by combining in different proportions.
  - These combinations can be represented as a ratio.

- 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.

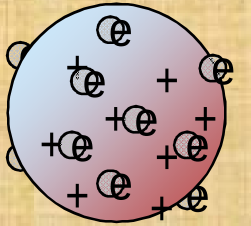


تطور مفهوم الذرة عبر التاريخ

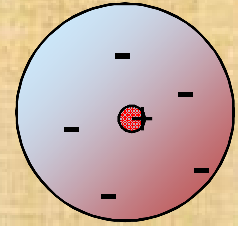
Historical Perspective



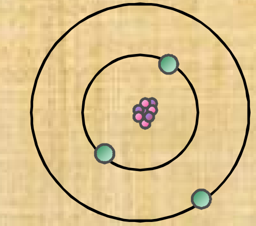
نموذج الاغريق
(400 B.C.)



نموذج طمسون
(1897)



نموذج رذرفورد
(1909)



نموذج بور
(1913)

الكثافة الإلكترونية

1803 John Dalton pictures **atoms as tiny, indestructible particles, with no internal structure.**

1897 J.J. Thomson, a British scientist, discovers the electron, leading to his "plum-pudding" model. He pictures electrons embedded in a sphere of positive electric charge.

1911 New Zealander Ernest Rutherford states that an atom has a dense, positively charged nucleus. Electrons move randomly in the space around the nucleus.

1913 In Niels Bohr's model, the electrons move in spherical orbits at fixed distances from the nucleus.

1926 Erwin Schr?dinger develops mathematical equations to describe the motion of electrons in atoms. His work leads to the electron cloud model.

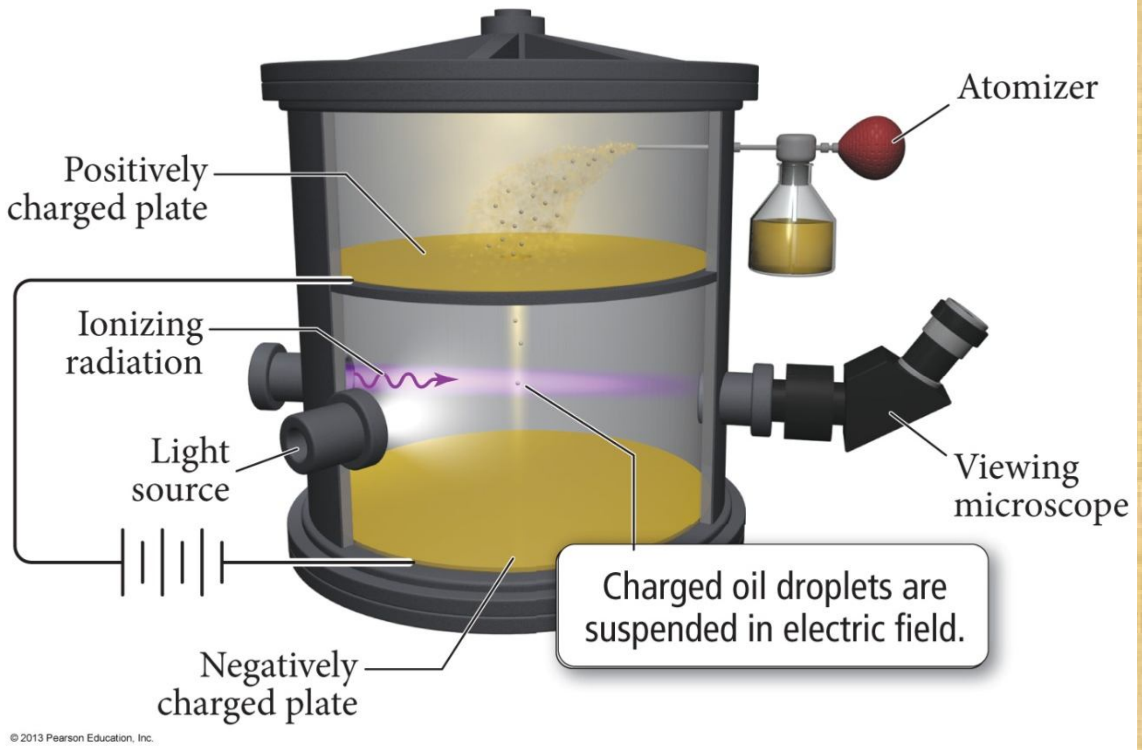
.... 1895 1900 1905 1910 1915 1920 1925 1930 1935

Dorin, Demmin, Gabel, Chemistry The Study of Matter , 3rd Edition, 1990, page 125

Atomic Structure: Discovery of the Electron

Discovery of electron:

- ***Millikan oil drop experiment***
  - **Investigation led to determining the <u>charge of the electron.</u>**



Charged oil droplets are suspended in electric field.

Atomic Structure: Electron

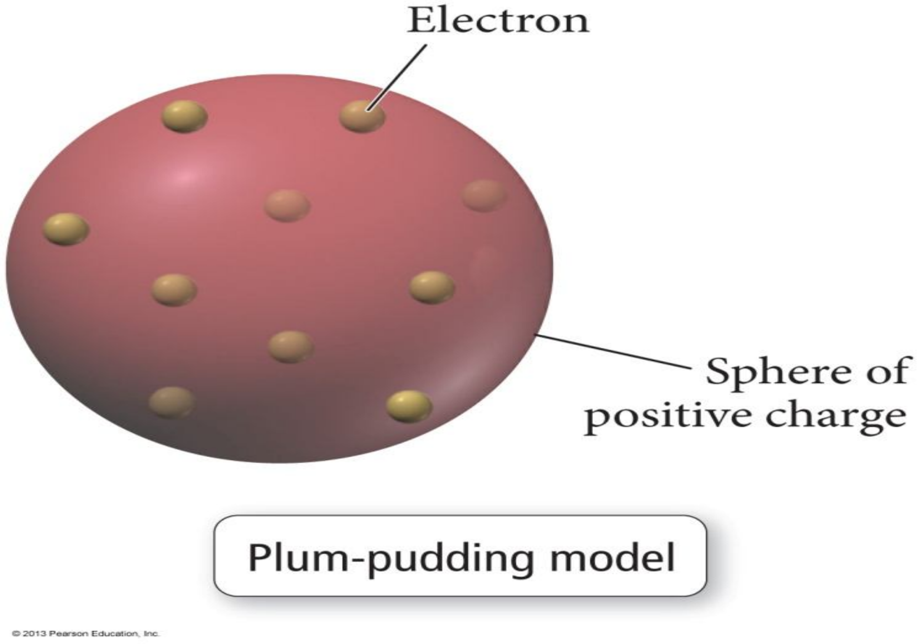
- **Electrons are particles found in all atoms.**
  - **One of the fundamental pieces of matter**

- **The electron has a charge of -1.60×10^{-19} C.**
- **The electron has a mass of 9.1×10^{-28} g.**

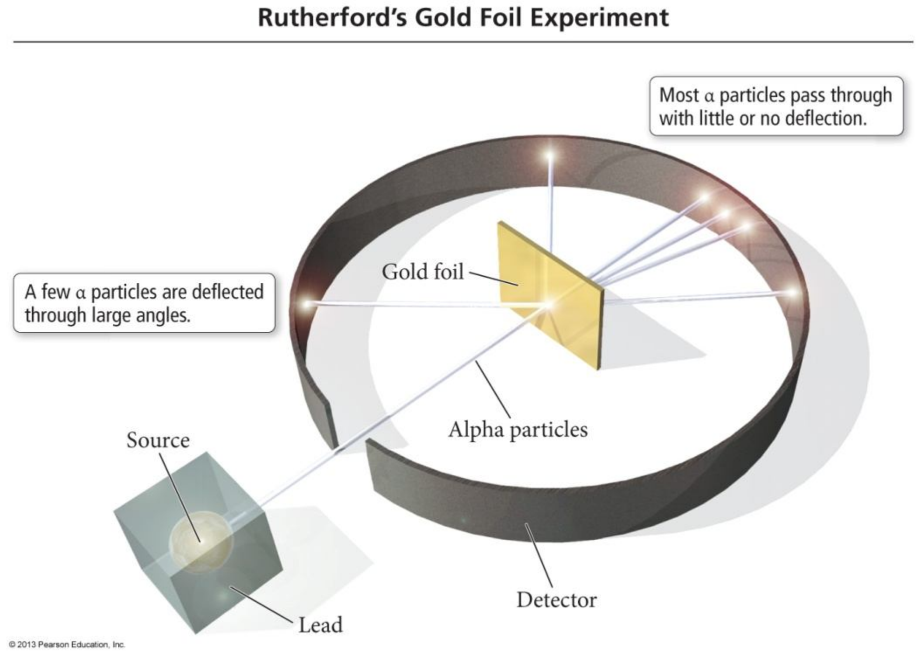
- If the particle has the same amount of charge as a hydrogen ion, then it must have a mass almost 2000x smaller than hydrogen atoms

Atomic Structure: Plum-Pudding Model

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



Rutherford’s Gold Foil Experiment Setup



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Gold foil experiment:

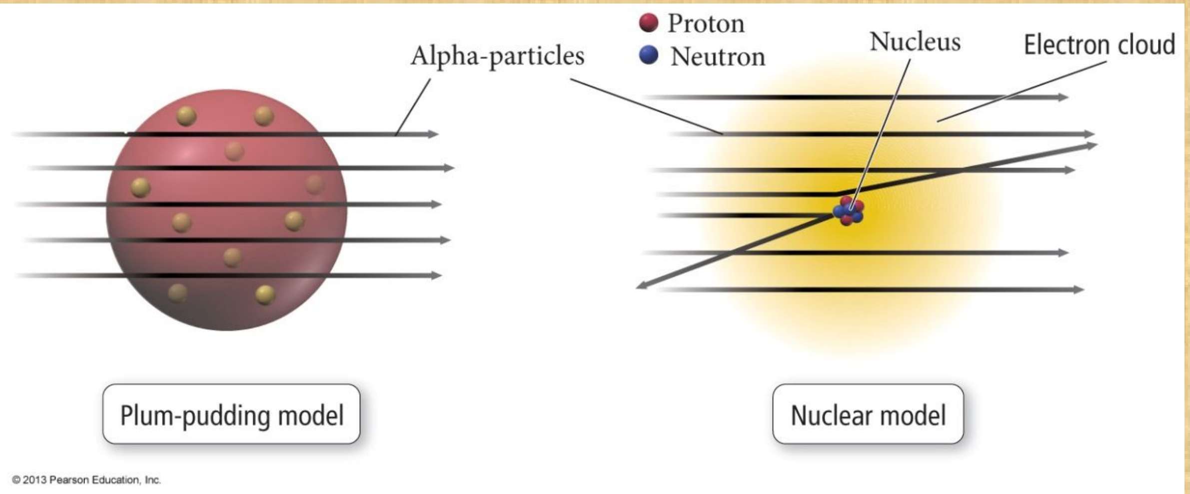
Could not explain Thomson’s plum-pudding atom model.

Led to the discovery of the atom’s nucleus.

Rutherford & the Nucleus: Gold Foil Experiment

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 balances the negative charge of the electrons.**
  - **The electrons are dispersed in the empty space of the atom surrounding the nucleus.**



Atomic Structure: Historic Perspective

- Rutherford's model (solar system)
  - The atom is mostly empty space with a **DENSE** center of mass (nucleus) and circling electrons.
  - It proposed that the nucleus had a particle that had the same amount of charge as an electron but opposite sign.
    - These particles are called **protons**.
      - charge = $+1.60 \times 10^{-19}$ C
      - mass = 1.67262×10^{-24} g

  - Since protons and electrons have the same amount of charge, for the ***atom to be neutral, there must be equal numbers of protons and electrons***.

Elements

Structure of the Atom

All matter is composed of the same basic building blocks called atoms.

Atoms are composed of three subatomic particles:

