Office Hours: M 1:30pm – 3:00pm
(in Park 169) W 11:00am – 12:00pm
or by appointment (email me)
or just email me with questions!

TA Hours w/ Dana: Th 6:00pm – 7:00pm
(in Park 273) F 4:30pm – 6:00pm

Q-center: https://www.brynmawr.edu/inside/offices-services/q-center/q-center-schedule

Brief History of Atomic Theory

Atoms and the Periodic Table

9/9/2026

Achieve HW #2 due Monday 9/14 at 10:00am

Quiz #1 in Recitation Friday (9/11) – all material up through unit conversions

Chemistry ice cream social

This Friday 4pm in Park atrium courtyard

Learning Objectives

- Outline milestones in the development of modern atomic theory
- Summarize and interpret the results of the experiments of Thomson, Millikan, Rutherford, and Chadwick

- Write and interpret symbols that depict the atomic number, mass number, and charge of an atom or ion
- Describe the three subatomic particles that compose atoms
- Define isotopes and give examples for several elements

- Write and interpret symbols that depict the atomic number, mass number, and charge of an atom or ion
- Define the atomic mass unit and average atomic mass
- Calculate average atomic mass and isotopic abundance

- Explain the organization of elements in the periodic table
- Predict the general properties of elements based on their location within the periodic table
  - # of valence electrons, and most common ion formed
- Identify metals, nonmetals, and metalloids by their properties and/or location on the periodic table

Atomic Theory

1808 CE: John Dalton's Atomic Theory expanded on the centuries-old Greek philosophies…

1. **All matter** is composed of **atoms**. → uncuttable pieces.

2. For a given element, all atoms have the same mass and other properties.

3. Different elements have atoms with different properties.

4. A compound consists of atoms of two or more elements combined in whole-number ratios.

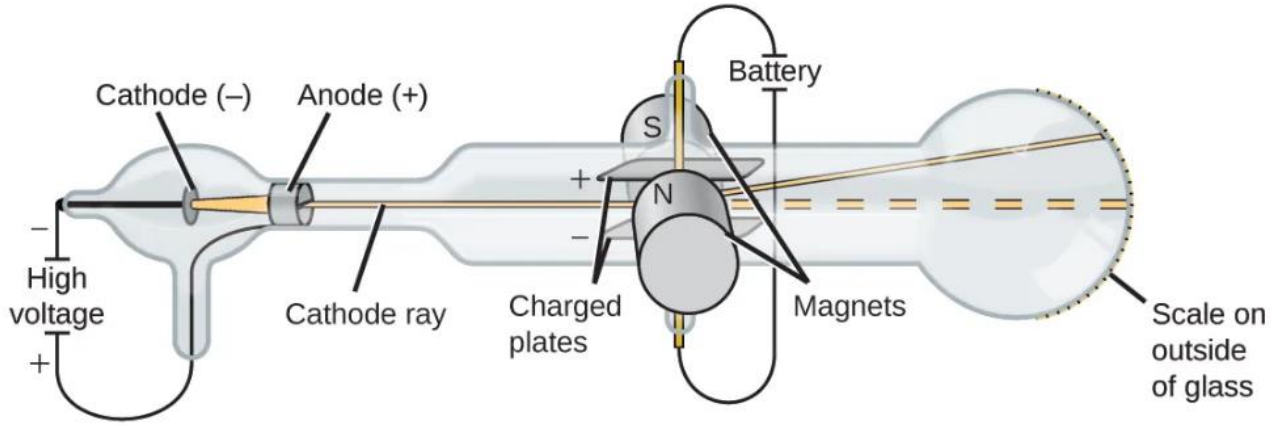
   Example: Water H_2O has 2 atoms of hydrogen for every 1 atom of ~~water~~ O oxygen

5. Atoms are not created or destroyed in a chemical change. They rearrange to make new compounds (with new ratios of atoms).

Revising Dalton's Atomic Theory

→ physicist

<u>1897 CE</u>: J.J. Thompson built on (and challenged) Dalton's atomic theory



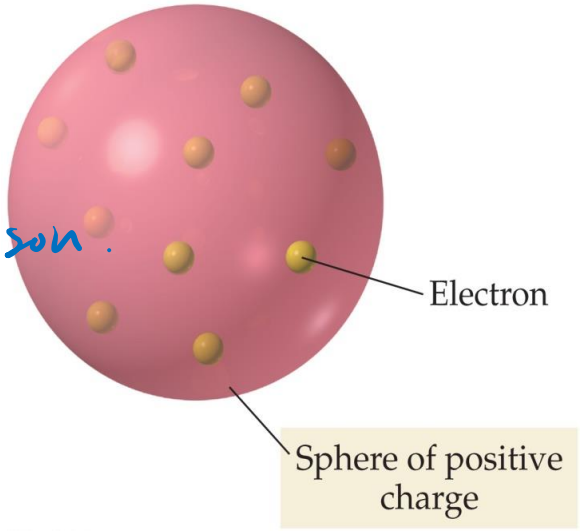
- Electricity applied in a vacuum generates a beam of negatively charged particles.
- Experimentation allowed calculation of the mass-to-charge ratio of these particles.
- They were negatively charged, and much lighter than any atoms that had been observed!
- Atomic theory was revised to account for new observations.

new "subatomic" particle

Plum Pudding Model of the Atom

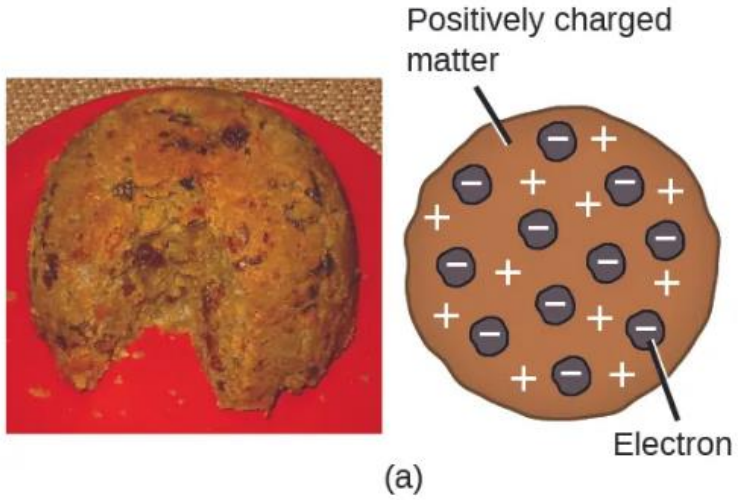
Atoms were known to be neutral so there must be positive charges to balance these negatively charged "electrons"

↳ from J.J. Thompson.



