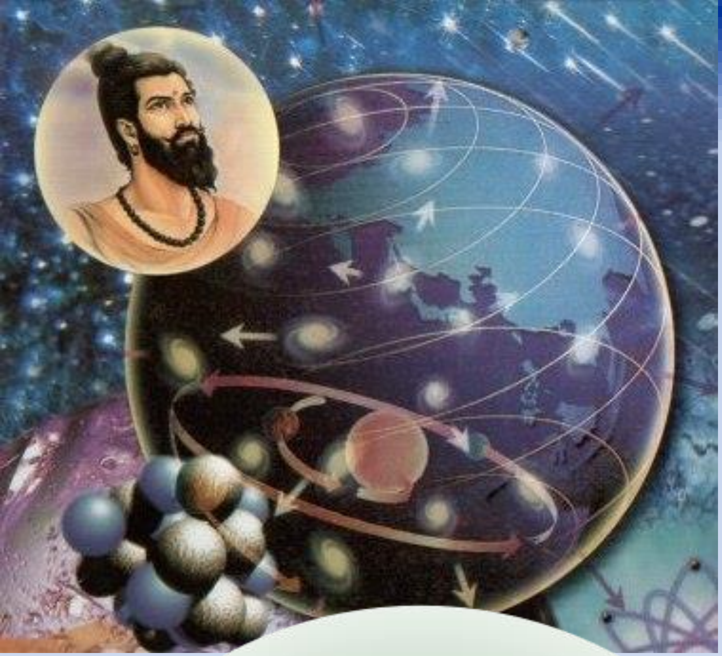
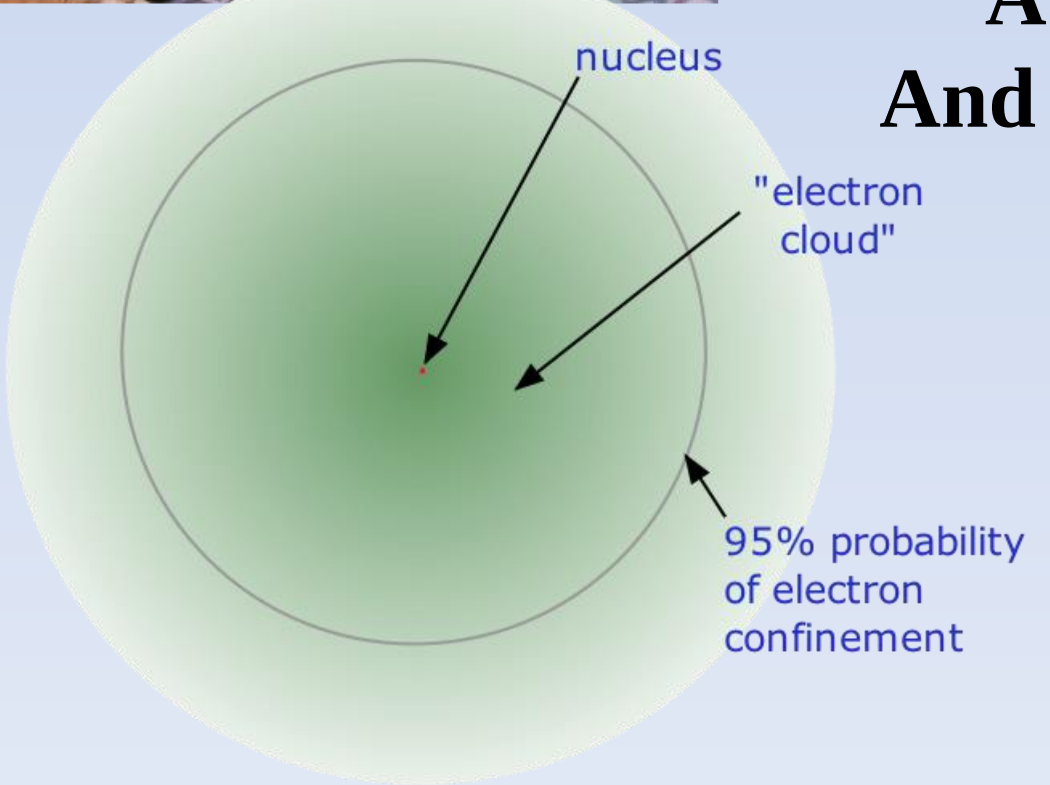


Atomic Structure
And The Periodic Table

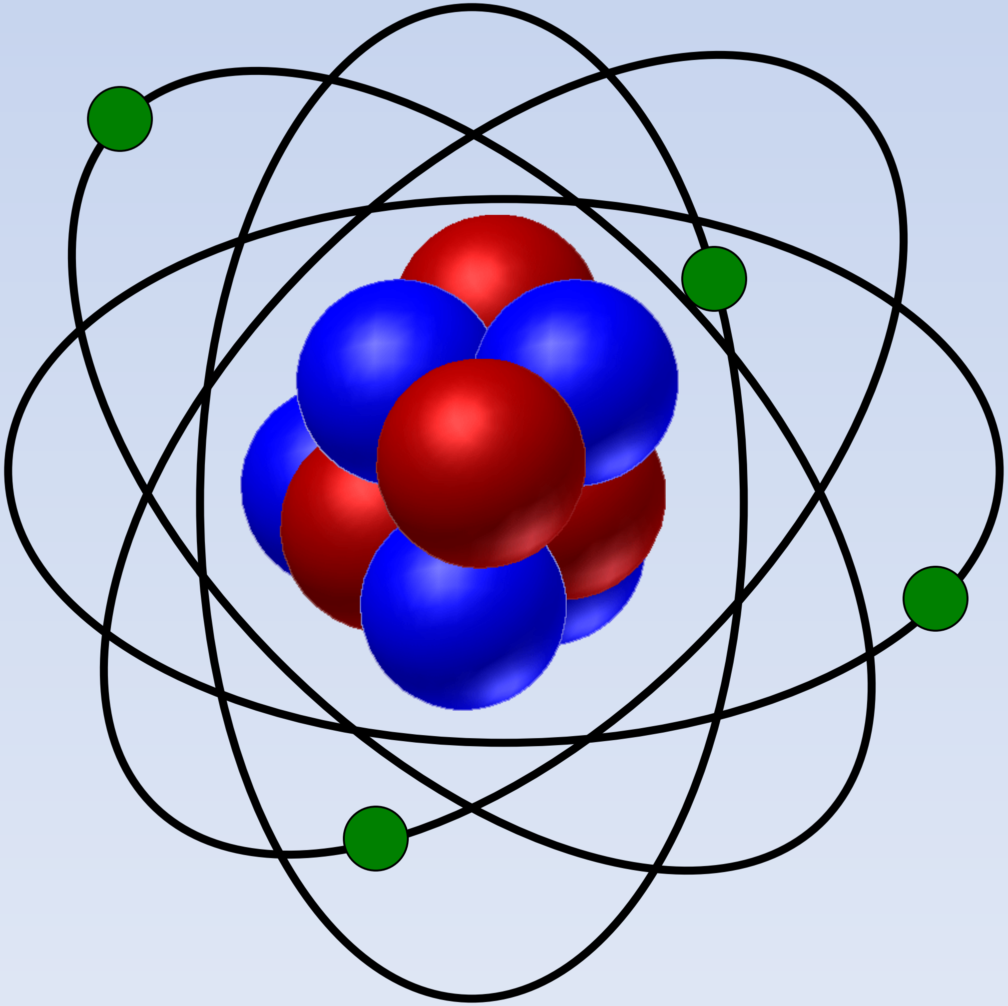


Elements

The simplest form of matter

Atoms

The smallest piece of an element that contains all properties of that element



Timeline of Atomic Theory

Greek Model
400 BC
Democristus

(Aristotle's 4 Elements)

Dalton Model
1803

Thomson Model
1897

Rutherford
Model
1911

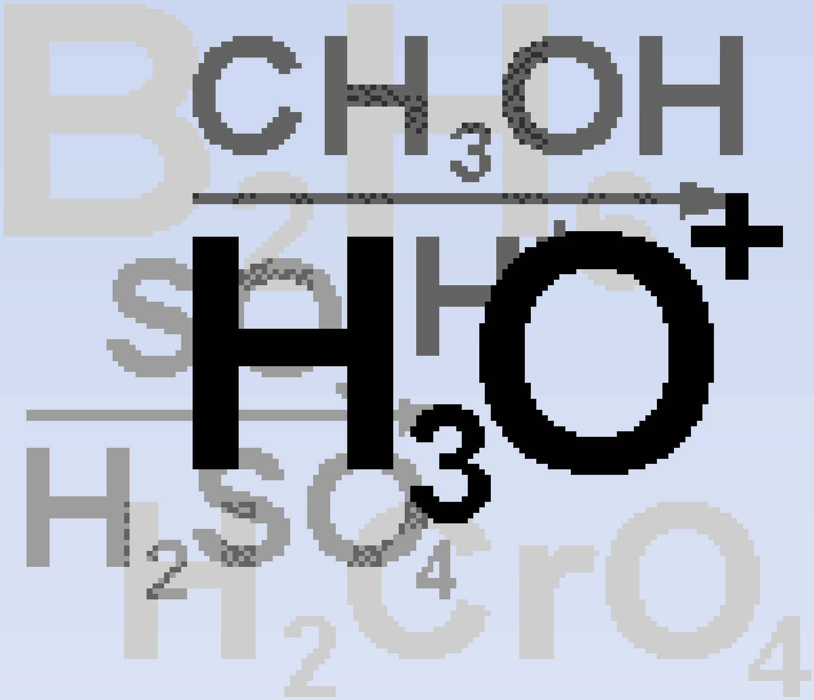
Bohr Model
1922

Wave
Model
Modern

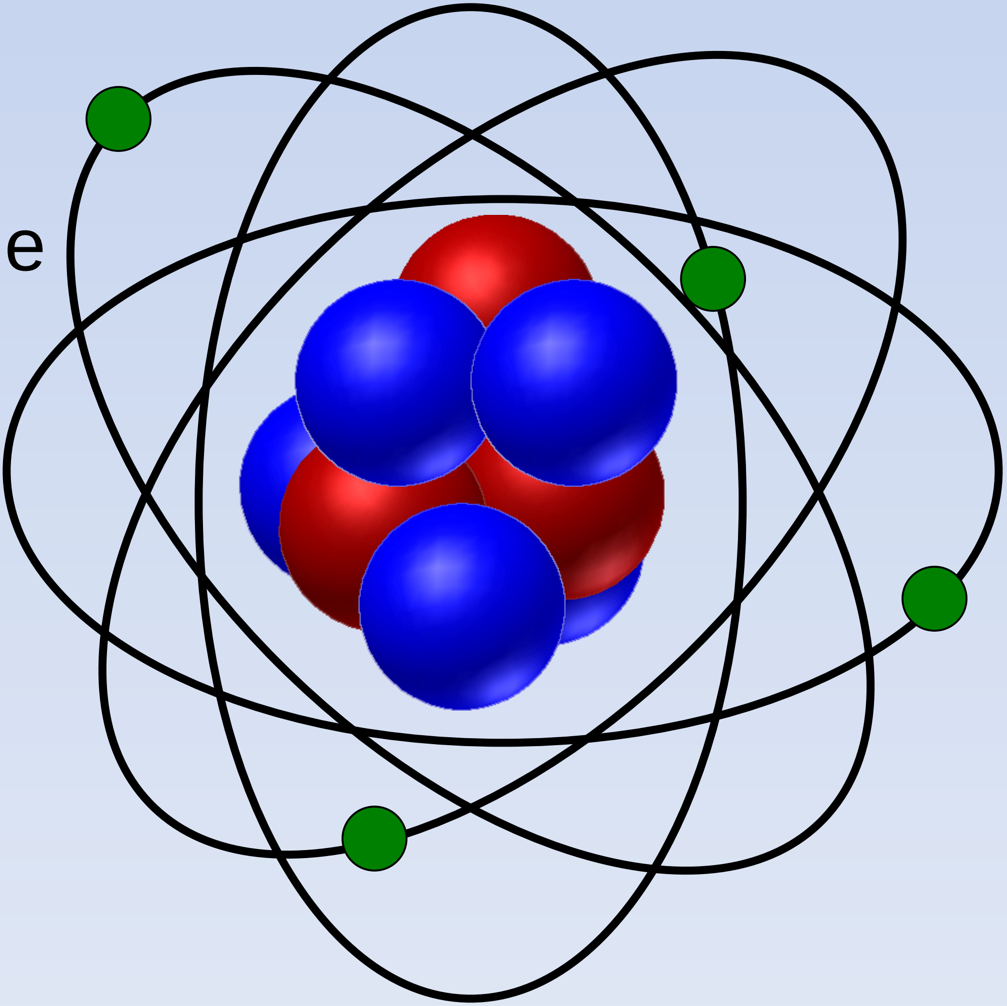
Dalton’s Model

- In the early 1800s, the English Chemist John <u>Dalton</u> performed a number of experiments that eventually led to the acceptance of the idea of atoms.

