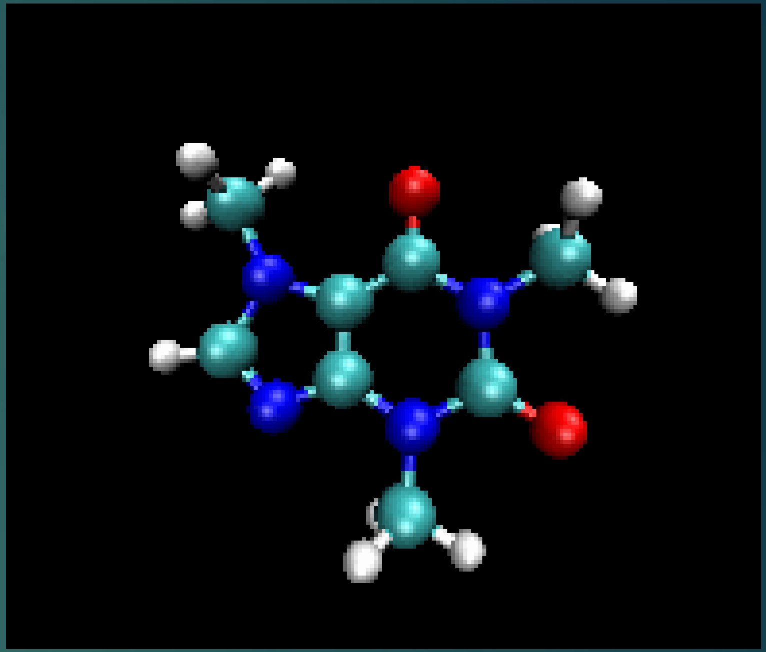
Science 10 Chemistry Unit



THE PERIODIC TABLE, ATOMS, COMPOUNDS AND CHEMICAL REACTIONS

TEXTBOOK REFERENCE: CHAPTERS 5 & 6

Brief History of Chemistry

The Creation of Chemistry - The Fundamental Laws: Crash Course Chemistry #3

CrashCourse

Chemistry

- The study of matter, its properties, and its changes or transformations.

- ***Matter***: Anything that has mass and takes up space.

The Atom

- The smallest part of an element.
  - If you could zoom in on elements, like iron, oxygen, helium, plutonium, etc., you would see the atoms that make up that element.

- Theorized by Democritus around 2500 years ago.
  - Not based on a scientific investigation.
  - Could not explain chemical properties of matter.
  - Would remain undeveloped until the early 1800s.

John Dalton’s Atomic Theory: 1803

1. All elements are composed of tiny indivisible particles called atoms.

2. Atoms of the same element are identical. The atoms of any one element are different from those of any other element.

John Dalton's Atomic Theory: 1803

3. Atoms of different elements can physically mix together or chemically combine in simple whole-number ratios to form compounds (like H_2O, CO_2).

4. Chemical reactions occur when atoms are separated, joined, or rearranged. Atoms of one of the element, however, are never changed to atoms of another element as a result of a chemical reaction. (Nuclear reactions change atoms from one type to another – happens in the naturally in the Sun and on Earth.