1904 CE: "Plum Pudding Model"

Atom is a sphere of positive charge with negative particles (electrons) mixed in

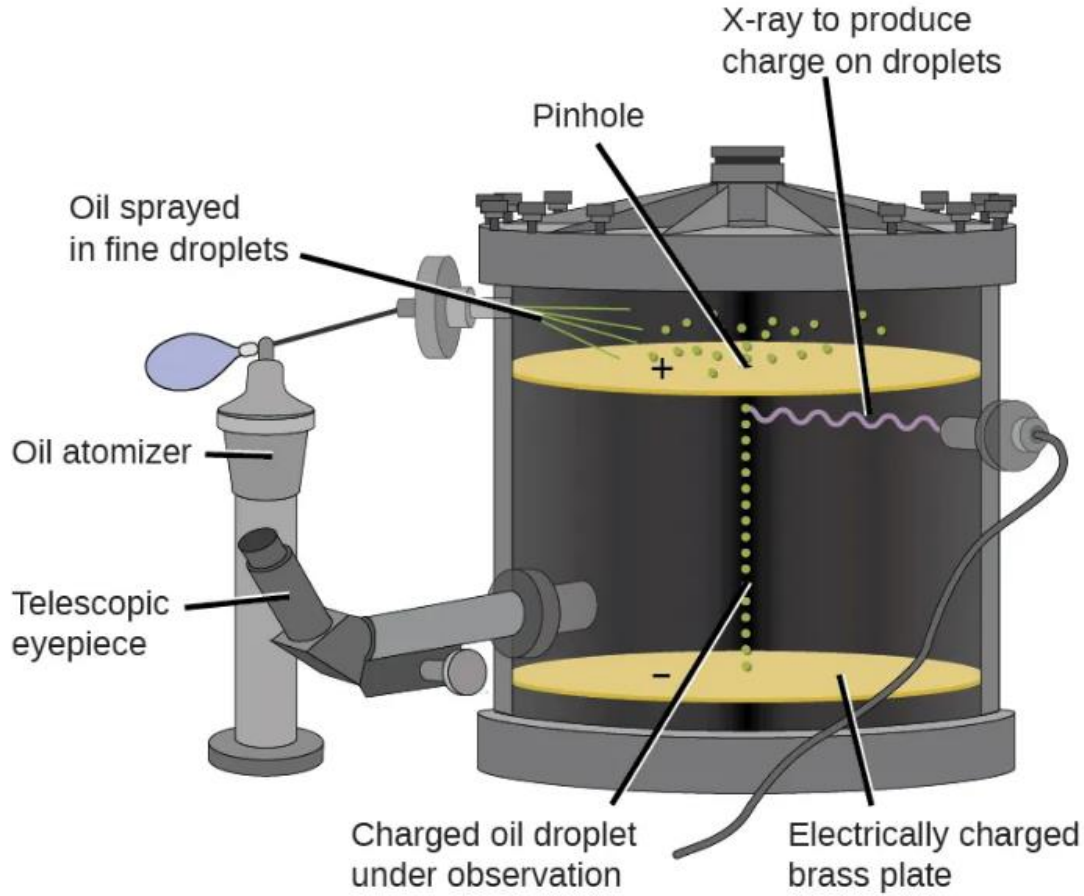


Millikan's Oil Drop Experiment

1909 CE: Further investigation into negatively charged electrons.

Experiments produced tiny oil droplets with slight negative charges

Observed that charge on droplets was always a multiple of the same number!



Oil drop	Charge in coulombs (C)
A	4.8×10^{-19} C
B	3.2×10^{-19} C
C	6.4×10^{-19} C
D	1.6×10^{-19} C
E	4.8×10^{-19} C

Determined the smallest unit of negative charge must be 1.6×10^{-19} C

Combined with Thompson's data, allowed calculation of mass of one electron

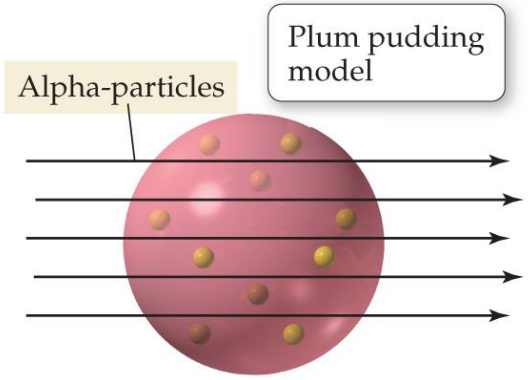
$$\frac{\text{mass (kg)}}{\text{charge (C)}} \times \text{charge (C)}$$

= mass (kg) of one e^-

Rutherford's Gold Foil Experiment

support

ca. 1909 CE: Ernest Rutherford and co. tried to ~~confirm~~ Thompson's Plum Pudding Model



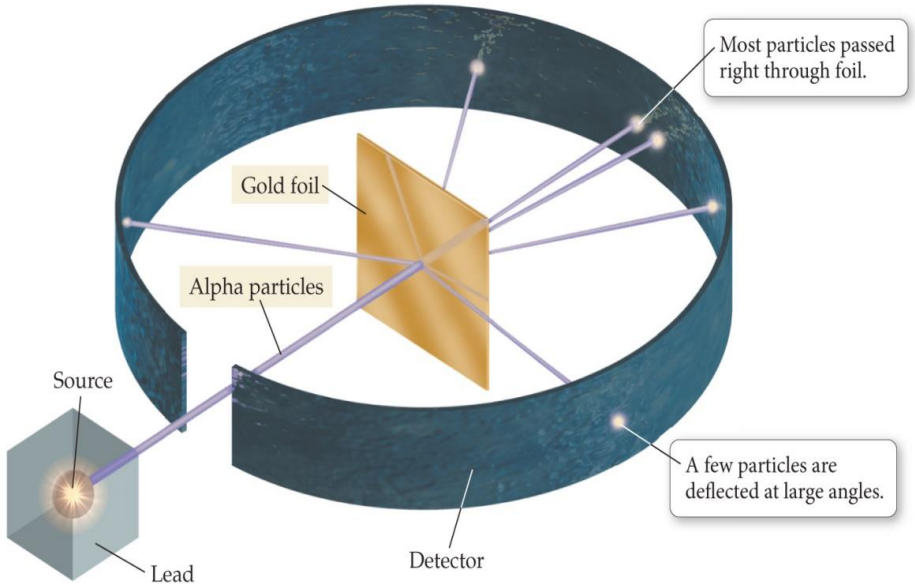
(a) Rutherford's expected result

Tiny alpha-particles were directed at a thin sheet of gold foil.

↳ He atom nucleus

Most of the particles passed directly through the foil.

A few were deflected, some of them at sharp angles.

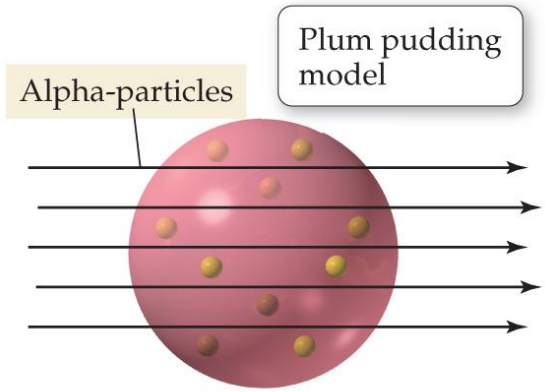


Nuclear Theory of the Atom (1911 CE)

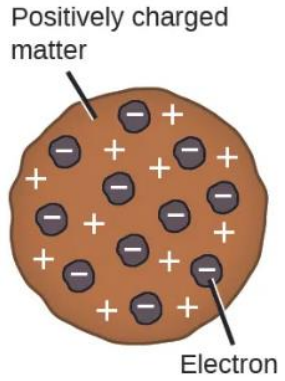
a) Rutherford's Expected result:

Alpha-particles should pass through gold foil with minimal deflection.

Deflections should be roughly uniform.



(a) Rutherford's expected result

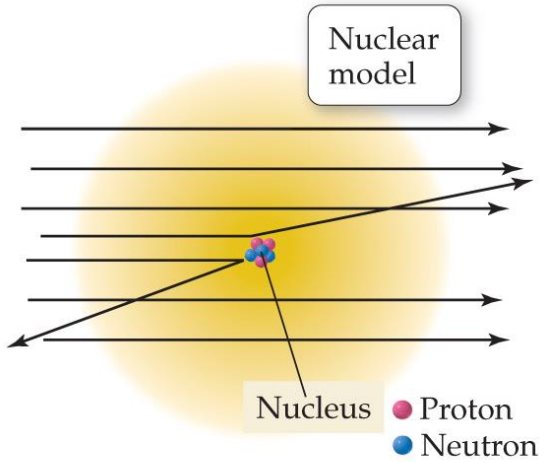


"Plum Pudding" model

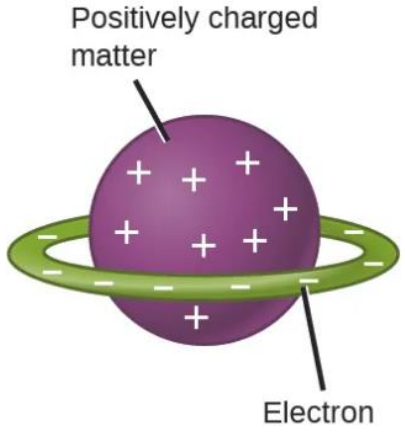
(b) Actual result:

A small number of alpha-particles were deflected or bounced back.