- *This theory became one of the <u>foundations of modern chemistry.</u>*



Components of an Atom



Nucleus

The center portion of an atom containing the protons and neutrons

Protons

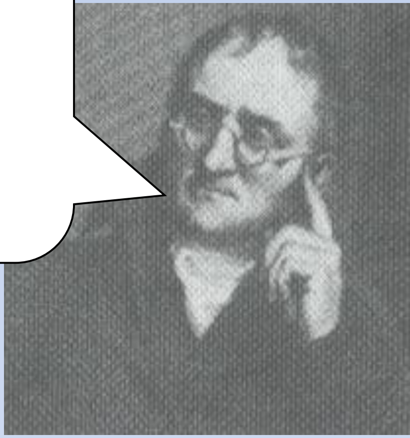
Positively charged atomic particles

Neutrons

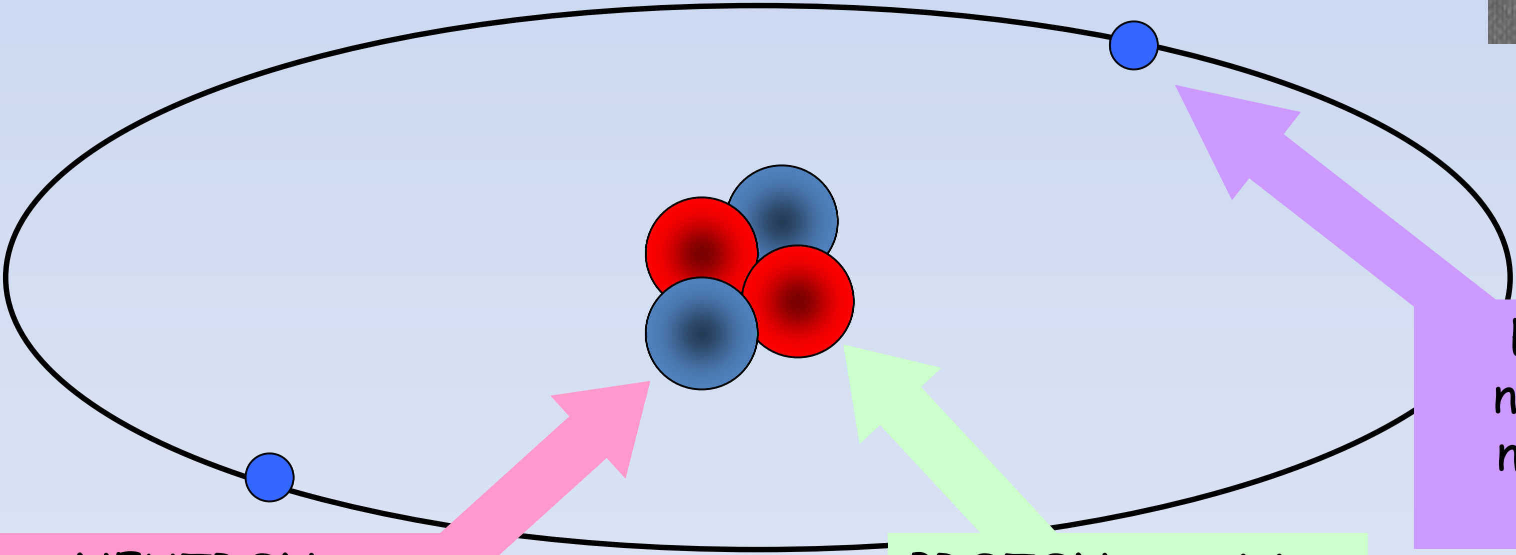
Uncharged atomic particles

The structure of the atom

The Ancient Greeks used to believe that everything was made up of very small particles. I did some experiments in 1808 that proved this and called these particles ATOMS:



Dalton



NEUTRON – neutral, same mass as proton (“1”)

PROTON – positive, same mass as neutron (“1”)

ELECTRON – negative, mass nearly nothing

Subatomic Particles

- **protons** (p), charge +1, mass ≈ 1amu
- **neutrons** (n), charge 0, mass ≈ 1amu

exist in the atomic **nucleus**

- **electrons (ē),** charge -1, mass ≈ 0 amu

exist ***outside*** the nucleus

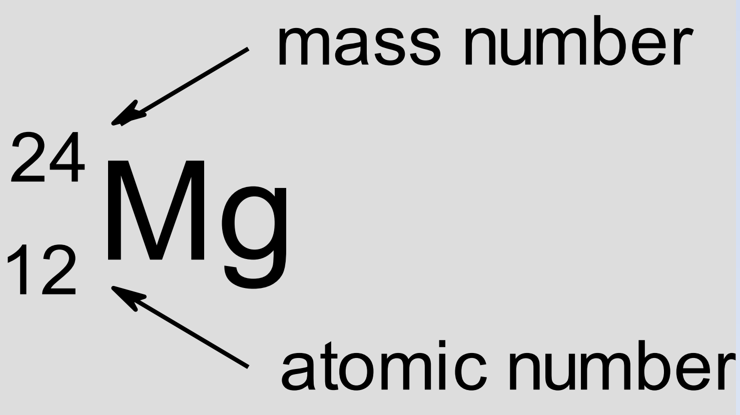
The nucleus (plural, nuclei) is incredibly small:

10000 × ***diameter of nucleus*** = ***diameter of atom***

The nucleus does not change during any ordinary chemical reaction

(for a neutral atom)

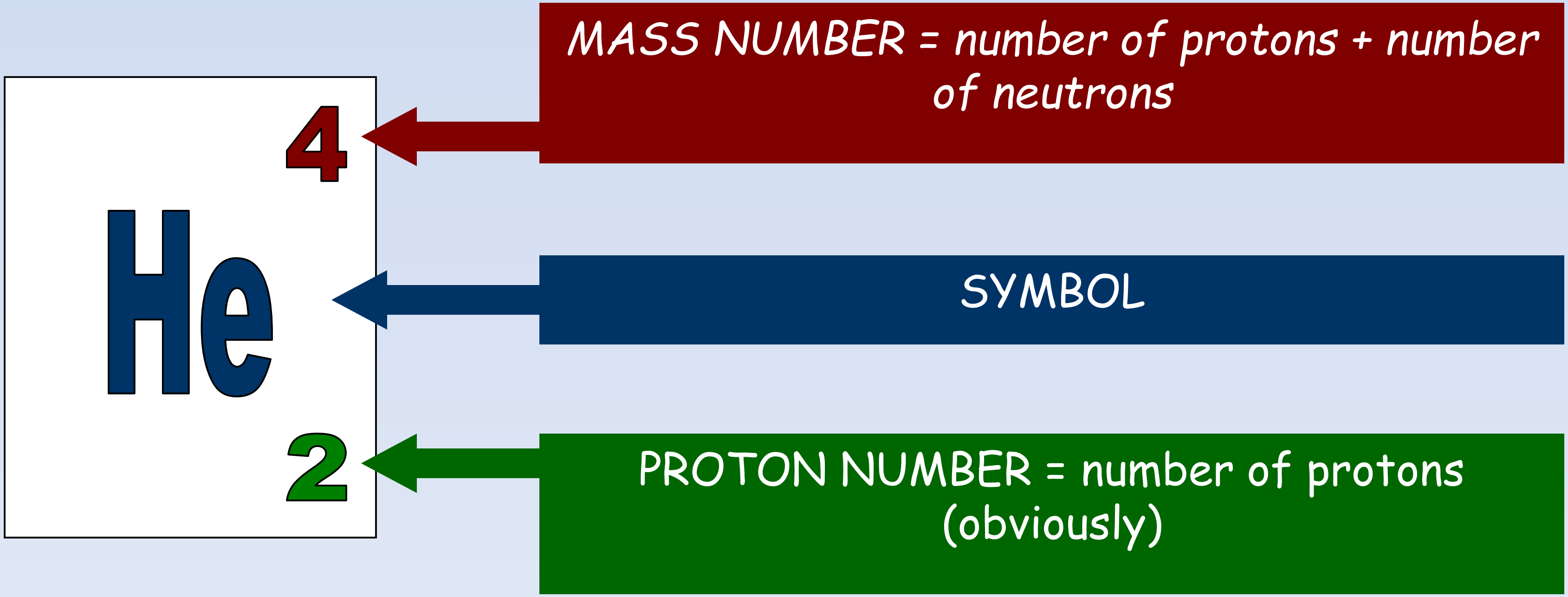
- **number of protons = number of electrons**
- **atomic number (Z)** of the element: **Z = p**
- **mass number (A)** : **A = p + n = Z + n**



$$^{24}_{12}\mathrm{Mg}$$

Mass and atomic number

Particle	**Relative Mass**	**Relative Charge**
Proton	1	1
Neutron	1	0
Electron	0	-1



MASS NUMBER = number of protons + number of neutrons

SYMBOL

PROTON NUMBER = number of protons (obviously)

- The absolute masses of atoms of elements is very small

  mass (O) = 2.667×10^{-26} kg

  mass (H) = 1.674×10^{-27} kg

  mass (C) = 1.993×10^{-26} kg

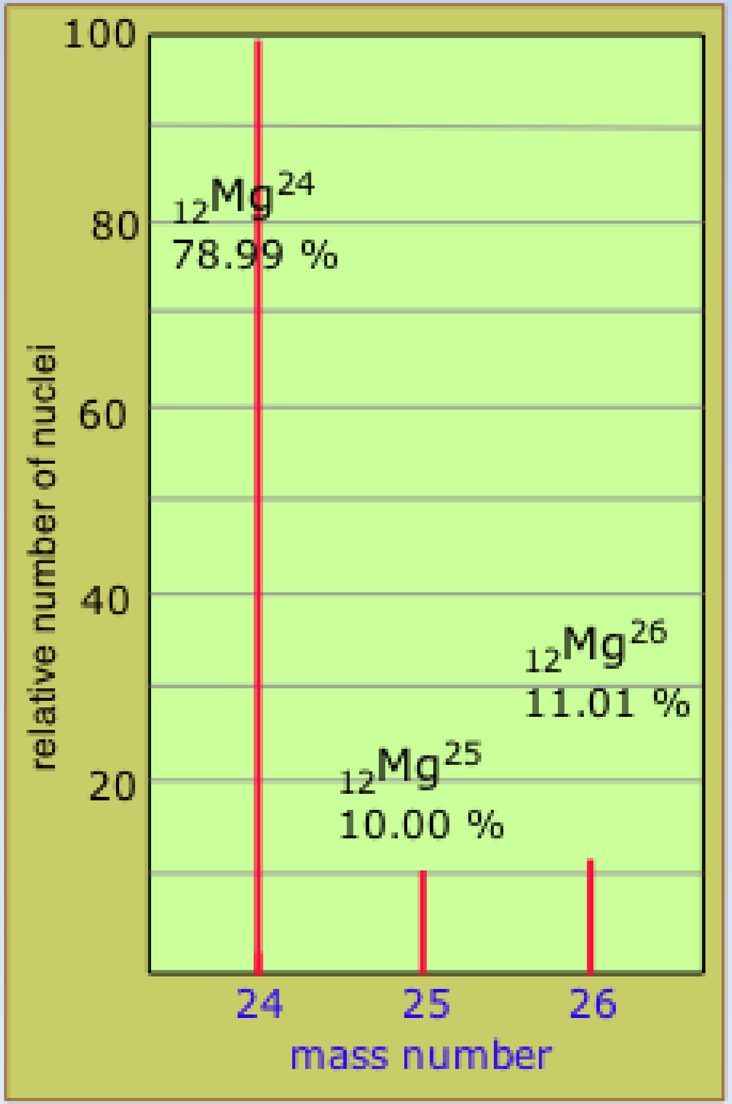
- To avoid using terms like 10^{-27} to describe the mass of atom, scientists have to define a much smaller unit of mass called the **atomic mass unit,** which is abbreviated ***amu***

atomic mass unit 1 amu = 1.66054×10^{-27} kg

- It is the *1/12* parts of mass of atom of an isotope of carbon ^{12}C

Isotopes. Relative mass of atoms

Isotopes have the ***same*** atomic number but ***different*** mass numbers



The atomic masses that are listed in tables are *weighted averages* of these isotopic mixtures.