Incorporation of Electrons

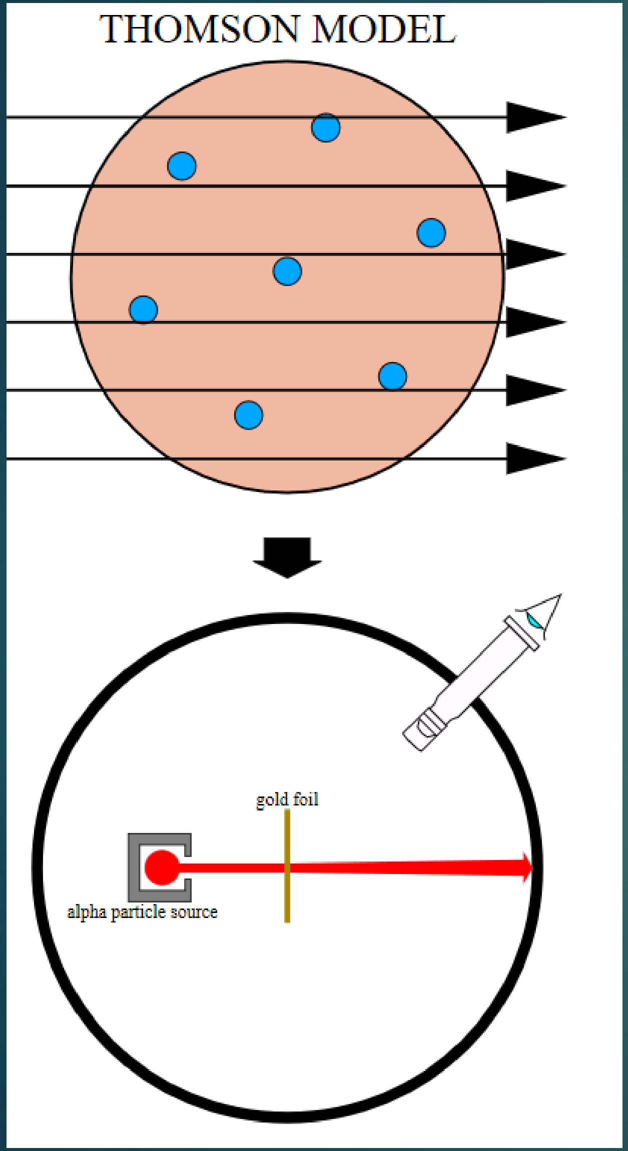
- Electrons, as particles, were first theorized in 1897 by English physicist J.J. Thompson. He invented the cathode ray tube to test for charges. That work eventually became the CRT television.

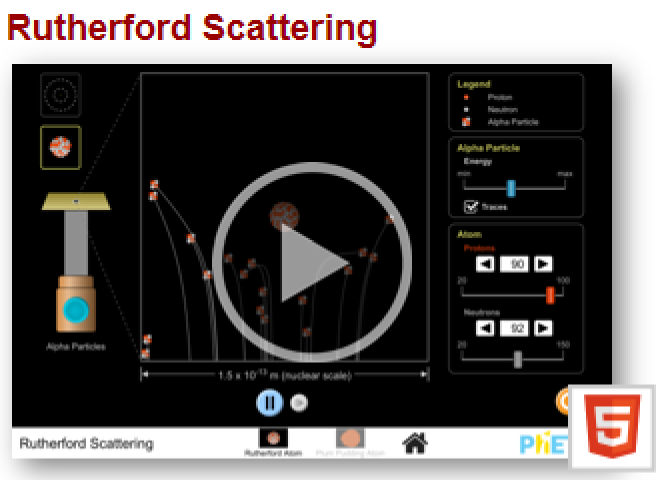
- Thompson adjusted the model of the atom to incorporate electrons; he proposed the atom is a lump of positive charge with electrons evenly spaced within it – dubbed the "plum pudding" model of the atom.

