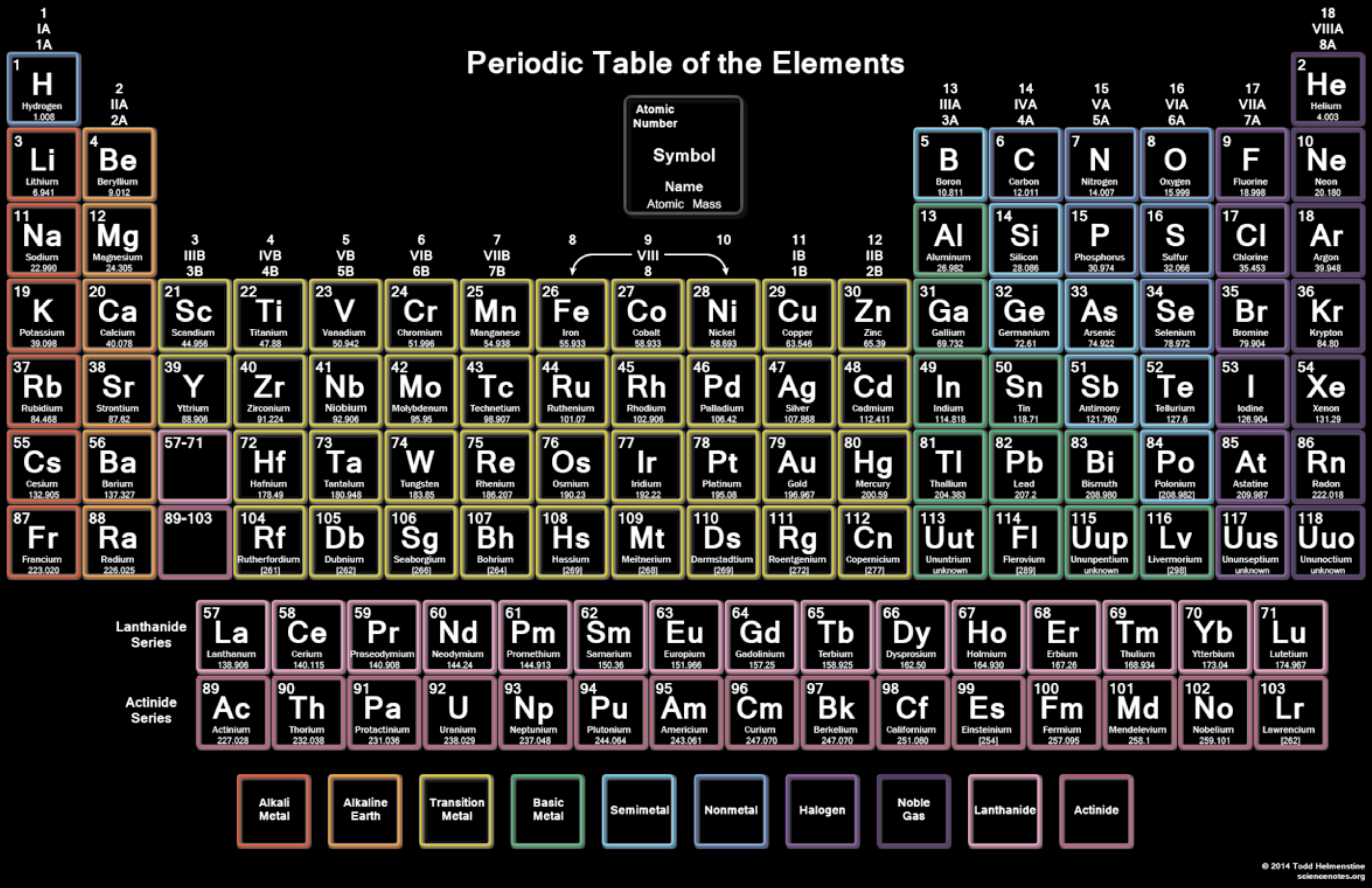
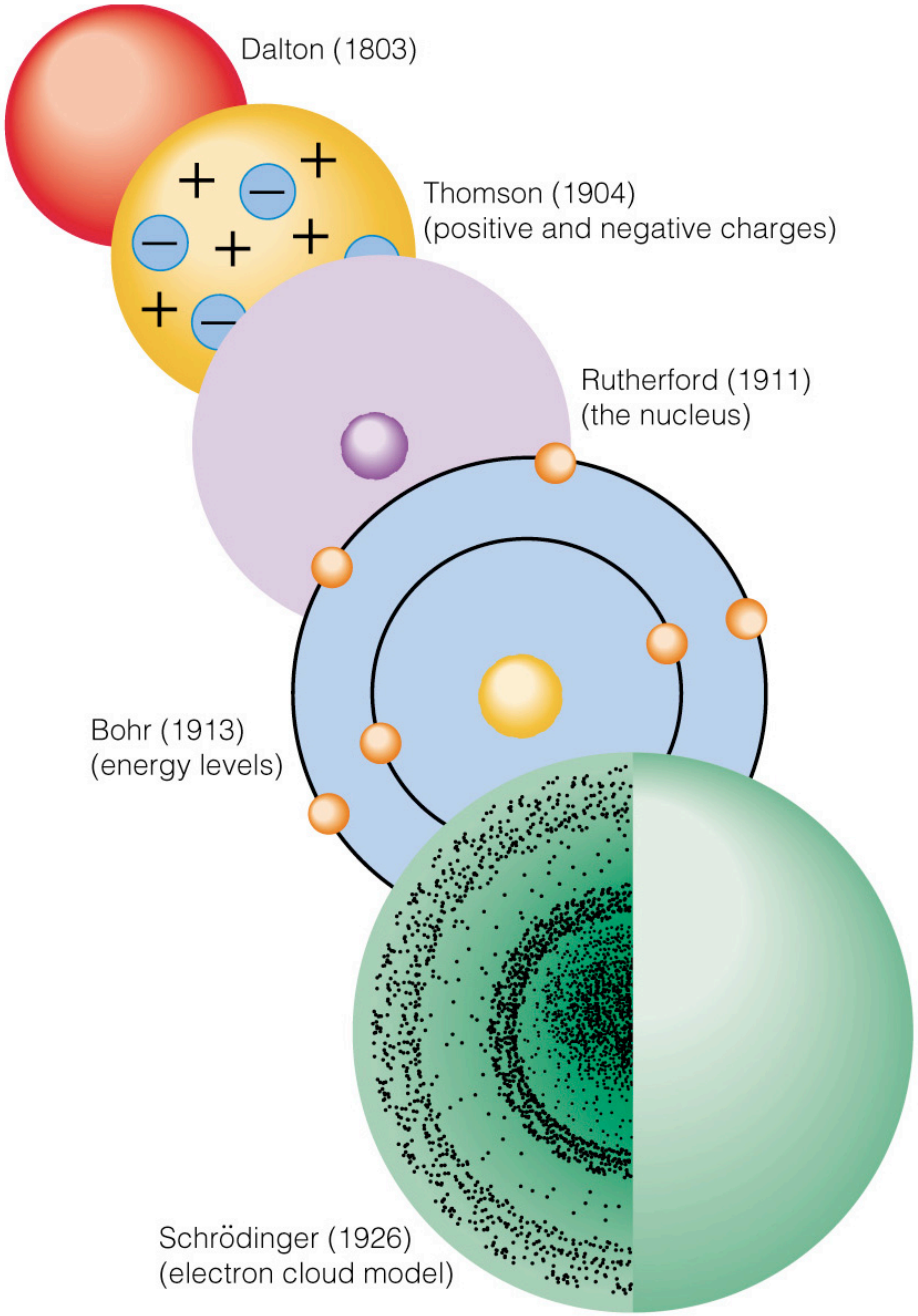
Chapter 08 Atoms and Periodic Properties



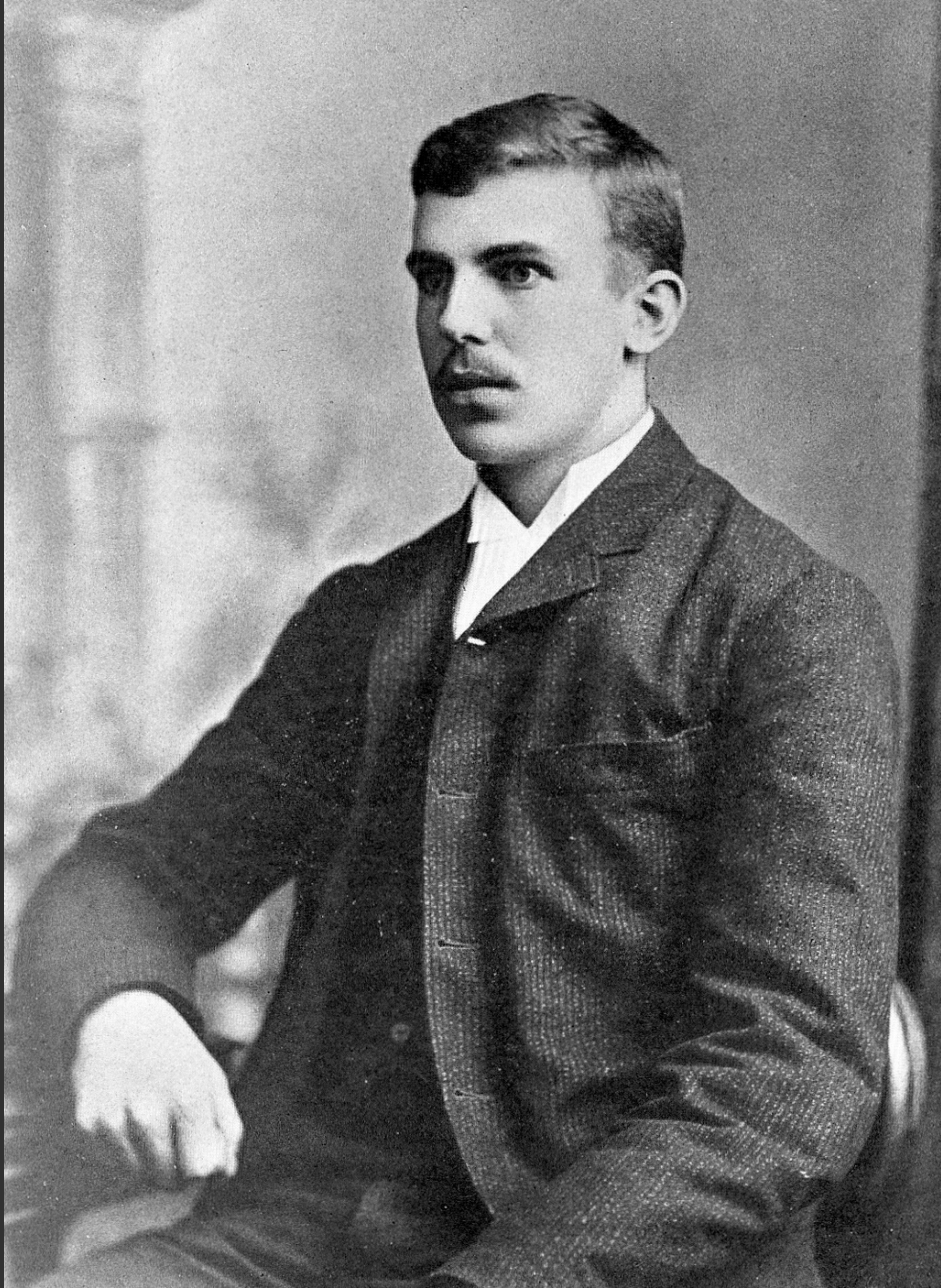


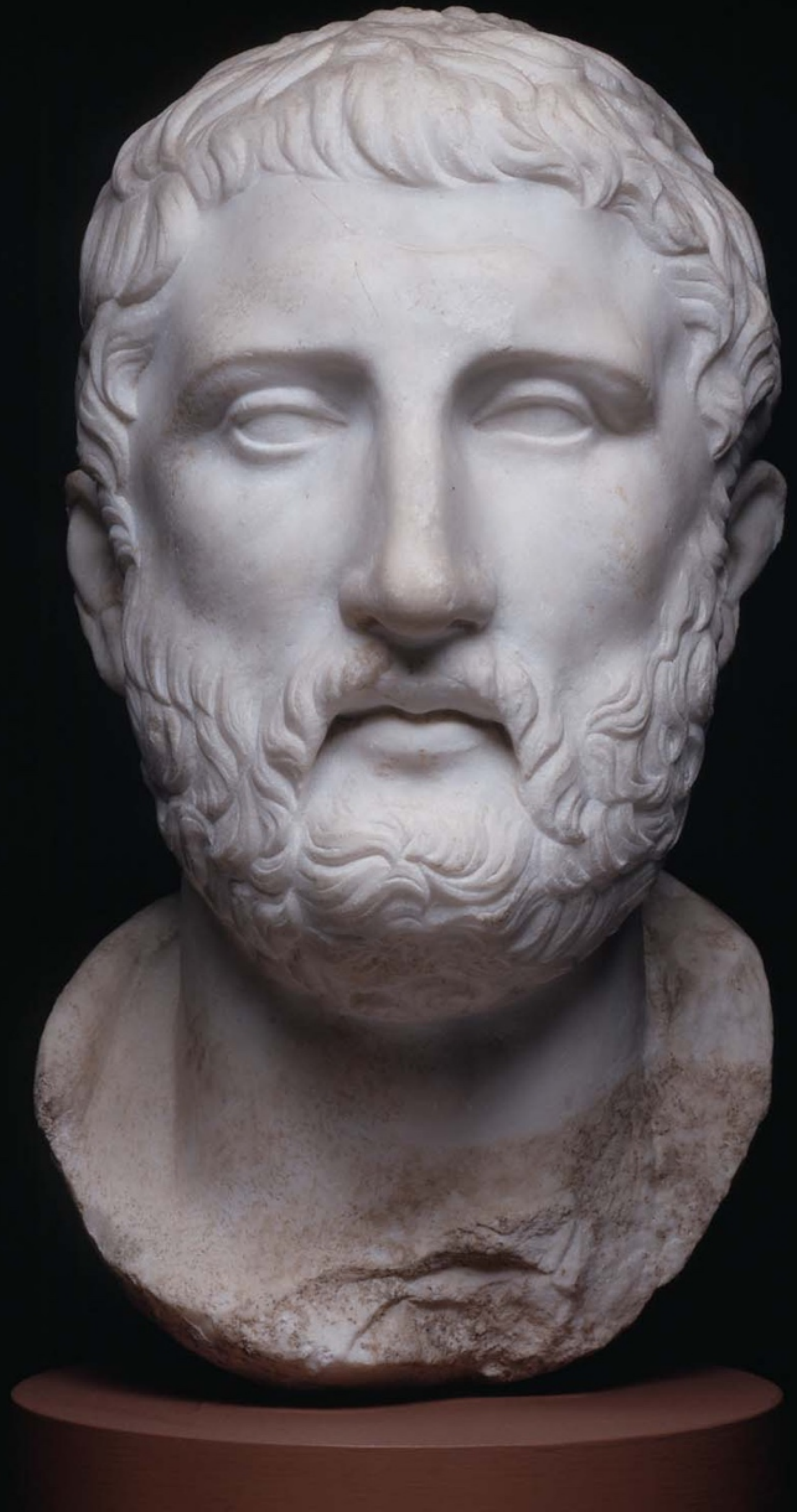
Section 8.1

Atomic Structure Discovered

True or false:

The idea that all matter is composed of atoms is pretty new, dating back only as far as the early 1900s.