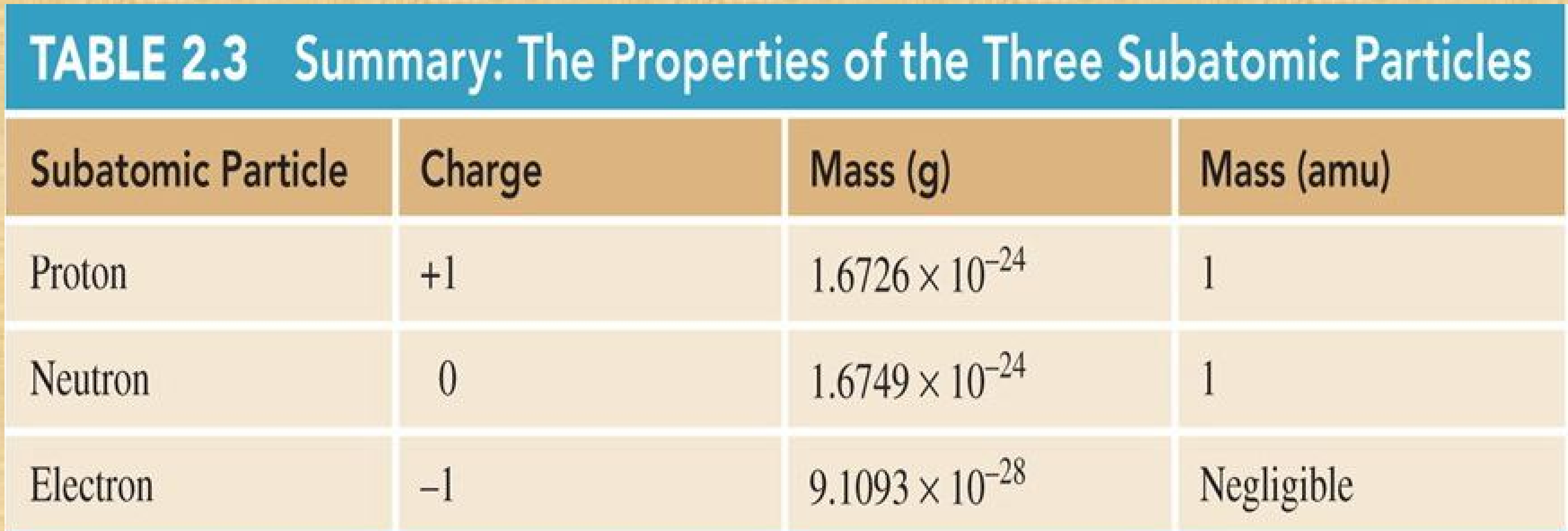
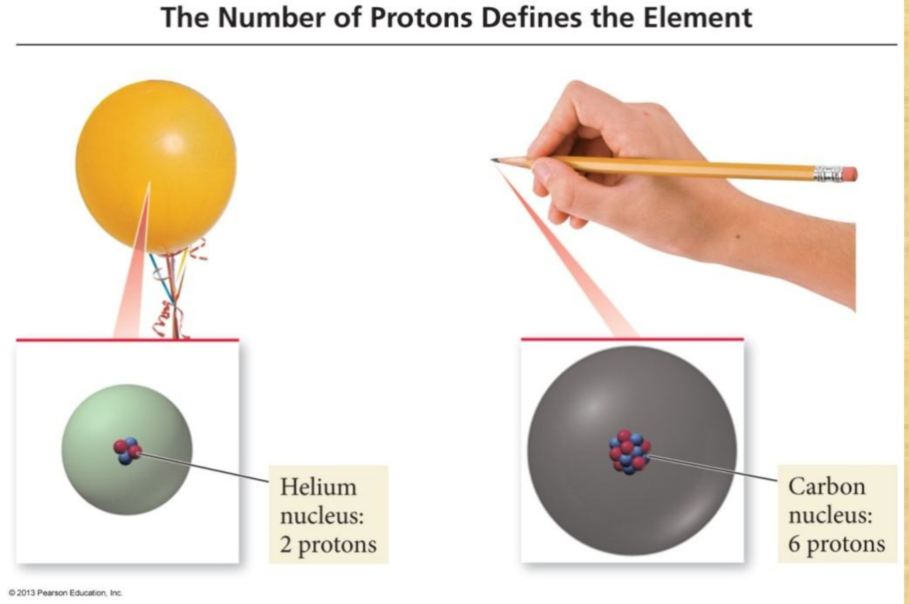


TABLE 2.3 Summary: The Properties of the Three Subatomic Particles

Subatomic Particle	Charge	Mass (g)	Mass (amu)
Proton	+1	1.6726×10^{-24}	1
Neutron	0	1.6749×10^{-24}	1
Electron	–1	9.1093×10^{-28}	Negligible

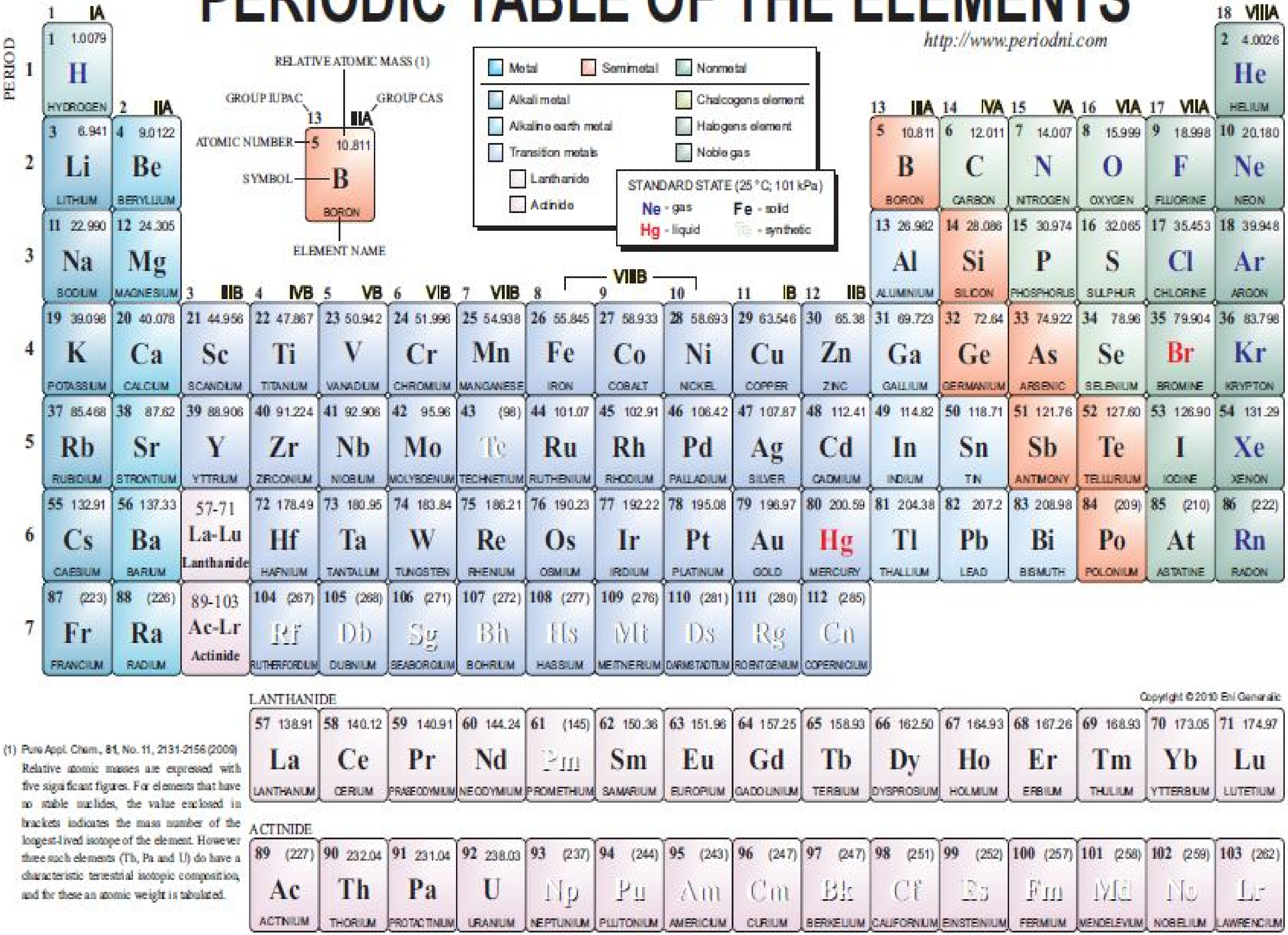
Elements

- **The number of protons located in an atom's nucleus determines the element's identity.**
  - **The number of <u>protons</u> in the nucleus of an atom is called the atomic number.**



PERIODIC TABLE OF THE ELEMENTS

http://www.periodni.com



PERIOD	GROUP 1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 1.0079 H HYDROGEN																	2 4.0026 He HELIUM
2	3 6.941 Li LITHIUM	4 9.0122 Be BERYLLIUM											5 10.811 B BORON	6 12.011 C CARBON	7 14.007 N NITROGEN	8 15.999 O OXYGEN	9 18.998 F FLUORINE	10 20.180 Ne NEON
3	11 22.990 Na SODIUM	12 24.305 Mg MAGNESIUM											13 26.982 Al ALUMINIUM	14 28.086 Si SILICON	15 30.974 P PHOSPHORUS	16 32.065 S SULPHUR	17 35.453 Cl CHLORINE	18 39.948 Ar ARGON
4	19 39.098 K POTASSIUM	20 40.078 Ca CALCIUM	21 44.956 Sc SCANDIUM	22 47.867 Ti TITANIUM	23 50.942 V VANADIUM	24 51.996 Cr CHROMIUM	25 54.938 Mn MANGANESE	26 55.845 Fe IRON	27 58.933 Co COBALT	28 58.693 Ni NICKEL	29 63.546 Cu COPPER	30 65.38 Zn ZINC	31 69.723 Ga GALLIUM	32 72.64 Ge GERMANIUM	33 74.922 As ARSENIC	34 78.96 Se SELENIUM	35 79.904 Br BROMINE	36 83.798 Kr KRYPTON
5	37 85.468 Rb RUBIDIUM	38 87.62 Sr STRONTIUM	39 88.906 Y YTTRIUM	40 91.224 Zr ZIRCONIUM	41 92.906 Nb NIOBIUM	42 95.96 Mo MOLYBDENUM	43 (98) Tc TECHNETIUM	44 101.07 Ru RUTHENIUM	45 102.91 Rh RHODIUM	46 106.42 Pd PALLADIUM	47 107.87 Ag SILVER	48 112.41 Cd CADMIUM	49 114.82 In INDIUM	50 118.71 Sn TIN	51 121.76 Sb ANTIMONY	52 127.60 Te TELLURIUM	53 126.90 I IODINE	54 131.29 Xe XENON
6	55 132.91 Cs CAESIUM	56 137.33 Ba BARIUM	57-71 La-Lu Lanthanide	72 178.49 Hf HAFNIUM	73 180.95 Ta TANTALUM	74 183.84 W TUNGSTEN	75 186.21 Re RHENIUM	76 190.23 Os OSMIUM	77 192.22 Ir IRIDIUM	78 195.08 Pt PLATINUM	79 196.97 Au GOLD	80 200.59 Hg MERCURY	81 204.38 Tl THALLIUM	82 207.2 Pb LEAD	83 208.98 Bi BISMUTH	84 (209) Po POLONIUM	85 (210) At ASTATINE	86 (222) Rn RADON
7	87 (223) Fr FRANCIUM	88 (226) Ra RADIUM	89-103 Ac-Lr Actinide	104 (267) Rf RUTHERFORDIUM	105 (268) Db DUBNIUM	106 (271) Sg SEABORGIUM	107 (272) Bh BOHRIUM	108 (277) Hs HASSIUM	109 (276) Mt MEITNERIUM	110 (281) Ds DARMSTADTIUM	111 (280) Rg ROENTGENIUM	112 (285) Cn COPERNICIUM						

RELATIVE ATOMIC MASS (1)
GROUP IUPAC 13 — GROUP CAS IIIA
ATOMIC NUMBER 5 — 10.811
SYMBOL B
BORON
ELEMENT NAME

☐ Metal ☐ Semimetal ☐ Nonmetal
☐ Alkali metal
☐ Alkaline earth metal
☐ Transition metals
☐ Lanthanide
☐ Actinide
☐ Chalcogens element
☐ Halogens element
☐ Noble gas

STANDARD STATE (25 °C; 101 kPa)
Ne - gas
Fe - solid
Hg - liquid
Tc - synthetic

LANTHANIDE

57 138.91 La LANTHANUM	58 140.12 Ce CERIUM	59 140.91 Pr PRASEODYMIUM	60 144.24 Nd NEODYMIUM	61 (145) Pm PROMETHIUM	62 150.36 Sm SAMARIUM	63 151.96 Eu EUROPIUM	64 157.25 Gd GADOLINIUM	65 158.93 Tb TERBIUM	66 162.50 Dy DYSPROSIUM	67 164.93 Ho HOLMIUM	68 167.26 Er ERBIUM	69 168.93 Tm THULIUM	70 173.05 Yb YTTERBIUM	71 174.97 Lu LUTETIUM

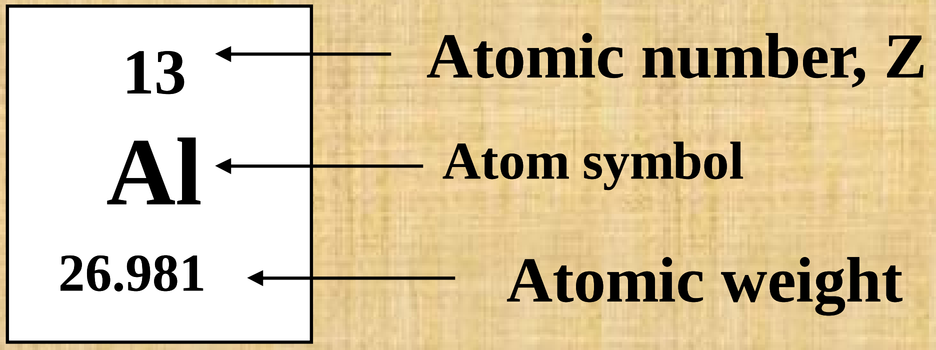
ACTINIDE

89 (227) Ac ACTINIUM	90 232.04 Th THORIUM	91 231.04 Pa PROTACTINIUM	92 238.03 U URANIUM	93 (237) Np NEPTUNIUM	94 (244) Pu PLUTONIUM	95 (243) Am AMERICIUM	96 (247) Cm CURIUM	97 (247) Bk BERKELIUM	98 (251) Cf CALIFORNIUM	99 (252) Es EINSTEINIUM	100 (257) Fm FERMIUM	101 (258) Md MENDELEVIUM	102 (259) No NOBELIUM	103 (262) Lr LAWRENCIUM

(1) Pure Appl. Chem., 81, No. 11, 2131-2156 (2009)
Relative atomic masses are expressed with five significant figures. For elements that have no stable nuclides, the value enclosed in brackets indicates the mass number of the longest-lived isotope of the element. However three such elements (Th, Pa and U) do have a characteristic terrestrial isotopic composition, and for these an atomic weight is tabulated.