↳ Supported Hantaro Nagaoka's "planetary" model of the atom (1903)



(b) Rutherford's actual result



"planetary" model

Nuclear Theory of the Atom (1911 CE)

↳ about the nuclear (center of atom)

1. Most of atom's mass and positive charge is contained in a small central core called the nucleus

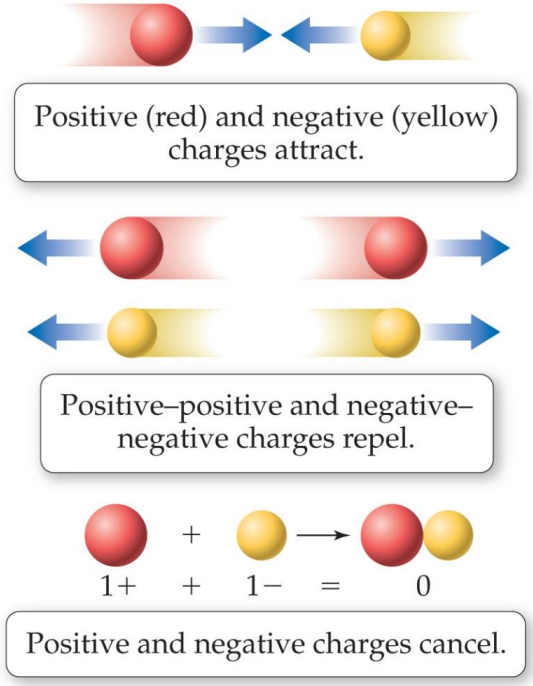
2. Atom is mostly empty space, with small negatively charged electrons interspersed.

3. Amount of positive charge = amount of negative charge (atom is neutral)

# of protons = # of electrons

Experiments showed that nuclei of **all elements** contained Hydrogen nuclei as "building blocks"

Rutherford named this building block of nuclear mass and positive charge the "**proton**."



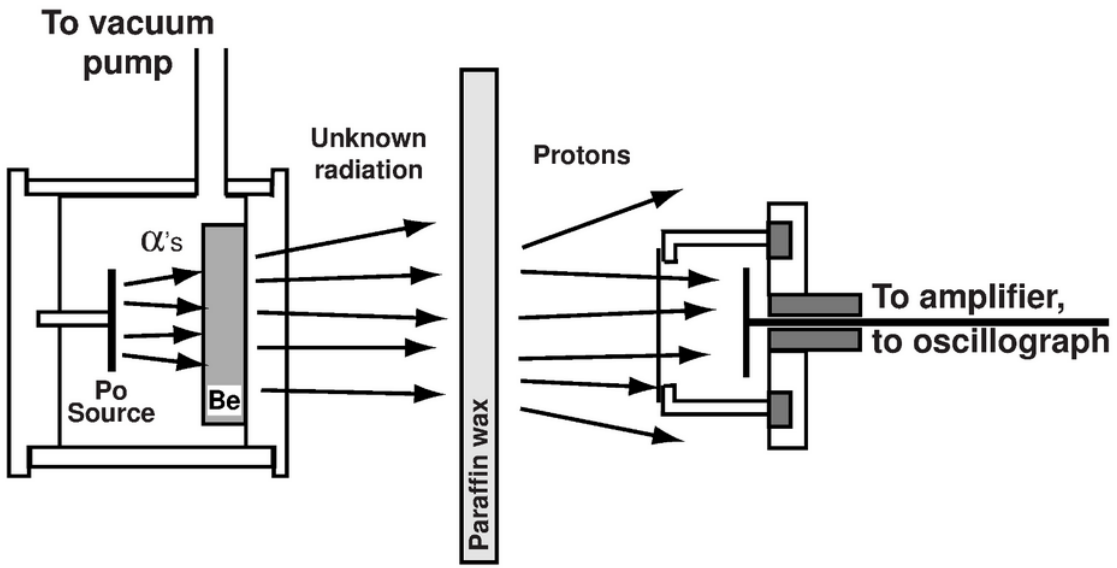
James Chadwick and the Neutron

Experiments to this point had determined:

- Atoms contain negatively (-) charged e^- which have very little mass.
- Most of an atom's mass is in the nucleus
- The nucleus contains positive (+) charged protons
- An atom of Helium is 4x as heavy as Hydrogen, but the nucleus only has 2x as much (+) charge.

Data suggested there was a neutral particle contributing to the nucleus mass!

↑ uncharged



1932 CE: James Chadwick discovered a new type of radiation.

Particles had similar mass to a Hydrogen nucleus, but no charge.

"neutrons"

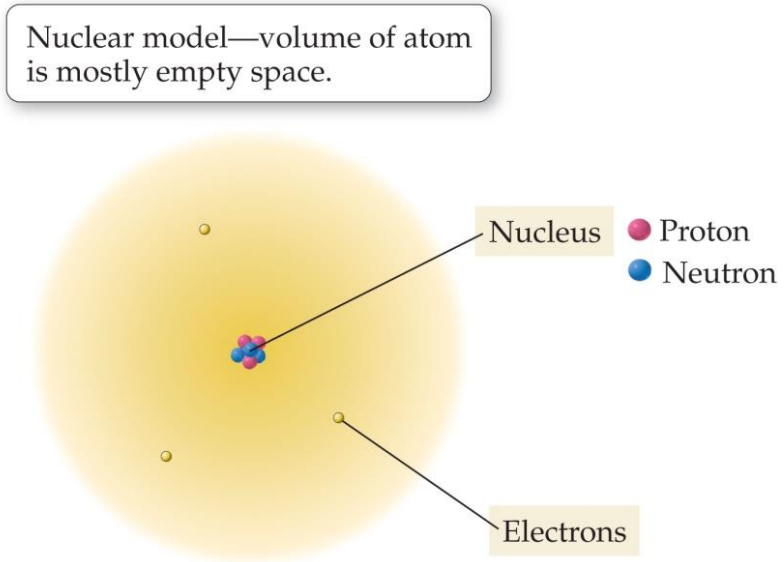
Protons, Neutrons, & Electrons (Subatomic Particles)

Protons (+ charge) and neutrons (neutral) make up nucleus (99.9% of atom's mass)

Electrons (- charge): cloud of negatively charged particles around nucleus; very little mass

Atomic mass unit (amu): convenient unit for the mass of subatomic particles

$$1 \text{ amu} = 1.66 \times 10^{-27} \text{ kg}$$



Based on 1/12 mass of a ^{12}C atom.

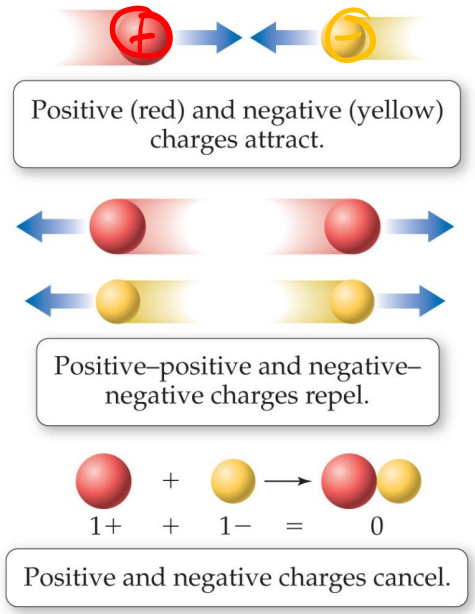
amu is also called Dalton, especially by biochemists.

Subatomic Particles

~ 1 amu.