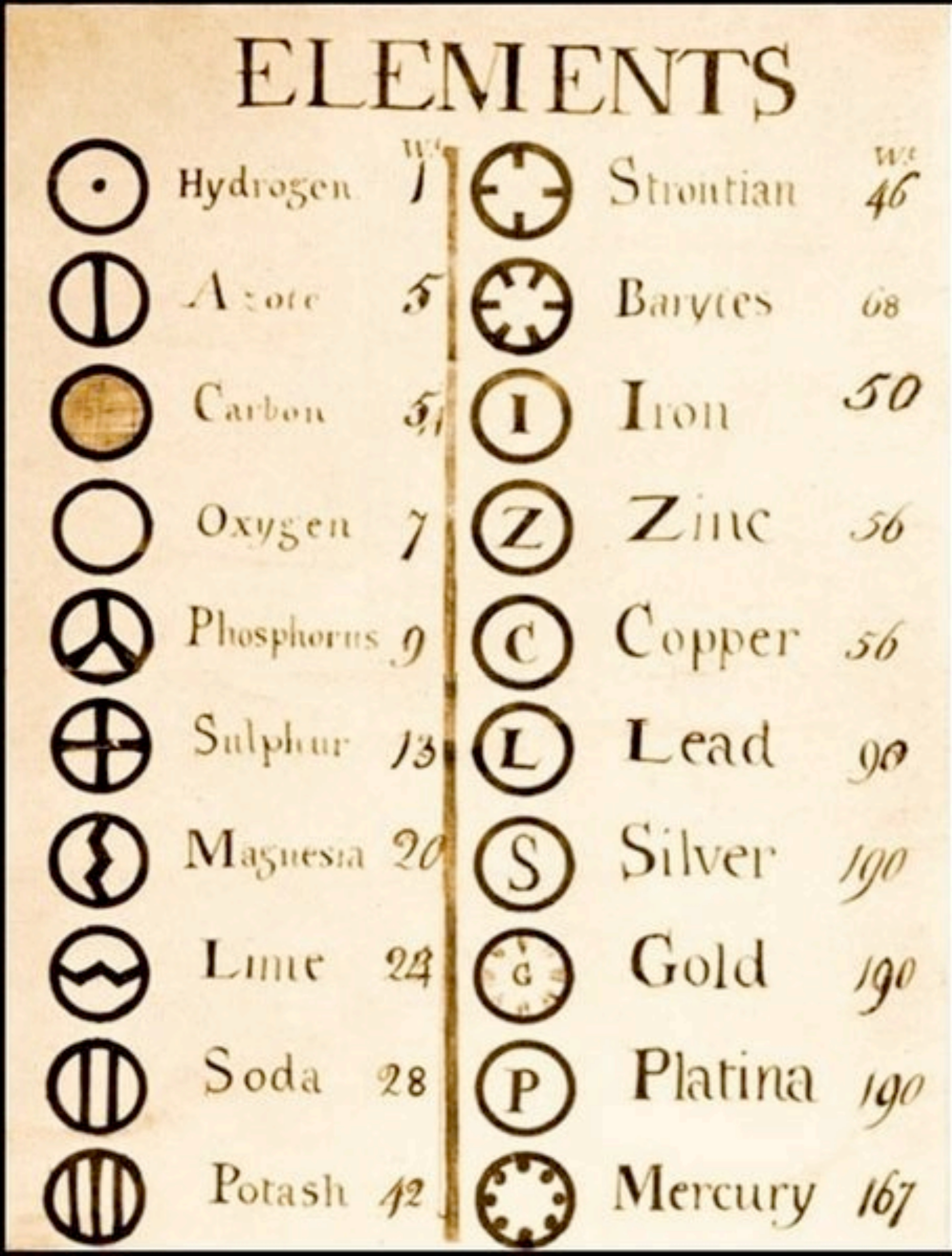


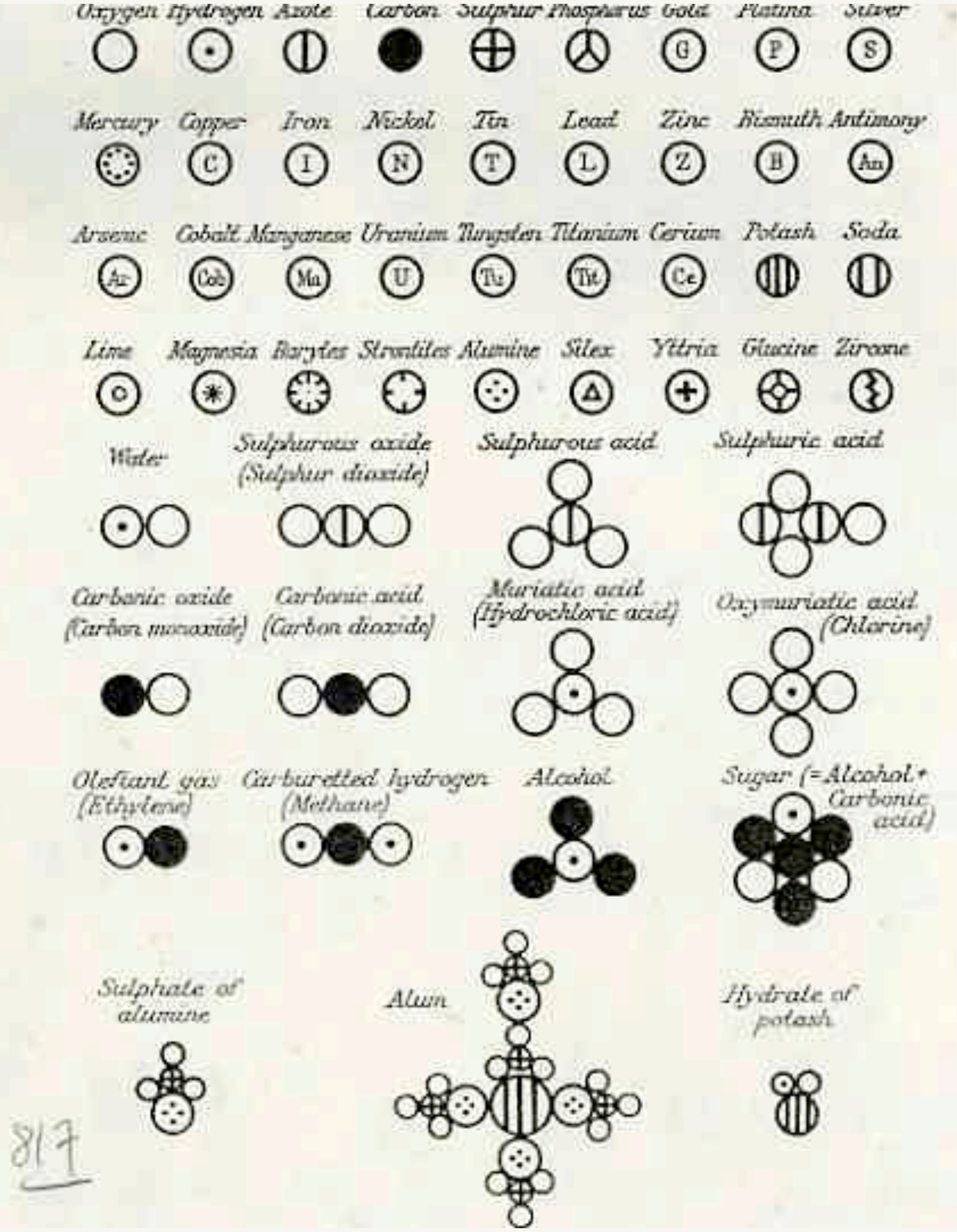
Democritus' Atom

- Ancient Greek, about 350 BCE
- Postulated that everything was made of tiny, invisible, indivisible particles
- Typically visualized as tiny, uniform hard spheres
- Thought experiment: If you break a bar of gold in half, it's still gold. Break it again, keep breaking it in half. How far can you go and have it still be gold?

Dalton's Experiments

- Early 1800s: Up to this point, no real progress on atoms—no context or framework for making progress
- Performed chemical experiments that proved that substances existed in specific, measurable ratios
- Was able to measure atomic weights relative to hydrogen for oxygen, nitrogen, carbon, sulfur, and phosphorus





Dalton's Atomic Hypotheses

- Indivisible minute particles called atoms make up all matter
- All the atoms of an element are exactly alike in shape and mass (Disclaimers apply, see isotopes)
- The atoms of different elements differ from one another in their masses
- Atoms chemically combine in definite whole-number ratios to form chemical compounds
- Atoms are neither created nor destroyed in chemical reactions

By 1803, Dalton

A) had discovered all 117 elements on the periodic table.

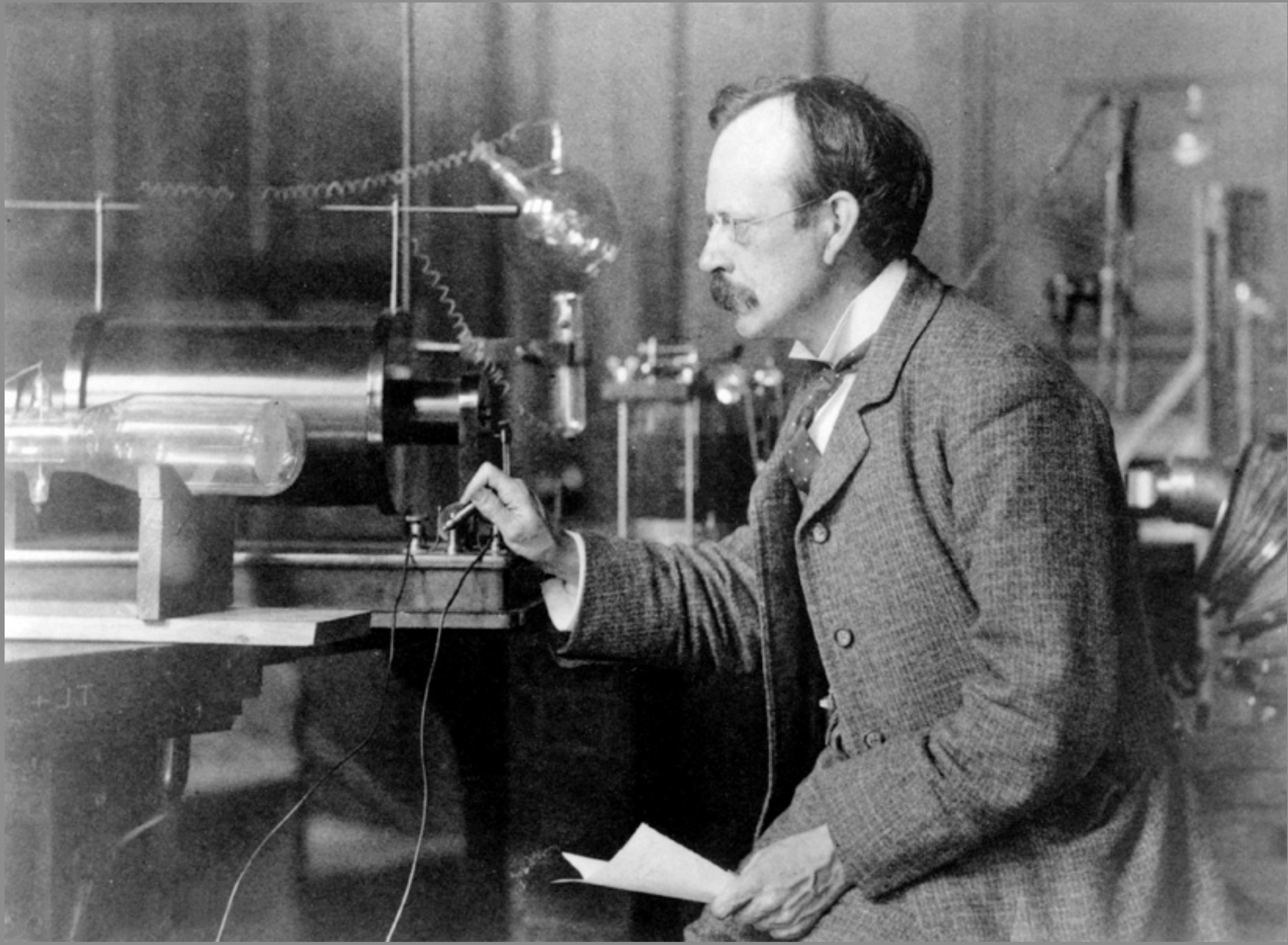
B) had proven that electrons were tiny, and located within the nuclei of atoms.

C) had measured the charge and mass of both electrons and protons.

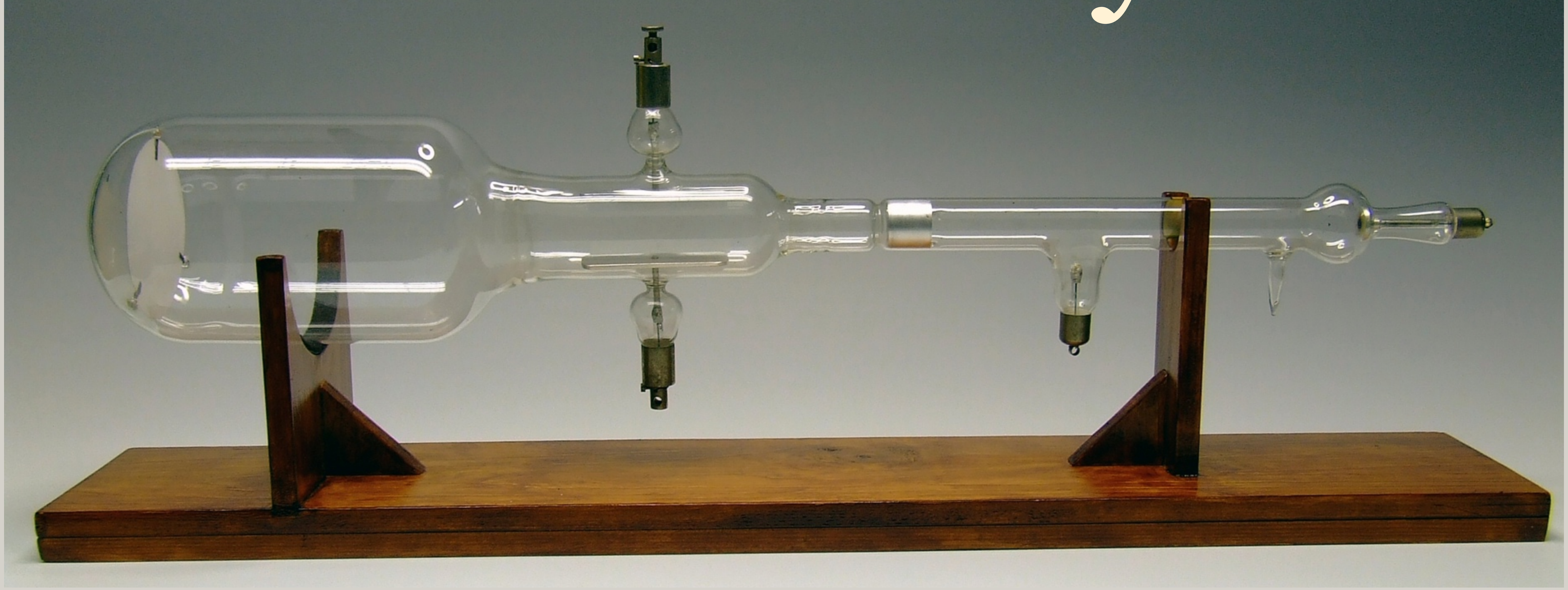
D) was still far ahead of his time, and unable to measure the absolute mass of the atoms he postulated, or discover the structure of the atoms themselves.

Electrons…Finally!

- Existence of charge well understood, but e– not actually discovered until 1897
- If you understand gravity (thanks, Newton!) and E and B fields (thanks, Maxwell!), then you can manipulate them
- You can also follow a recipe, even if you don't understand why it works…



Cathode Rays



- Never-fail cathode ray recipe: Place two metallic terminals in a vacuum tube. Evacuate tube. Apply high voltage. Watch a mysterious green ray shoot across the tube.
- Easy to show that the beam of cathode rays has a negative charge (opposites attract, likes repel)
- Shoot a beam of cathode rays through a pair of known E and B fields, and measure how the path changes in response to controlled field changes
- Adjust E and B fields to balance our gravity, and you just figured out the charge/mass ratio of your particles

Cathode ray tube?

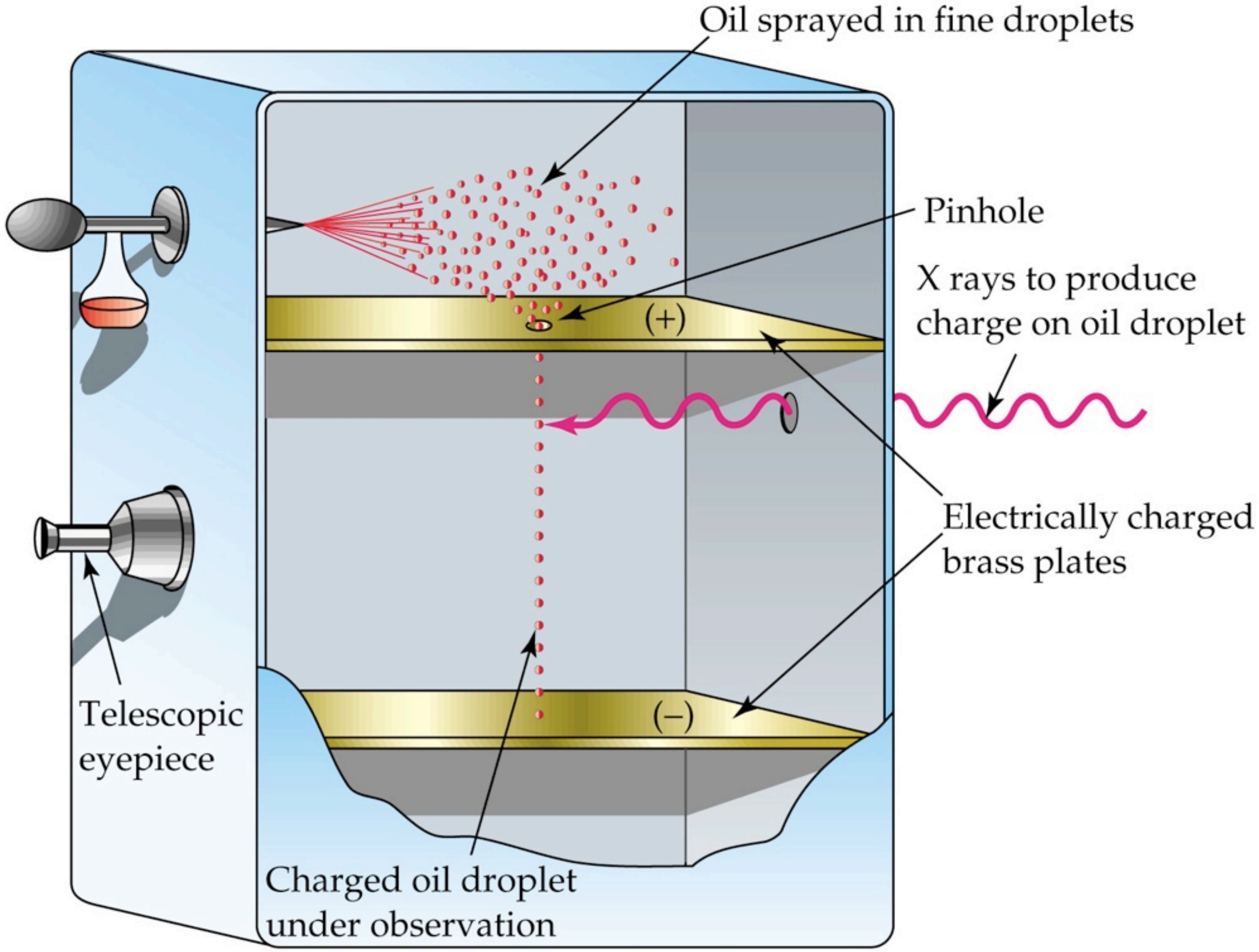
A) Old-school televisions!