Isotopes

- Isotopes are elements whose atoms differ in mass only.
  - They differ in mass because these elemental atoms have different number of neutrons.
  - They are the same element because they have the same number of protons (atomic number).
  - They are chemically identical.
- Isotopes are identified by their mass numbers.
  - Protons + neutrons = mass number
- Isotopic symbol

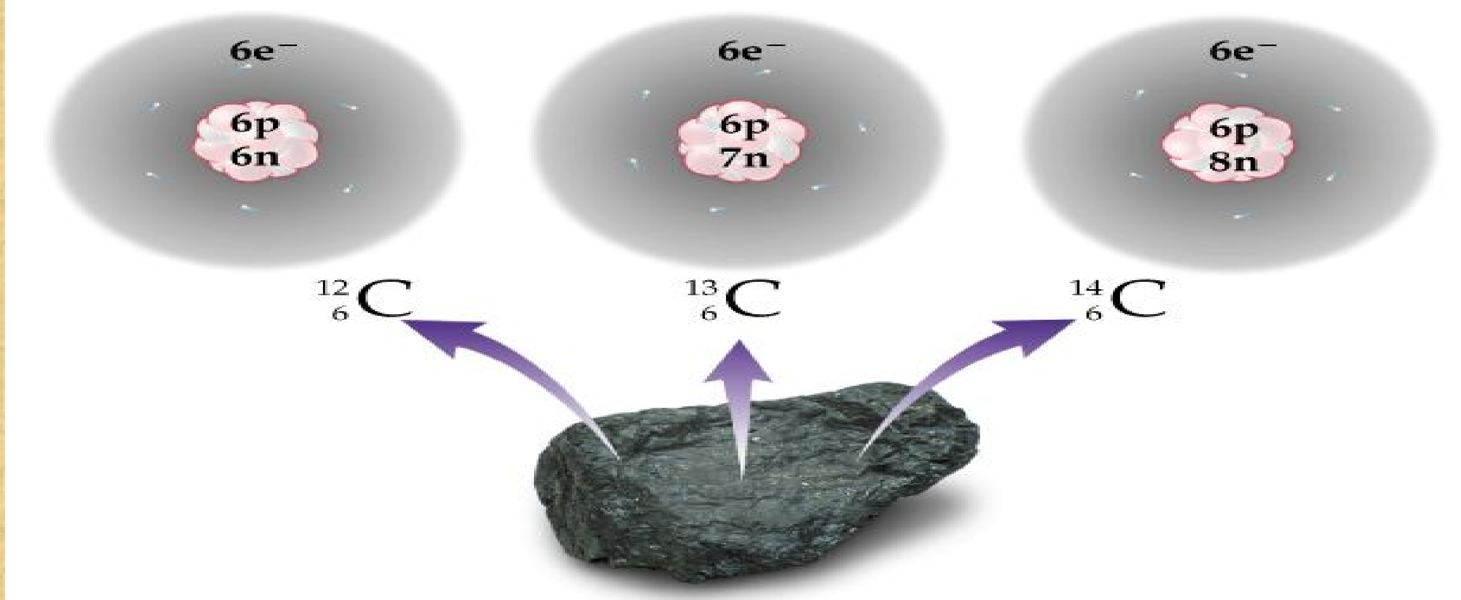


13 ← Atomic number, Z

Al ← Atom symbol

26.981 ← Atomic weight

Carbon - 12 Carbon - 13 Carbon - 14



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

Ions: Losing and Gaining Electrons

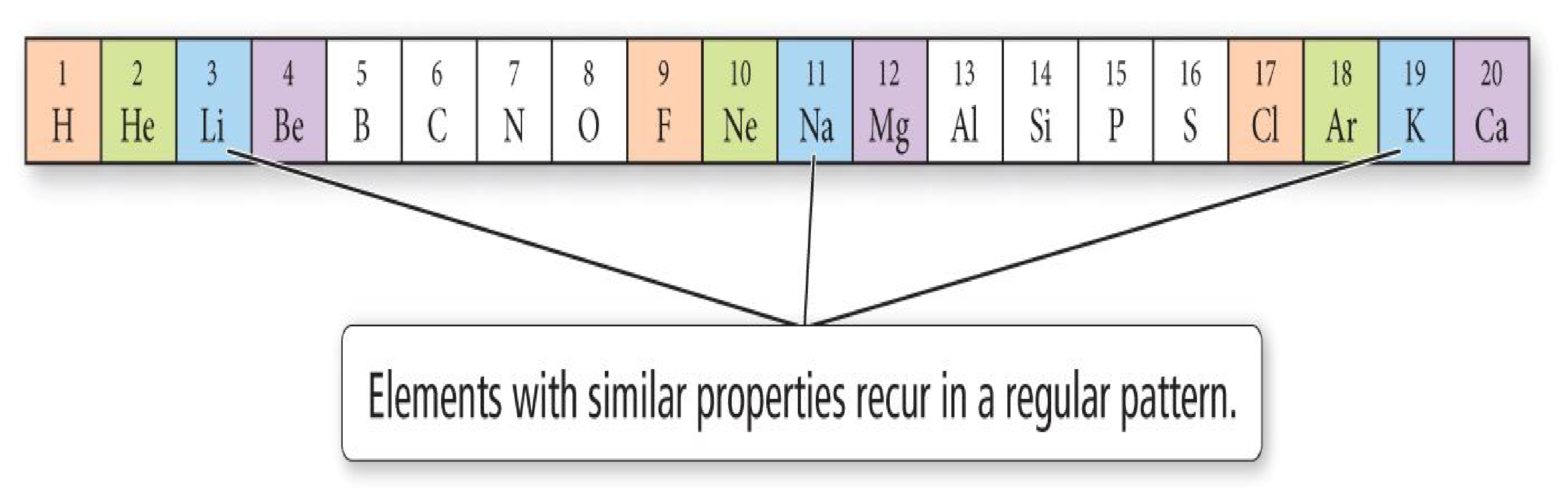
- The number of electrons in a neutral atom is equal to the number of protons in its nucleus (designated by its atomic number *Z*).

- In a chemical changes, however, atoms can lose or gain electrons and become charged particles called **ions**.
  - Positively charged ions, such as Na^+, are called **cations.**
  - Negatively charged ions, such as F^-, are called **anions**.

Finding Patterns: The Periodic Law and the Periodic Table

- In 1869, Mendeleev noticed that certain groups of elements had similar properties.
- He found that when elements are listed in order of increasing mass, these similar properties recurred in a periodic pattern.
  - To be periodic means to exhibit a repeating pattern.

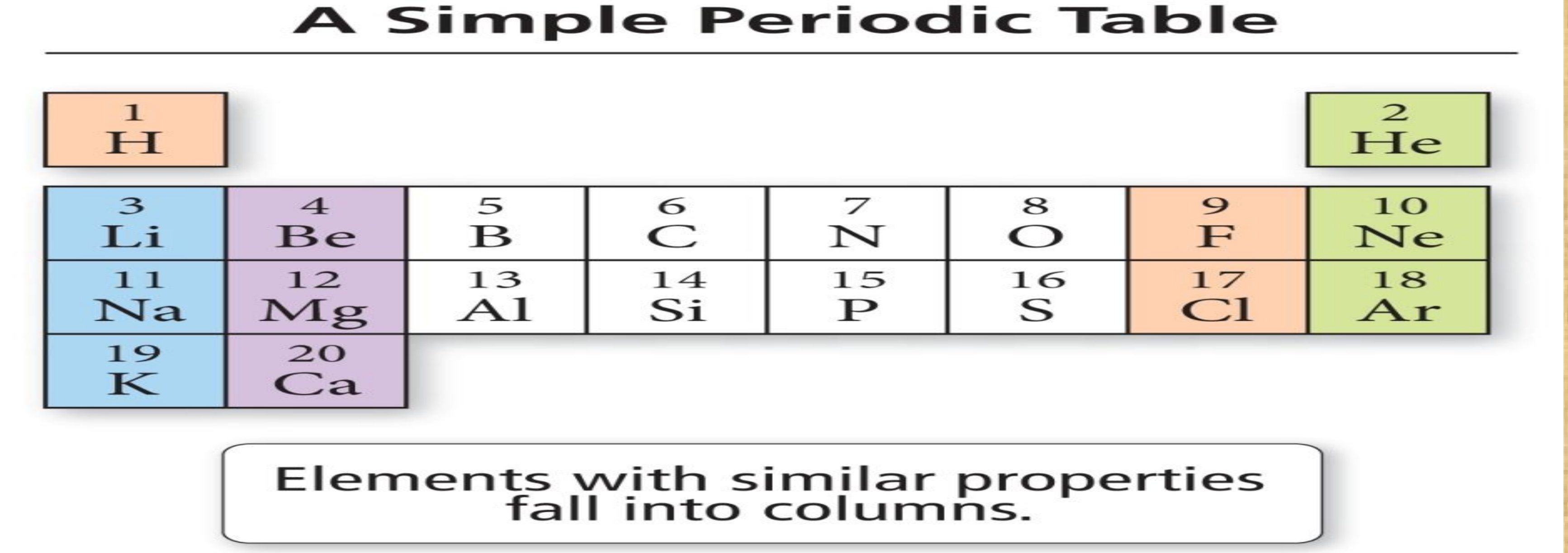
The Periodic Law



- Mendeleev summarized these observations in the **periodic law**:
  - **When the elements are arranged in order of increasing mass, certain sets of properties recur periodically.**

Periodic Table

- Mendeleev organized the known elements in a table.
- He arranged the rows so that elements with similar properties fall in the same vertical columns.



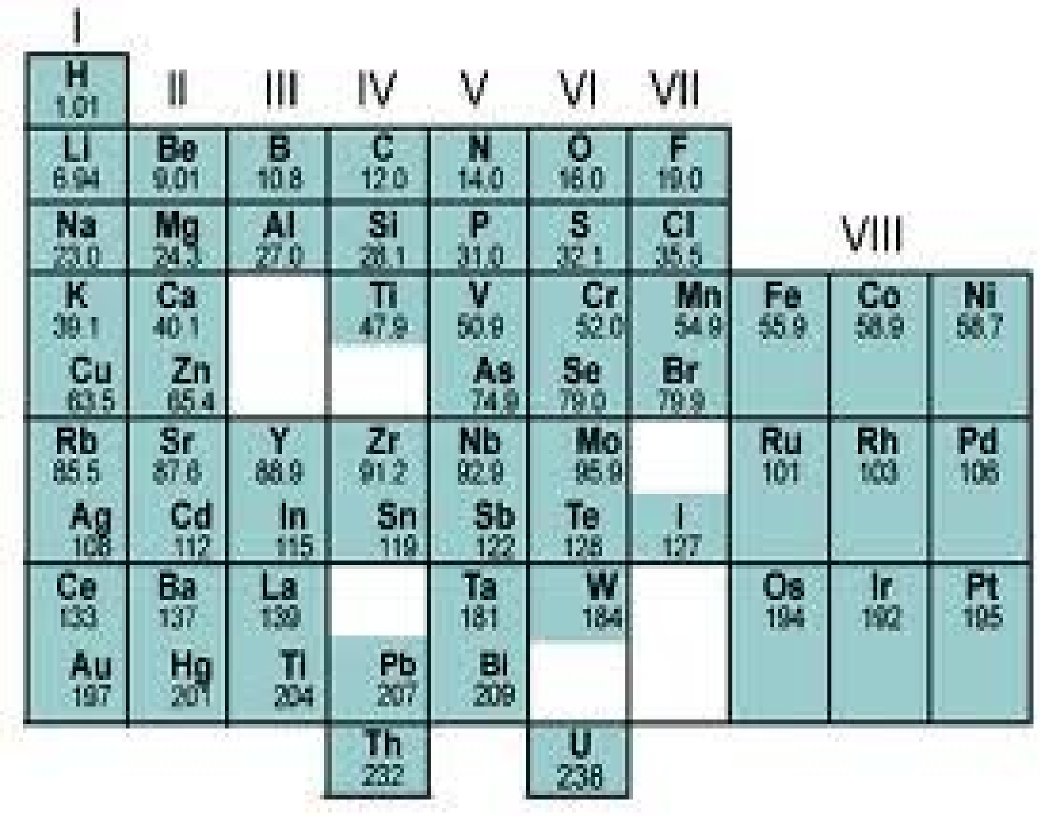
A Simple Periodic Table

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						

Elements with similar properties fall into columns.

Periodic Table

- Mendeleev's table contained **some gaps**, which allowed him to predict the existence (and even the properties) of yet undiscovered elements.
  - Mendeleev predicted the existence of an element he called **eka-silicon**.
  - In 1886, eka-silicon was discovered by German chemist Clemens Winkler (1838–1904), who named it **germanium**.