	Mass (kg)	Mass (amu)	Charge
Proton	1.67273×10^{-27}	1.0073	+1
Neutron	1.67493×10^{-27}	1.0087	0
Electron	9.1×10^{-31}	0.00055	-1

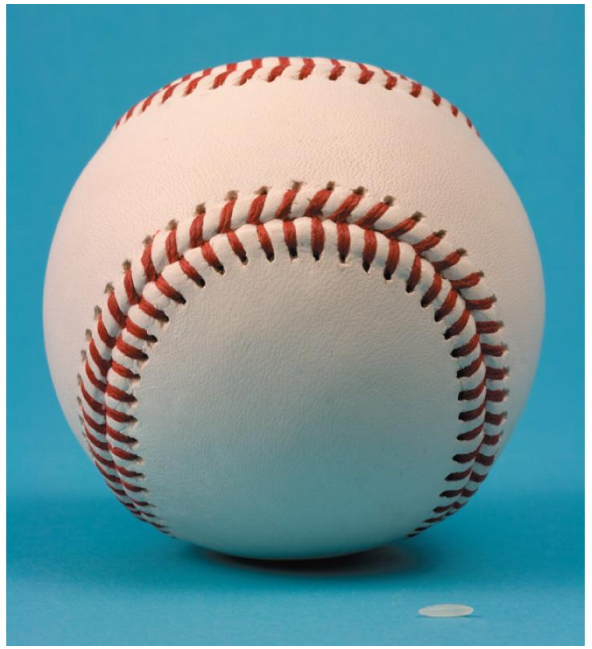
~0 amu



If a proton had the mass of a baseball,

an electron would have the mass of a grain of rice.

A proton is nearly 2000 times more massive than an electron.

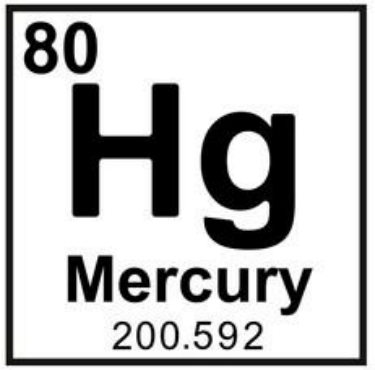


Atomic Numbers & Symbols

Elements are defined by their number of protons

↳ nickname for nucleus w/ 1 proton → Hydrogen.

Atomic Number (Z) = number of protons



Chemical symbol = abbreviation for an element

Example: Carbon (C), Silicon (Si), Lead (Pb), Mercury (Hg)

Some symbols are based on the Latin names

lead – plumbum

mercury – hydrargyrum

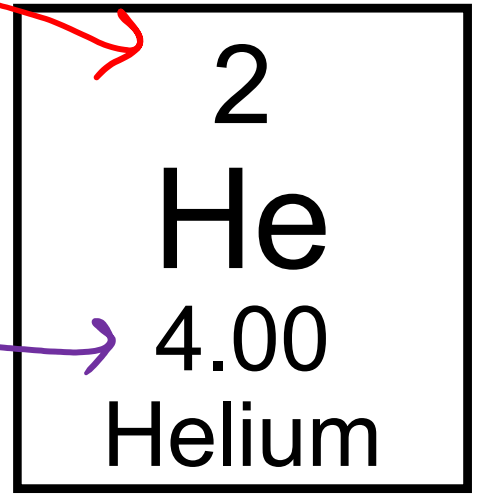
Atomic Numbers & Symbols

2 = Atomic Number (Z)

He = Chemical Symbol

4.00 = Atomic Mass (in amu)

Helium = Name



# of protons

protons + # neutrons

not always shown

Periodic Table

Organized by Dmitri Mendeleev (1869)

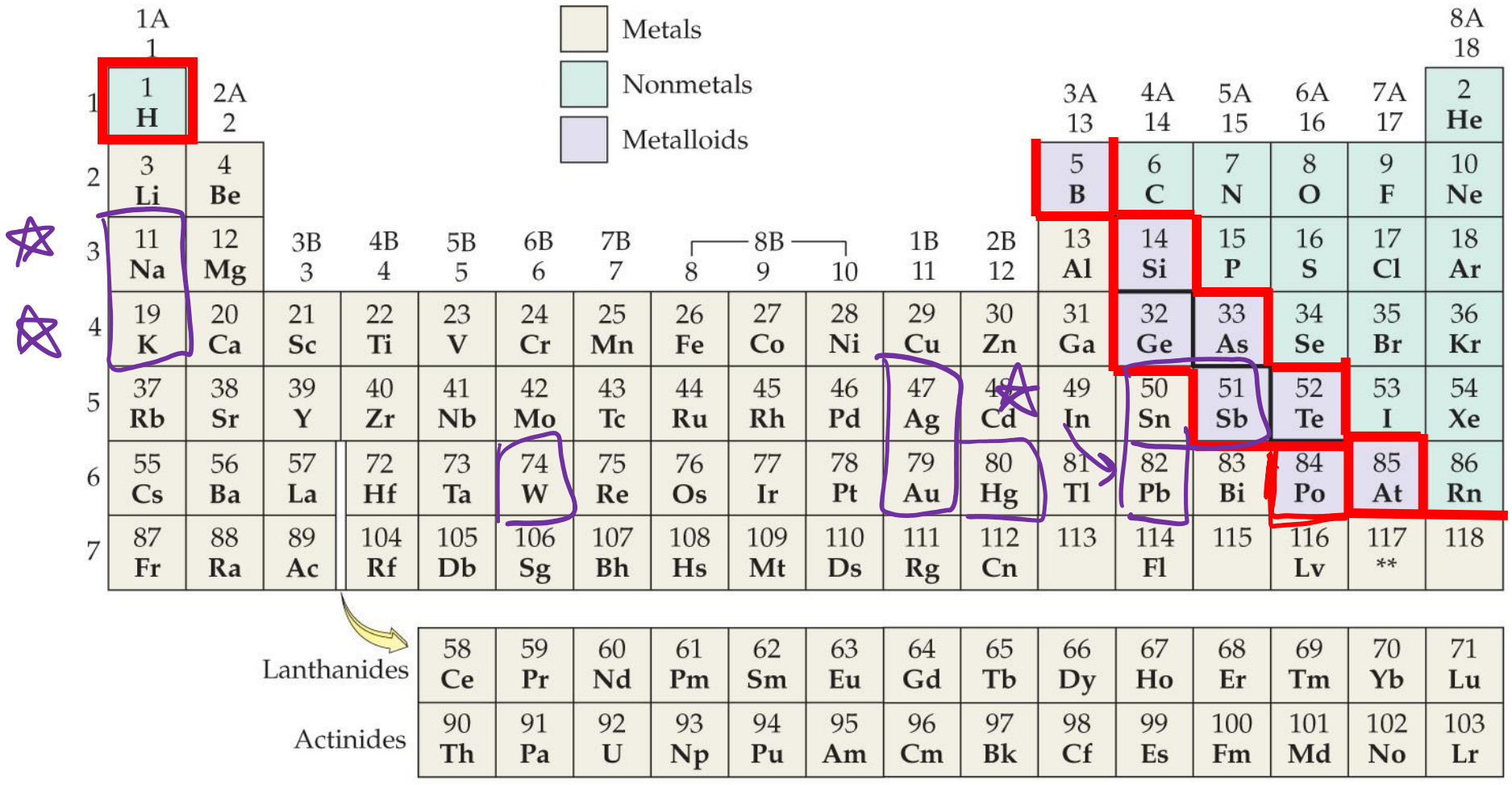
When elements are arranged in order of increasing mass, certain sets of properties recur periodically

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

Arranged elements in rows so vertical columns share similar properties.