B) Old-school computer monitors!

C) Old-school oscillo-scopes!

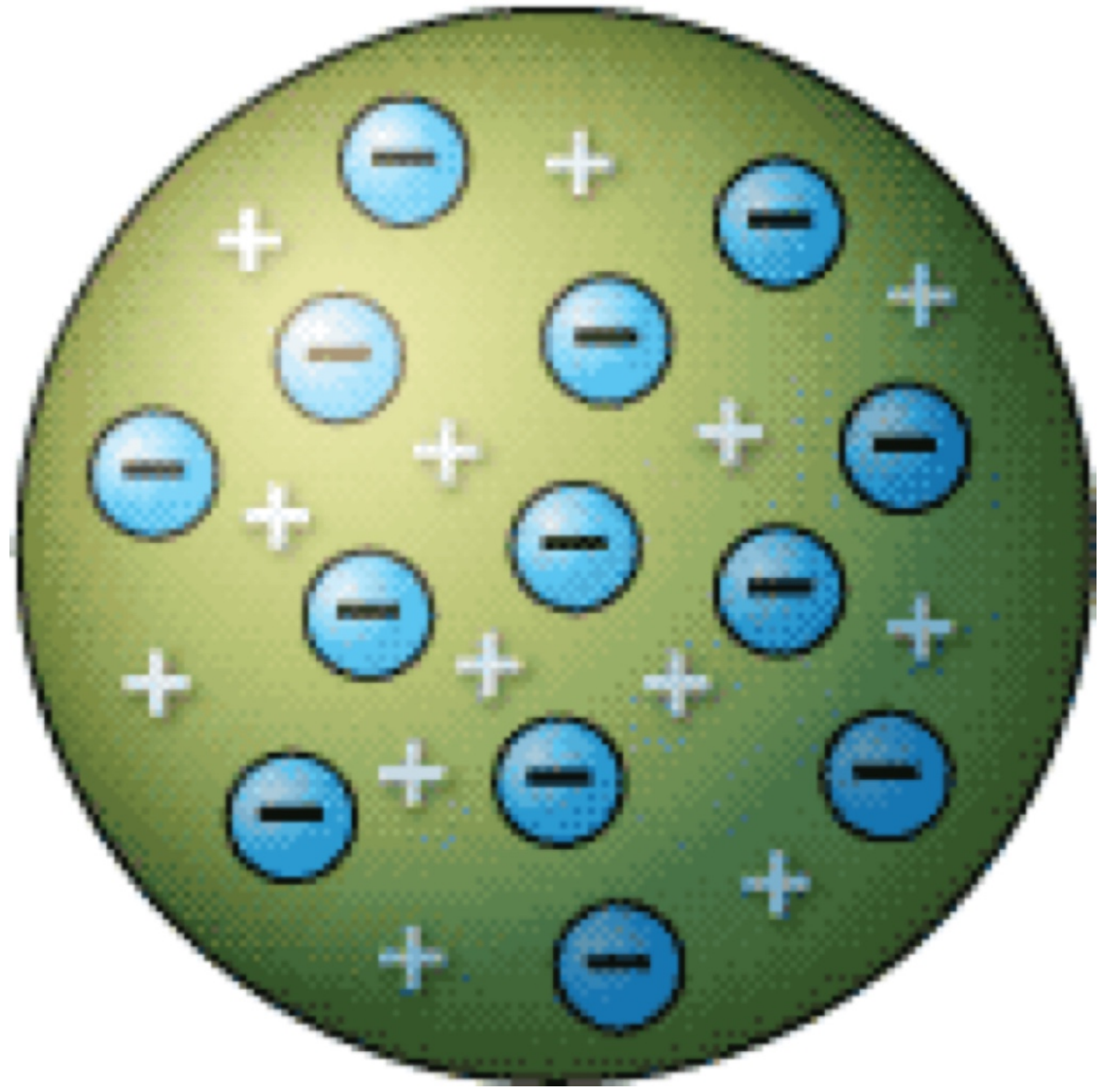
D) All of the above!

Electrons…Finally!



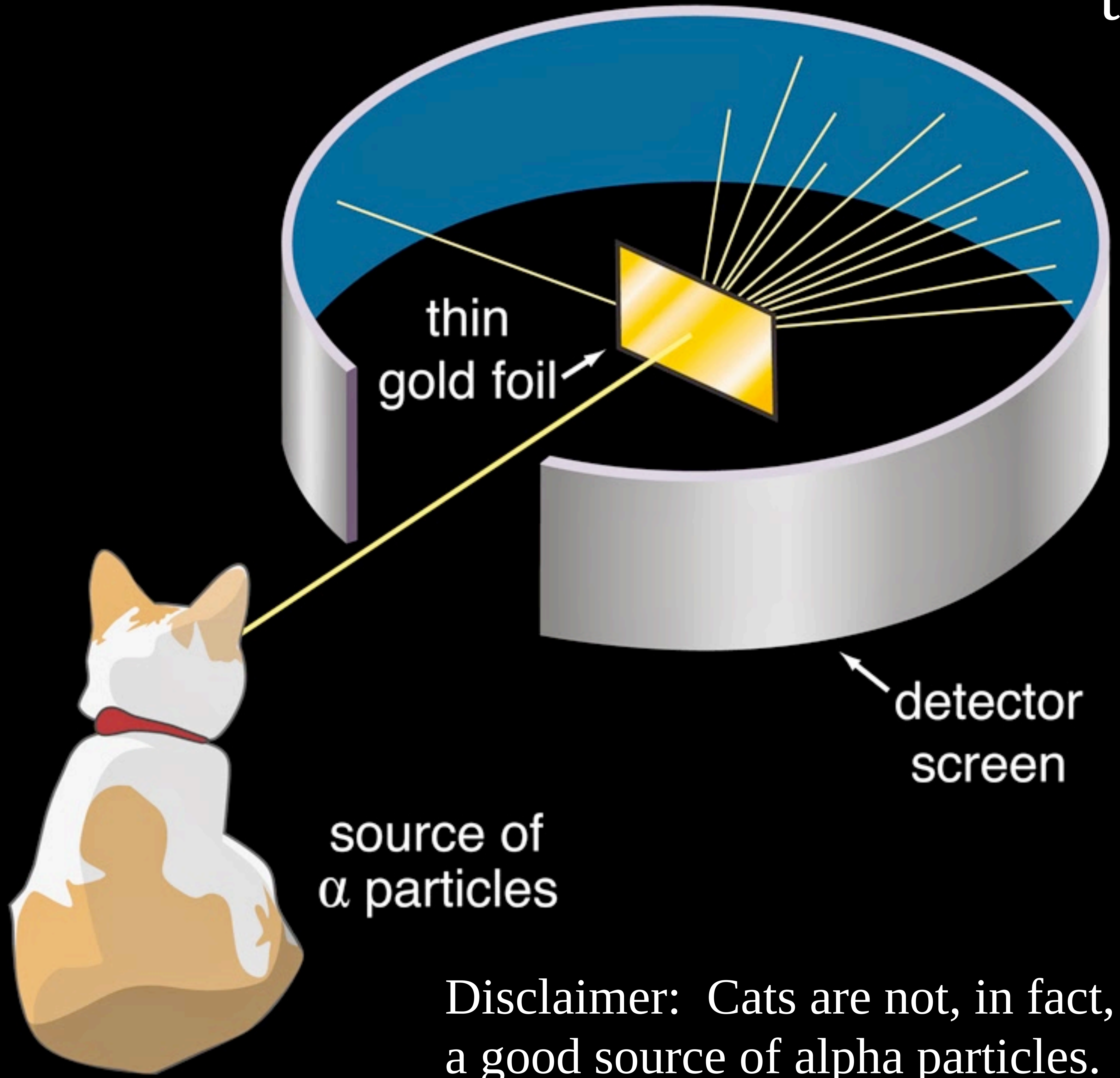
- Thomson knew particles were waaaaay too small to be the ions which had already been measured
- Millikan Oil Drop: Finally able to measure actual values (not just the ratio) for charge and mass
- Electrons are tiny compared to hydrogen: It would take almost 2000e^- to equal mass of a hydrogen atom

And Now, Dessert: Plum Pudding



- So if atoms are huge (by comparison) and neutral, and electrons are tiny (by comparison) and negative: What the heck?
- Thomson modifies the Greek idea of a uniform homogeneous sphere: Plum pudding (blueberry muffin) model
- An atom is like a tiny blueberry muffin: A positive matrix (the cake) randomly embedded with e– (the blueberries)
- Disclaimer: Atoms are not actual muffins, although muffins are, in fact made of atoms. Go figure.

If you throw a tennis ball at the wall, you expect it to bounce back (and you also expect that it will obey the law of reflection, $\theta_i=\theta_r$). What happens if you throw an alpha particle at a thin gold foil?



Disclaimer: Cats are not, in fact, a good source of alpha particles.

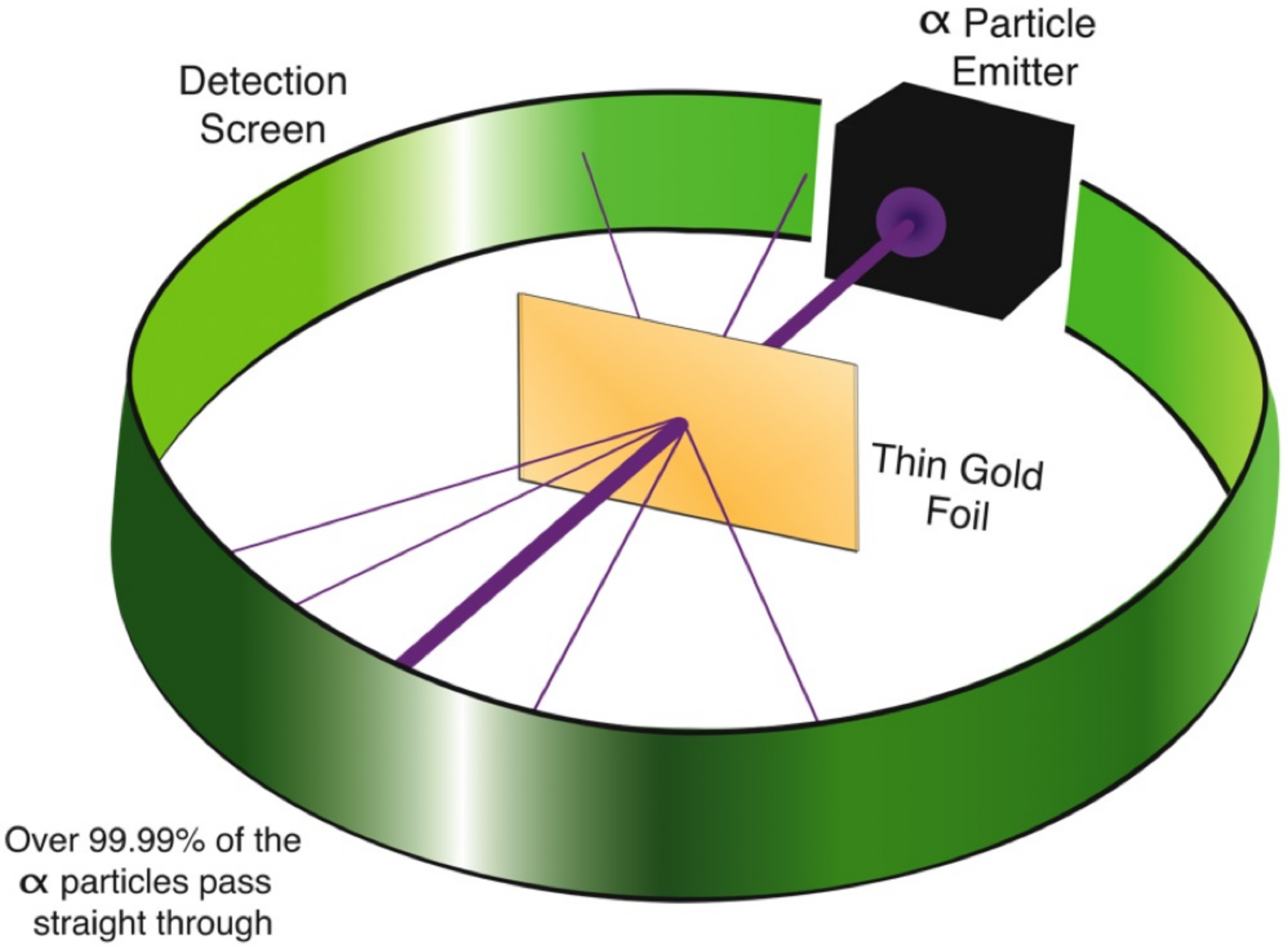
A) Your cat gets mad. He was saving those alpha particles to kill you in your sleep.

B) Just like a tennis ball, an alpha particle will bounce back off any solid surface (it doesn't *have* to be gold, but the cat insisted).

C) Totally *unlike* a tennis ball, most of the time an alpha will pass right through the foil like there was nothing in the way (cat laughs maniacally then stalks away with his tail in the air).

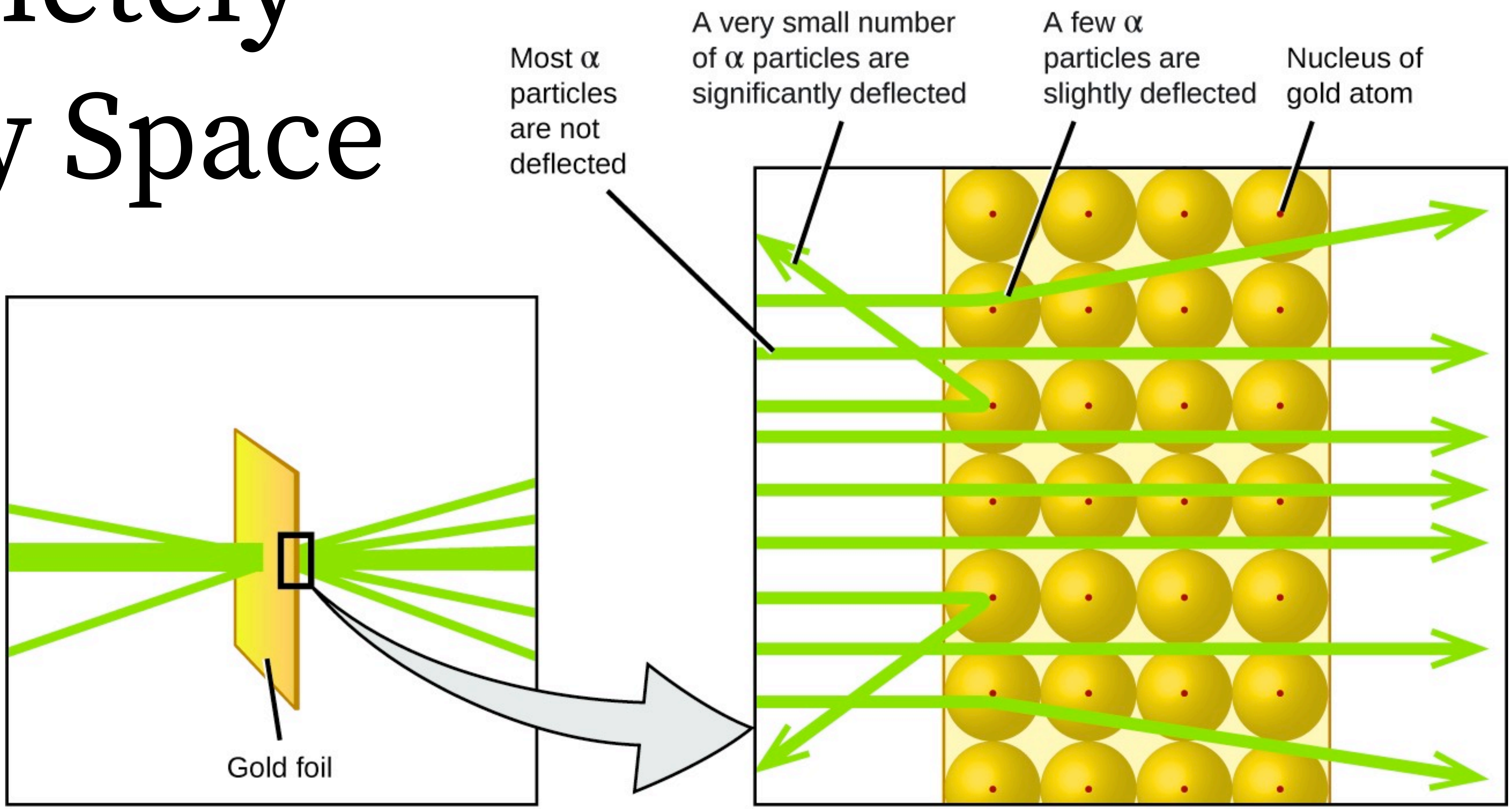
The Nucleus

- Discovered by Ernest Rutherford (one of Thomson's students) in 1911
- Recipe for alpha particles: Find an alpha source (multiple possibilities, but radium is a better choice than a cat) . Use E and B fields to herd them into a stream and accelerate them.
- Shoot these alphas at a thin gold foil, surrounded by a detector (that goes ping! when hit with an alpha)
- Surprise! Most of the particles pass through the foil like nothing's even there
- Surprise! Some of the particles get ricocheted off in random directions, or slammed backwards