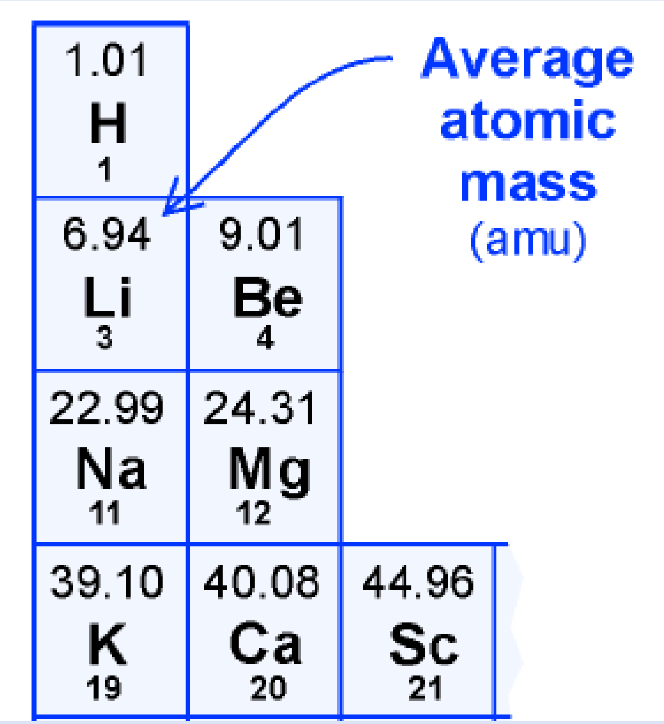
The average atomic mass of **magnesium** :

$(0.7899 \times 24) + (0.1000 \times 25) + (0.1101 \times 26) = \mathbf{24.305}$

All three isotopes are present in all compounds of magnesium in the same proportions

Approximately 290 isotopes occur in nature

Average atomic mass



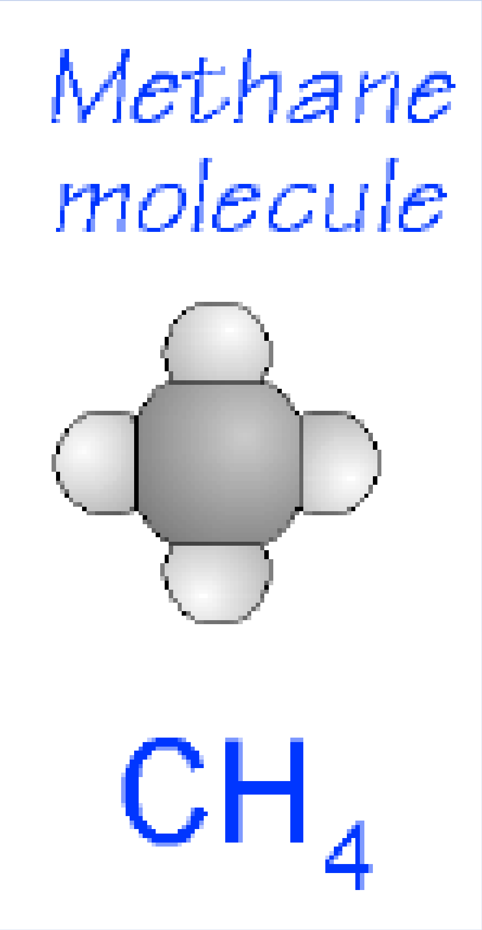
- The average atomic mass give the proportion of each isotope by mass.
- For example, the periodic table lists an atomic mass of 6.94 for lithium.
- On average, 94% of lithium atoms are Li^7 and 6% are Li^6.

Atomic Mass for Stable Isotopes of Elements 1 - 26

Atomic Number	Element Symbol	Element Name	Mass Numbers of Stable Isotopes	Average Atomic Mass (amu)
1	H	Hydrogen	1, 2	1.008
2	He	Helium	3, 4	4.003
3	Li	Lithium	6, 7	6.941
4	Be	Beryllium	9	9.012
5	B	Boron	10, 11	10.81
6	C	Carbon	12, 13	12.01
7	N	Nitrogen	14, 15	14.07
8	O	Oxygen	16, 17, 18	16.00
9	F	Fluorine	19	19.00
10	Ne	Neon	20, 21, 22	20.18
11	Na	Sodium	23	22.99
12	Mg	Magnesium	24, 25, 26	24.31
13	Al	Aluminum	27	26.98
14	Si	Silicon	28, 29, 30	28.09
15	P	Phosphorus	31	30.97
16	S	Sulfur	32, 33, 34, 36	32.06
17	Cl	Chlorine	35, 37	35.45
18	Ar	Argon	36, 38, 40	39.95
19	K	Potassium	39, 41	39.10
20	Ca	Calcium	40, 42, 43, 44, 46, 48	40.08
21	Sc	Scandium	45	44.96
22	Ti	Titanium	46, 47, 48, 49, 50	47.88
23	V	Vanadium	51	50.94
24	Cr	Chromium	50, 52, 53, 54	52.00
25	Mn	Manganese	55	54.94
26	Fe	Iron	54, 56, 57, 58	55.85

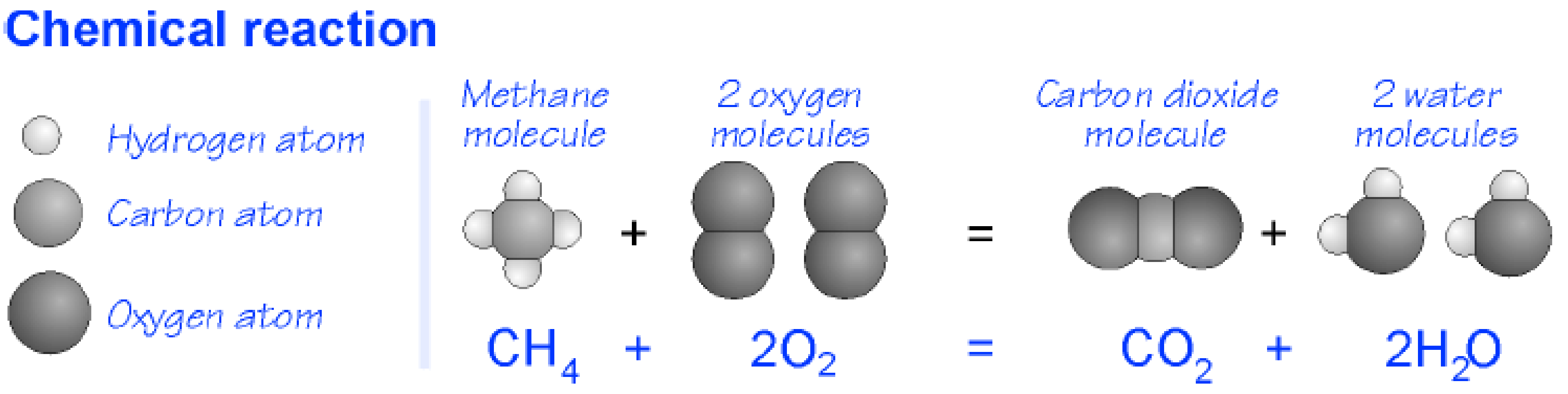
Reactions inside and between atoms

- Most atoms in nature are found combined with other atoms into molecules.
- A molecule is a group of atoms that are chemically bonded together.



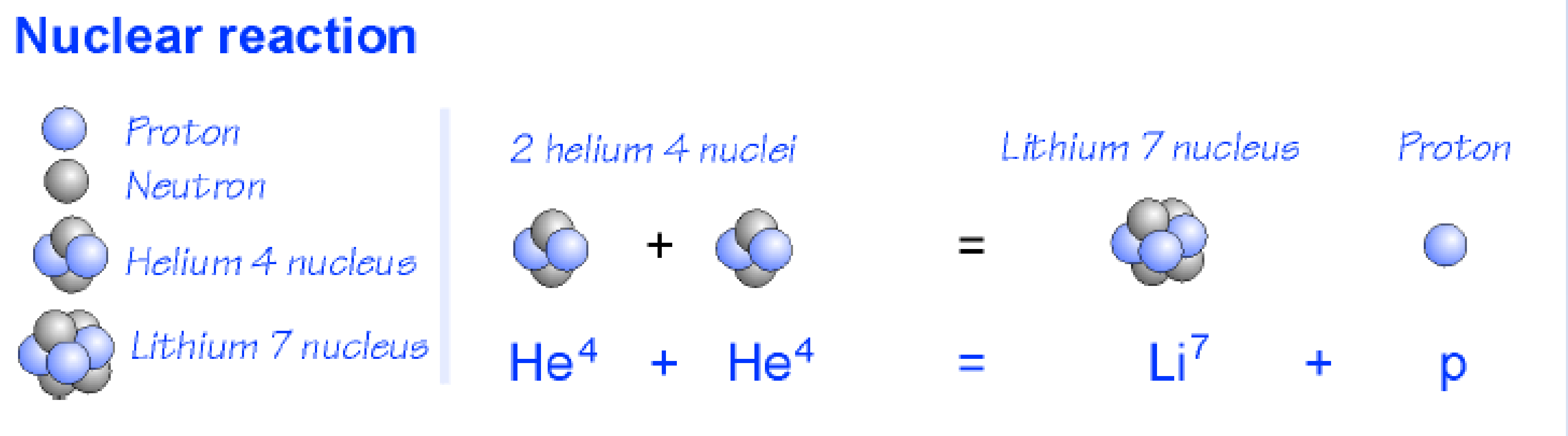
Reactions between atoms

- A chemical reaction rearranges the same atoms into different molecules.
- Chemical reactions rearrange atoms into new molecules but do not change atoms into other kinds of atoms.



Reactions inside atoms

- A nuclear reaction is any process that changes the nucleus of an atom.
- A nuclear reaction can change atoms of one element into atoms of a different element.