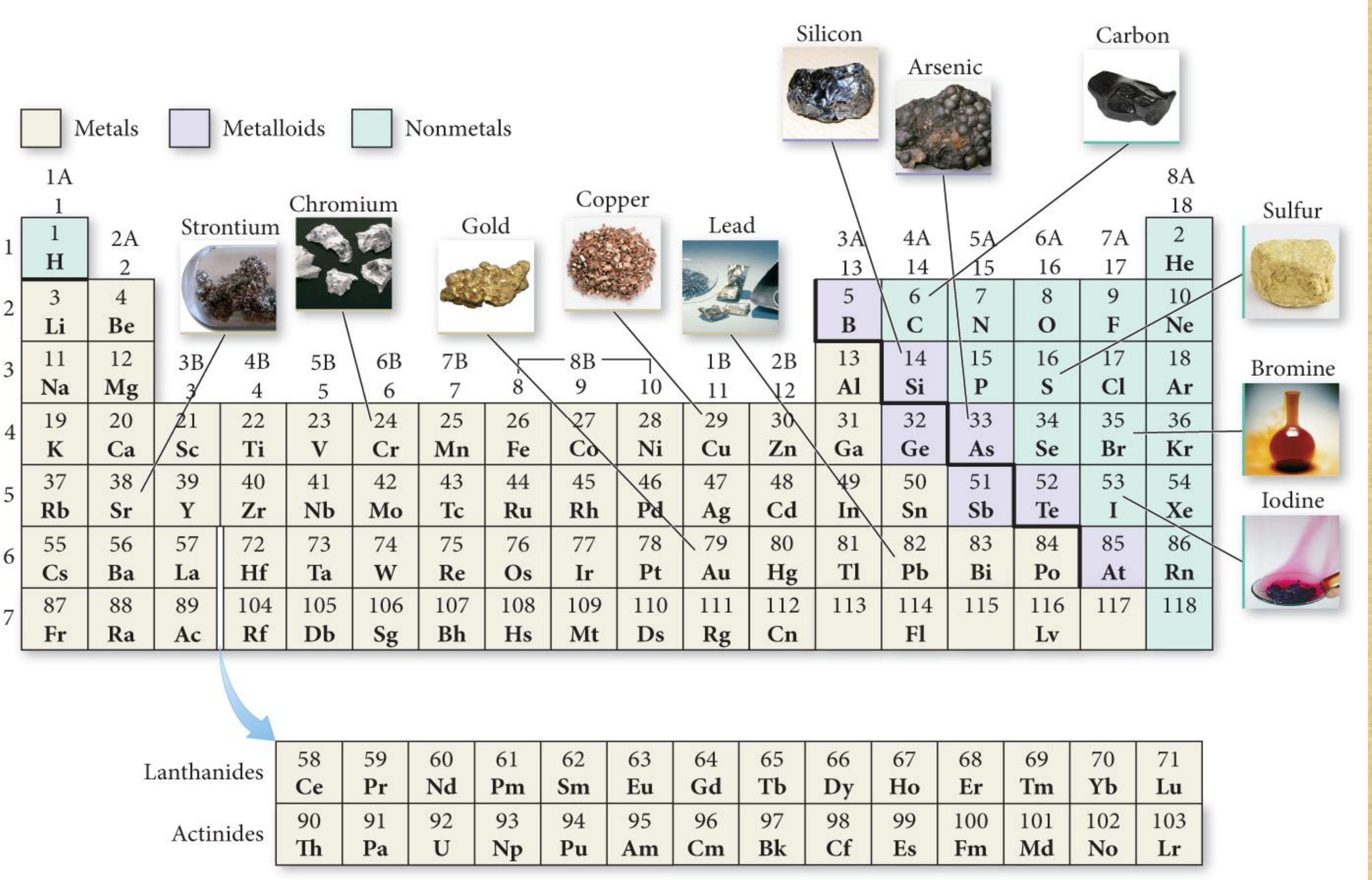


I	II	III	IV	V	VI	VII	VIII		
H 1.01									
Li 6.94	Be 9.01	B 10.8	C 12.0	N 14.0	O 16.0	F 19.0			
Na 23.0	Mg 24.3	Al 27.0	Si 28.1	P 31.0	S 32.1	Cl 35.5			
K 39.1	Ca 40.1		Ti 47.9	V 50.9	Cr 52.0	Mn 54.9	Fe 55.9	Co 58.9	Ni 58.7
Cu 63.5	Zn 65.4			As 74.9	Se 79.0	Br 79.9			
Rb 85.5	Sr 87.6	Y 88.9	Zr 91.2	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	Tl 204	Pb 207	Bi 209					
			Th 232		U 238				

Modern Periodic Table

- In the modern table, elements are listed in order of increasing atomic number rather than increasing relative mass.
- The modern periodic table also contains more elements than Mendeleev’s original table because more have been discovered since his time.

Modern Periodic Table



Major Divisions of the Periodic Table

🟨 Metals 🟪 Metalloids 🟩 Nonmetals

	1A	2A	3B	4B	5B	6B	7B		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**																	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 **Fl**	115	116 **Lv**	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

- Elements in the periodic table are classified as the following:
    - Metals
    - Nonmetals
    - Metalloids

Metals

- **Metals** lie on the lower left side and middle of the periodic table and share some common properties:
    - They are good conductors of heat and electricity.
    - 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.
  - Chromium, copper, strontium, and lead are typical metals.

Nonmetals

- **Nonmetals** lie on the upper right side of the periodic table.
- There are a total of 17 nonmetals:
  - Five are solids at room temperature (C, P, S, Se, and I)
  - One is a liquid at room temperature (Br)
  - Eleven are gases at room temperature (H, He, N, O, F, Ne, Cl, Ar, Kr, Xe, and Rn)

Nonmetals

- Nonmetals as a whole tend to
  - be poor conductors of heat and electricity.
  - be not ductile and not malleable.
  - gain electrons when they undergo chemical changes.

Oxygen, carbon, sulfur, bromine, and iodine are nonmetals.

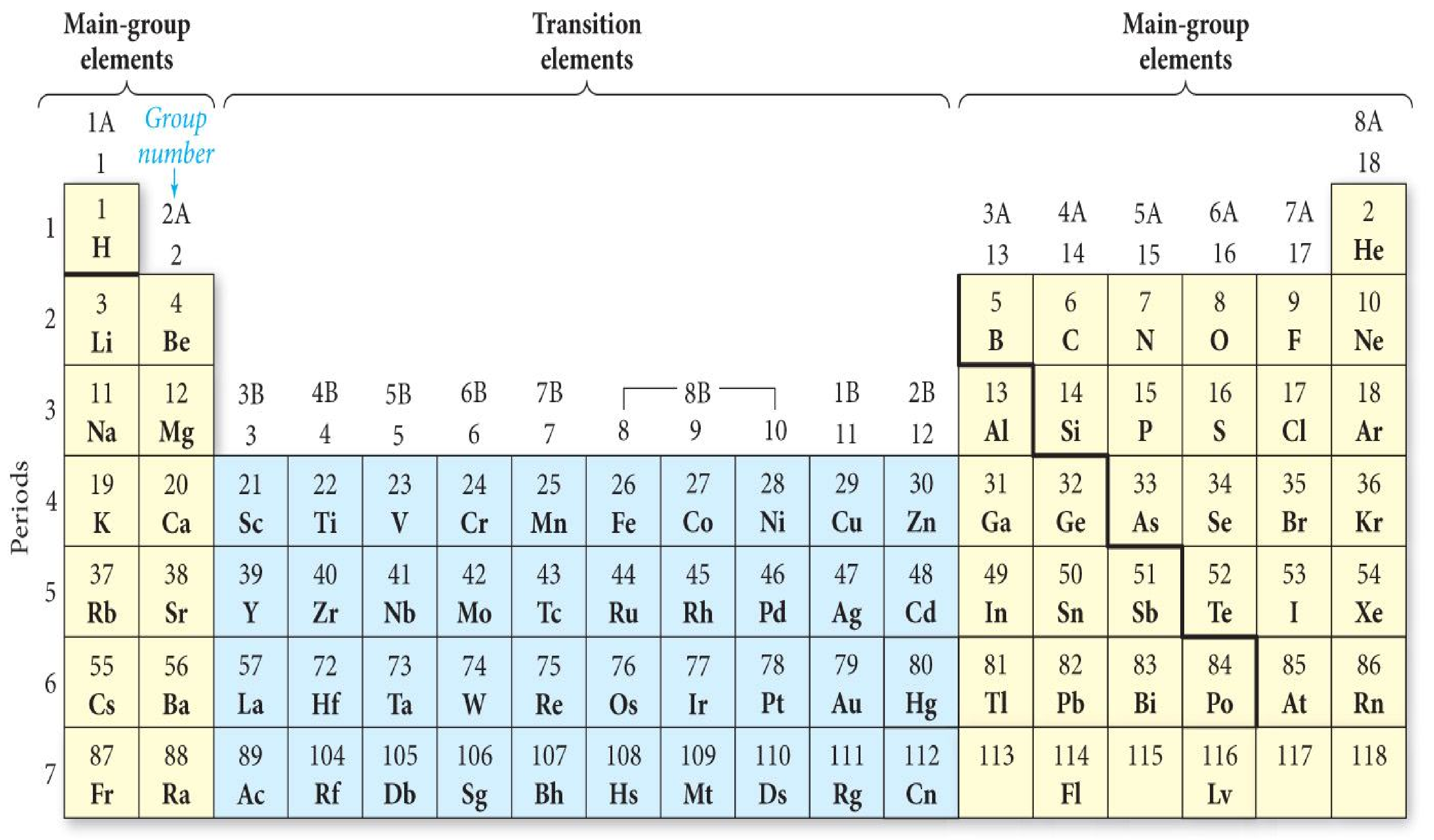
Metalloids

- Metalloids are sometimes called semimetals.
- They are elements that lie along the zigzag diagonal line that divides metals and nonmetals.
- They exhibit mixed properties.
- Several metalloids are also classified as **semiconductors** because of their intermediate (and highly temperature-dependent) electrical conductivity.

Periodic Table

- The periodic table can also be divided into
  - **main-group elements**, whose properties tend to be largely predictable based on their position in the periodic table.
  - **transition elements** or **transition metals**, whose properties tend to be less predictable based simply on their position in the periodic table.

Periodic Table



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

Group number

Periods	1A	2A	3B	4B	5B	6B	7B		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																	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 Fl	115	116 Lv	117	118

Periodic Table

- The periodic table is divided into vertical columns and horizontal rows.
  - Each vertical column is called **a group (or family).**
  - Each horizontal row is called **a period**.
- There are a total of 18 groups and 7 periods.
- The groups are numbered 1–18 (or the A and B grouping).

Periodic Table

- Main-group elements are in columns labeled with a number and the letter A (1A–8A or groups 1, 2, and 13–18).

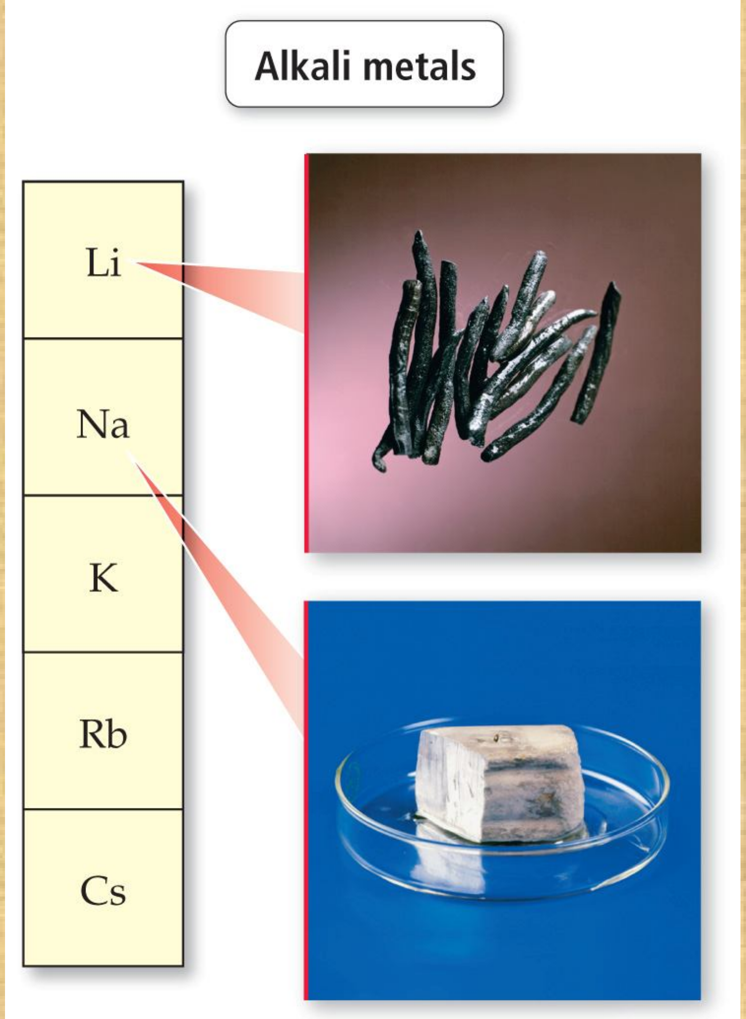
- Transition elements are in columns labeled with a number and the letter B (or groups 3–12).

Noble Gas

- The elements within a group usually have similar properties.
- The group 8A elements, called the **noble gases**, are mostly unreactive.
  - The most familiar noble gas is probably helium, used to fill buoyant balloons. Helium is chemically stable—it does not combine with other elements to form compounds—and is therefore safe to put into balloons.
  - Other noble gases are neon (often used in electronic signs), argon (a small component of our atmosphere), krypton, and xenon.

Alkali metals

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

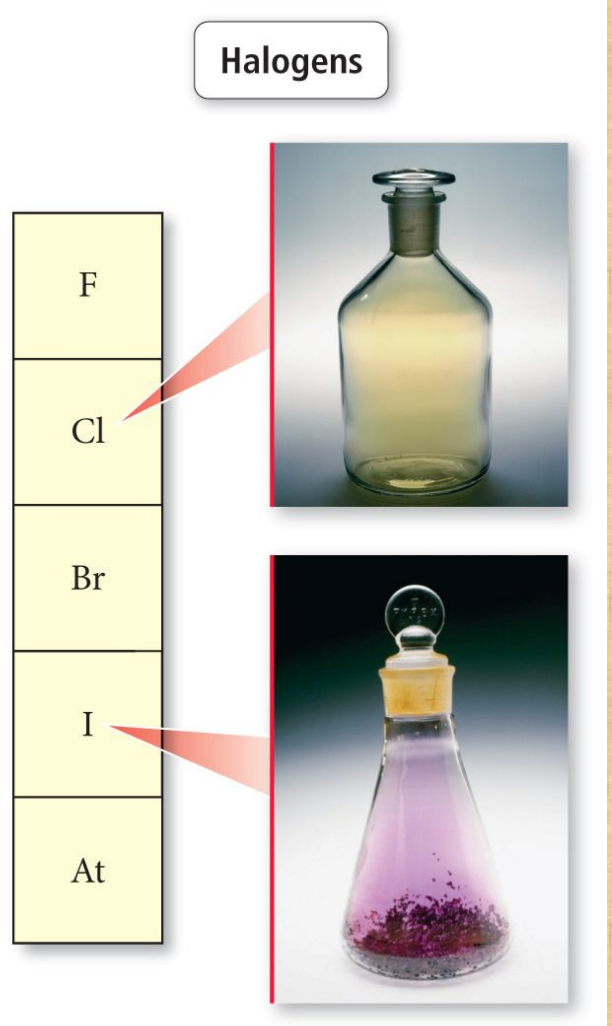


Alkaline Earth Metals

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

Halogens

- 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



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.