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

Locating Metals, Nonmetals, Metalloids



Common confusing symbols

Metals

Electrical - Conductive

Thermal - Conductive

Metallic gray luster

Most are solids**

Malleable - sheets

Ductile – wires

Tend to form (+) ions

Nonmetals

Electrical - Insulator

Thermal - Insulator

No luster (not shiny)

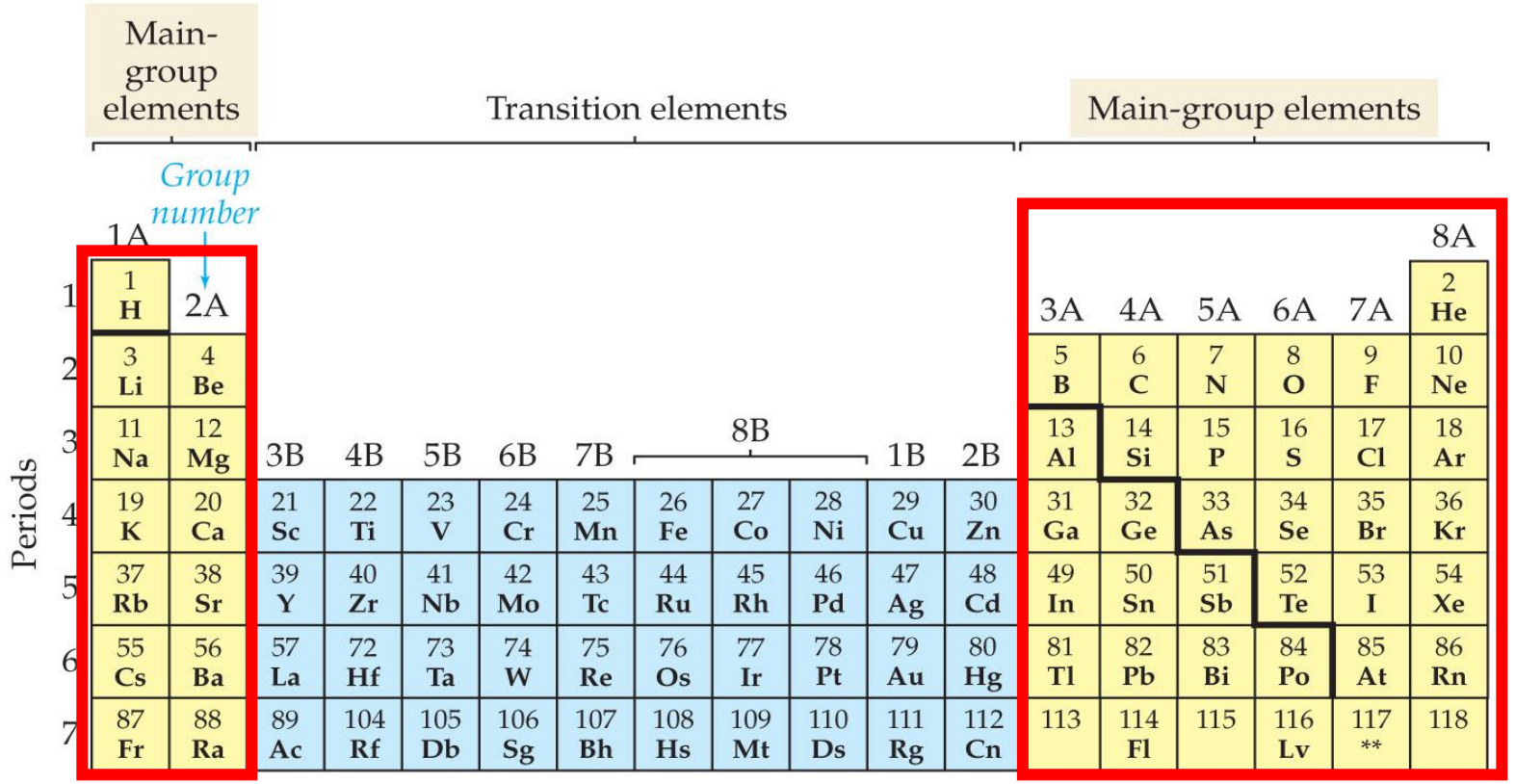
All phases

Brittle

Nonductile

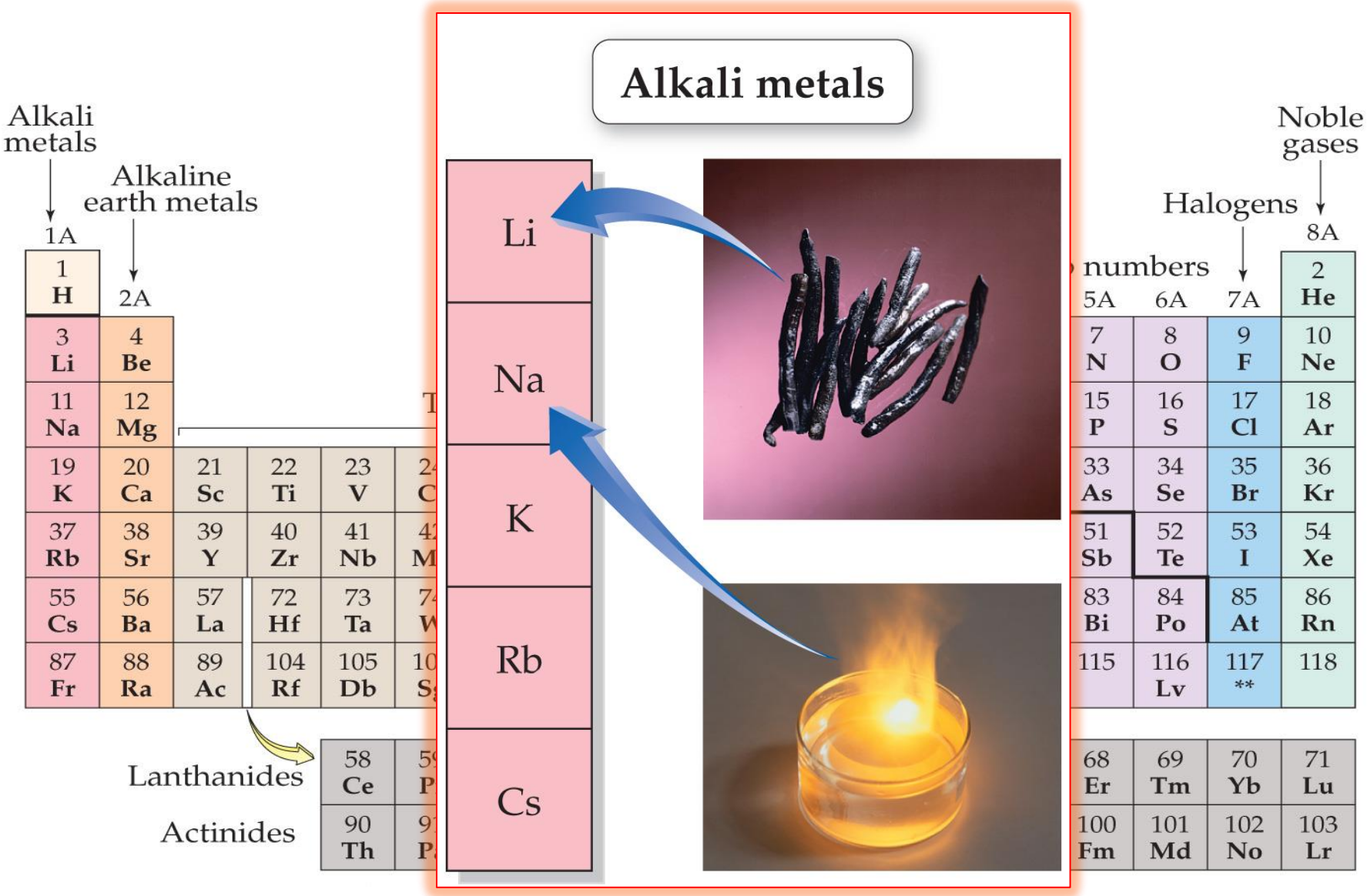
Tend to form (-) ions, but
(+) ions are not uncommon.