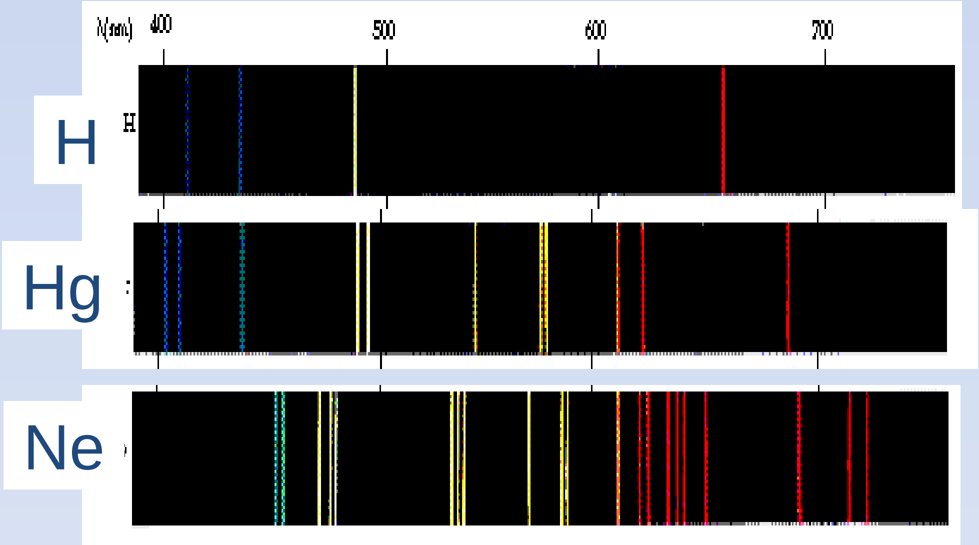


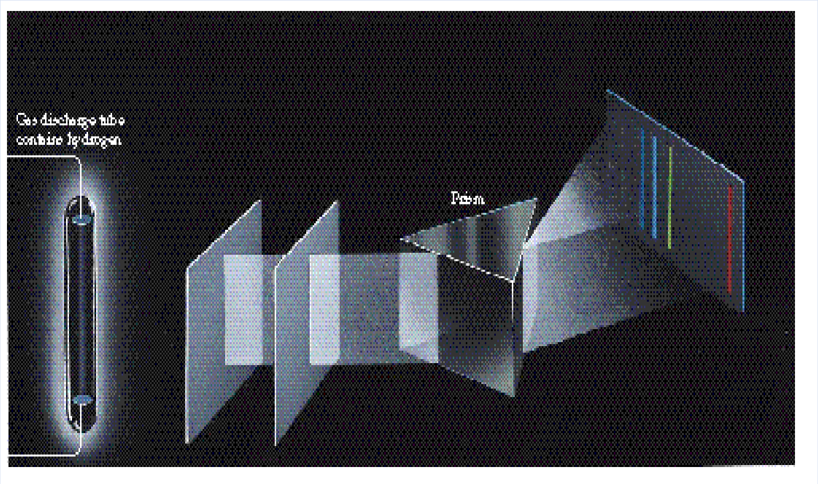
Electrons and Quantum Theory

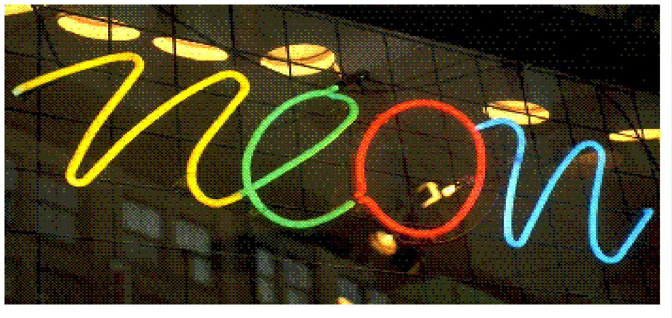
- Quantum physics is the branch of science that deals with extremely small systems such as an atom.
- A brilliant scientist, Neils Bohr is often called the father of quantum physics.
- Niels Bohr was the first person to put the clues together correctly and in 1913 proposed a theory that described the electrons in an atom.

Line Spectra of Excited Atoms

- Excited atoms emit light of only certain wavelengths
- The wavelengths of emitted light depend on the element.

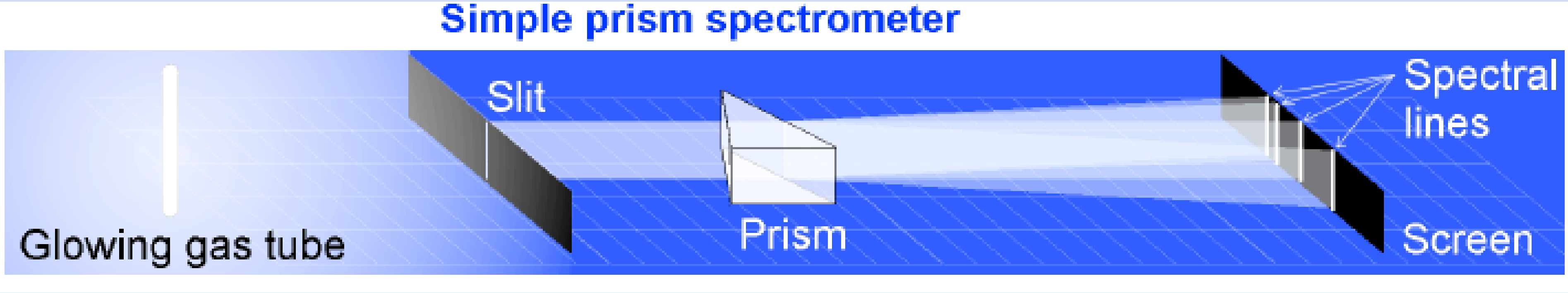


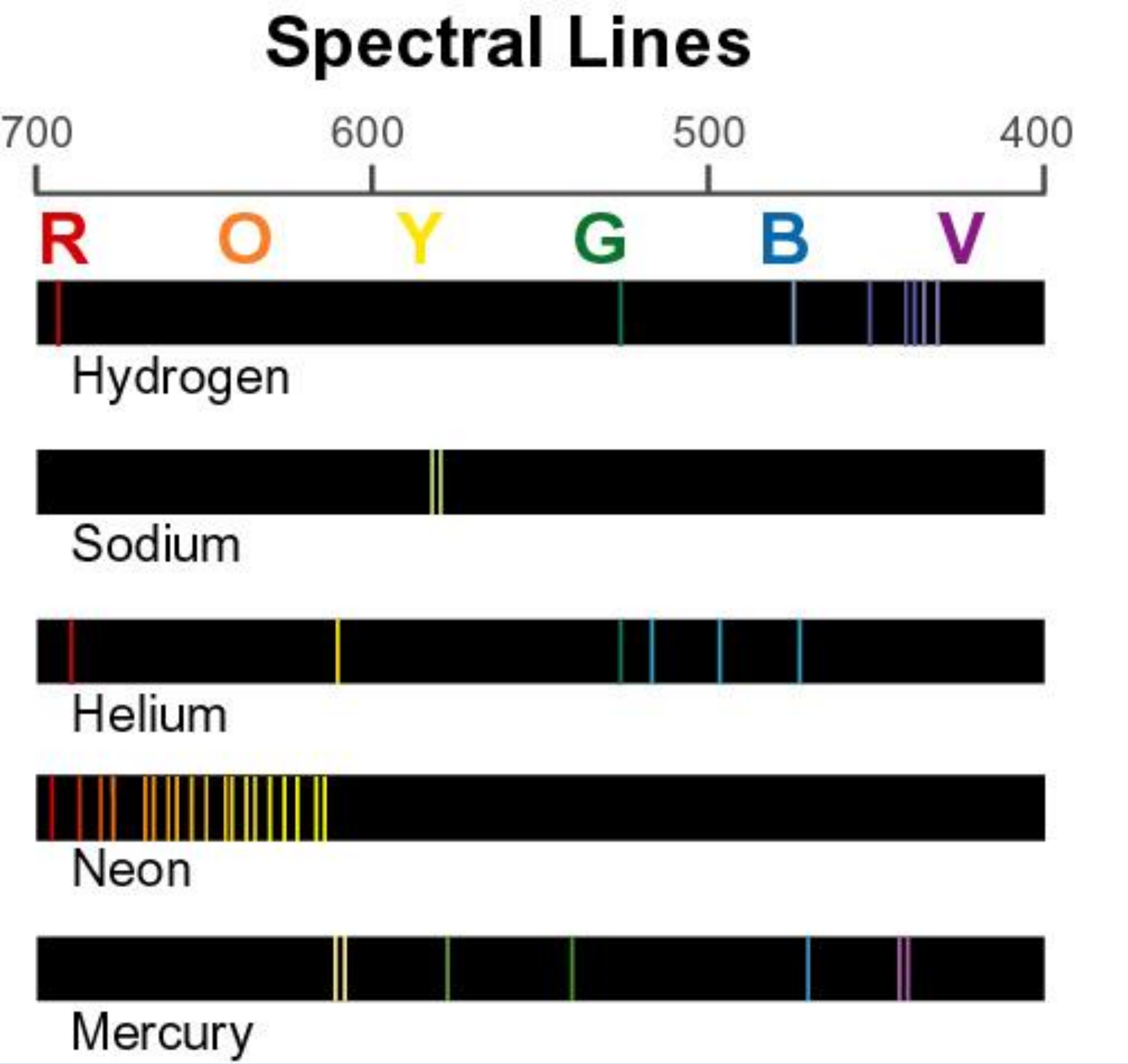




Electrons and Quantum Theory

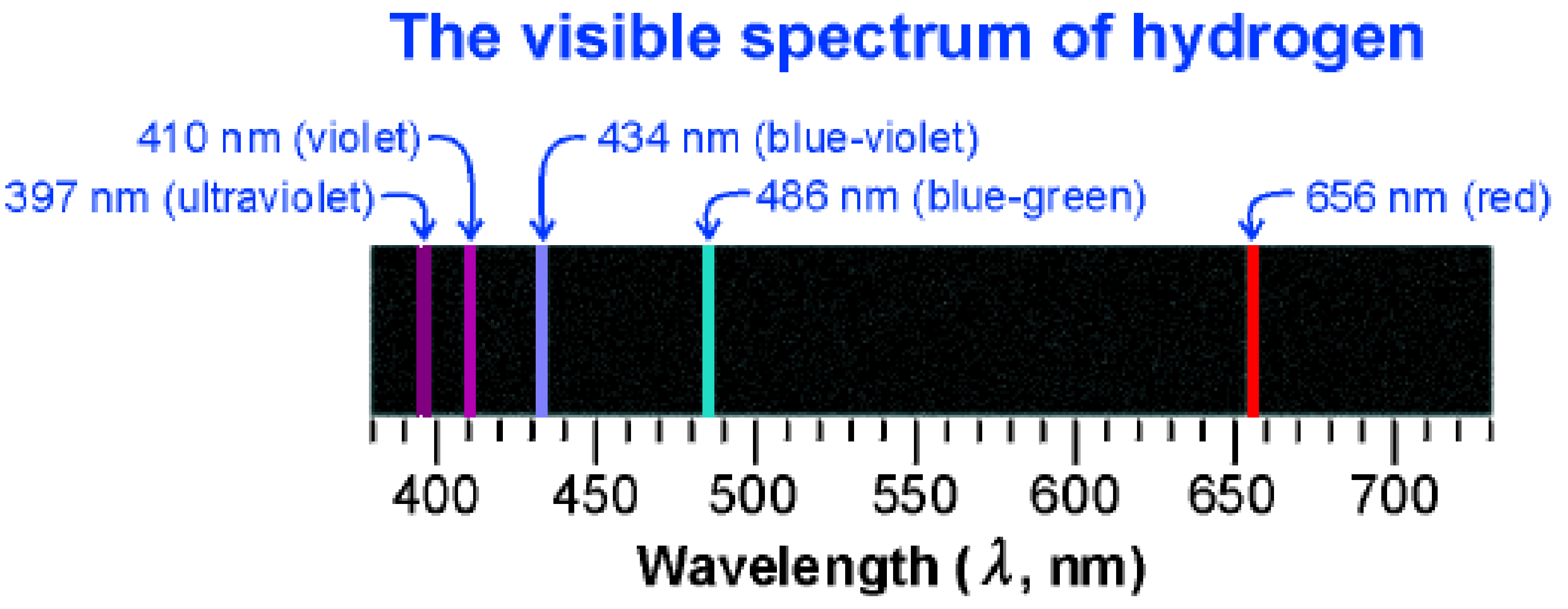
- Each individual color is called a spectral line because each color appears as a line in a spectrometer.
- A spectrometer is a device that spreads light into its different wavelengths, or colors.




Spectral Lines
700
600
500
400
R
O
Y
G
B
V
Hydrogen
Sodium
Helium
Neon
Mercury

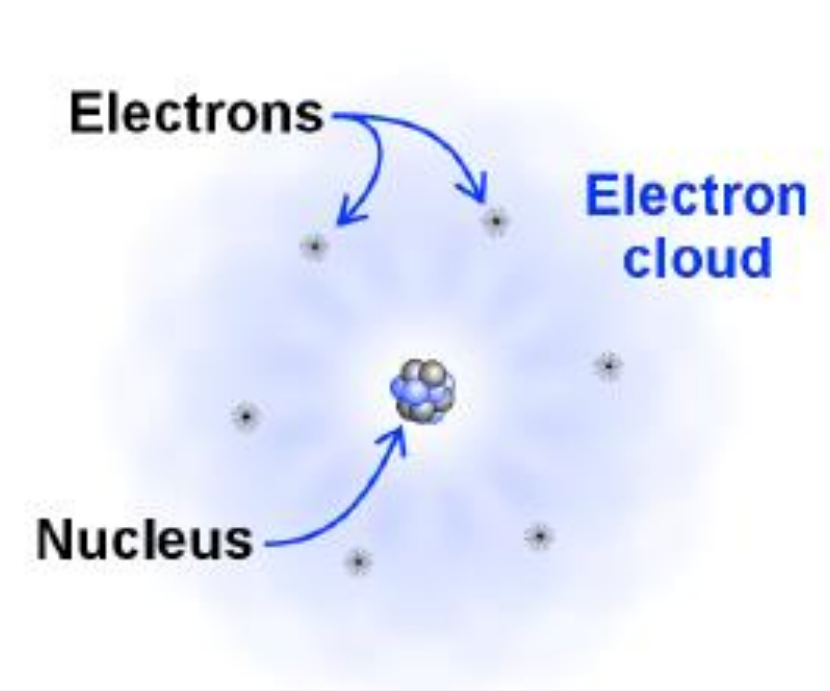
Balmer's formula

- The first serious clue to an explanation of the atom was discovered in 1885 by Johann Balmer, a Swiss high school teacher.
- He showed that the wavelengths of the light given off by hydrogen atoms could be predicted by a mathematical formula (Balmer’s formula).