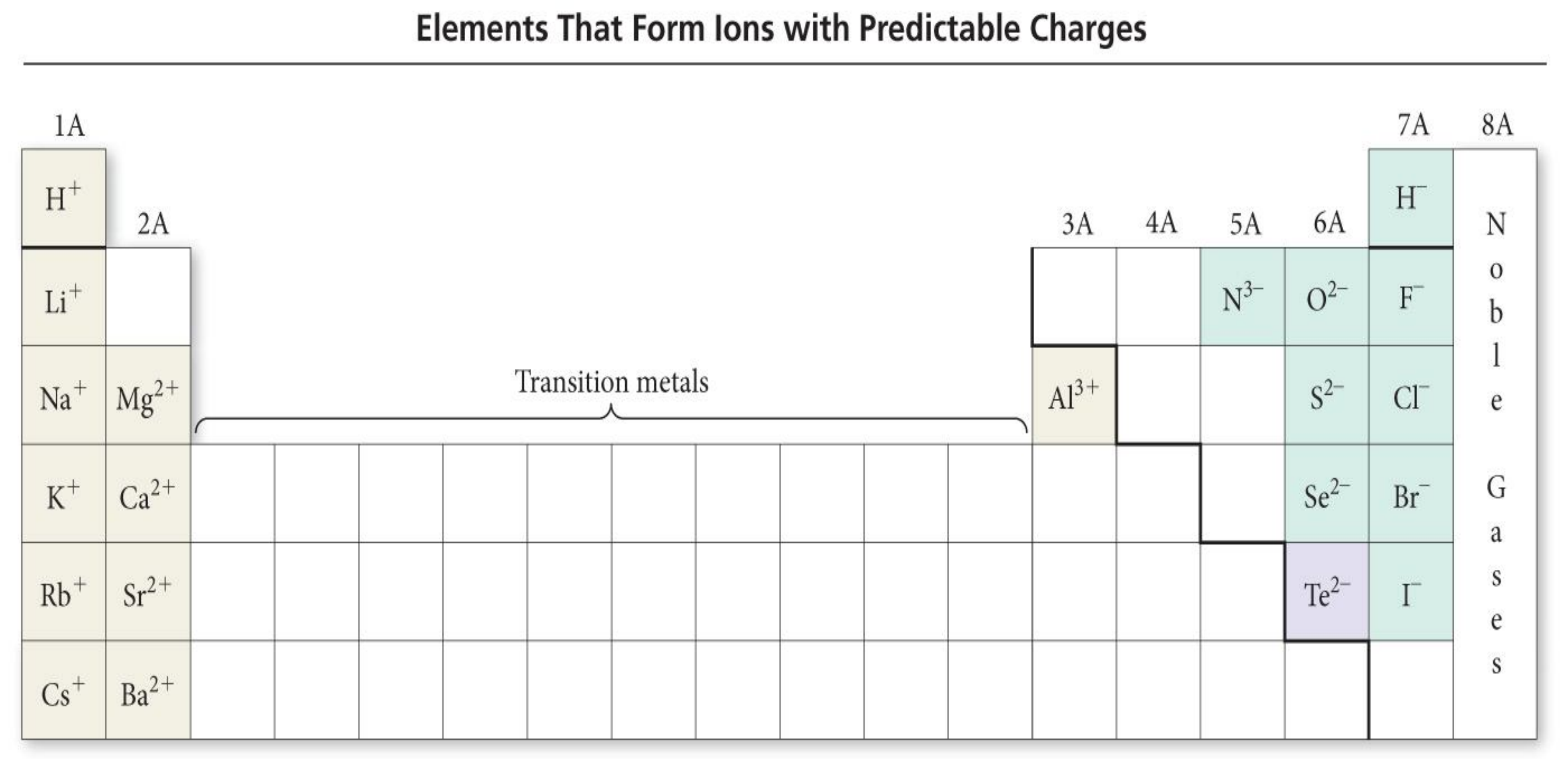
Ions and the Periodic Table

- In general, the alkali metals (group 1A) have a tendency to lose one electron and form 1+ ions.
- The alkaline earth metals (group 2A) tend to lose two electrons and form 2+ ions.
- The halogens (group 7A) tend to gain one electron and form 1– ions.
- The oxygen family nonmetals (group 6A) tend to gain two electrons and form 2– ions.

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 main-group elements that form anions with predictable charge, the charge is equal to the group number minus eight.
- Transition elements may form various different ions with different charges.

Ions and the Periodic Table



Elements That Form Ions with Predictable Charges

1A	2A	Transition metals										3A	4A	5A	6A	7A	8A
H^+																H^-	Noble Gases
Li^+														N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}											Al^{3+}			S^{2-}	Cl^-	
K^+	Ca^{2+}														Se^{2-}	Br^-	
Rb^+	Sr^{2+}														Te^{2-}	I^-	
Cs^+	Ba^{2+}																

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

- Atomic mass is sometimes called *atomic weight* or *standard atomic weight.*
- The atomic mass of each element is directly beneath the element's symbol in the periodic table.
- It represents the average mass of the isotopes that compose that element, *weighted according to the natural abundance of each isotope*.

Atomic Mass

- Naturally occurring chlorine consists of 75.77% chlorine-35 atoms (mass 34.97 amu) and 24.23% chlorine-37 atoms (mass 36.97 amu). We can calculate its atomic mass:

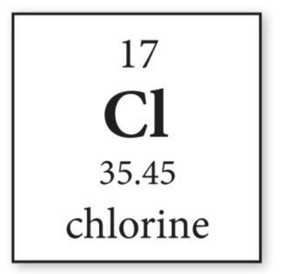
- Solution:
  - Convert the percent abundance to decimal form and multiply it with its isotopic mass:

    Cl-37 = 0.2423(36.97 amu) = 8.9578 amu

    Cl-35 = 0.7577(34.97 amu) = 26.4968 amu

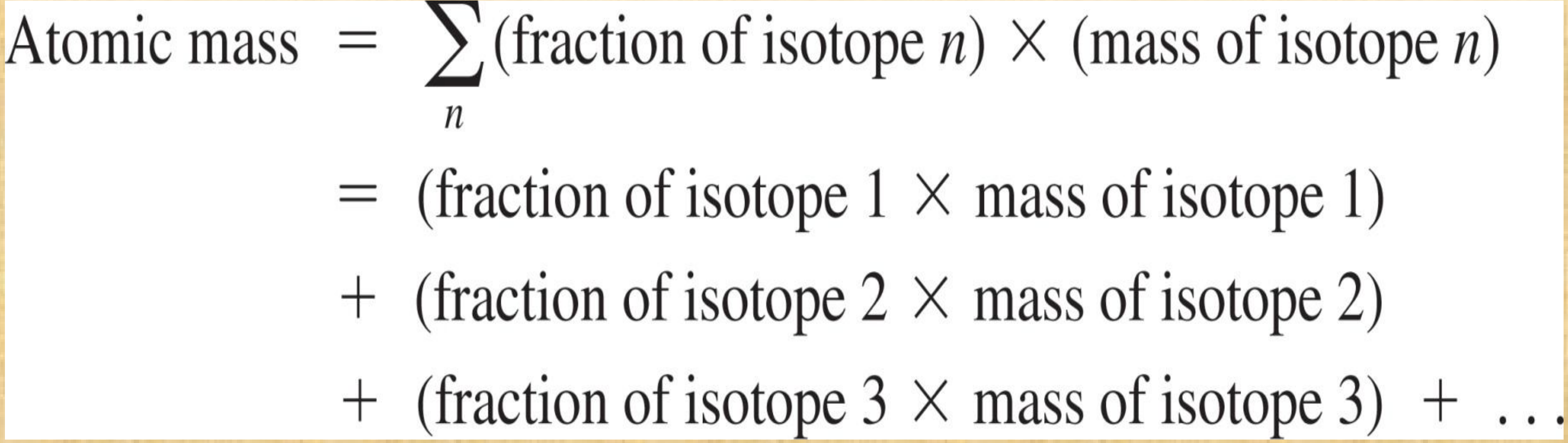
    Atomic Mass Cl = 8.9578 + 26.4968 = 35.45 amu

Atomic Mass



17
Cl
35.45
chlorine

- In general, we calculate the atomic mass with the equation:



$$\begin{aligned}\text{Atomic mass} &= \sum_{n} (\text{fraction of isotope } n) \times (\text{mass of isotope } n)\\ &= (\text{fraction of isotope 1} \times \text{mass of isotope 1})\\ &+ (\text{fraction of isotope 2} \times \text{mass of isotope 2})\\ &+ (\text{fraction of isotope 3} \times \text{mass of isotope 3}) + \ldots\end{aligned}$$

Molar Mass: Counting Atoms by Weighing Them

- As chemists, we often need to know the number of atoms in a sample of a given mass. Why? *Because chemical processes happen between particles*.

- Therefore, if we want to know the number of atoms in anything of ordinary size, we count them by weighing.

The Mole: A Chemist’s “Dozen”

- When we count large numbers of objects, we often use units such as
  - 1 dozen objects = 12 objects.
  - 1 gross objects = 144 objects.
- The chemist’s “dozen” is the **mole** (abbreviated mol). A mole is the measure of material containing 6.02214 × 10^{23} particles:

  1 mole = 6.02214 × 10^{23} particles
- This number is **Avogadro’s number.**

The Mole

- First thing to understand about the mole is that it can specify Avogadro’s number of anything.

- For example, 1 mol of marbles corresponds to 6.02214×10^{23} marbles.
- 1 mol of sand grains corresponds to 6.02214×10^{23} sand grains.

- *One mole of anything is* 6.02214×10^{23} *units of that thing*.

The Mole

- The second, and more fundamental, thing to understand about the mole is how it gets its specific value.
- **The value of the mole is equal to the number of atoms in exactly 12 grams of pure C-12.**
- **12 g C = 1 mol C atoms = 6.022×10^{23} C atoms**

Converting between Number of Moles and Number of Atoms

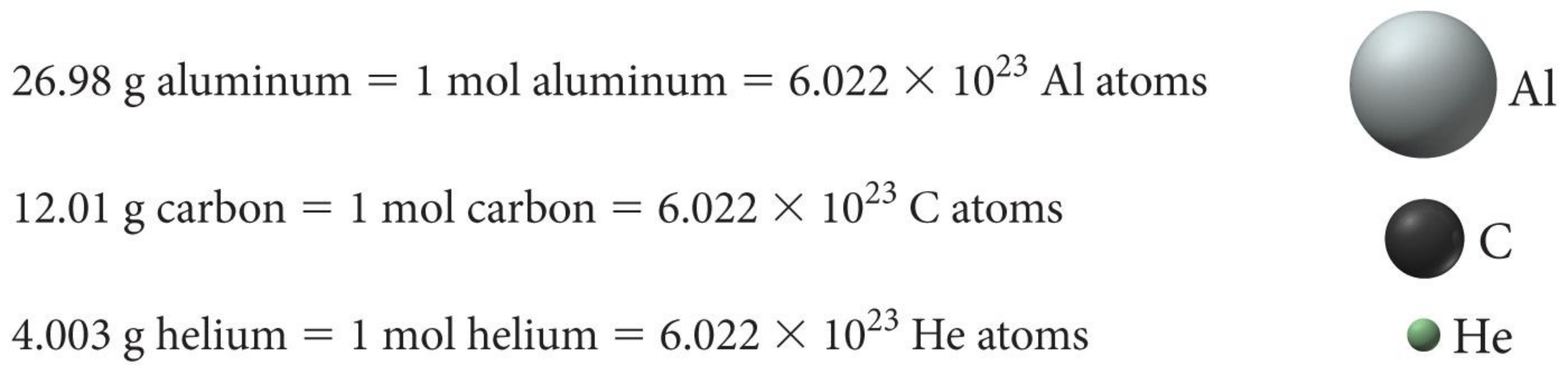
- Converting between number of moles and number of atoms is similar to converting between dozens of eggs and number of eggs.
- For atoms, you use the conversion factor 1 mol atoms = 6.022×10^{23} atoms.
- The conversion factors take the following forms:

$$\frac{1 \text{ mol atoms}}{6.022 \times 10^{23} \text{ atoms}} \quad \text{or} \quad \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol atoms}}$$

Converting between Mass and Amount (Number of Moles)

- To count atoms by weighing them, we need one other conversion factor—the mass of
  1 mol of atoms.
- The mass of 1 mol of atoms of an element is the **molar mass**.
- **An element's molar mass in grams per mole is numerically equal to the element's atomic mass in atomic mass units (amu).**

Converting between Mass and Moles



26.98 g aluminum = 1 mol aluminum = 6.022×10^{23} Al atoms

12.01 g carbon = 1 mol carbon = 6.022×10^{23} C atoms

4.003 g helium = 1 mol helium = 6.022×10^{23} He atoms

- The lighter the atom, the less mass in 1 mol of atoms.