Main Group Elements and Transition Elements

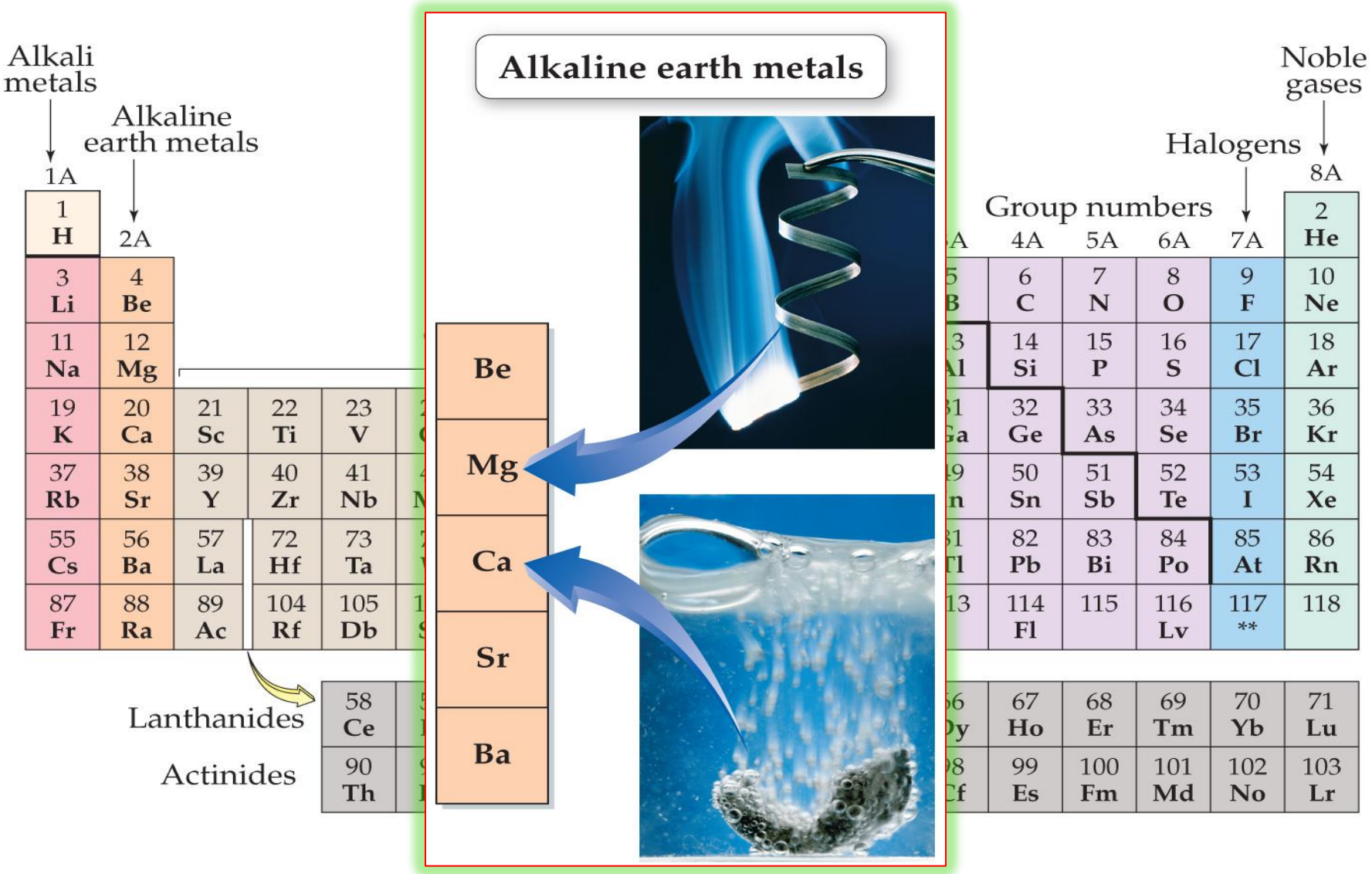


For main group elements, properties can generally be predicted based on the element's position. (column #)

Groups 1A, 2A, 7A, 8A Highlighted



Groups 1A, 2A, 7A, 8A Highlighted



Groups 1A, 2A, 7A, 8A Highlighted

Alkali metals

Alkaline earth metals

Noble gases

Halogens

1A	2A											3A	4A	5A	6A	7A	8A
1 H																	2 He
3 Li	4 Be													7 N	8 O	9 F	10 Ne
11 Na	12 Mg													15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 M								33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 T								51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 R								83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	10 B								115	116 Lv	117 **	118