The structure of the atom

- Electrons are outside the nucleus in the electron cloud.
- Because electrons are so fast and light, physicists tend to speak of the "electron cloud" rather than talk about the exact location of each electron.



Quantum states

- Every quantum state in the atom is identified by a unique combination of the four quantum numbers

Quantum Numbers (q.n.)

- *principal (energetic) q.n., n*
- *orbital (azimuthal) q.n., ℓ*
- *magnetic q.n., m*
- *spin q.n.,m_s*

principal q.n. n

It can have any positive integer values in the ***exiting state*** of atom

$$n = 1, 2, 3, 4 \ldots + \infty$$

For the ***ground state*** of an atom
(the most stable energetic states of an atom)

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

- atoms have a series of energy levels called **principal energy levels**
- level is defined as a group of electrons with the same principal q.n.
- the energy increases as the value of n increases

Orbital q.n. ℓ

- It cannot be negative and it cannot be any large than $n - 1$

$$\ell = 0, 1, 2, 3, \dots, n - 1$$

- In atom each principal energy level contains one or more types of orbitals called ***sublevel***

Types of sublevels:

$\ell = 0$ (s - sublevel)
$\ell = 1$ (p - sublevel)
$\ell = 2$ (d - sublevel)
$\ell = 3$ (f - sublevel)
$\ell = 4$ (g - sublevel)

magnetic q.n. *m*

- Its value may be positive or negative, and may rang from – ℓ through zero to + ℓ in integral steps.

$$m = -\ell, \ldots 0, \ldots +\ell$$

- It does determine the orientation in space of the volume that can contain the electron

ℓ= 0 (s) m= 0 s –sublevel has only 1 orbital

ℓ= 1 (p) m= -1, 0, +1 p –sublevel has 3 orbitals

ℓ= 2 (d) m= -2, -1, 0, +1, +2 d –sublevel has 5 orbitals

ℓ= 3 (f) m= -3, -2, -1, 0, +1, +2, +3 f –sublevel has 7 orbitals

The number of the sublevels is equal to the number of energy level

The 1-st energy level

$n = 1$ $\ell = 0$ (s)

The 2-nd energy level

$n = 2$ $\ell = 0, 1$

$\ell = 0$ (s)
$\ell = 1$ (p)

The 3-rd energy level

$n = 3$ $\ell = 0, 1, 2$

$\ell = 0$ (s)
$\ell = 1$ (p)
$\ell = 2$ (d)

The 4 energy level

$n = 4$ $\ell = 0, 1, 2, 3$

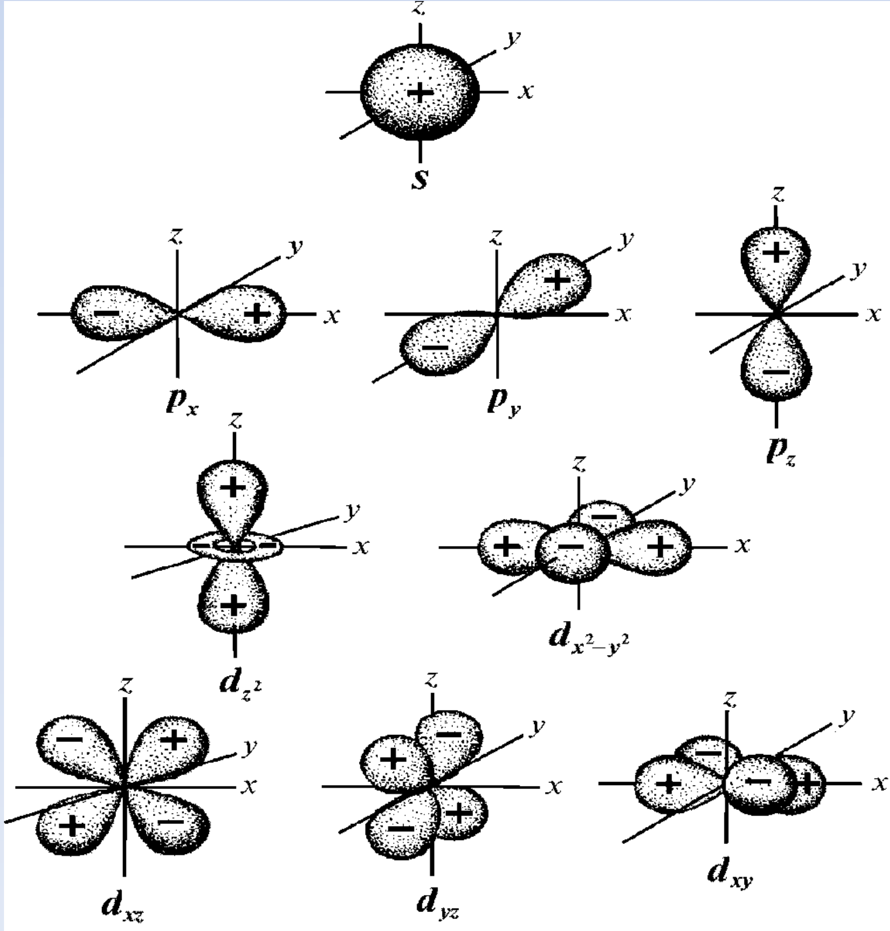
$\ell = 0$ (s)
$\ell = 1$ (p)
$\ell = 2$ (d)
$\ell = 3$ (f)

- According to wave mechanical model electron is the particle from one hand and the wave from the other hand.

- It is impossible to know exactly both the location and the momentum of an electron in an atom at the same time. This fact is known as ***Heisenberg uncertainty principle***.

Shapes of Orbitals

- Therefore, scientists describe the probable locations of electrons. These locations describe the ***orbital shapes***.



spin q.n. m_s

- may have a value of - ½ or + ½ only.
- the spin value indicates that the electron is spinning on its axis in one direction or the opposite.
- we often represent spin with an arrow:

either ↑ or ↓

- maximum of two electrons can occupy any given orbital

in an atom (Pauli exclusion principle)

Electron Configurations of the Elements

The electronic configuration (formulae) of argon: Ar $\underbrace{1s^2}_{K}\underbrace{2s^22p^6}_{L}3s^23p^6$

If the energy level complete with electrons the capital letters can be used KLM…..
The short electronic configuration (formulae) of argon:

Ar $KL3s^23p^6$

K $\underbrace{1s^22s^22p^63s^23p^6}_{Ar}4s^1$

The short electronic configuration (formulae) of potassium:

K $[Ar]4s^1$

The symbols of inert gases (VIIIA group)
can be used in the short electronic configuration

Task 1. Calculate the maximum numbers of electrons in electronic shells 4, 5 (n = 4; 5) and subshells d- , f- (l = 2; 3).

$$N = 2n^2$$

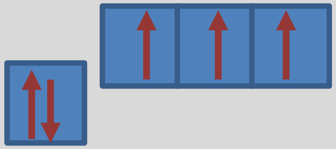
n = 4 $N = 2\times4^2 = 32$ n = 5 $N = 2\times5^2 = 50$

l = 2 (d) 5 orbitals×2electrons = 10 electrons

l = 3 (f) 7 orbitals×2electrons = 14 electrons

The principle of maximum multiplicity – **the Hund's rule**. When orbital with identical energy is available, electrons occupying these singly rather than in pairs.

N $1s^2\underline{2s^2 2p^3}$



Klechkovsky rules

sum ***n* + ℓ**

1s = 1+0 = 1

2s = 2+0 = 2

2p = 2+1 = 3

3s = 3+0 = 3

3p = 3+1 = 4

3d = 3+2 = 5

4s = 4+0 = 4

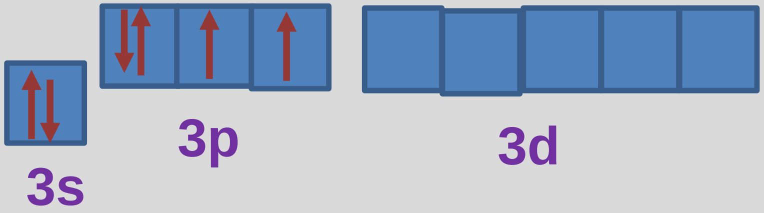
4p = 4+1 = 5

4d = 4+2 = 6

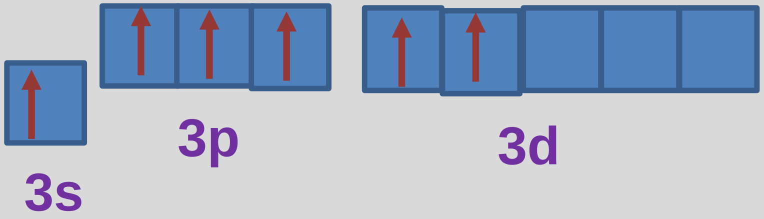
1s 2s 2p 3s 3p 4s 3d 4p………

Task 2. Write the electronic formula and orbital diagram of the sulfur atom in basic and excited states. Determine the number of protons and neutrons in the nucleus of its atom.

$\mathrm{S}\quad 1s^2 2s^2 2p^6 \underline{3s^2 3p^4 3d}$

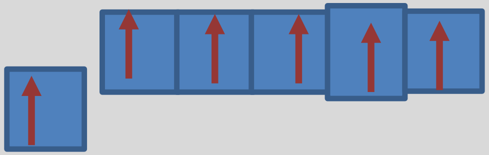


$\mathrm{S}^*\quad 1s^2 2s^2 2p^6 \underline{3s^1 3p^3 3d^2}$

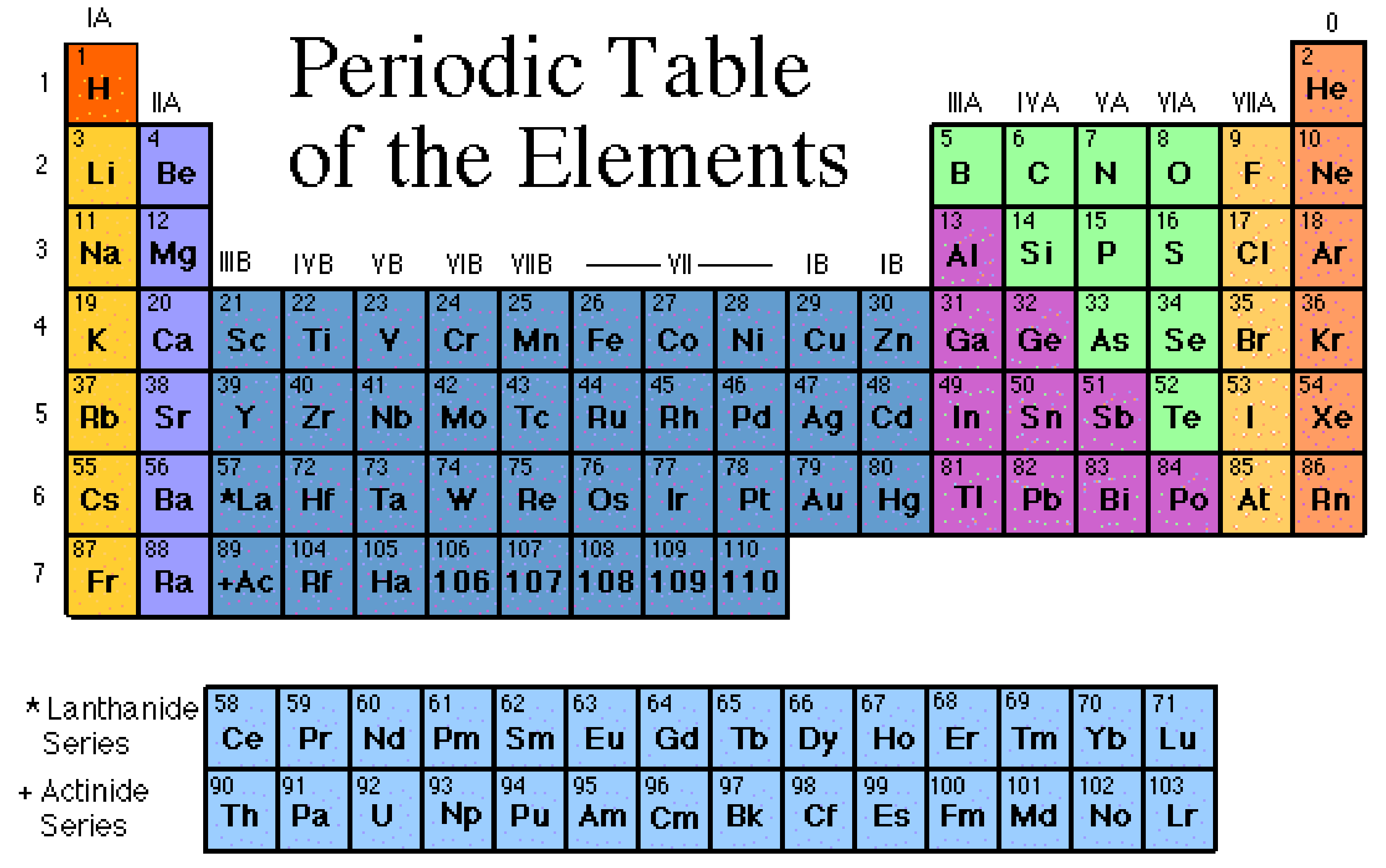


Task 3. Draw the orbital (box) diagram of the valence electrons of chromium atom and copper atom.