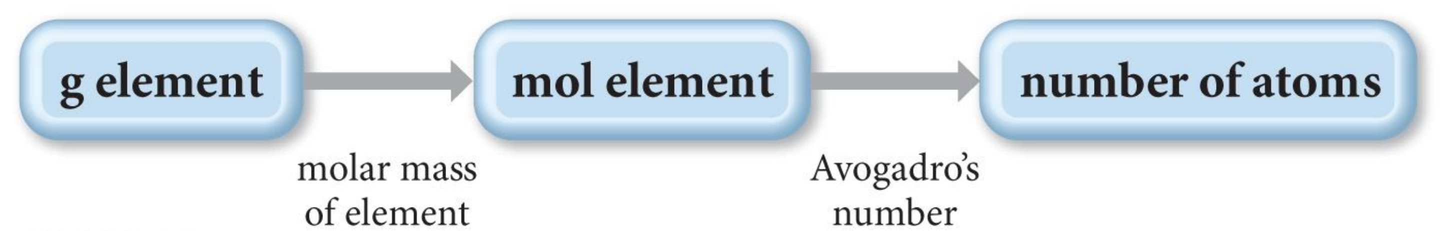
Converting between Mass and Moles

- The molar mass of any element is the conversion factor between the mass (in grams) of that element and the amount (in moles) of that element. For carbon,

$$12.01\ \text{g C} = 1\ \text{mol C} \ \text{or}\ \frac{12.01\ \text{g C}}{\text{mol C}}\ \text{or}\ \frac{1\ \text{mol C}}{12.01\ \text{g C}}$$

Conceptual Plan

- We now have all the tools to count the number of atoms in a sample of an element by weighing it.
  - First, we obtain the mass of the sample.
  - Then, we convert it to the amount in moles using the element's molar mass.
  - Finally, we convert it to the number of atoms using Avogadro's number.
- The conceptual plan for these kinds of calculations takes the following form:

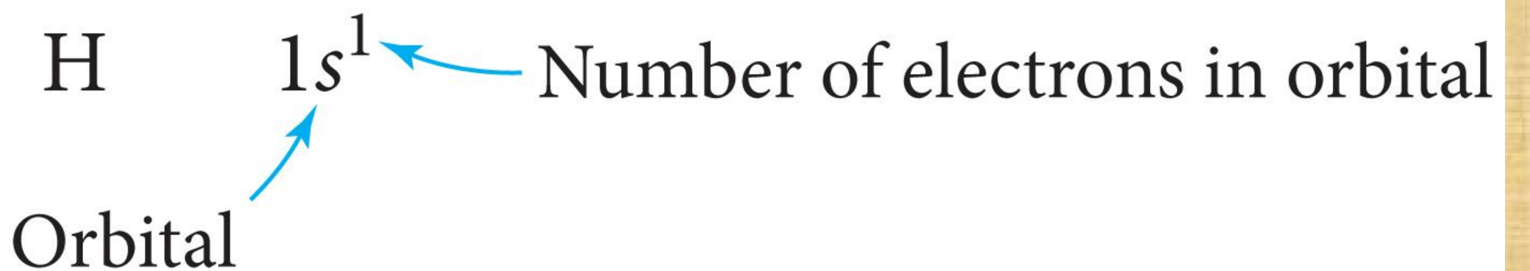


Electron Configurations

- Quantum-mechanical theory describes the behavior of electrons in atoms.

- The electrons in atoms exist in orbitals.

- A description of the orbitals occupied by electrons is called an **electron configuration**.



Principal Quantum Number, *n*

- The principal quantum number, *n*, describes the energy level on which the orbital resides.
- The values of *n* are integers ≥ 0.

Azimuthal 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*

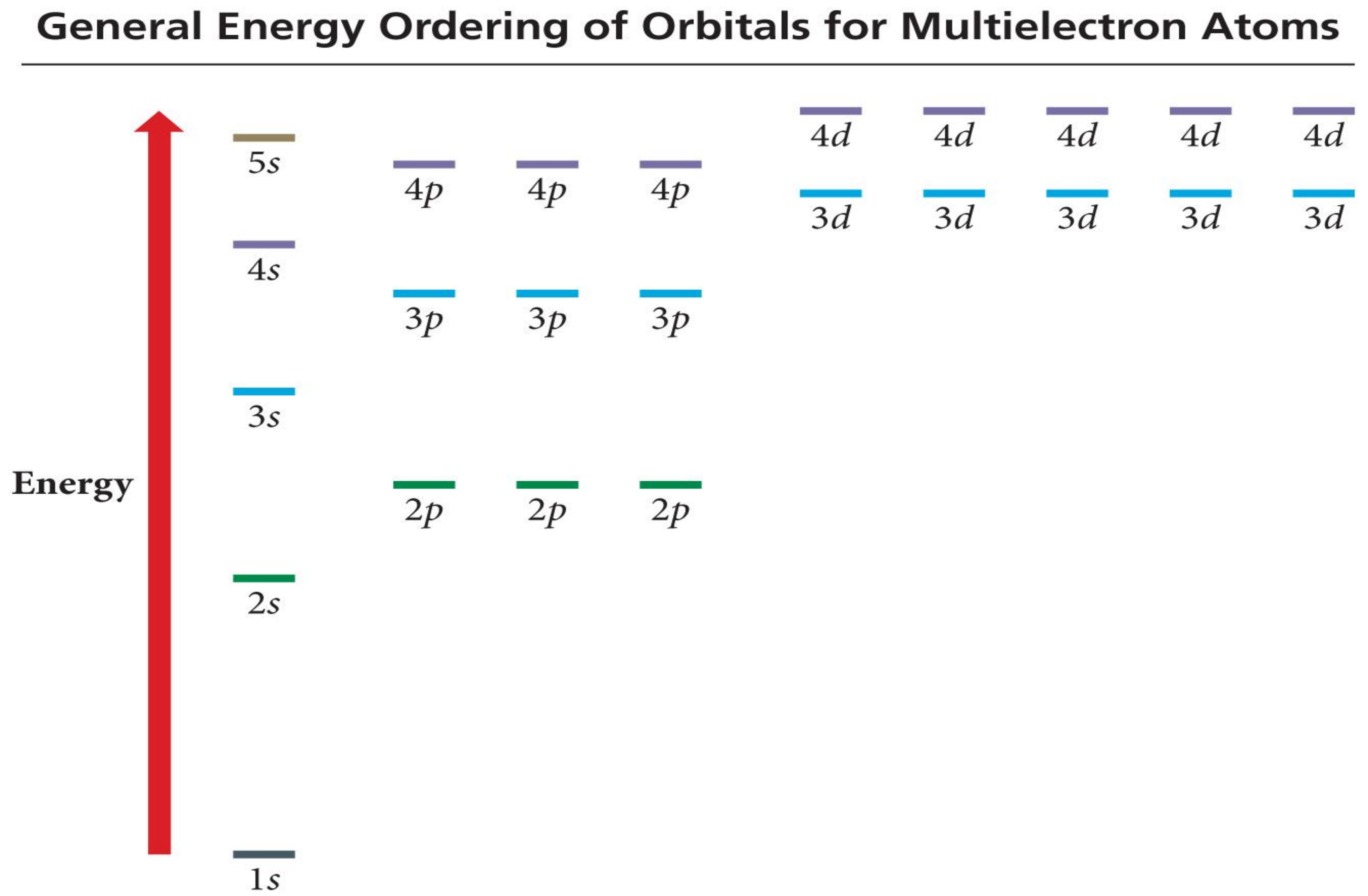
Magnetic Quantum Number, m_l

- Describes the three-dimensional orientation of the orbital.
- Values are integers ranging from -*l* to *l*:

$$-l \leq m_l \leq l.$$

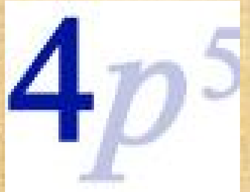
- Therefore, on any given energy level, there can be up to 1 *s* orbital, 3 *p* orbitals, 5 *d* orbitals, 7 *f* orbitals, etc.

Energies of Orbitals



General Energy Ordering of Orbitals for Multielectron Atoms

Electron Configurations



- Distribution of all electrons in an atom
- Consist of
  - Number denoting the energy level

Electron Configurations

$$4p^5$$

- Distribution of all electrons in an atom
- Consist of
  - Number denoting the energy level
  - Letter denoting the type of orbital

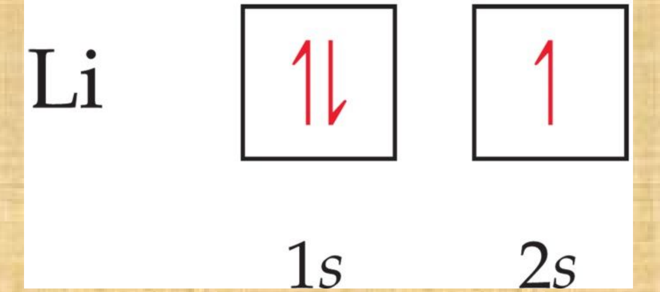
Electron Configurations

$4p^5$

- Distribution of all electrons in an atom.
- Consist of
  - Number denoting the energy level.
  - Letter denoting the type of orbital.
  - Superscript denoting the number of electrons in those orbitals.

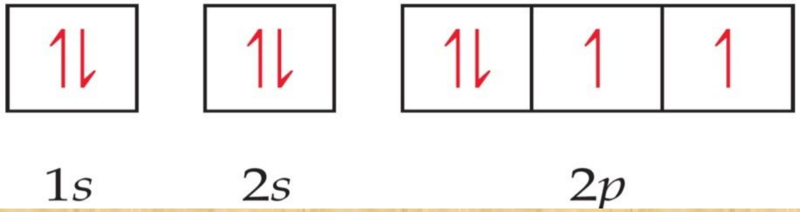
Orbital Diagrams

- Each box represents one orbital.
- Half-arrows represent the electrons.
- The direction of the arrow represents the spin of the electron.

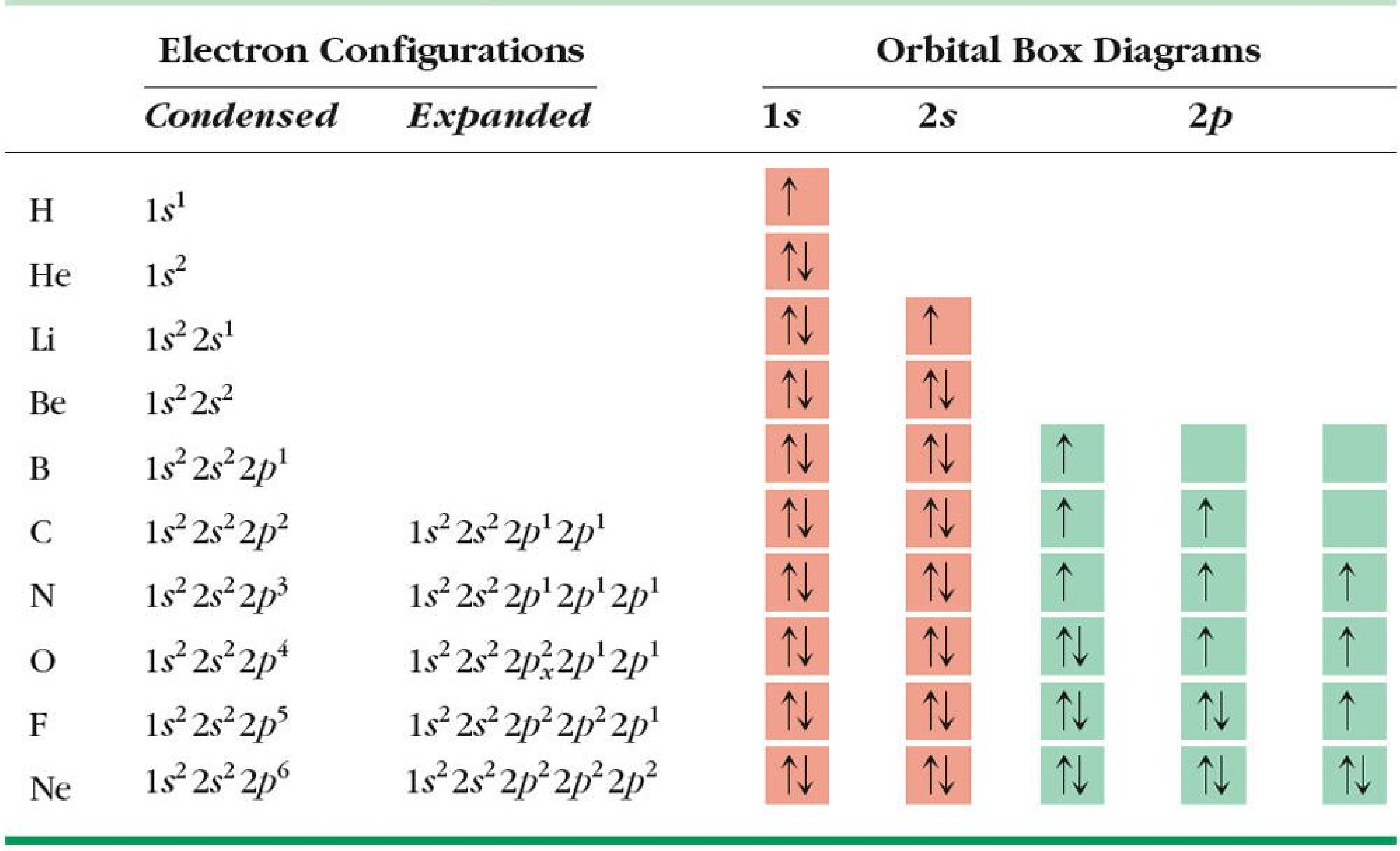


Filling the Orbitals with Electrons

- Energy levels and sublevels fill from lowest energy to high:
  - ✓ $s \rightarrow p \rightarrow d \rightarrow f$
  - ✓ **Aufbau principle**
- Orbitals that are in the same sublevel have the same energy.
- No more than two electrons per orbital.
  - ✓ **Pauli exclusion principle**
- When filling orbitals that have the same energy, place one electron in each before completing pairs.
  - ✓ **Hund's rule**



- **The electron configurations of the first ten elements illustrate this point.**