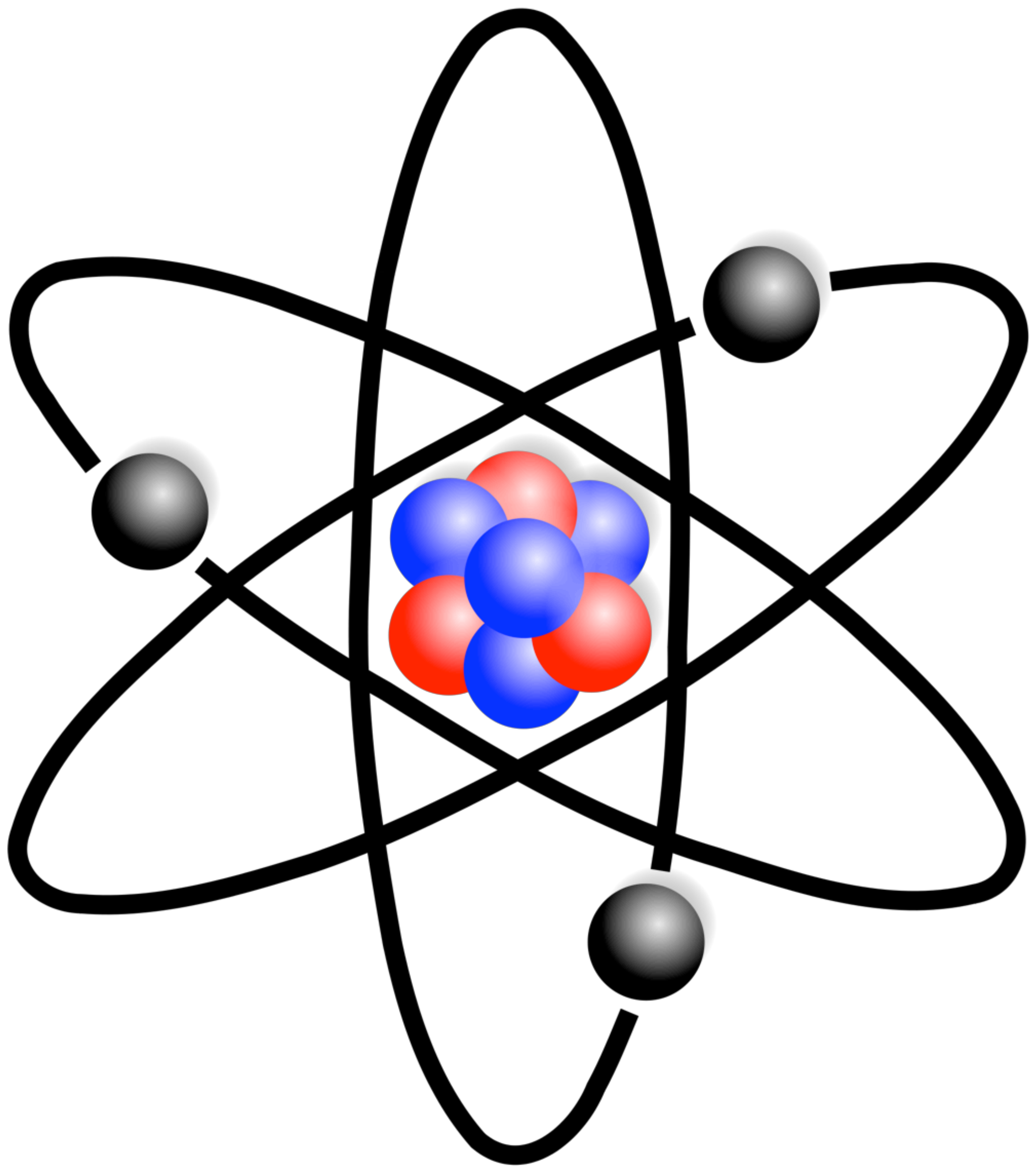


Not Quite Completely Empty Space

True or false: Rutherford calculated that electrons were moving outside the nucleus at a distance of about 100,000 times the radius of the nucleus.

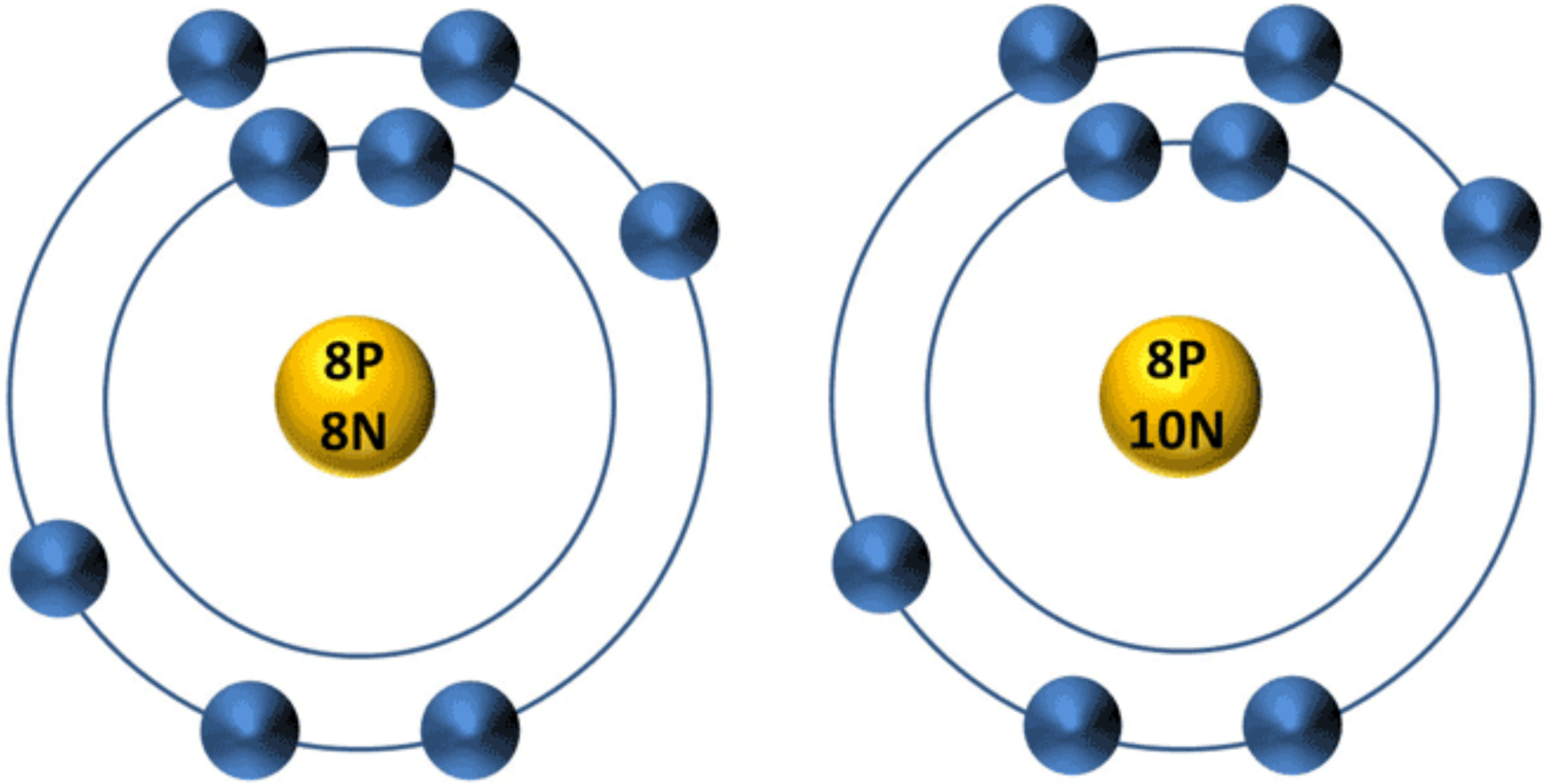


The Nucleus We Know Now



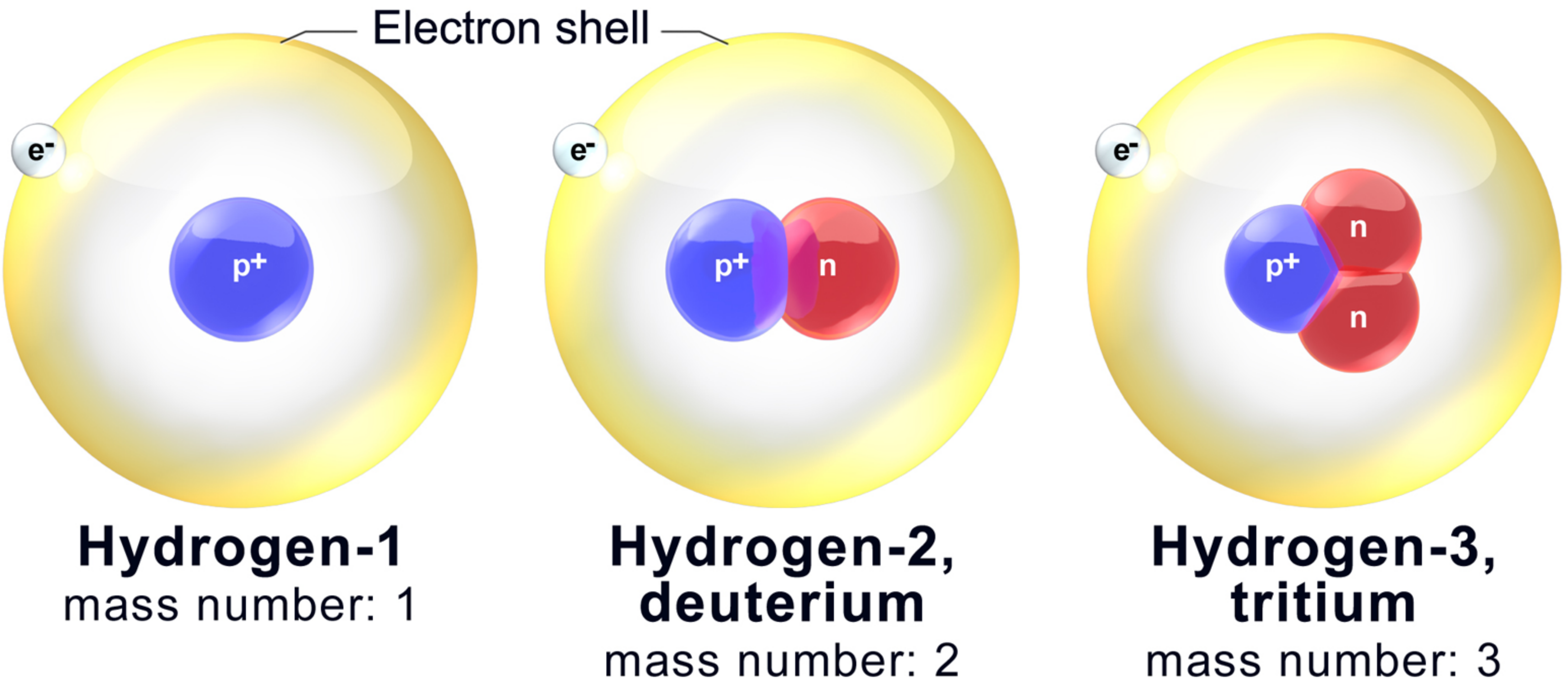
- Rutherford suspected neutrons, but they were not actually discovered until 1932!
- The nucleus contains positive protons and neutral neutrons: Huge compared to an e^-, but still a tiny part of the atomic volume
- The number of protons within a nucleus identifies it specifically and uniquely: Atomic number = number p^+ in nucleus

True or false:



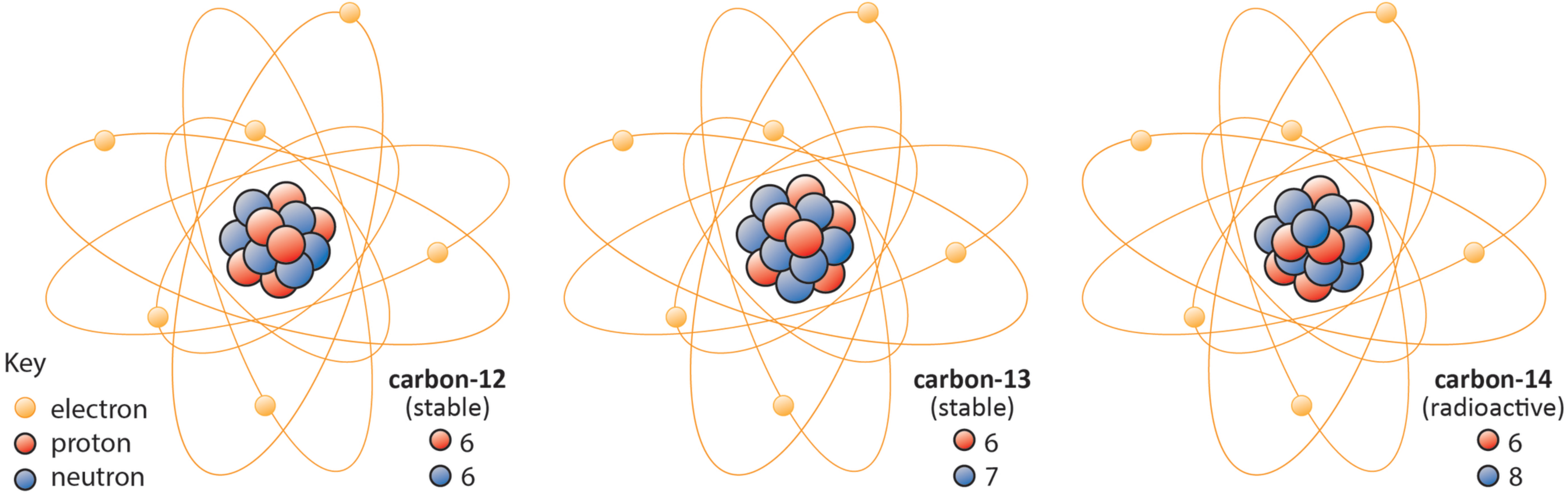
Both of these nuclei are oxygen!