$_{24}\mathrm{Cr}\quad 1s^2 2s^2 2p^6 3s^2 3p^6 \underline{3d^{\overset{5}{\cancel{4}}} 4s^{\overset{1}{\cancel{2}}}}$



Periodic Table of Elements



Periodic Table of the Elements

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VII			IB	IB	IIIA	IVA	VA	VIA	VIIA	0
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 Ha	106 106	107 107	108 108	109 109	110 110								

* Lanthanide Series	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
+ Actinide Series	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

Elements

- Science has come along way since Aristotle's theory of Air, Water, Fire, and Earth.
- Scientists have identified 90 naturally occurring elements, and created about 28 others.

Dmitri Mendeleev (1834-1907)

- Russian chemist
- Credited as being the creator of the first version of the periodic table of elements
- Arranged his periodic table according to atomic mass so that elements with similar properties were in the same group

Periodic Law

When the elements are arranged in order of increasing atomic number, there is a periodic repetition of their physical and chemical properties

Symbols

C

Carbon

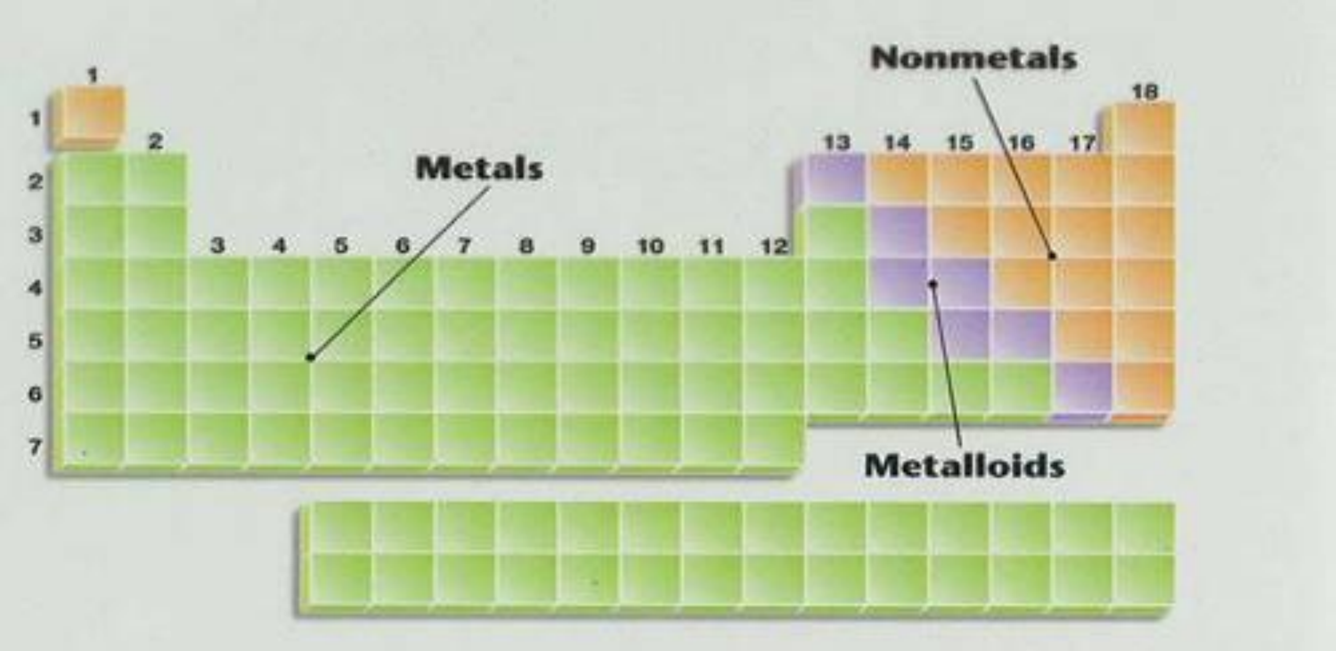
Cu

Copper

- All elements have their own unique symbol.
- It can consist of a single capital letter, or a capital letter and one or two lower case letters.

Classifying the Elements

- Three main classifications for the elements
  - Metals
  - Nonmetals
  - Metalloids



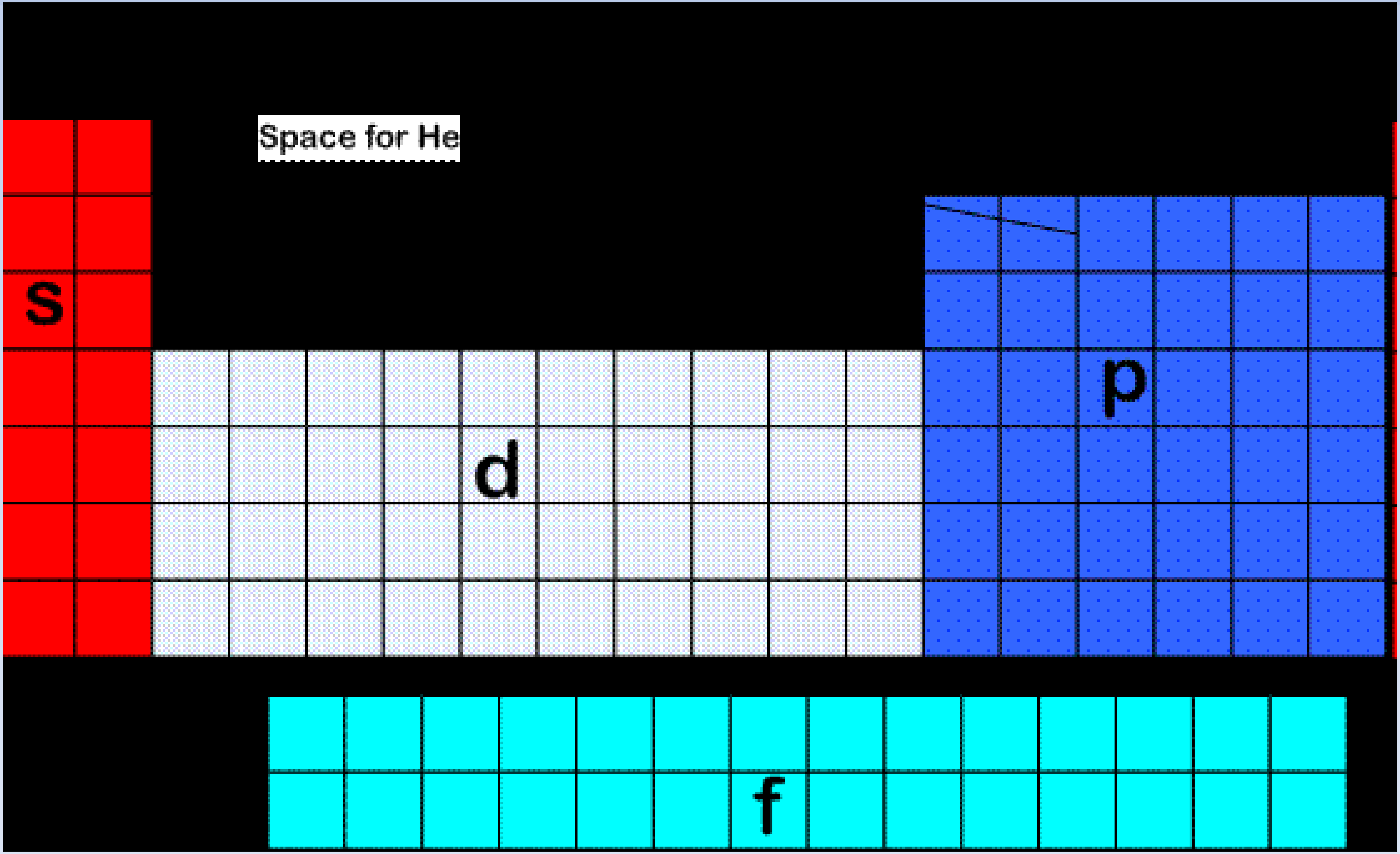
The Structure of the Periodic Table

- Periods are the horizontal rows of the elements (7 periods)
- Groups are the vertical columns of the elements (8 groups)

  the main – group elements (subgroup A) only for s- and p-elements

  the transition elements (subgroup B) only for d- and f- elements
- There are four electronic families of the elements:

  s-, *p*-, *d*-, *f*- in the periodic table.


Space for He
s
d
p
f

Traditional families of elements

Example of electronic configuration

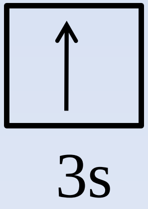
Sodium is the element of the period 3

S-electronic family element

One valence electron
on s-sublevel - IA group

Na $1s^2 2s^2 2p^6 \underline{3s^1}$

The electrons in the outermost (highest) energy level of an atom are ***valence electrons***



Electron-grafic configuration (orbital (box) diagram) only for ***valence electrons***

Task 1. Explain the similarity and the difference between chromium and sulphur.

S $1s^2 2s^2 2p^6 \underline{3s^2 3p^4}$

Cr [Ar] $\underline{4s^1 3d^5}$

6 valence electrons

+6 o.d. in the compounds

CrO_3 SO_3 are *acidic* oxides

Different valence electronic configuration

S – element of VIA group

Cr – element of VIB group

Simple compounds:
Sulfur is typical nonmetal

Chromium is typical metal

Elements of the same period have the same number of filled energy shells.

Task 2. Enter the group number and period of elements with numbers 35, 41.

At Number = 35

Br Bromine

Group VIIA
Period 4

At Number = 41

Nb Niobium

Group VB
Period 5

Task 3. The electronic configuration of anion E^{2-} is $1s^2 2s^2 2p^6 3s^2 3p^6$.

Define the period, the group and the character of the chemical element.

E^{2-} $1s^2 2s^2 2p^6 3s^2 3p^6$

E^0 $1s^2 2s^2 2p^6 3s^2 3p^4$

At Number = 16

S Sulfur

Group VIA
Period 3

Task 4. The electronic configuration of cation E^{2+} is $1s^2 2s^2 2p^6$.

Define the period, the group and the character of the chemical element.