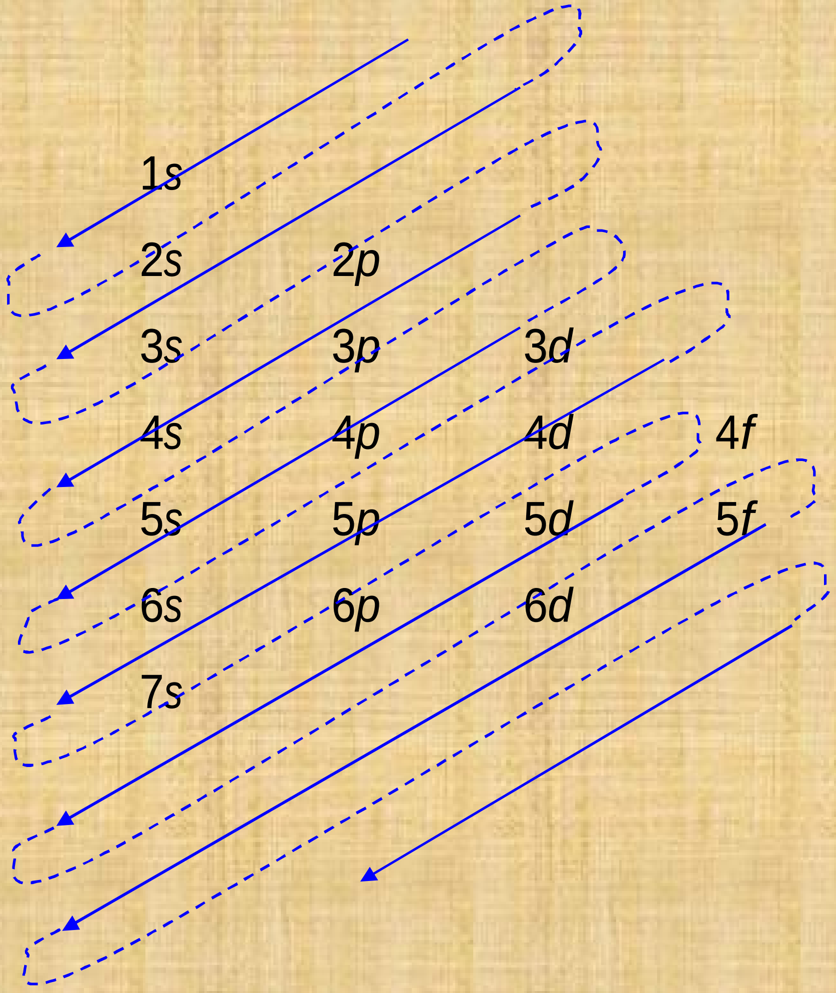


	Electron Configurations		Orbital Box Diagrams				
	Condensed	*Expanded*	1*s*	2*s*	2*p*		
H	$1s^1$		↑				
He	$1s^2$		↑↓				
Li	$1s^2 2s^1$		↑↓	↑			
Be	$1s^2 2s^2$		↑↓	↑↓			
B	$1s^2 2s^2 2p^1$		↑↓	↑↓	↑		
C	$1s^2 2s^2 2p^2$	$1s^2 2s^2 2p^1 2p^1$	↑↓	↑↓	↑	↑	
N	$1s^2 2s^2 2p^3$	$1s^2 2s^2 2p^1 2p^1 2p^1$	↑↓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

Order of Sublevel Filling in Ground State Electron Configurations

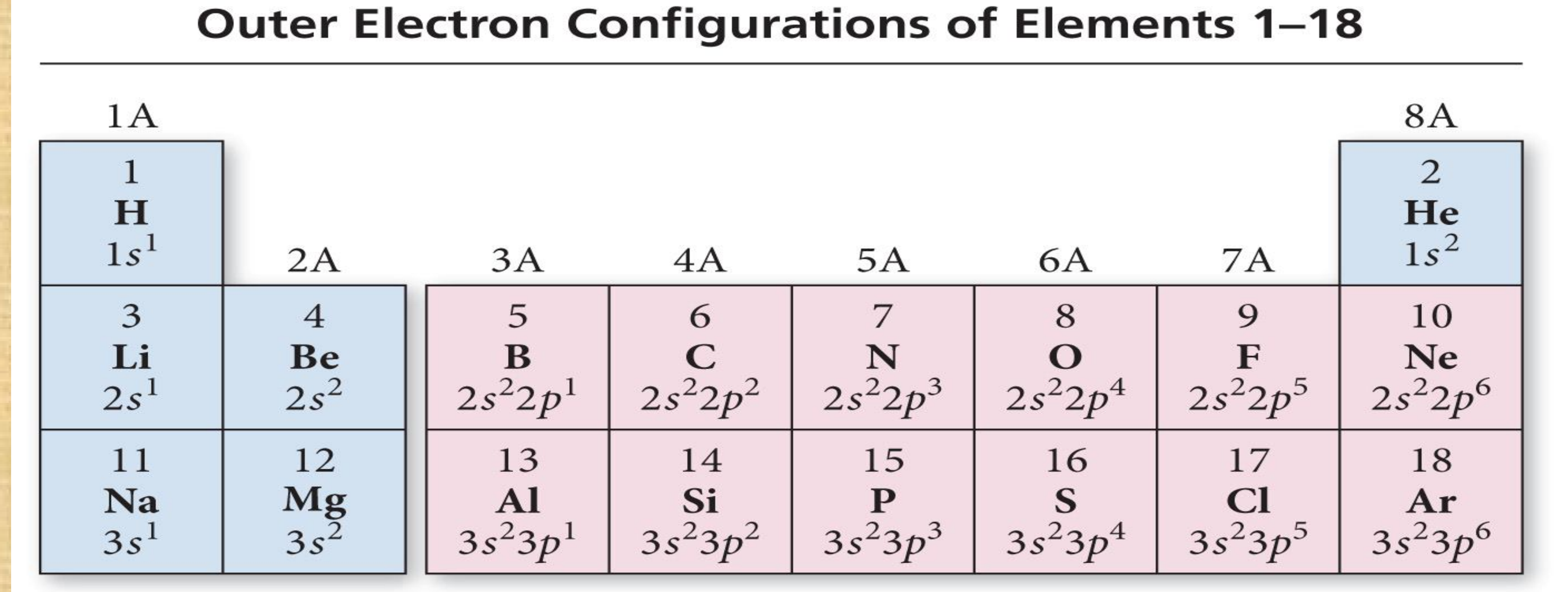
Start by drawing a diagram, putting each energy shell on a row and listing the sublevels (*s, p, d, f*) for that shell in order of energy (from left to right).

Next, draw arrows through the diagonals, looping back to the next diagonal each time.



Valence Electrons

- The electrons in all the sublevels with the highest principal energy shell are called the **valence electrons**.
- Electrons in lower energy shells are called **core electrons**.
- One of the most important factors in the way an atom behaves, both chemically and physically, is the number of valence electrons.

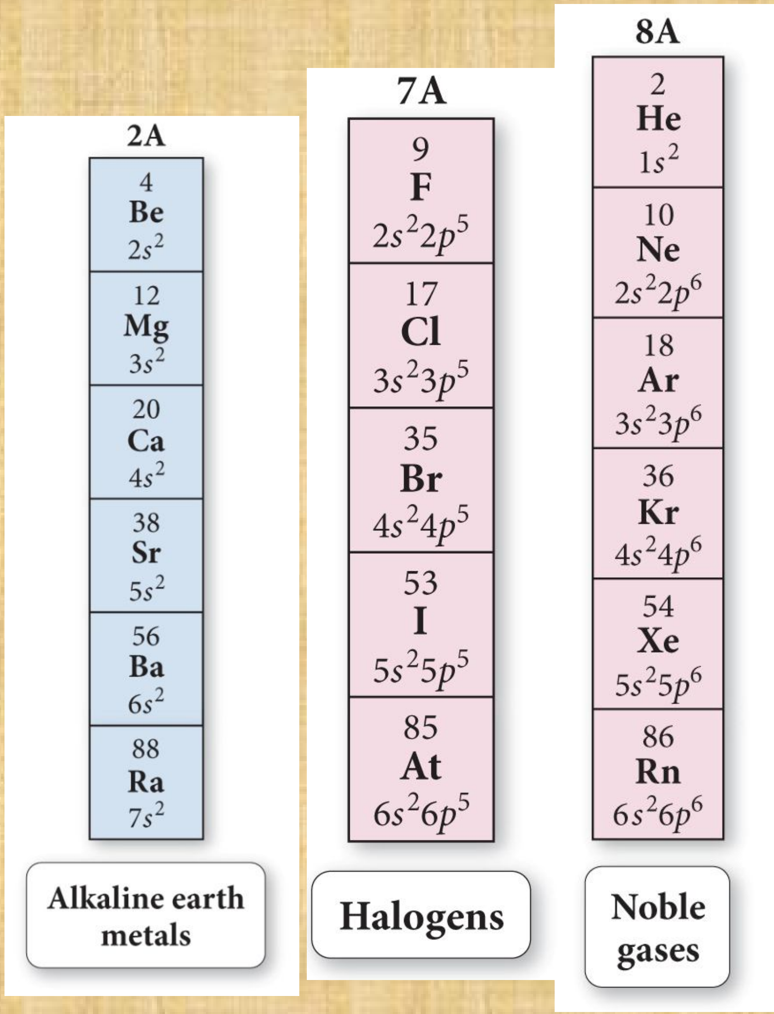


Outer Electron Configurations of Elements 1–18

1A	2A	3A	4A	5A	6A	7A	8A
1 **H** $1s^1$							2 **He** $1s^2$
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** $2s^22p^6$
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** $3s^23p^6$

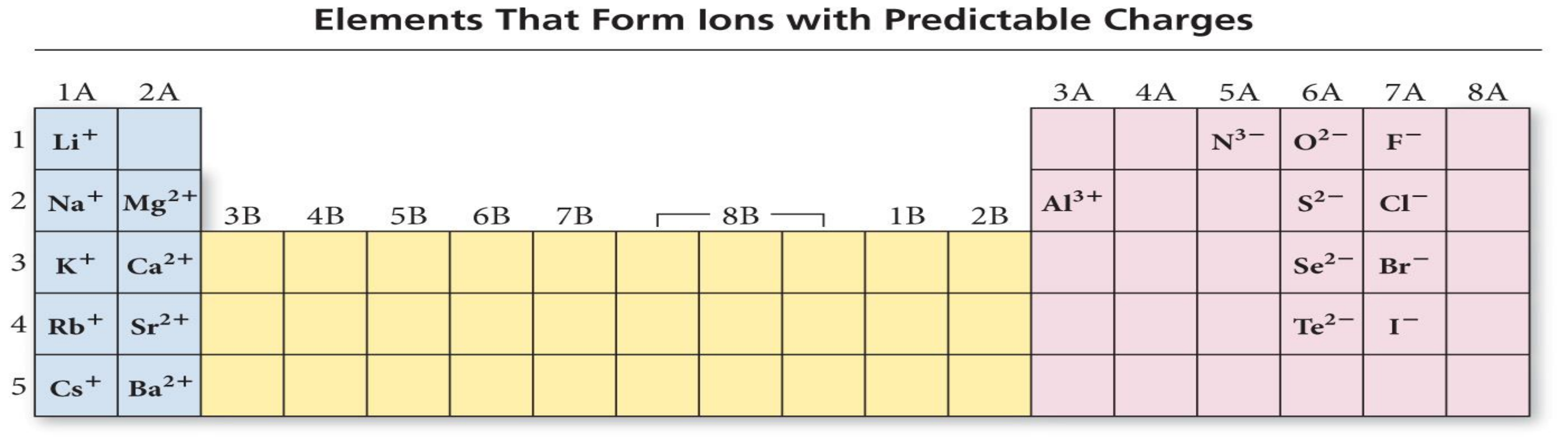
Properties and Electron Configuration

- The properties of the elements follow a periodic pattern.
  - Elements in the same column have similar properties.
  - The elements in a period show a pattern that repeats.
- The quantum-mechanical model explains this because the number of valence electrons and the types of orbitals they occupy are also periodic.



Electron Configuration and Ion Charge

- We have seen that many metals and nonmetals form one ion, and that the charge on that ion is predictable based on its position on the periodic table.
  - Group 1A = 1+, group 2A = 2+, group 7A = 1–, group 6A = 2–, etc.
- These atoms form ions that will result in an electron configuration that is the same as the nearest noble gas.



Elements That Form Ions with Predictable Charges

	1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
1	Li^+														N^{3-}	O^{2-}	F^-	
2	Na^+	Mg^{2+}											Al^{3+}			S^{2-}	Cl^-	
3	K^+	Ca^{2+}														Se^{2-}	Br^-	
4	Rb^+	Sr^{2+}														Te^{2-}	I^-	
5	Cs^+	Ba^{2+}																

- The sulfur atom has six valence electrons.

  S atom = $1s^22s^22p^63s^23p^4$

- To have eight valence electrons, sulfur must gain two more.

  S^{2-} anion = $1s^22s^22p^63s^23p^6$

- The magnesium atom has two valence electrons.

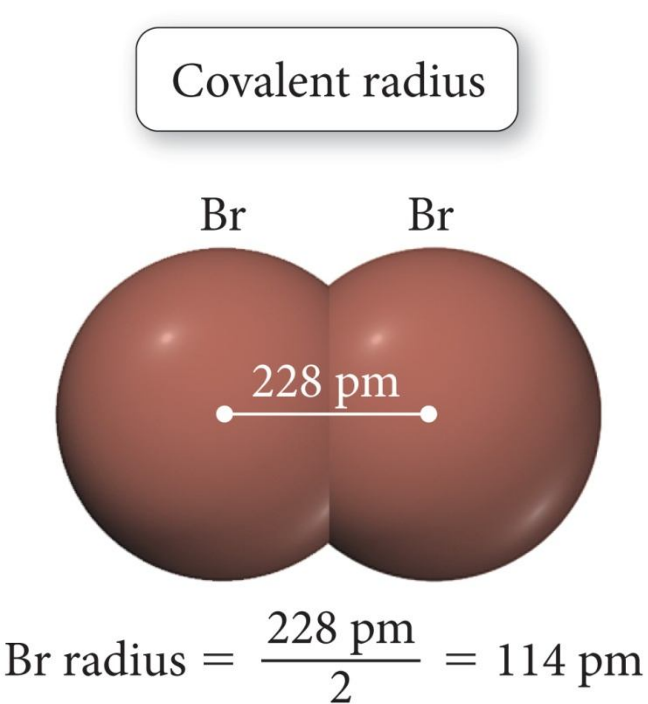
  Mg atom = $1s^22s^22p^63s^2$

- When magnesium forms a cation, it loses its valence electrons.