Isotopes



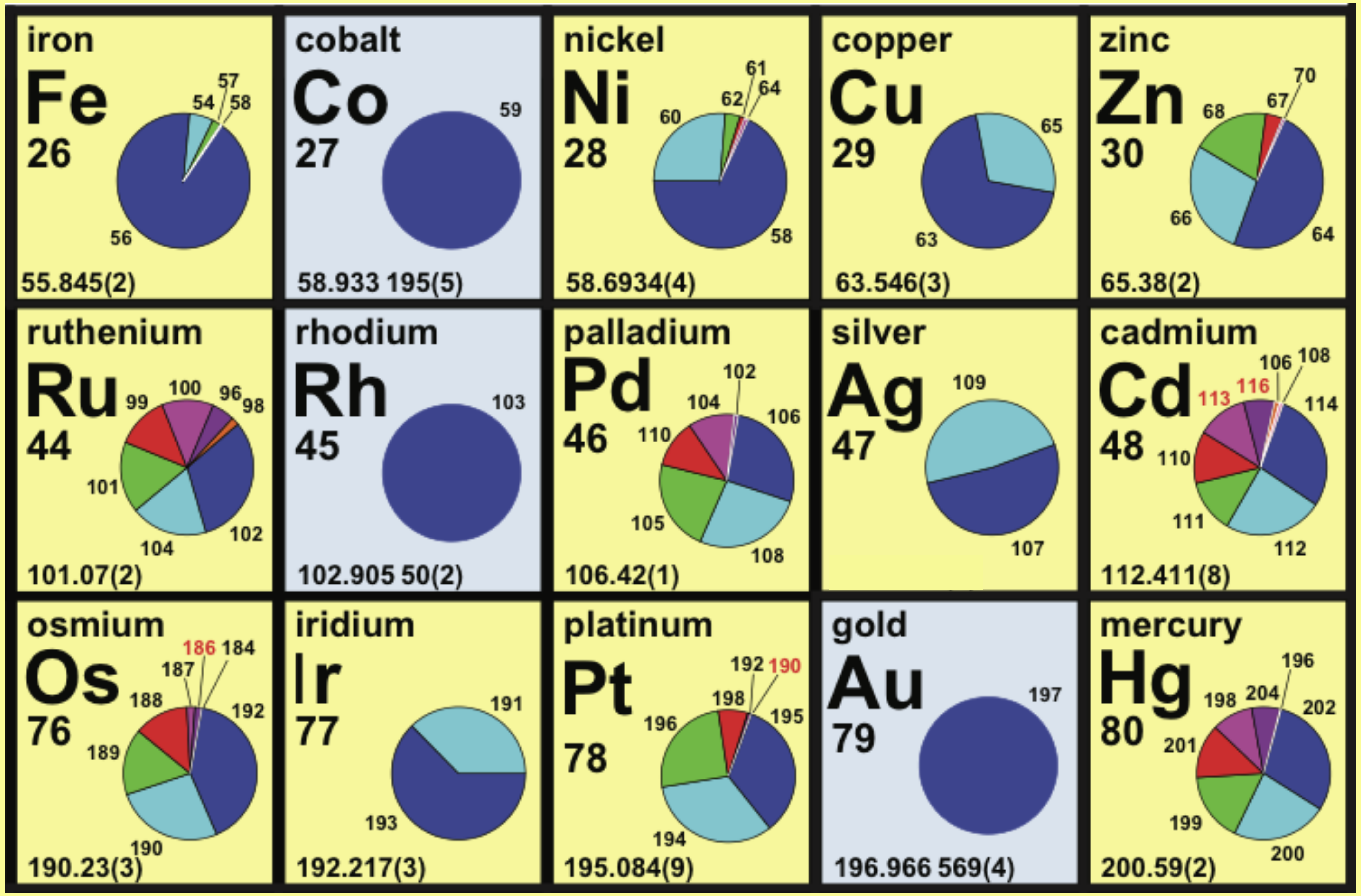
- You cannot build anything heavier than hydrogen without using neutrons
- The neutrons are necessary to bind multiple protons close together
- You might use different numbers of neutrons, even with the same number of protons
- An isotope of an element has the same atomic number, but different neutron number

Atomic Mass and Weight



- AMU: Atomic mass unit, exactly equal to $^1/_{12}$ the mass of isotope ^{12}C (not quite exactly the mass of one proton)
- Individual atoms have specific mass: Every ^1H masses exactly 1.007 amu, but not every hydrogen atom is a ^1H isotope!
- Atomic Weight: Normalized average accounts for relative abundances of different isotopes
- Example: ^{14}N = 14.003 amu and ^{15}N = 15.000 amu, but the atomic weight of nitrogen = 14.007, not 14.5—because there's way more ^{14}N than ^{15}N

Examine the element silver (Ag) on the periodic table below. What is the atomic weight? Will the actual value be slightly greater, or slightly less than your response?



A) 47

B) 94

C) 107

D) 108

E) 109

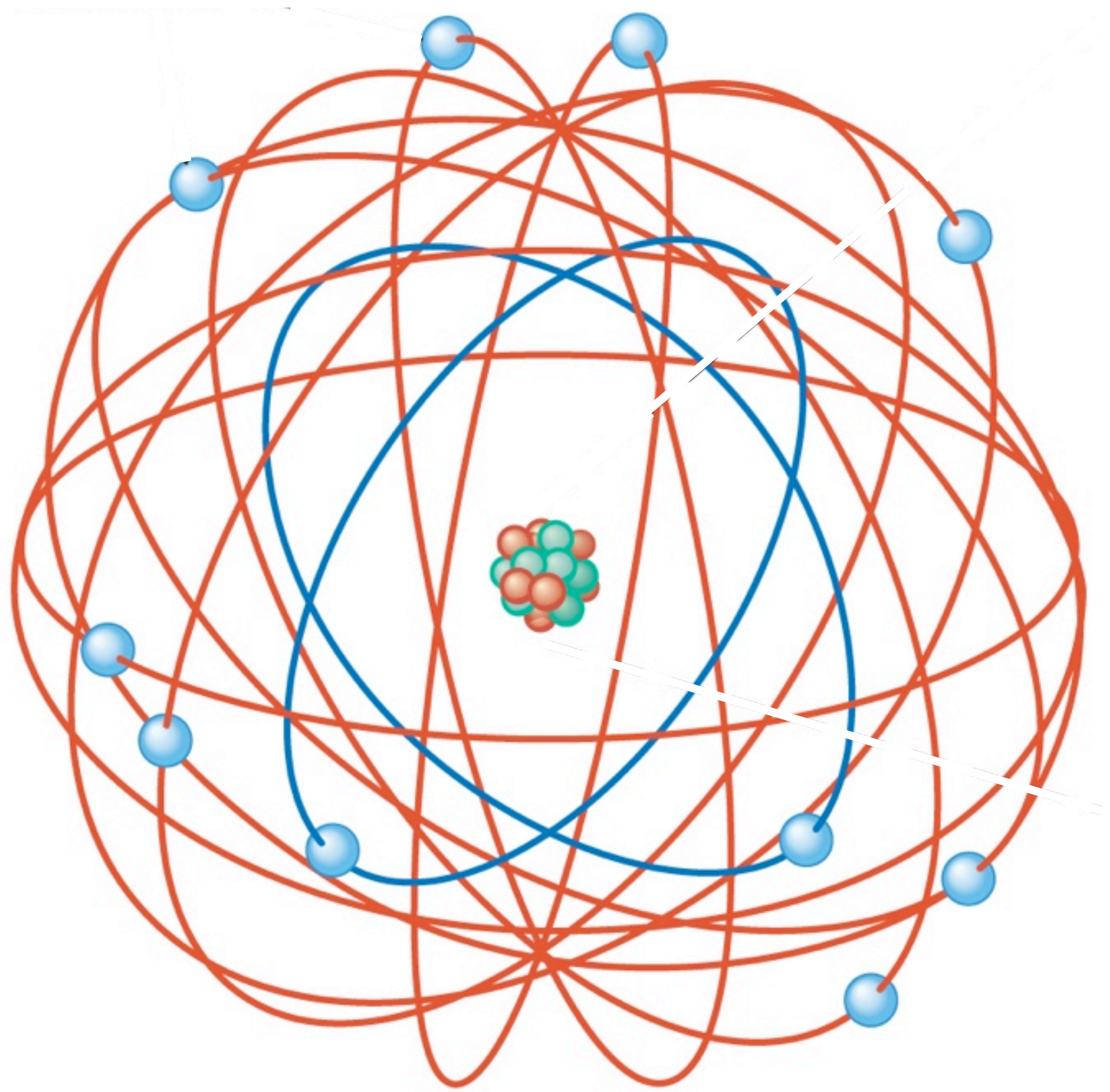

Rutherford and Bohr, circa 1930

Section 8.2

The Bohr Model

Tiny Baby Solar Systems: *Not*

- Rutherford speculated that electrons orbited like planets around the sun, but this falls apart (*literally*!) because it would result in the electrons collapsing into the nucleus
- Niels Bohr got tot talking with Rutherford (circa 1912), and wondered out loud why the heck those orbiting electrons did not emit e·m radiation???



As usual, those e^- are shown about 10,000× too big (compared to the size of the protons & neutrons)

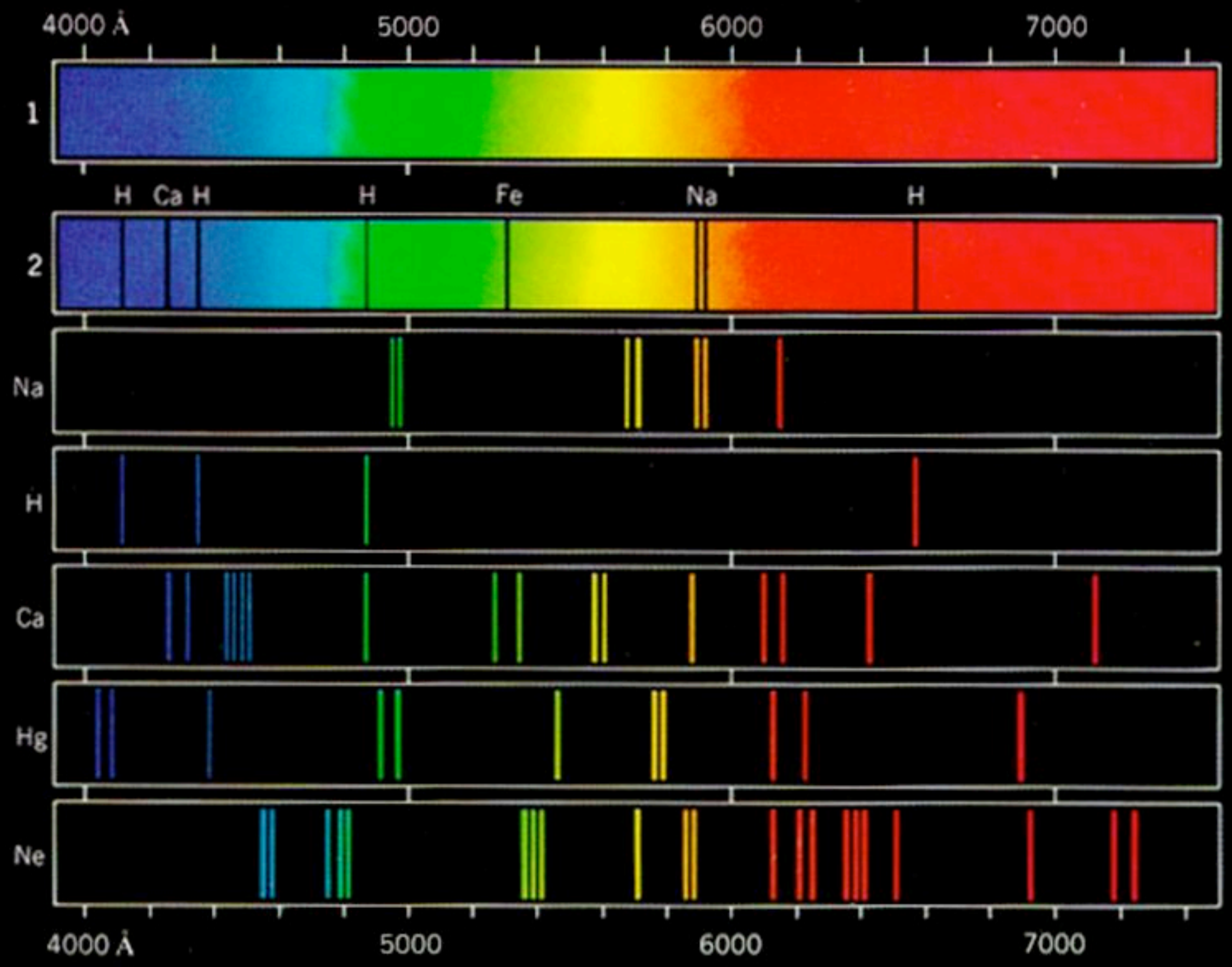
Bohr and Planck, circa 1930

Everything Is Connected To Everything Else

- JJ Balmer (≈ 1885) tries to make sense of hydrogen emission spectra: Why so many lines????
- Max Planck (≈ 1900) comes up with $E = hf$, showing the e·m radiation is quantized (no surprise—charge is quantized!)
- Einstein's Photoelectric Effect (1905) demonstrates that light is also quantized (again, should not surprise us, oscillating electrons and all)
- All of these ideas synthesize to explain how electrons have to be behaving

A Closer Look at Those Emission Spectra

- Recipe for emission spectrum: Seal some elemental gas in a tube with electrode ends. Apply a voltage. When tube glows, pass the light through a prism to separate it by color/ frequency.
- When you do this for hydrogen, you do not get a ROYGBIV continuum; you get only a few bright lines (red, teal, indigo, violet)
- You get the same lines for hydrogen all the time, but they are not the same as the lines for helium—or for any other element
- Color means wavelength mean frequency means energy: The hydrogen is emitting very specific quanta (there's that word again) of energy! Why???
- If a hydrogen atom only has one e^-, how can it be spitting out so many different frequencies?

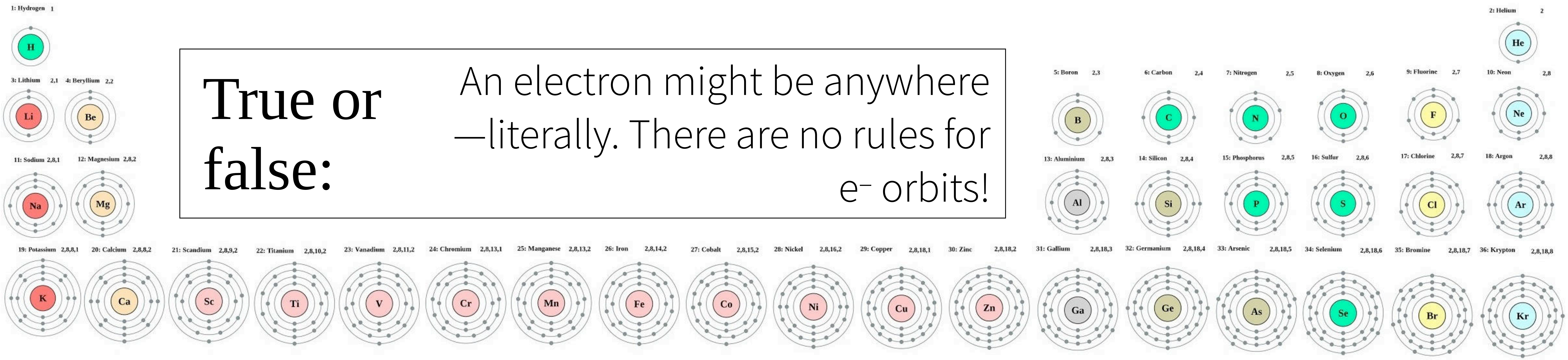



Einstein and Bohr, circa 1930

Back to Bohr

- Scientific method at work! Any atomic model has to be able to incorporate Balmer, Planck, Einstein, Rutherford
- Start with that "solar system atom" as a metaphor: What else has to be true?

Bohr #1: Allowed Orbits



True or false: An electron might be anywhere—literally. There are no rules for e^- orbits!