E^{2+} $1s^2 2s^2 2p^6$

E^0 $1s^2 2s^2 2p^6 3s^2$

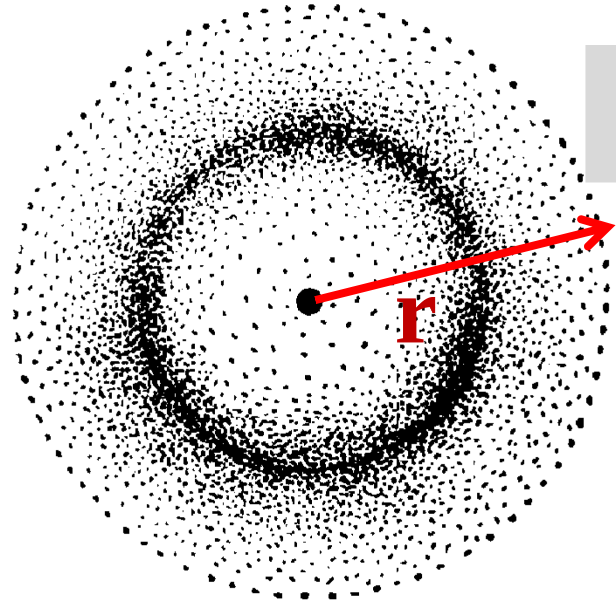
At Number = 12

Mg Magnesium

Group IIA
Period 3

Atomic Properties

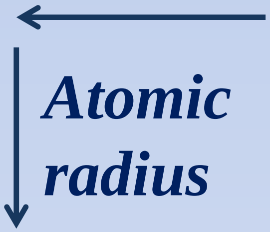
- Atomic Radius (size of atom)
- Ionization energy
- Electron affinity
- Electronegativity
- Metallic and nonmetallic character



Atomic Radius

- The sizes of atoms vary:

Atoms get ***large*** down a group
Atoms get ***smaller*** from left to right across a period



Atomic radius

Element	Li	Be	B	C	N	O	F	Ne
r, nm	0.155	0.113	0.091	0.077	0.074	0.066	0.064	0.030
Element	Na							
r, nm	0.189							

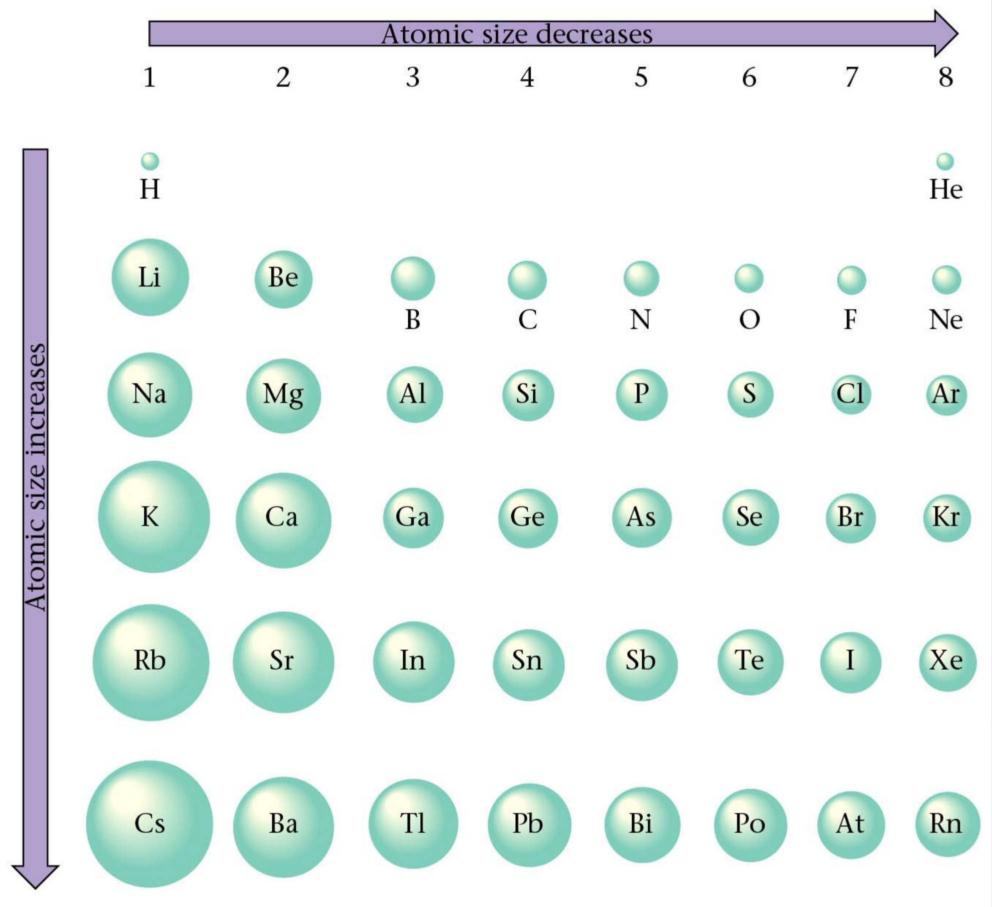
- The unit that has long been used to describe atomic size is the ***angstrom***, **Å**. Appropriate SI units are ***nanometer*** (**nm**) or ***picometer*** (**pm**).

$1Å = 1\times10^{-10}m$
$1nm = 1\times10^{-9}m$
$1\ pm = 1\times10^{-12}\ m$

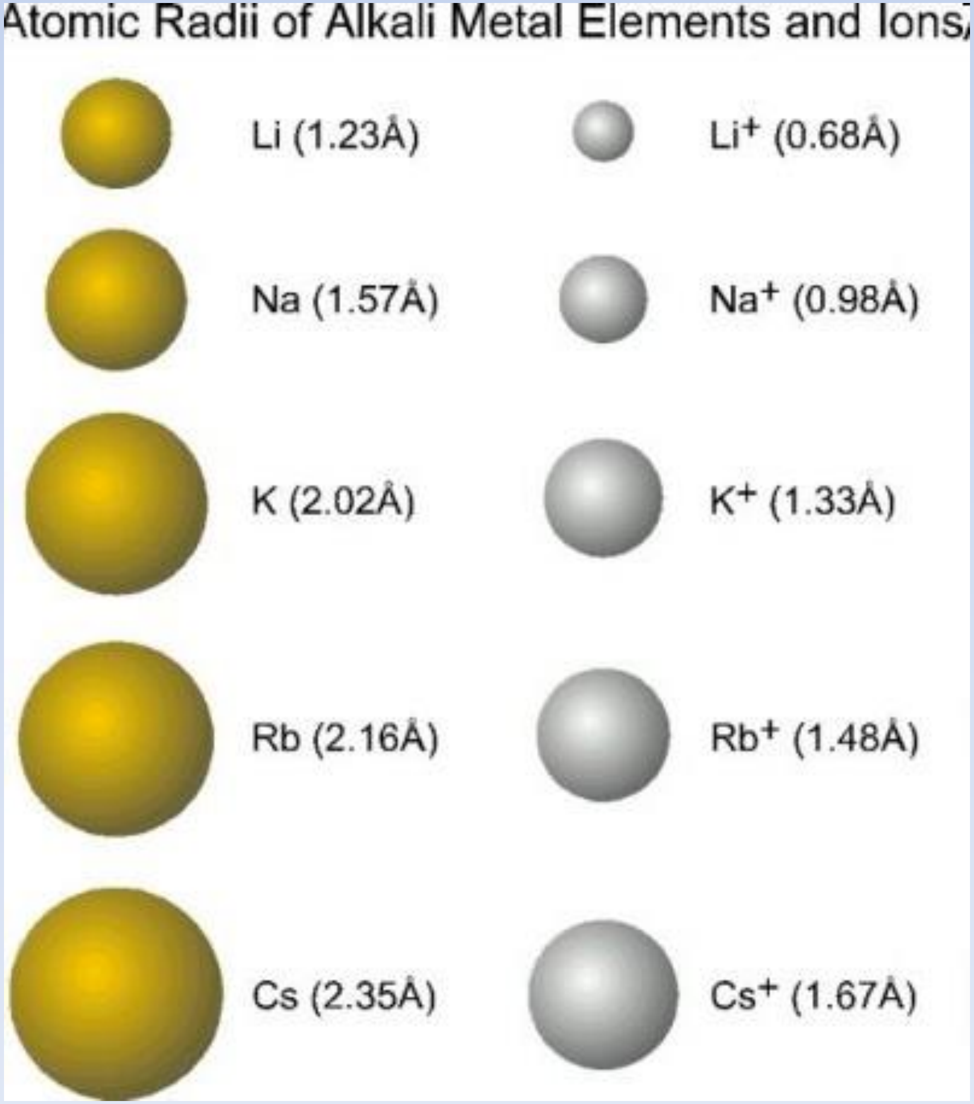
LectNote

Atomic Radii Trend

- Trends within periods
  - Generally decreases as you move left-to-right across a period (row)
- Trends within groups
  - Generally increases as you move down a group

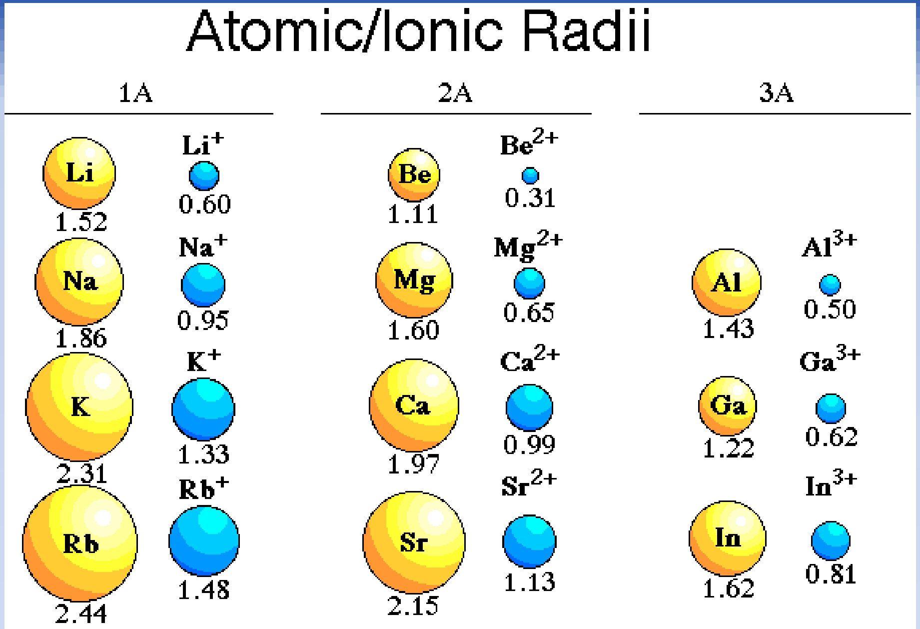


An ion is an atom or a bonded group of atoms that has a positive or negative charge



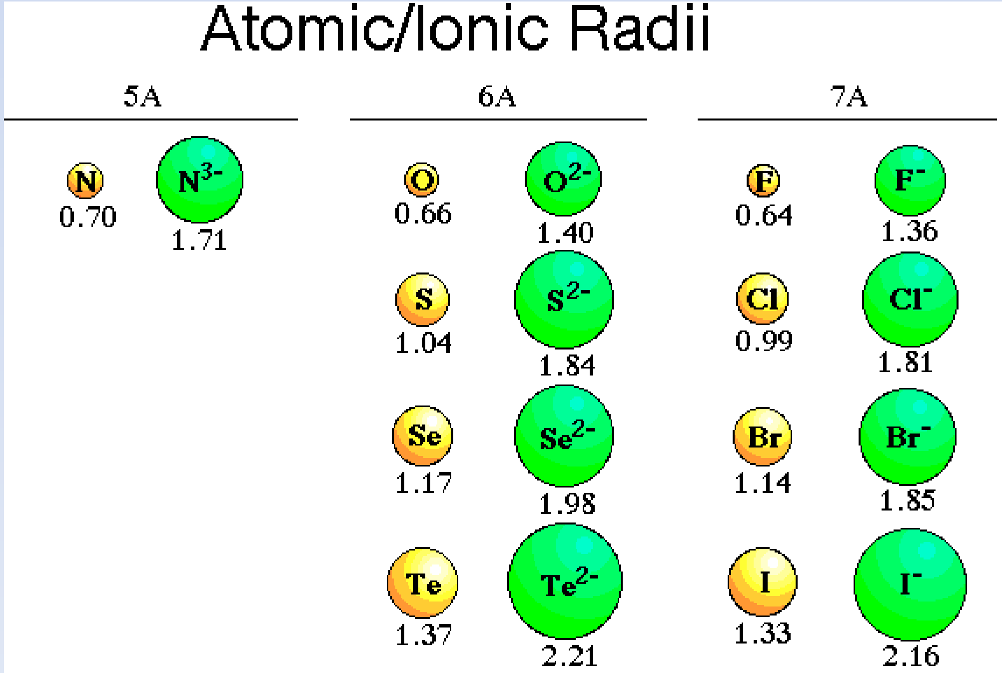
When atoms lose electrons and form positively charged ions, they always become smaller

When atoms gain electrons and form negatively charged ions, they always become larger



Lose Electrons → Smaller ionic radii

Gain Electrons → larger ionic radii



Ionization Energy (potential)

- ***Ionization energy (I)*** is the energy that the gaseous atom must absorb in order that its most loosely held electron may become completely separated from it.