Transi numbers

Lanthanides

Actinides

			58 Ce	59 Pr	6 N									68 Er	69 Tm	70 Yb	71 Lu
			90 Th	91 Pa	92 U									100 Fm	101 Md	102 No	103 Lr

Halogens

F

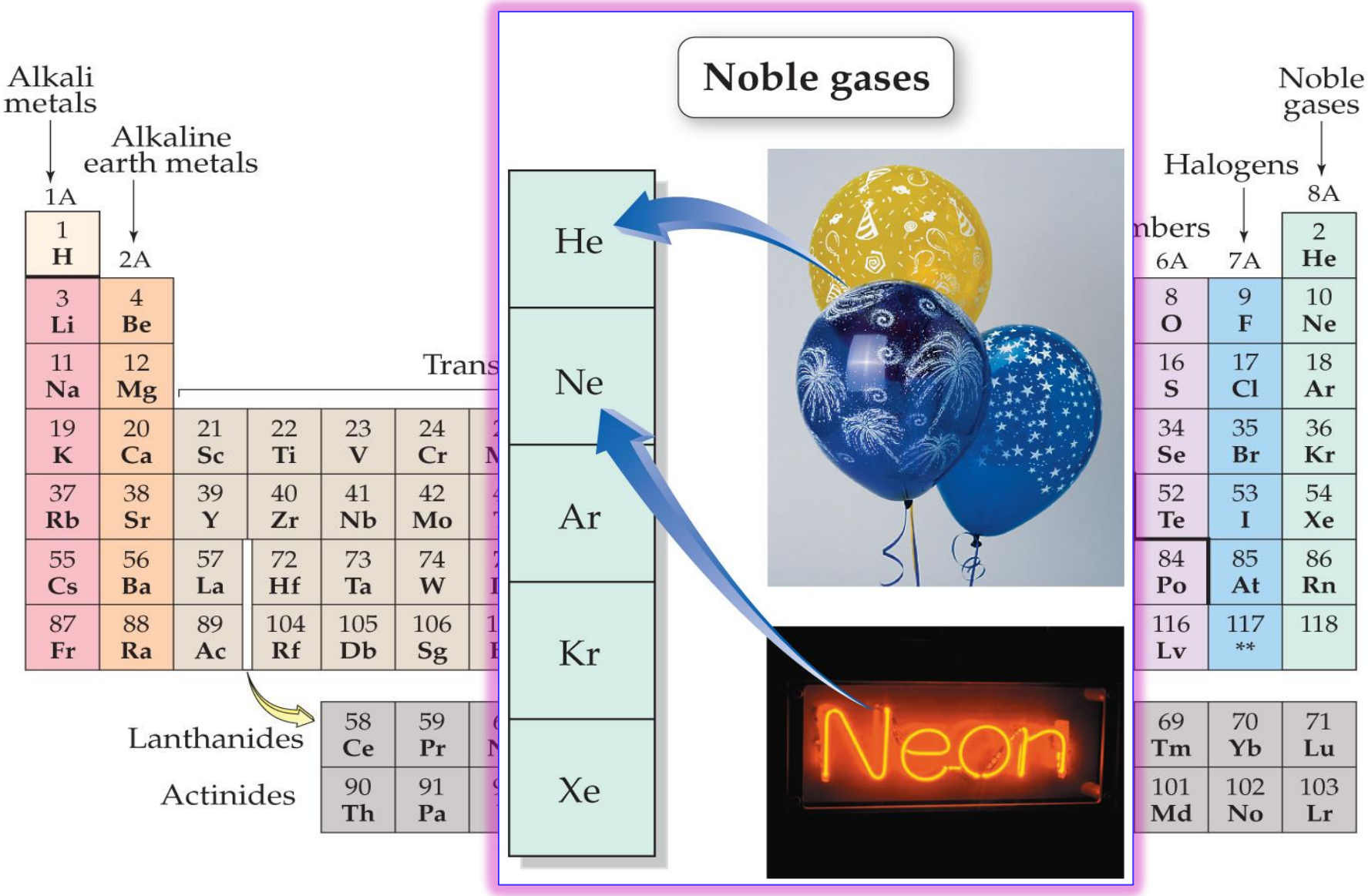
Cl

Br

I

At

Groups 1A, 2A, 7A, 8A Highlighted



Valence Electrons

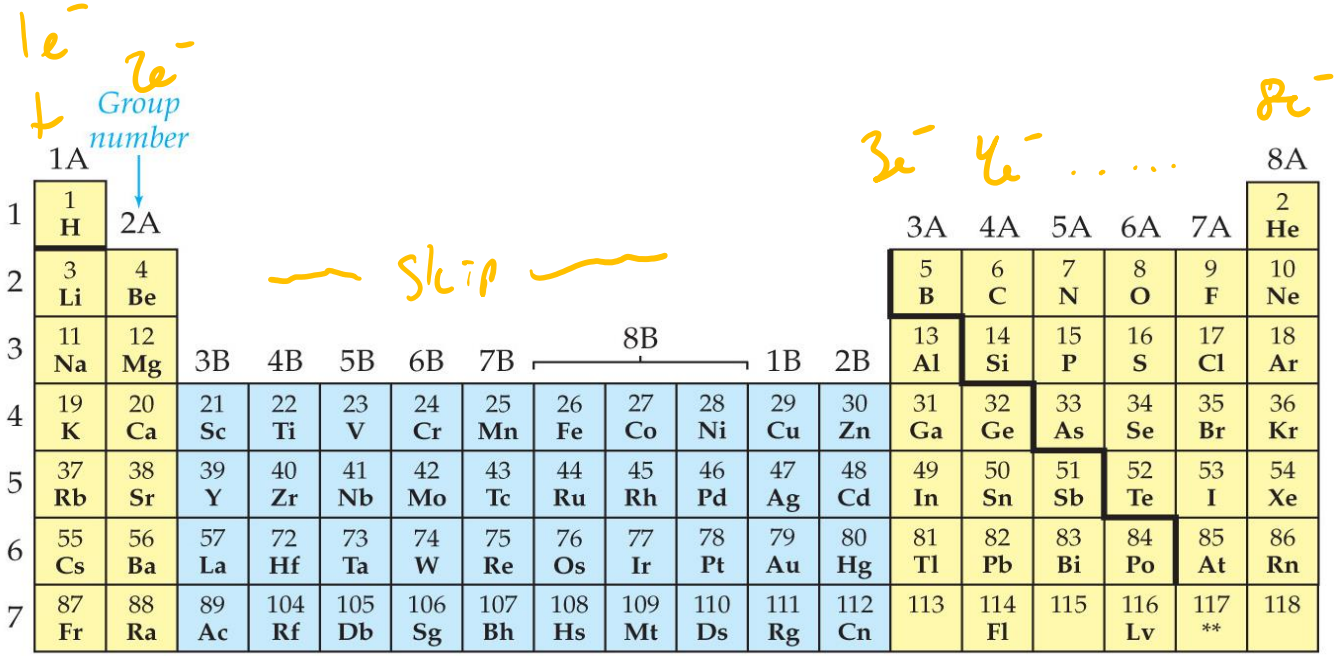
Valence electrons: located in the outer most shell of an atom.

have the highest energy of all electrons in an atom.

responsible for the **reactivity** and other properties of an atom.

Main Group Elements have same number of valence electrons as their group number (column)

Li has 1 valence e⁻, Be has 2 valence e⁻, B has 3 valence e⁻, …



Atom

Protons, neutron and electrons

Protons = # Electrons

P = e

overall, no charge

Atoms are neutral particles

Ion

Protons, neutron and electrons

Protons ≠ # Electrons

Ions are charged particles

Cations vs Anions

Cation – positively charged particle (atom has lost one or more electrons)

Ex: $Li \rightarrow Li^{+} + 1e^{-}$ Lithium ion charge = 1+

Anion – negatively charged particle (atom has gained one or more electrons)

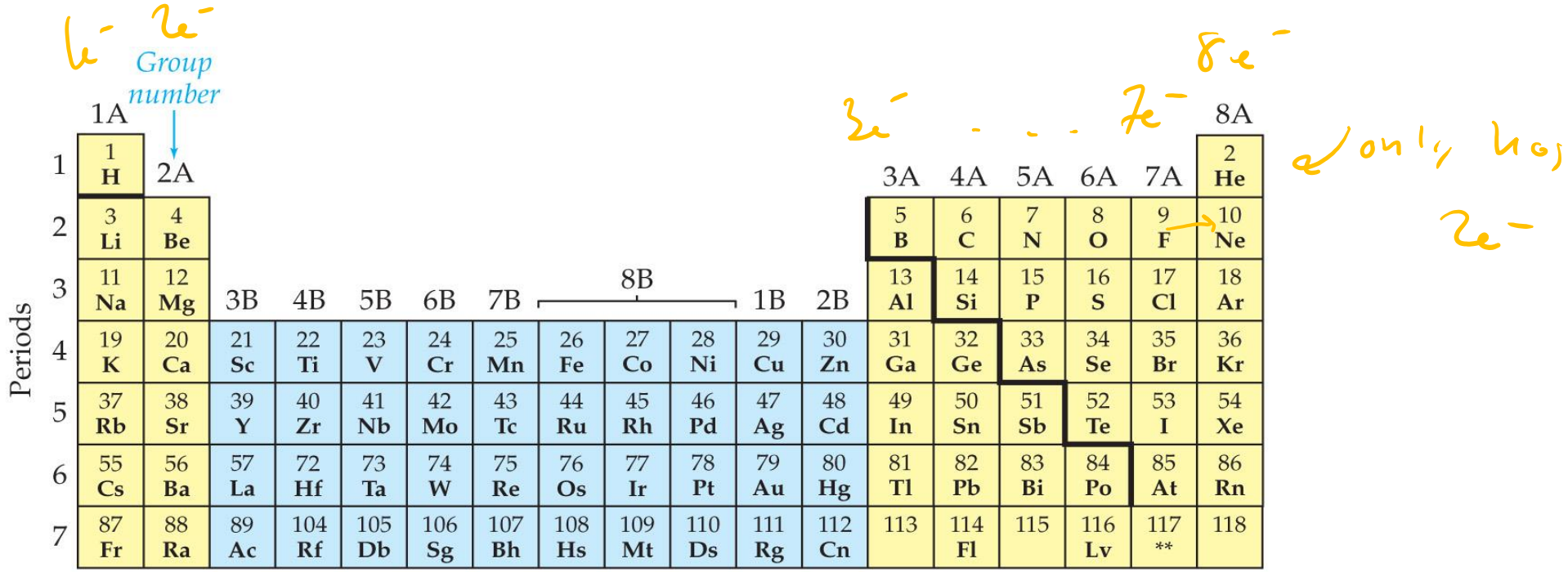
Ex: $F + 1e^{-} \rightarrow F^{-}$ Fluorine ion charge = 1-