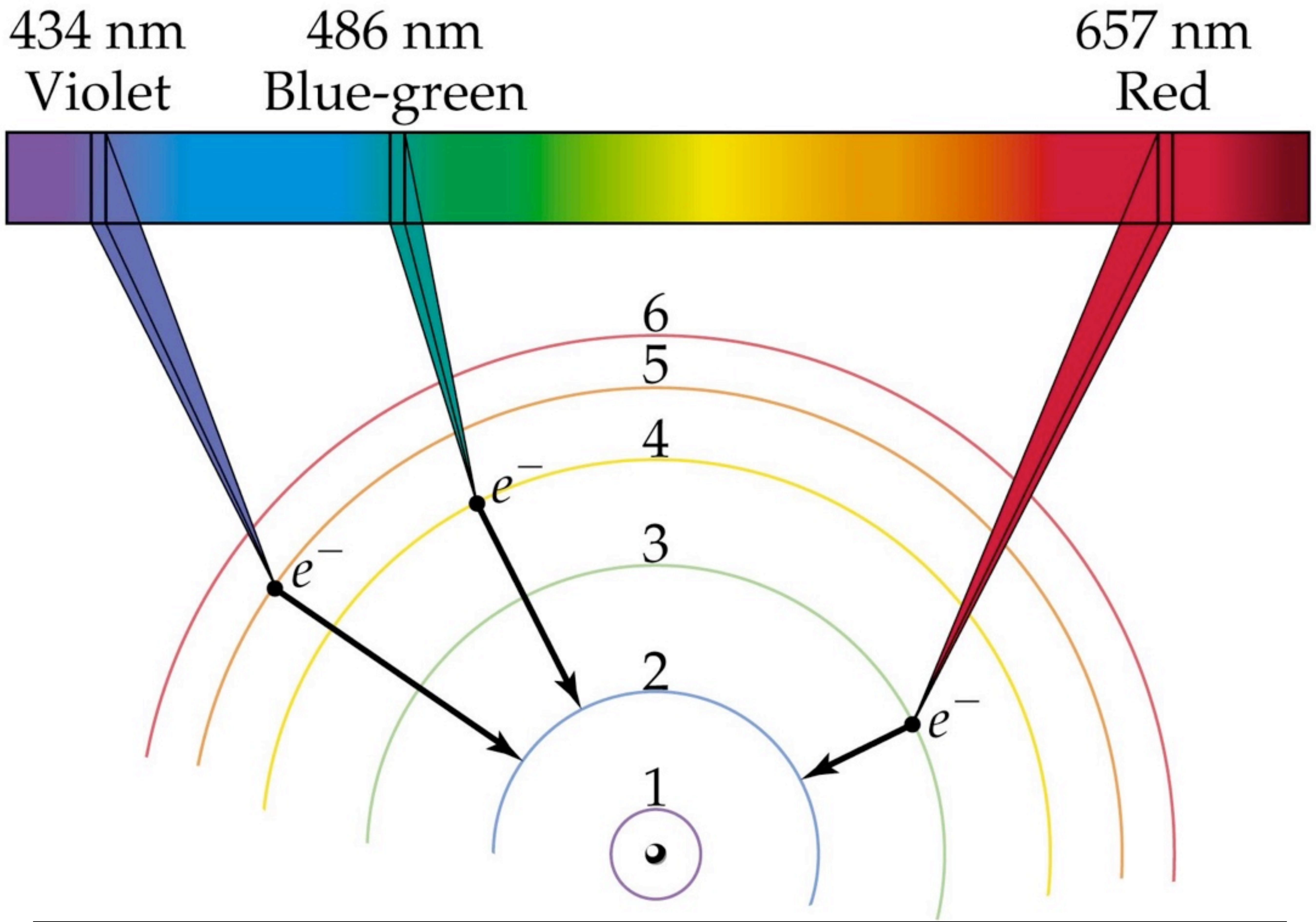
- Known that e^- is a particle with specific mass
- Apply ordinary Newtonian mechanics to figure out where orbits are allowed
- Electrons can only be in those places; cannot have an e^- anywhere except in an allowed orbit

Bohr #2: Radiationless Orbits

- As long as it orbits, an e^- does not emit any e·m radiation
- This is problematic; if an e^- is orbiting, it is accelerating and by definition should be emitting e·m energy
- However problematic it seems, it's demonstrably true—the trick is now to figure out why

Seriously, this comes up with keywords "radiationless orbit"

Bohr #3: Quantum Leaps

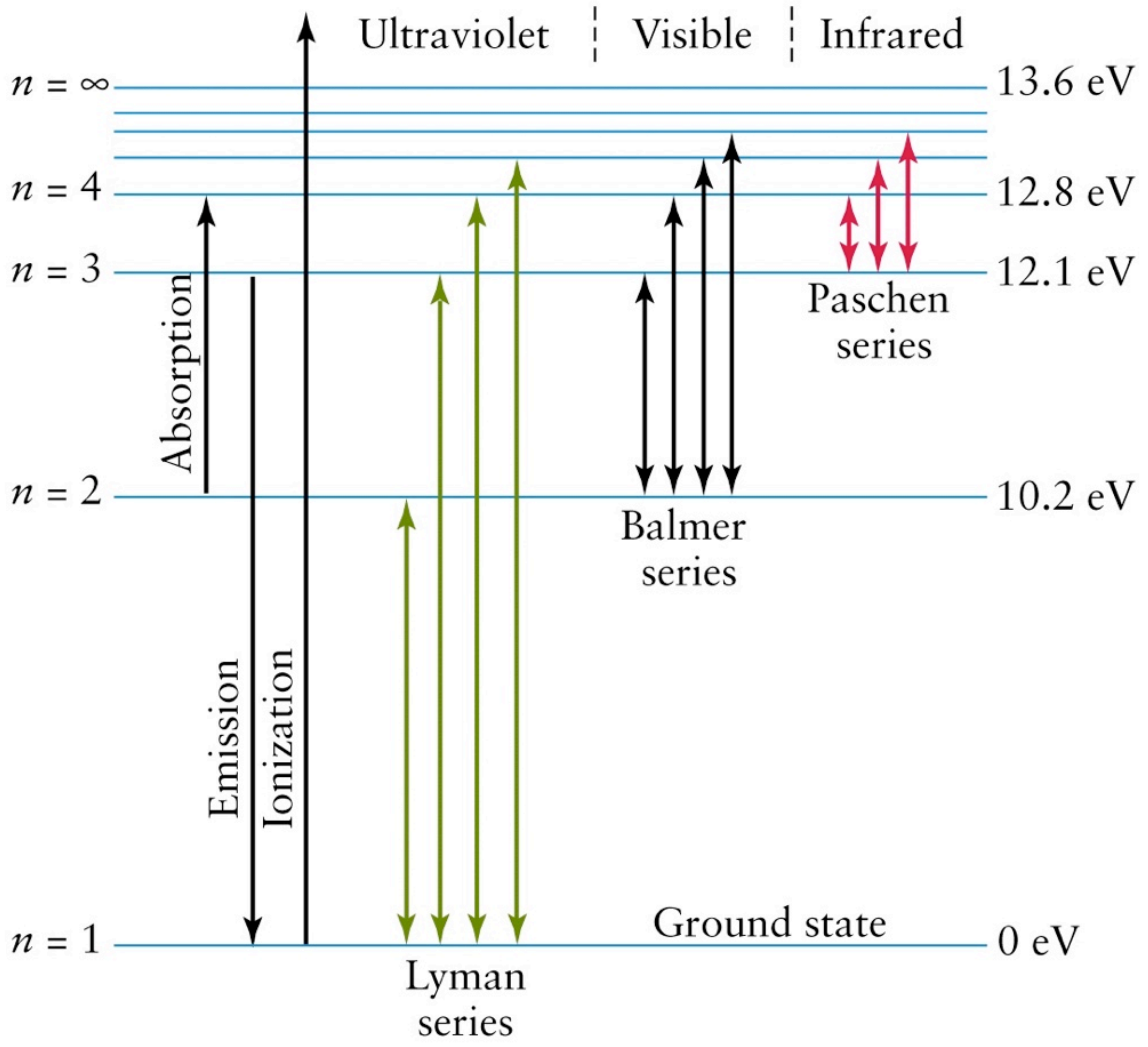


- So what's up with those emission spectra, then, if radiationless orbits?
- Possible for an e^- to move from one orbit to another
- If an e^- absorbs a photon with exactly the right quantum of energy, it can jump to a higher orbit
- Jumps have to be all-or-nothing: Must jump to an allowed orbit, no partial jumps
- Once it's there, it can emit a photon having precisely the same quantum of energy and fall back down

True or false: The 3→2 transition takes less energy than the 4→2 transition.

Ground State vs Excited State

- If the e^- is in the lowest orbit, it is in the ground state: Lowest energy, most stable
- An e^- in the ground state will not emit any e·m radiation (most matter, most of the time!)
- If an e^- absorbs a photon, jumps to a higher orbit, it is no longer as stable; e^- wants to be in that lower energy state

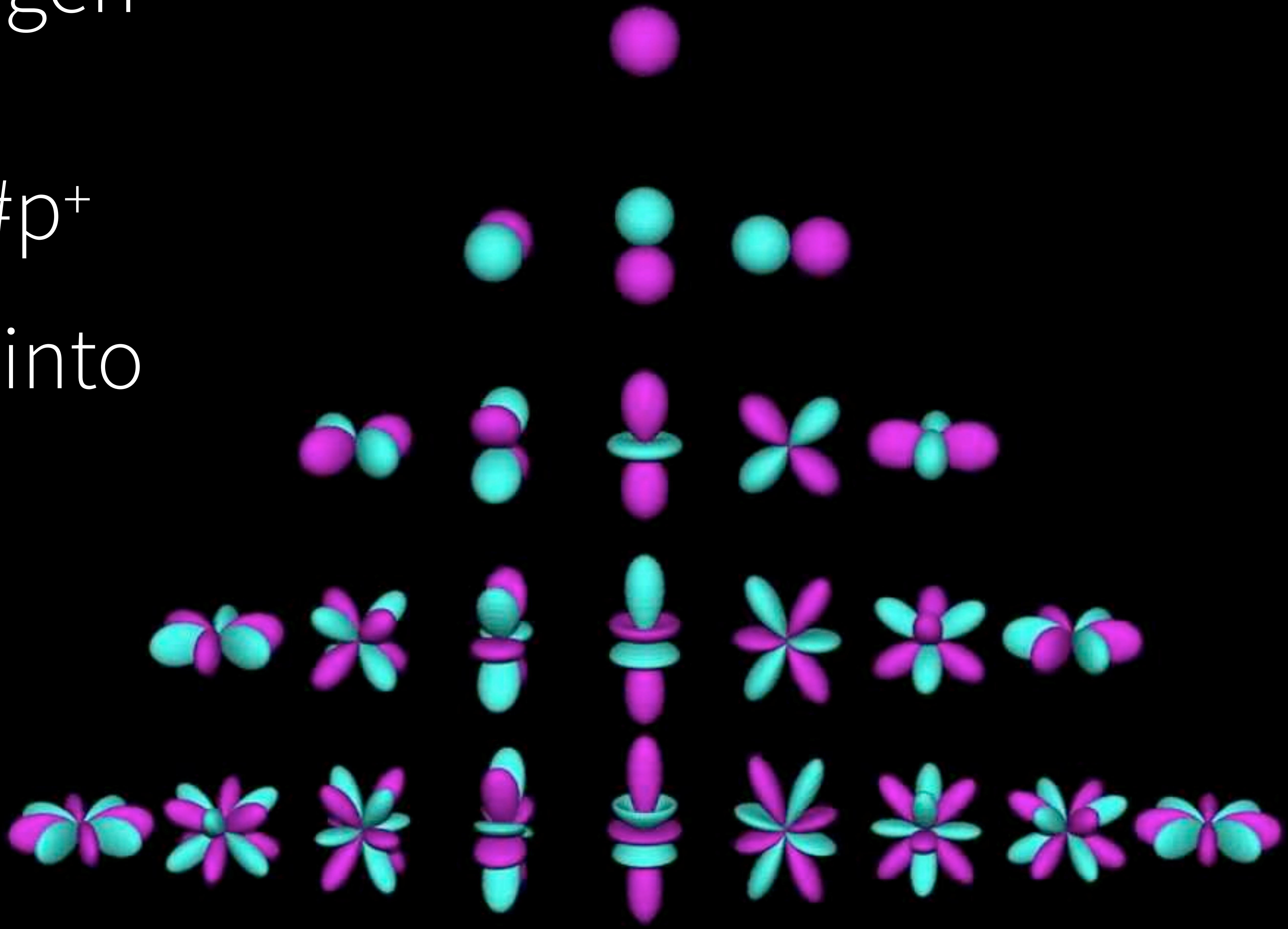


Section 8.4 Electron Configuration

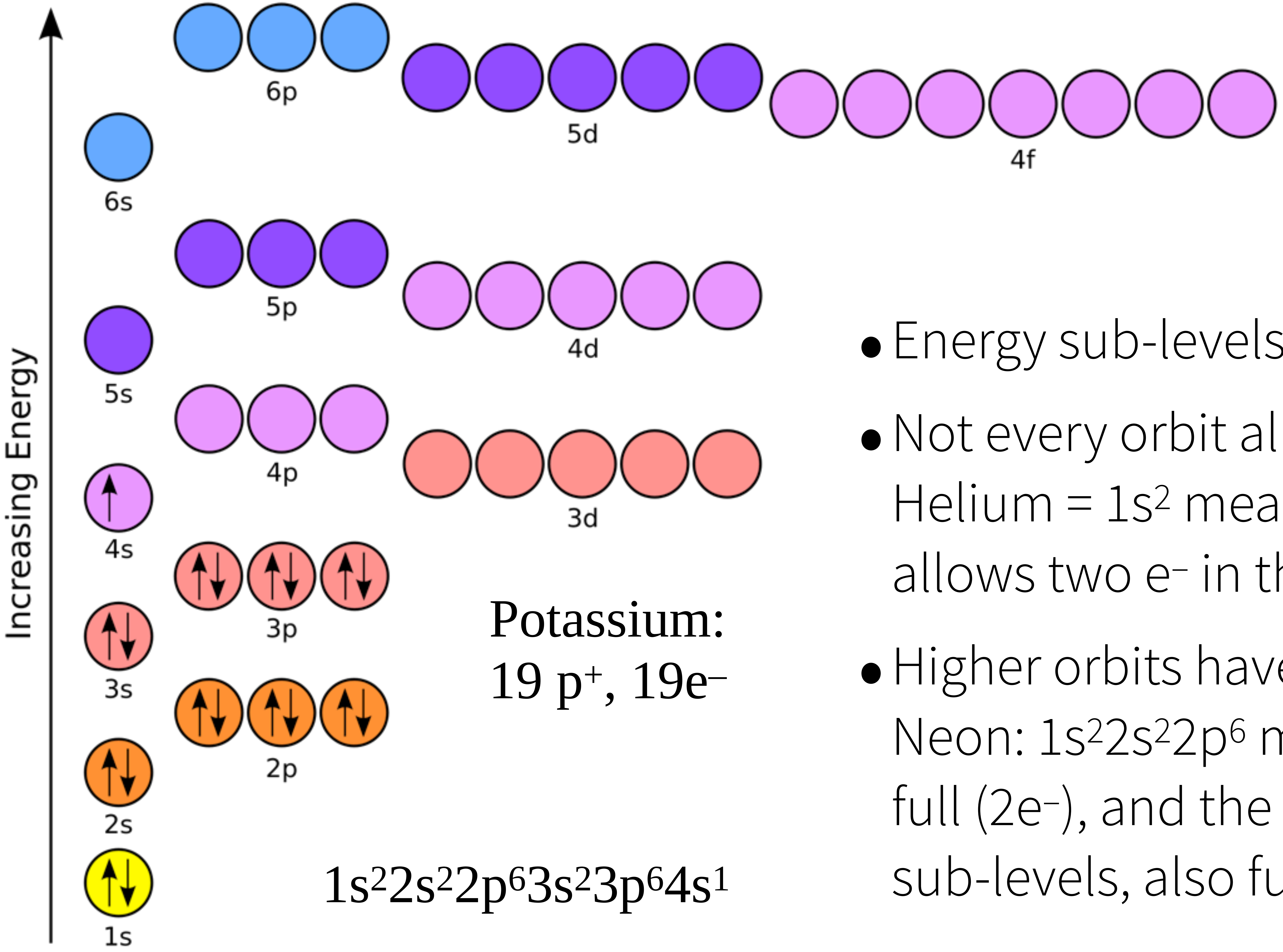
Solvay 1927: Quite possibly the biggest collection of brains ever photographed in one place.

Heavier Than Hydrogen

- Anything heavier than hydrogen has more than one electron: Neutral atoms mean # e^- = #p^+
- How do you put all of the e^- into lowest energy states?
- Multiple orbits, main energy level, energy sub-level



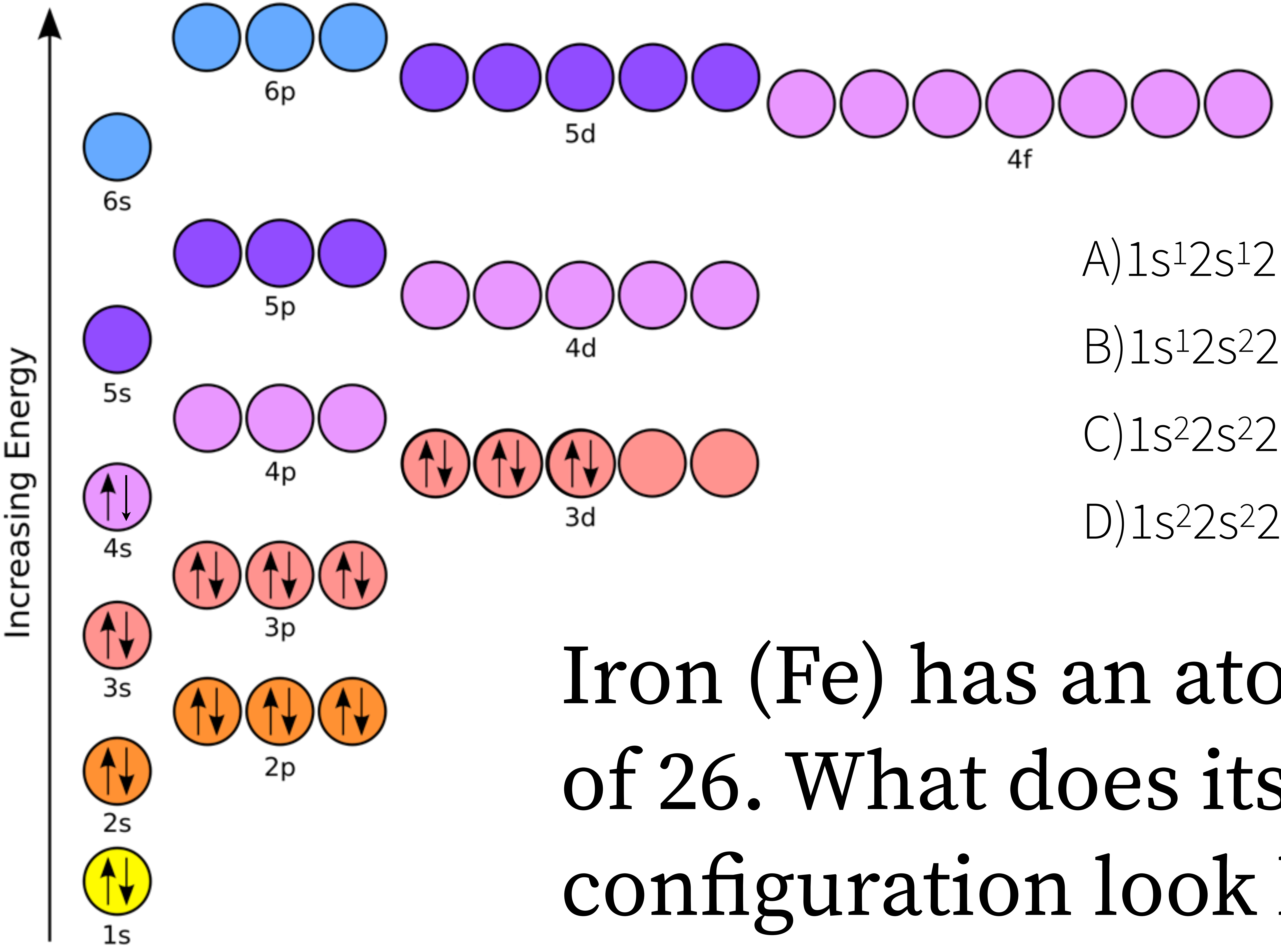
SPDF



Potassium:
19 p^+, 19e^-

$1s^22s^22p^63s^23p^64s^1$

- Energy sub-levels (low to high): s, p, d, f
- Not every orbit allows every sub-level: Helium = $1s^2$ means the first orbit only allows two e^- in the s sub-level
- Higher orbits have more sub-levels: Neon: $1s^22s^22p^6$ means the first orbit is full (2e^-), and the second orbit has two sub-levels, also full (total 8e^-)