It is the energy required to ***remove an electron*** from an individual atom

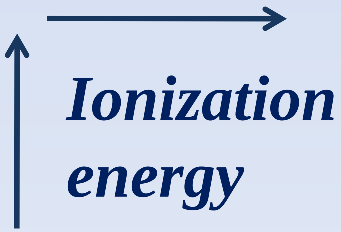
For example

$$\mathrm{Mg(g)} \longrightarrow \mathrm{Mg^{+}\,(g)} + \bar{e} \quad I_1 = 7.65\ \mathrm{eV/atom\ (kJ/mol)}$$

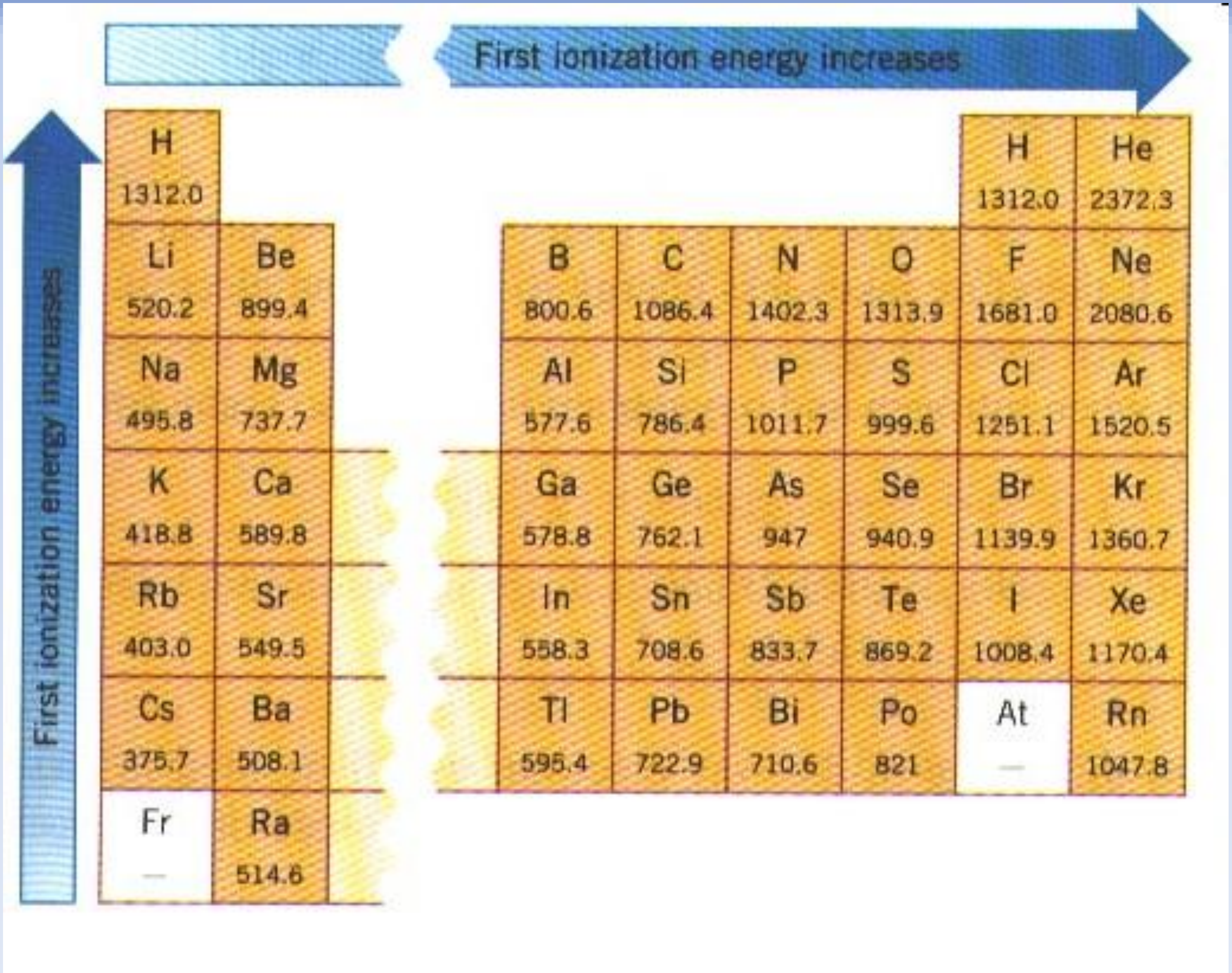
$$\mathrm{Mg^{+}\,(g)} \longrightarrow \mathrm{Mg^{2+}\,(g)} + \bar{e} \quad I_2 = 15.04\ \mathrm{eV/atom\ (kJ/mol)}$$

- Metals have relatively low ***I.*** Relatively small amount of energy is required to remove an electron from a typical metal.

- ***I*** tends to ***decrease*** in going from the top to the bottom of a group
- ***I*** tends to ***increase*** from left to right across a given period



Ionization energy



H 1312.0						H 1312.0	He 2372.3
Li 520.2	Be 899.4	B 800.6	C 1086.4	N 1402.3	O 1313.9	F 1681.0	Ne 2080.6
Na 495.8	Mg 737.7	Al 577.6	Si 786.4	P 1011.7	S 999.6	Cl 1251.1	Ar 1520.5
K 418.8	Ca 589.8	Ga 578.8	Ge 762.1	As 947	Se 940.9	Br 1139.9	Kr 1360.7
Rb 403.0	Sr 549.5	In 558.3	Sn 708.6	Sb 833.7	Te 869.2	I 1008.4	Xe 1170.4
Cs 375.7	Ba 508.1	Tl 595.4	Pb 722.9	Bi 710.6	Po 821	At —	Rn 1047.8
Fr —	Ra 514.6						

Electron Affinity

- ***Electron affinity (EA)*** is the energy what can be spent to change a neutral atom into a negative charge ion

For example

$Cl^0(g) + \bar{e} \rightarrow Cl^-(g)$ EA = -3.615 eV/atom

- Low (very negative) value of EA is a characteristic of active non-metals (acidic elements)

- EA tends to ***increase*** from left to right across a given period

- EA tends to ***decrease*** in going from the top to the bottom of a group

Electronegativity

- ***Electronegativity* (χ)** is the ability of an atom to lose or gain an electron.
- Electronegativity is related to ionization energy(I) and electron affinity(E)
- The most widely used electronegativity scale was devised by Linus Pauling

$$\chi = \mathbf{I + E}$$

- As a rule,

metals have ***electronegativities*** < 2;

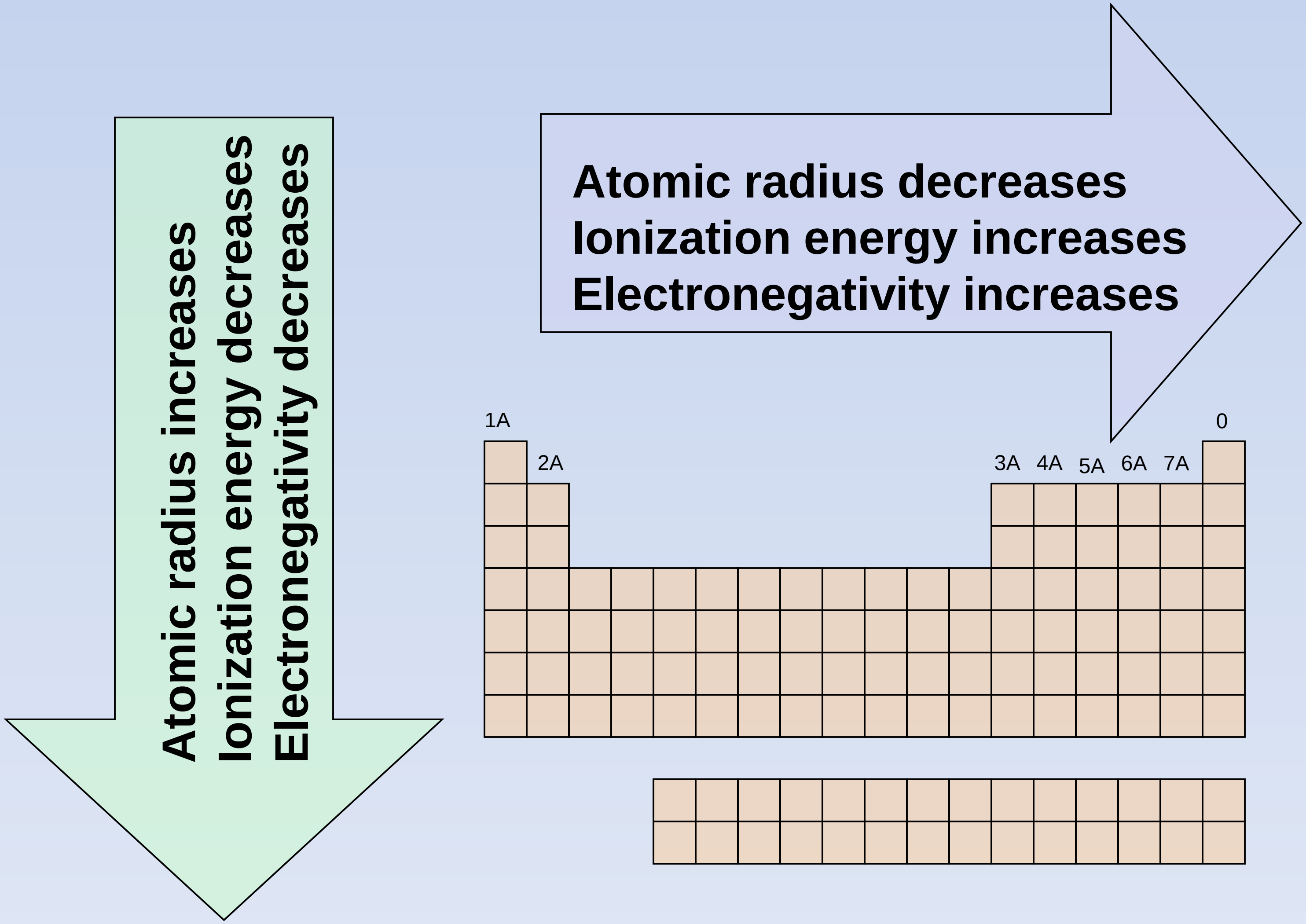
metalloids ≈ 2;

nonmetals (acidic elements) > 2.

Electronegativities of some elements

	I	II	III	IV	V	VI	VII	VIII		
I	**H** 2.1									
II	**Li** 1.0	**Be** 1.5	**B** 2.0	**C** 2.5	**N** 3.0	**O** 3.5	**F** 4.0			
III	**Na** 0.9	**Mg** 1.2	**Al** 1.5	**Si** 1.8	**P** 2.1	**S** 2.5	**Cl** 3.0			
IV	**K** 0.8	**Ca** 1.0	**Sc** 1.3	**Ti** 1.5	**V** 1.6	**Cr** 1.6	**Mn** 1.5	**Fe** 1.8	**Co** 1.8	**Ni** 1.8
	Cu 1.9	**Zn** 1.6	**Ga** 1.6	**Ge** 1.8	**As** 2.0	**Se** 2.4	**Br** 2.8			
V	**Rb** 0.8	**Sr** 1.0	**Y** 1.2	**Zr** 1.4	**Nb** 1.6	**Mo** 1.8	**Tc** 1.9	**Ru** 2.2	**Rh** 2.2	**Pd** 2.2
	Ag 1.9	**Cd** 1.7	**In** 1.7	**Sn** 1.8	**Sb** 1.9	**Te** 2.1	**I** 2.5			
VI	**Cs** 0.7	**Ba** 0.9	**La** 1.1	**Hf** 1.3	**Ta** 1.5	**W** 1.7	**Re** 1.9	**Os** 2.2	**Ir** 2.2	**Pt** 2.2
	Au 2.4	**Hg** 1.9	**Ti** 1.8	**Pb** 1.8	**Bi** 1.9	**Po** 2.0	**At** 2.2			
VII	**Fr** 0.7	**Ra** 0.9	**Ac** 1.1							

In Summary



Atomic radius increases
Ionization energy decreases
Electronegativity decreases

Atomic radius decreases
Ionization energy increases
Electronegativity increases