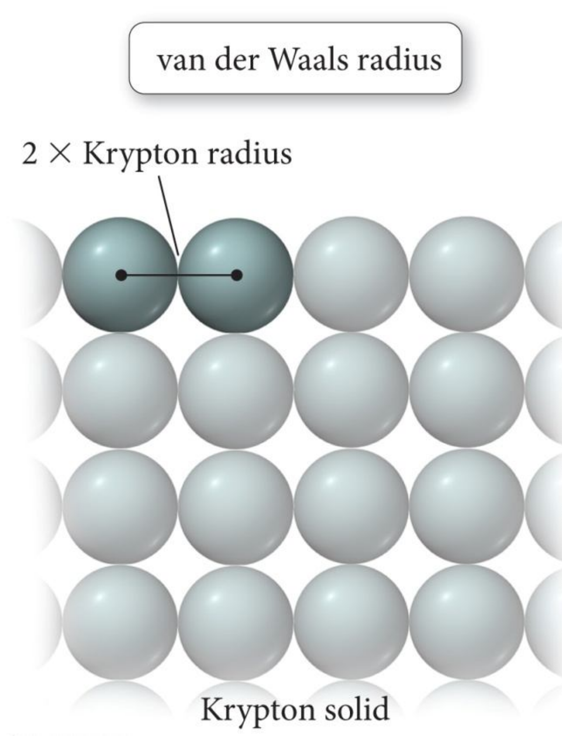
  Mg^{2+} cation = $1s^22s^22p^6$

Trend in Atomic Radius – Main Group

- 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
  - ✓ Atomic radius is an average radius of an atom based on measuring large numbers of elements and compounds.

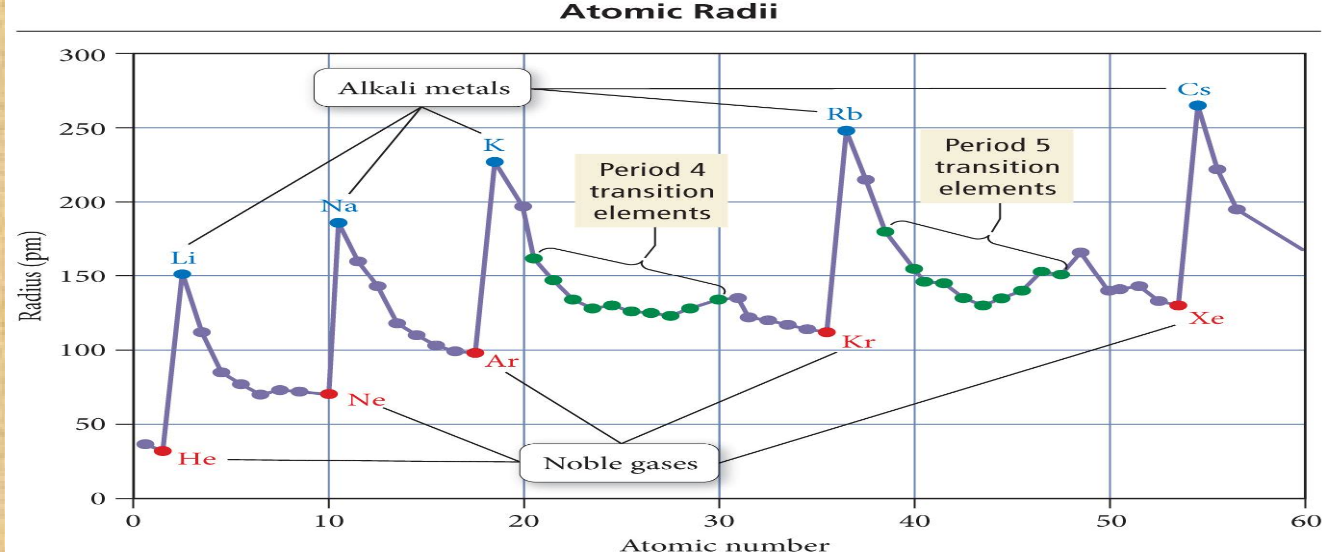


$$\text{Br radius} = \frac{228\ \text{pm}}{2} = 114\ \text{pm}$$



Trend in Atomic Radius – Main Group

- 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



Effective Nuclear Charge

- The **effective nuclear charge** is a net positive charge that is attracting a particular electron.
- **Z** is the nuclear charge, and **S** is the number of electrons in lower energy levels.
  - – Electrons in the same energy level contribute to screening but since their contribution is so small they are not part of the calculation.
  - – Trend is *s* > *p* > *d* > *f*.

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

Trends in Ionic Radius

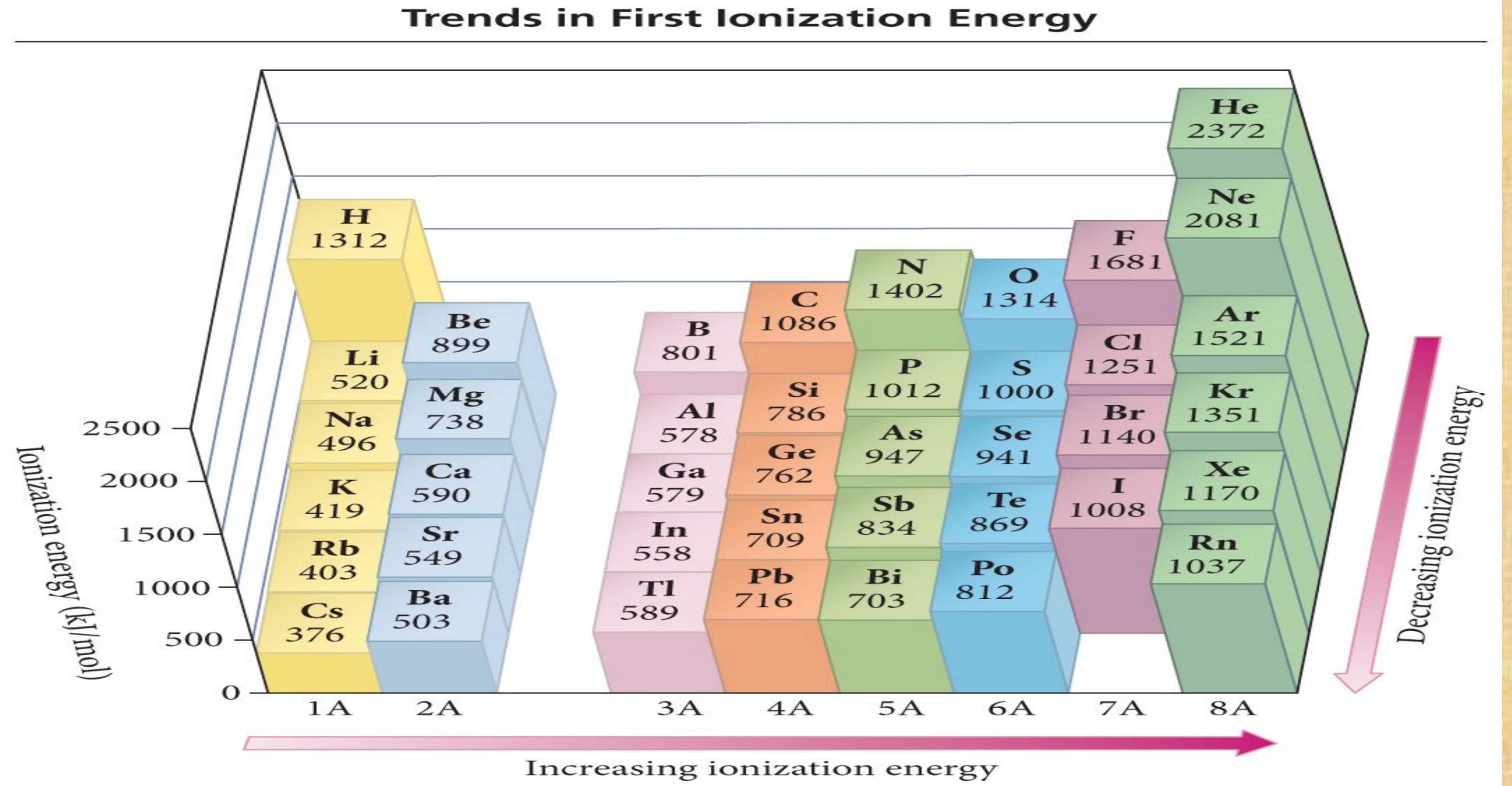
- Ions in the same group have the same charge.
- Ion size increases down the column.
  - ✓ Higher valence shell, larger
- Cations are smaller than neutral atoms; anions are larger than neutral atoms.
- Cations are smaller than anions.
  - ✓ Except Rb^+ and Cs^+ bigger or same size as F^- and O^{2-}.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

Ionization Energy (IE)

- Minimum energy needed to remove an electron from an atom or ion
  - ✓ Gas state
  - ✓ 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 = energy to remove electron from neutral atom, second IE = energy to remove from 1+ ion, etc.

General Trends in First Ionization Energy

- The larger the effective nuclear charge on the electron, the more energy it takes to remove it.
- The farther the most probable distance the electron is from the nucleus, the less energy it takes to remove it.
- First IE **decreases** down the group.
  - Valence electron farther from nucleus
- First IE generally **increases** across the period.
  - Effective nuclear charge increases



Trends in First Ionization Energy

Electron Affinity

- Energy is released when an neutral atom gains an electron.
  - ✓ Gas state
  - ✓ $M(g) + 1e^{-} \rightarrow M^{1-}(g) + EA$
- Electron affinity is defined as exothermic (−), but may actually be endothermic (+).
  - ✓ Some alkali earth metals and all noble gases are endothermic. Why?
- The more energy that is released, the larger the electron affinity.
  - ✓ The more negative the number, the larger the EA.

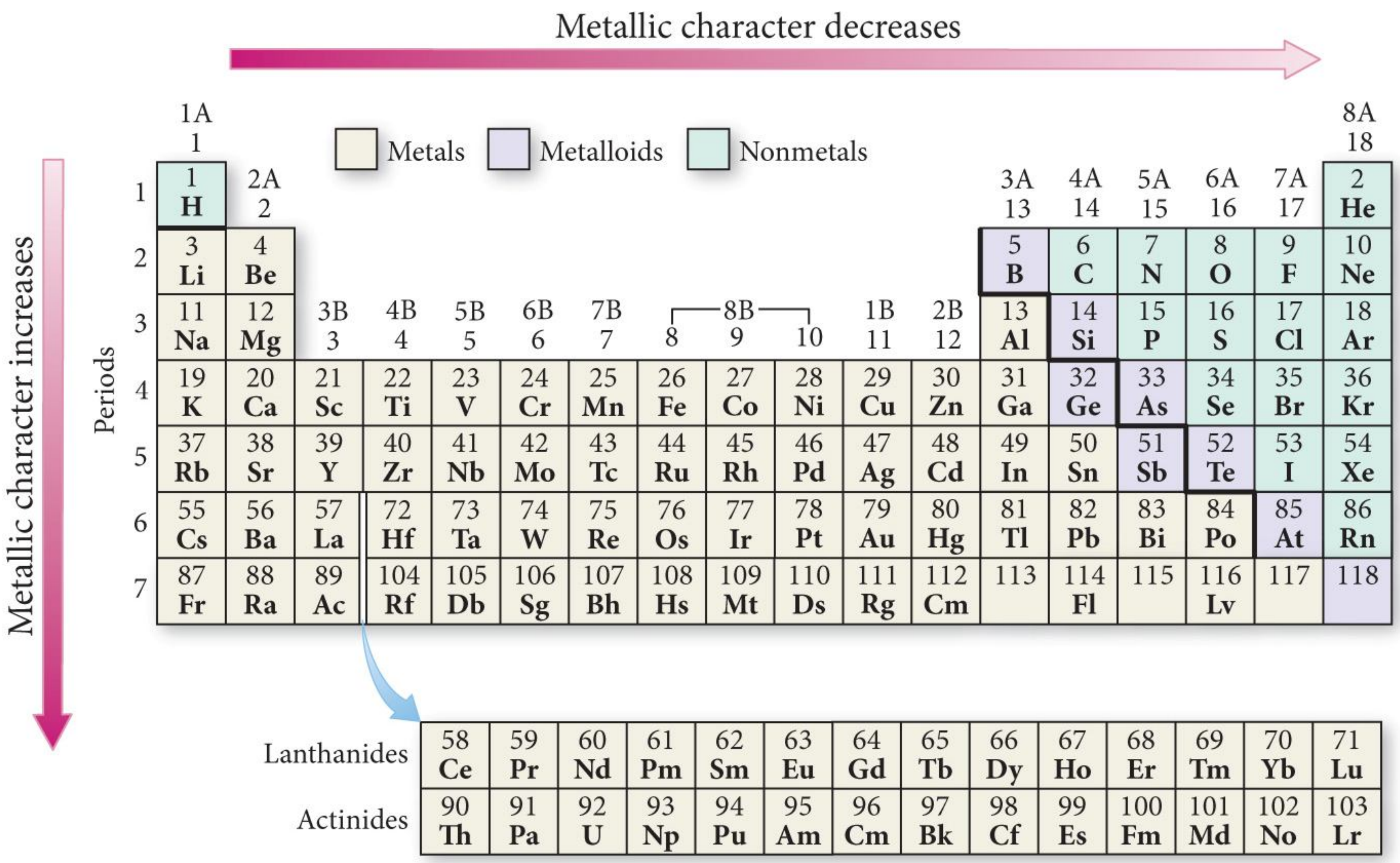
Trends in Electron Affinity

- Alkali metals decrease electron affinity down the column.
  - But not all groups do
  - Generally irregular increase in EA from second period to third period
- “Generally” increases across period
  - Becomes more negative from left to right
  - Not absolute
  - Group 5A generally lower EA than expected because extra electron must pair
  - Groups 2A and 8A generally very low EA because added electron goes into higher energy level or sublevel
- Highest EA in any period = halogen

Metallic Character

- **Metallic character** is how closely an element's properties match the ideal properties of a metal.
  - More malleable and ductile, better conductors, and easier to ionize
- Metallic character **decreases left to right** 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

Trends in Metallic Character



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 **Cm**	113	114 **Fl**	115	116 **Lv**	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**