A) $1s^1 2s^1 2p^1 3s^1 3p^1 4s^1 3d^1$

B) $1s^1 2s^2 2p^2 3s^3 3p^3 4s^4 3d^3$

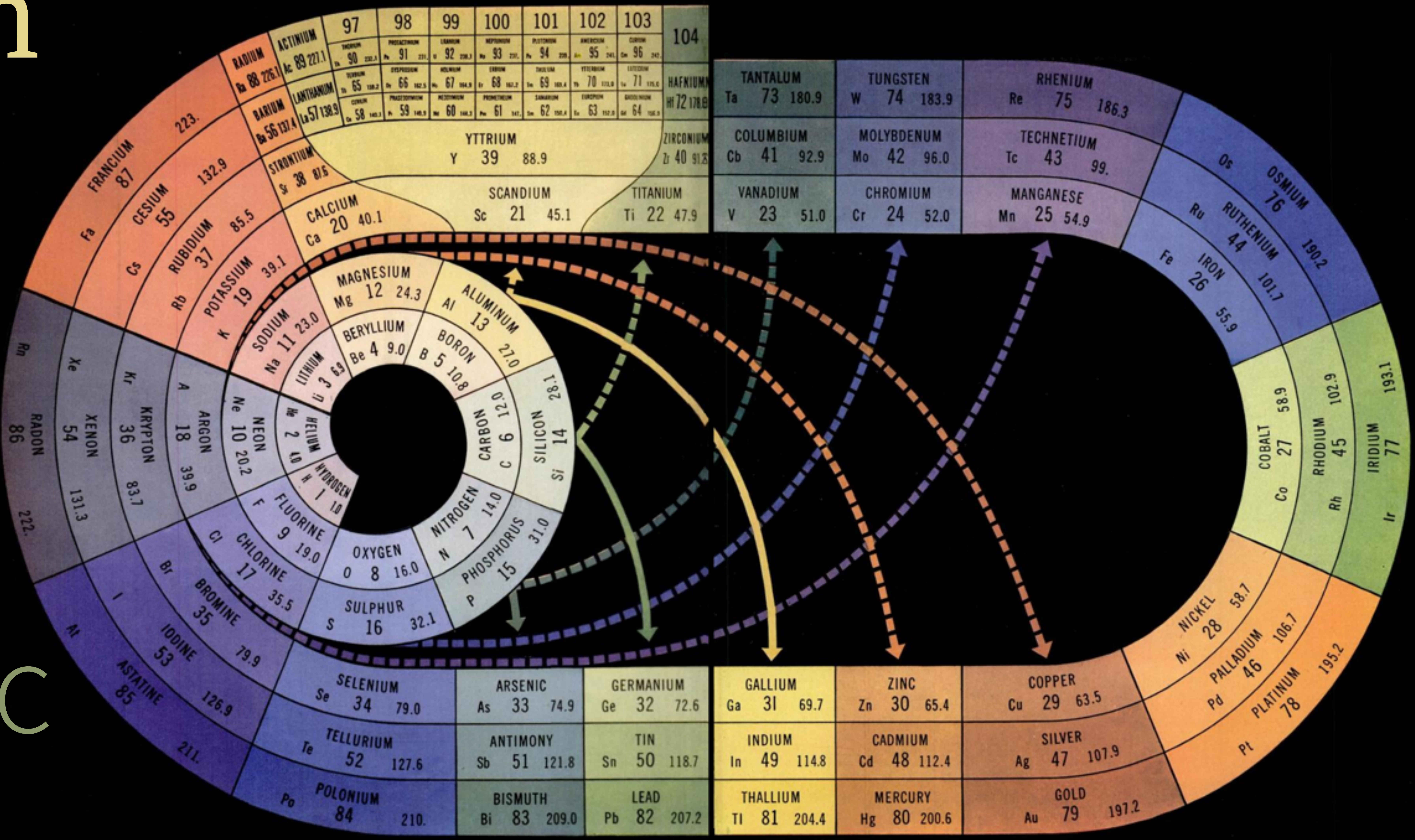
C) $1s^2 2s^2 2p^3 3s^4 3p^5 4s^6 3d^7$

D) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

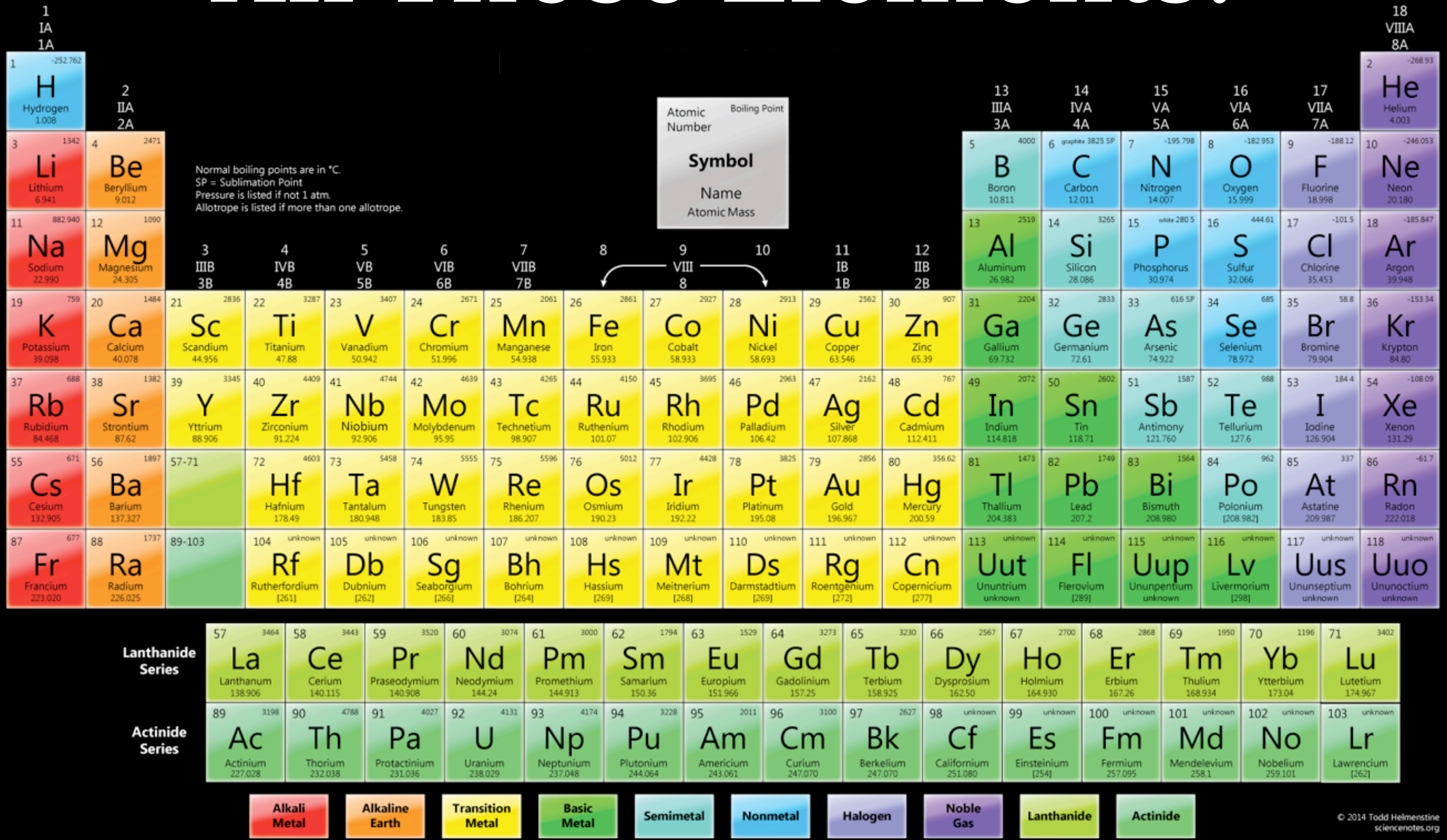
Iron (Fe) has an atomic number of 26. What does its electron configuration look like?

Section 8.5

The Periodic Table



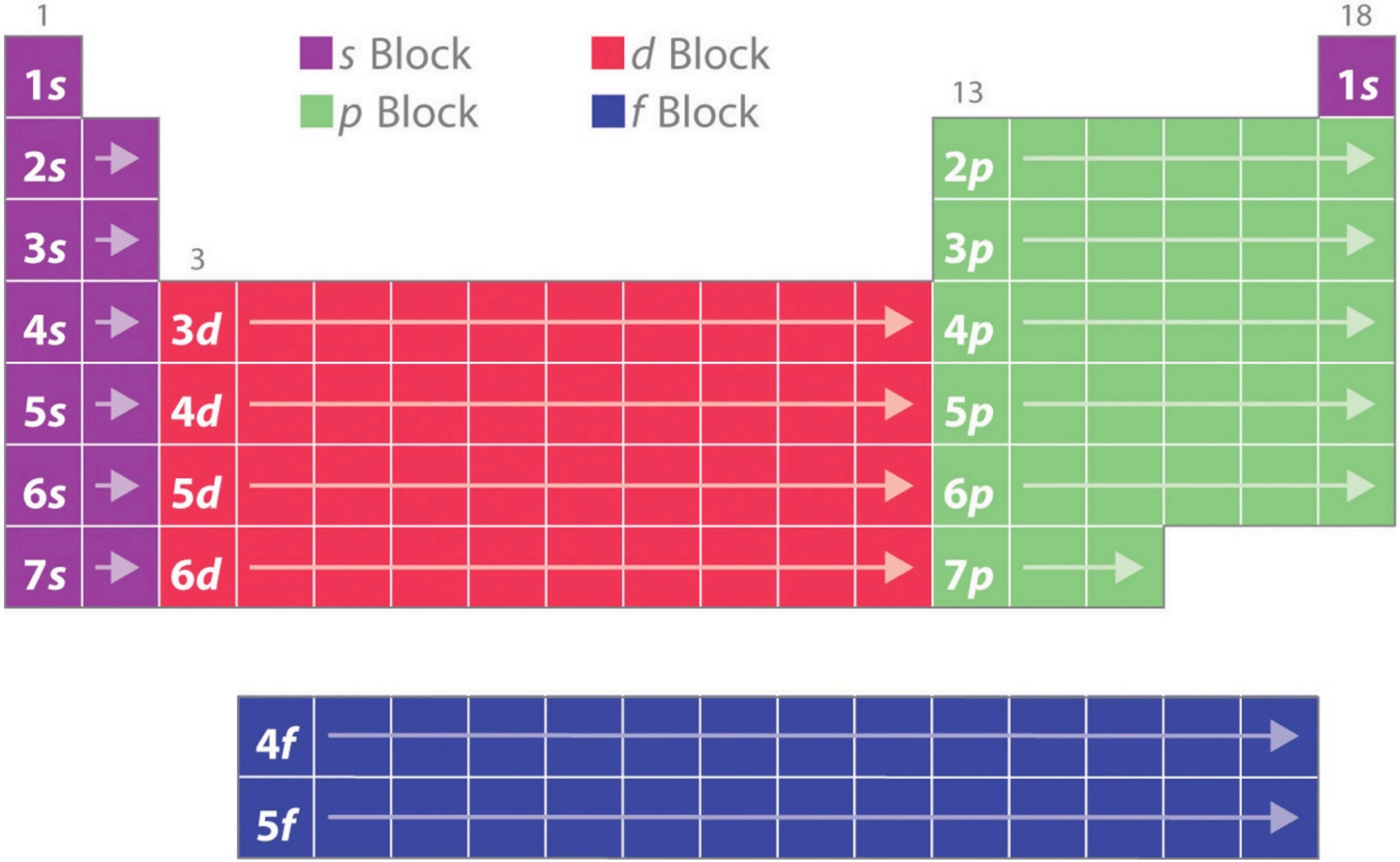
How Do You Keep Track of All Those Elements?



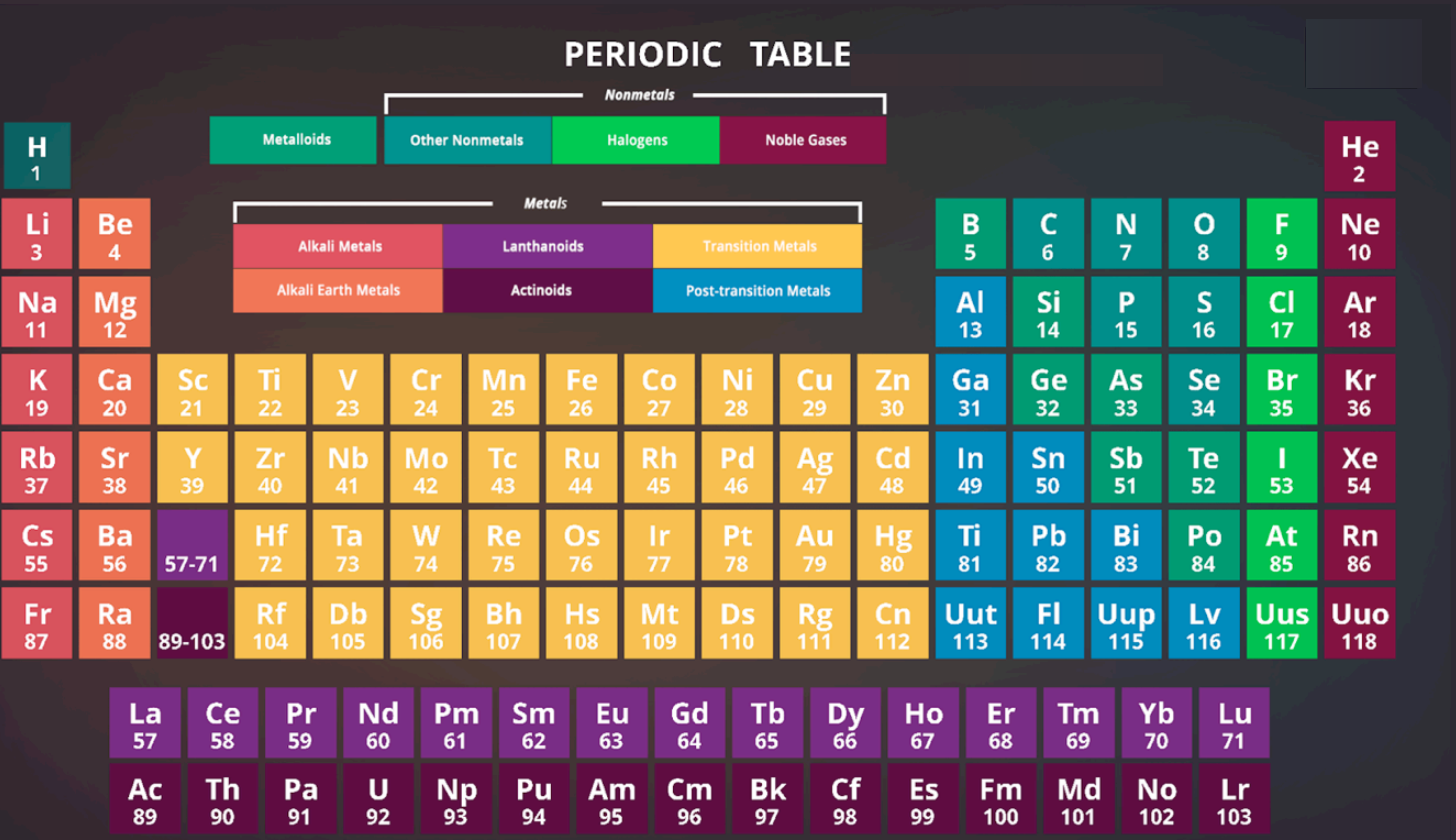
- Dalton tried to organize the few elements he knew, according to their weight: This is a pretty good place to start
- 1860s: Dmitri Mendeleev and Lothar Meyer independently publish tables that organize elements by weight and by chemical properties
- Mendeleev typically gets more credit: He was able to use his table to predict the existence of previously unknown elements

The Rows Mean Something!