Ion Charge = # protons - # electrons

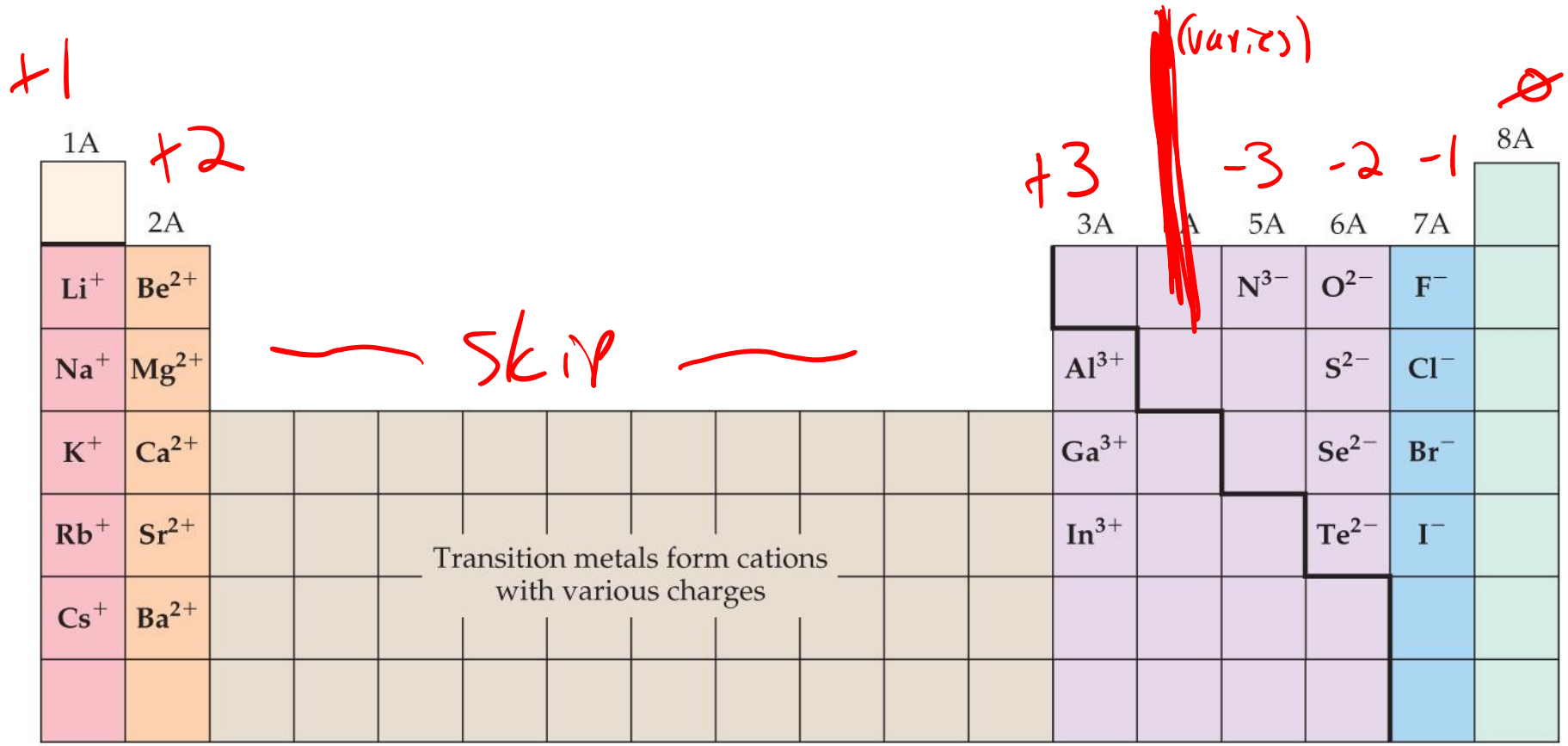
Ions & Periodic Table

Main group elements tend to form ions that have the same number of **valence electrons** as the nearest noble gas.

- Li tends to lose 1 electron (become Li^{+}) to become like He
- Be tends to lose 2 electrons (become Mg^{2+}) to become like He
- F tends to gain 1 electron (become F^{-}) to become like Ne
- O tends to gain 2 electrons (become O^{2-}) to become like Ne



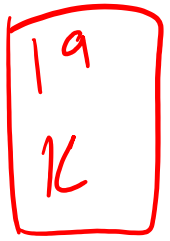
Ions and the Periodic Table



This is one of the most important slides in the entire class!!!

(memorize these charge rules ASAP, they will come up a lot)

(a) How many protons does potassium (K) have?



19

(b) What type of ion will K normally form?

+1

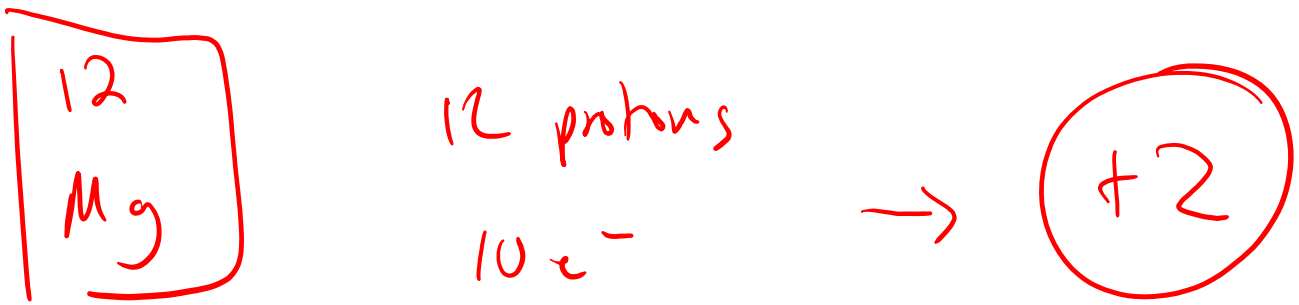
(c) Write a reaction showing a K atom forming this ion.

$K \rightarrow K^+ + e^-$

(d) What group/family does K belong to?

group 1 (alkali metals

Determine the charge of a magnesium ion with 10 electrons

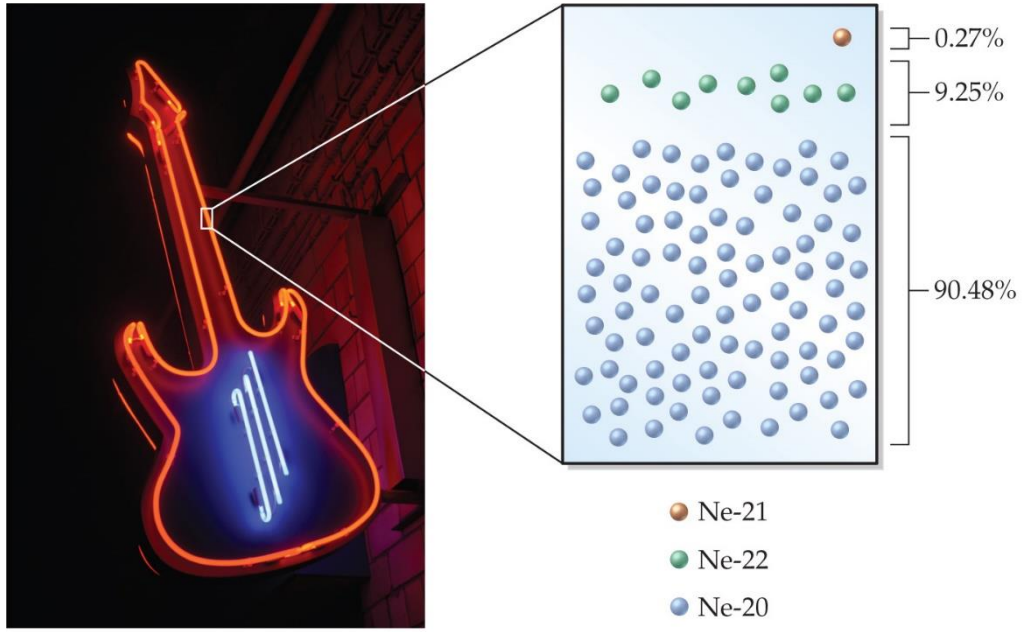


Isotopes: When the Number of Neutrons Varies

- All atoms of a given element have the same number of protons. atomic # (Z)
- They do not necessarily have the same number of neutrons. ∴ therefore mass
- Atoms with the same number of protons but different numbers of neutrons are called **isotopes**.

  e.g., Carbon-12 vs. carbon-13
- All elements have their own unique percent **natural abundance** of isotopes.

Isotopes: Natural Abundance of Isotopes in Neon

Naturally occurring neon contains three different isotopes:

	# protons	# neutrons	Mass Number (A)	Percent Abundance
Ne-20	10	10	20	90.48%
Ne-21	10	11	21	0.27%
Ne-22	10	12	22	9.25%