- The rows are determined by electron configuration: Top row, 1st orbit; next row, second orbit
- Left to right: Each subsequent element has one more proton in its nucleus, and one more electron to orbit
- By the time you get to the end of the row, the orbit is completely filled

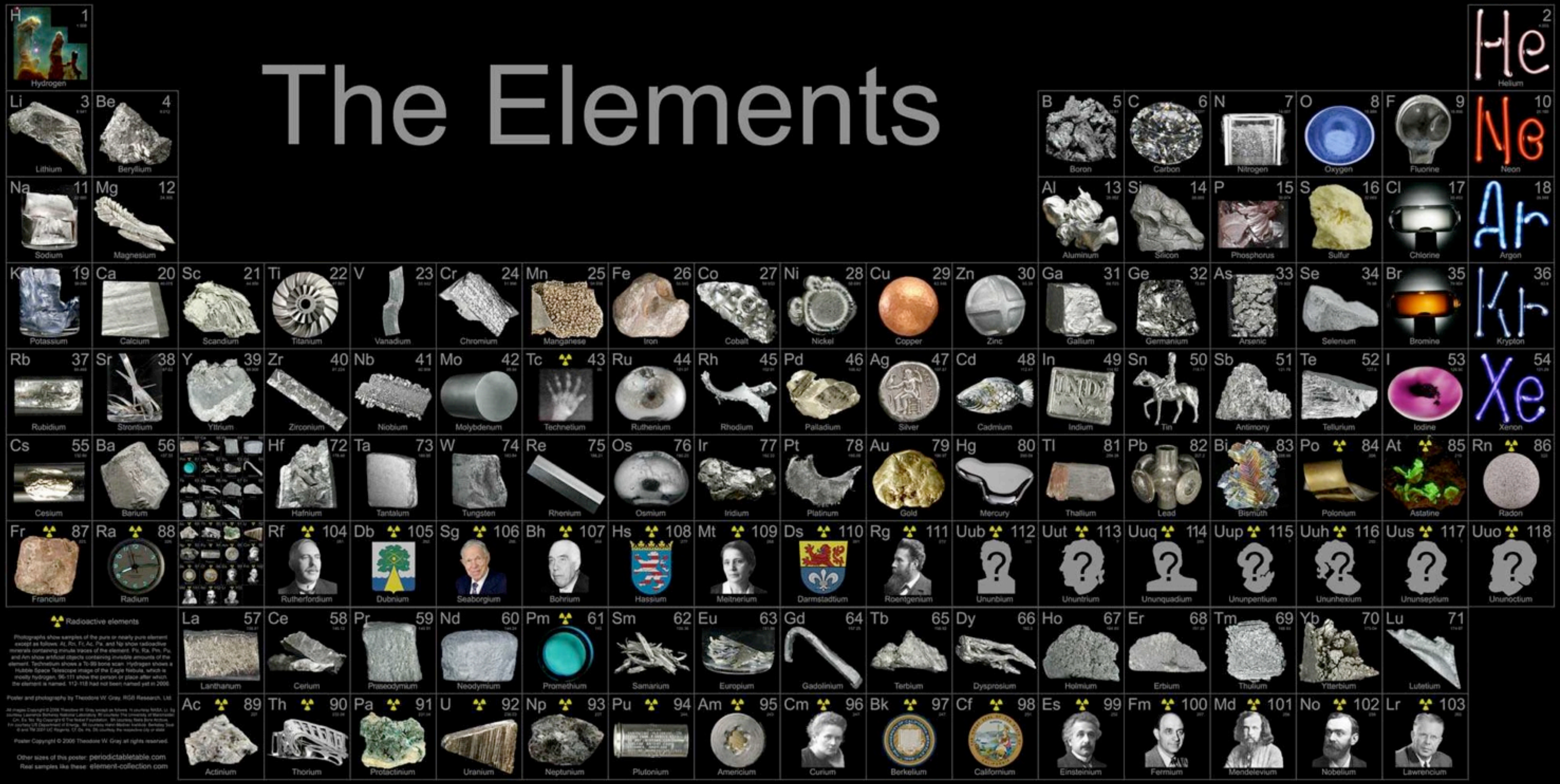


So Do The Columns!



- So every time you move one space to the right, you have a new element with one more proton in its nucleus
- The same column means the same valence: Same number of electrons in the outermost shell
- The elements right below (or right above), in the same column, share chemical properties

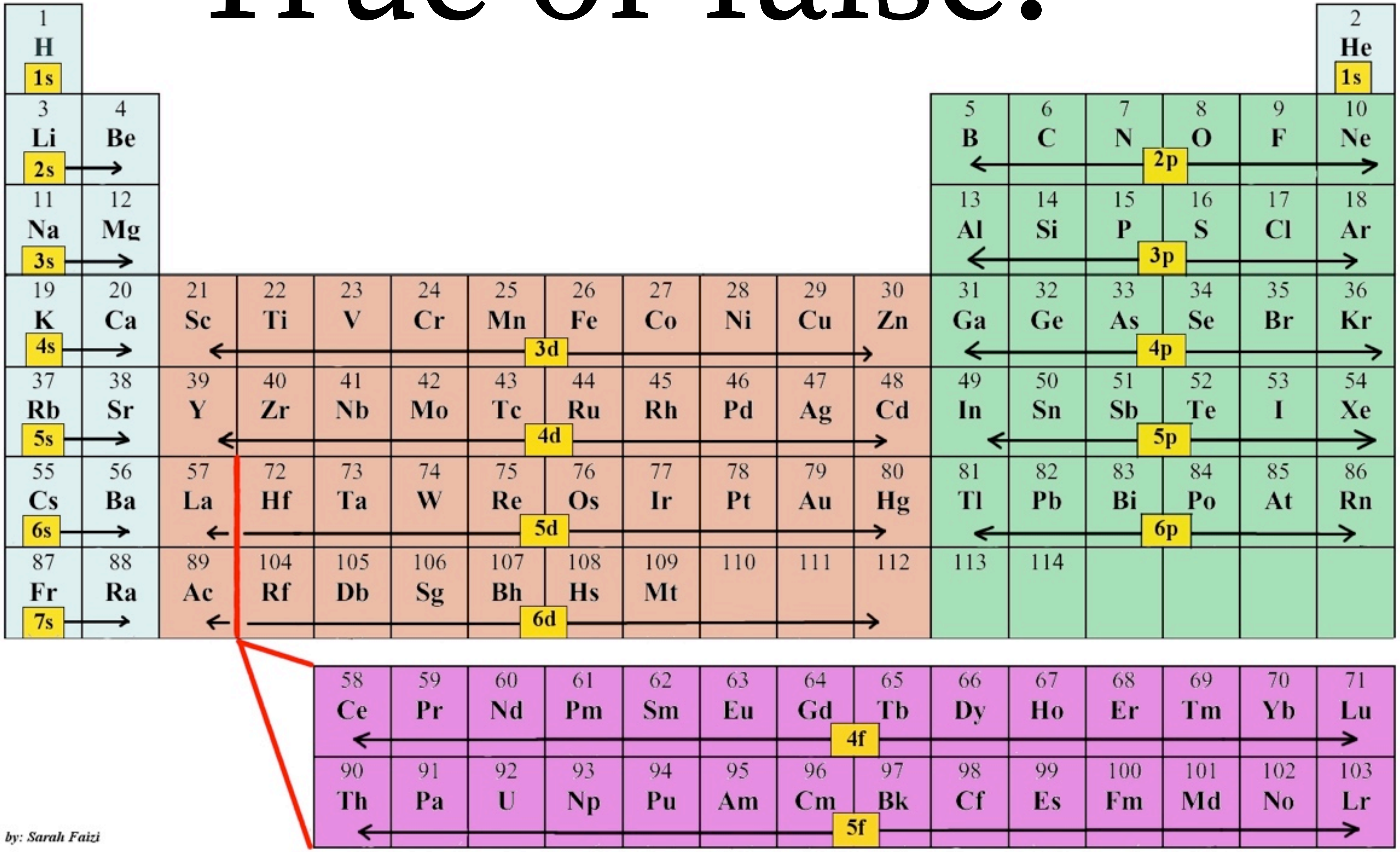
Section 8.6



Metals, Non-Metals, and Semiconductors

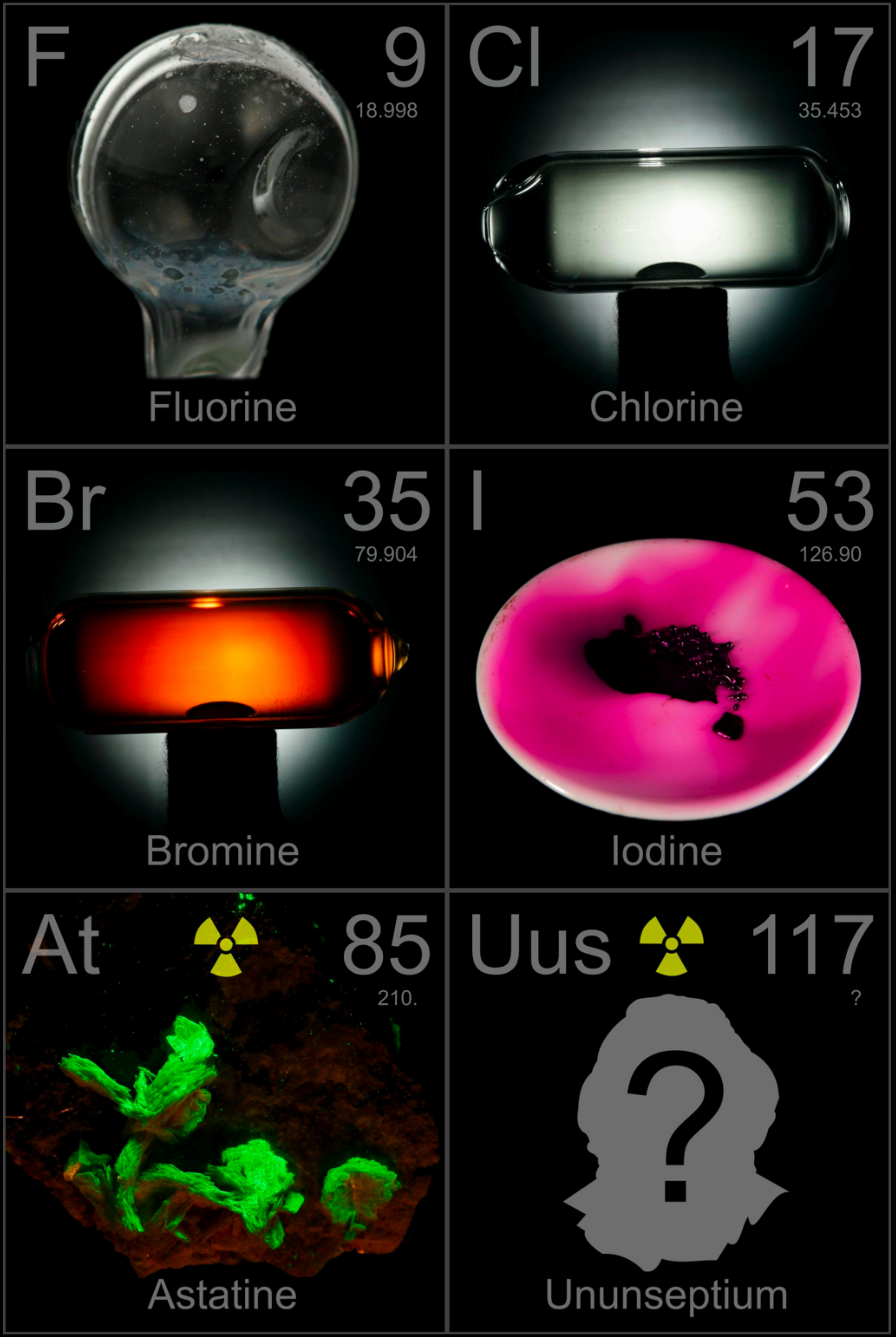
True or false:

Most elements have full valence shells, and as a result have strongly non-metallic characteristics.



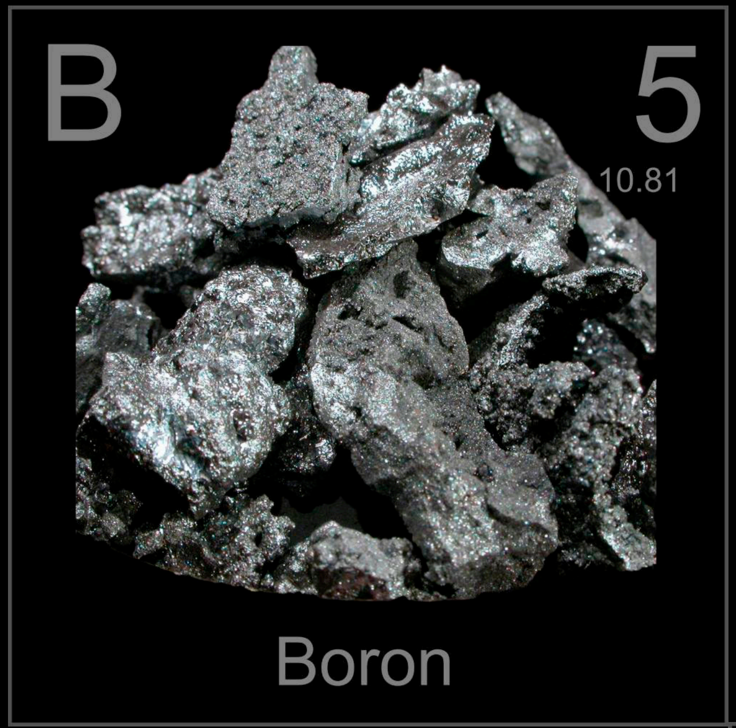
Metals

- Unfilled valence shell: Well, almost every element has an unfilled outer shell, but not every element is metallic
- Metals lose e^-: An element with 1, 2, or 3 valence e^- will give up those e^- to empty out the outer shell
- Metals become (+) ions when they lose valence e^-



Non-Metals

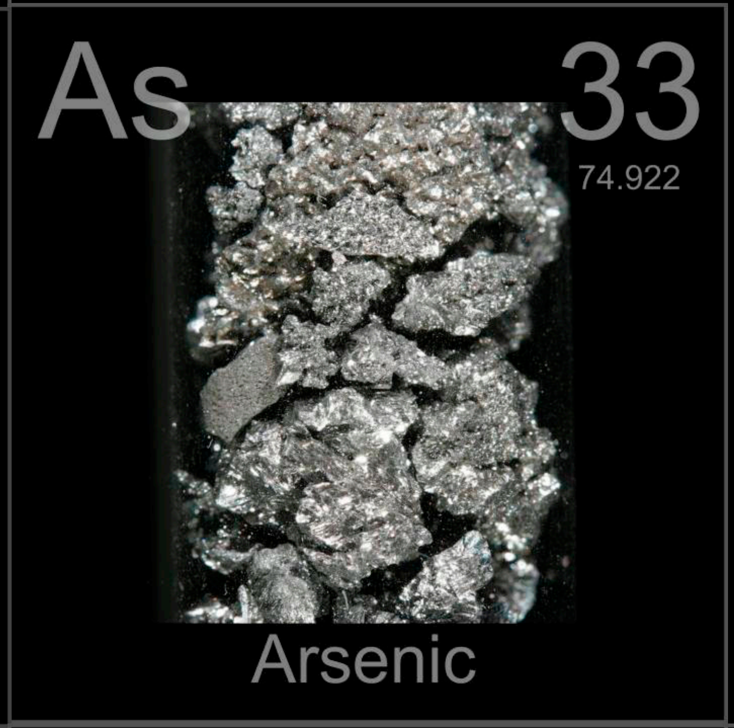
- Unfilled valence shell: Again, that's just about everything on the table
- Non-metals gain e^-: An element with 5, 6, or 7 valence e^- will gain e^- to fill up the outermost shell
- Non-metals become (–) ions when they gain valence e^-



Semiconductors

- Metalloid: Can conduct electricity (metallic property), but typically brittle/non-malleable (non-metallic property)

- Unfilled valence shell: Yes, yes, we know; just about everything on the table has an unfilled valence shell!



- Might gain or lose e^-; just as easy to empty the valence shell as to fill it

- They really are all silvery-grey sort-of metal- looking

Noble Gases

- Right-most column on periodic table: Completely full valence shell
- Colorless, odorless gases
- Emission spectra above are not colorless!!

True or false:
Noble gases are inert, and do not readily form compounds with other elements.