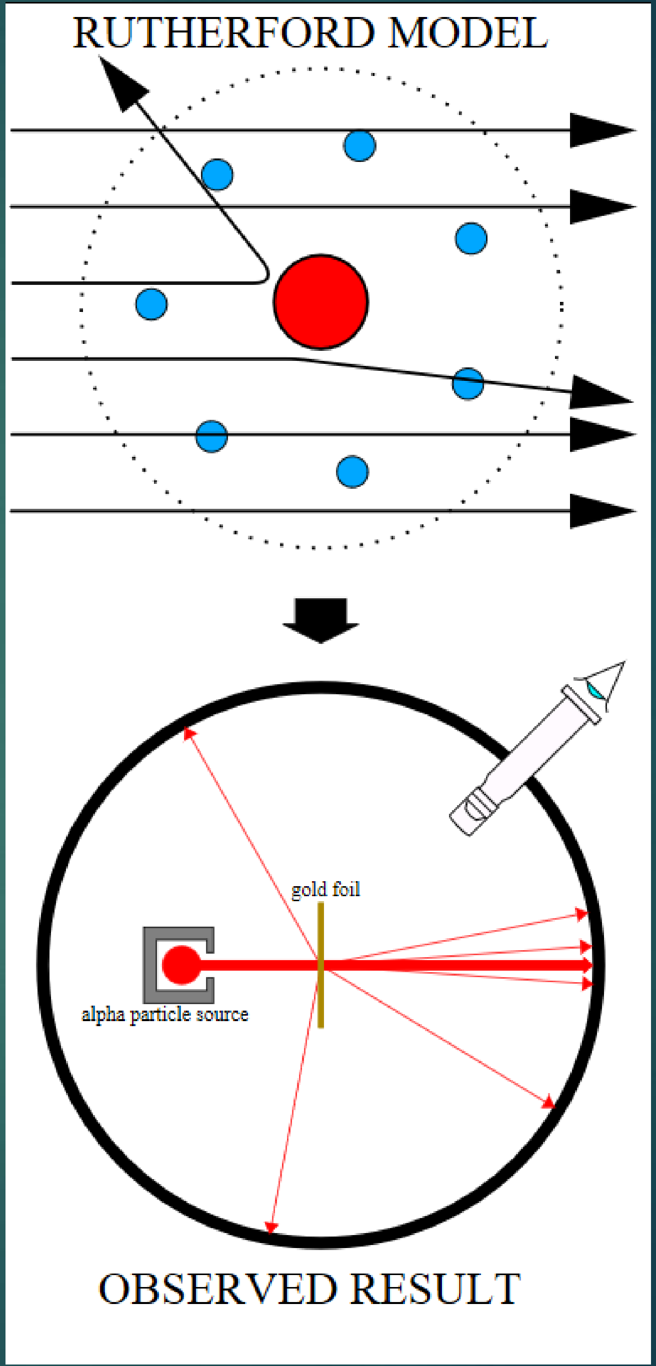
Discovery of the Nucleus

- Earnest Rutherford and coworkers at University of Manchester, England, were the first to theorize, based on experimental evidence, the existence of the atomic nucleus.
- In 1911 he performed the "Gold-Foil" experiment.
- His discovery changed the model of the atom significantly – the first evidence of the atomic nucleus and that atoms are mostly empty space.

Rutherford Experiment

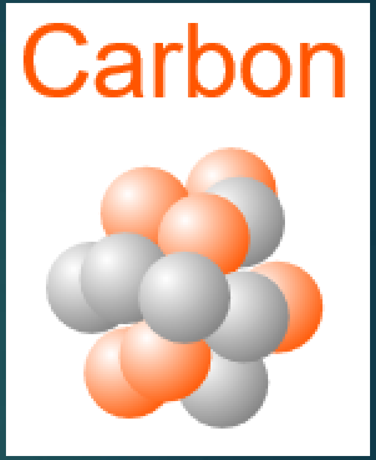


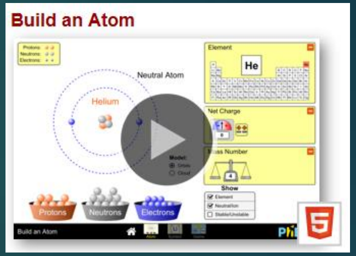




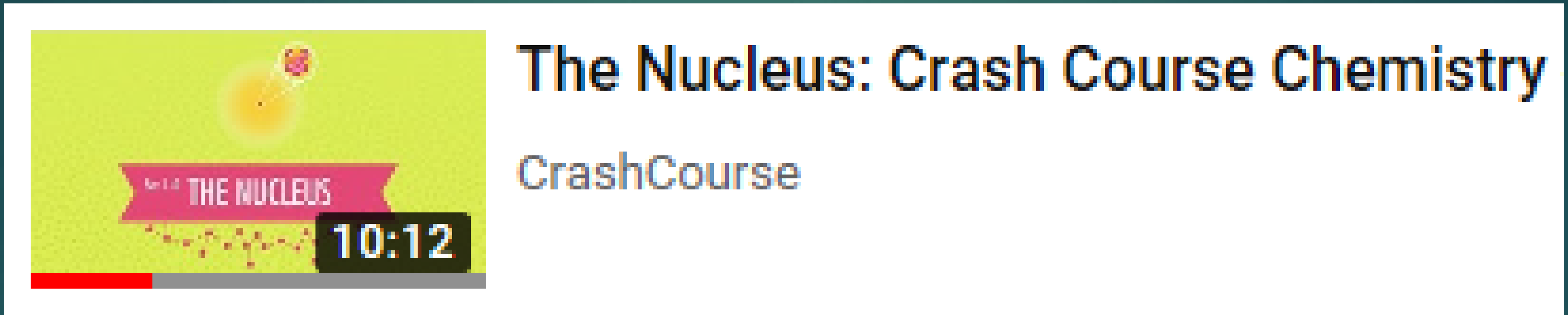
The Atomic Nucleus

- Rutherford's experiment confirmed the presence of a small, dense, area of positive charge. The term ***proton*** was used to name the unseen positive particles. Also, that atoms are mostly empty space.
- It would be 21 years later, in 1932, when physicist James Chadwick discovered the ***neutron***, which also exists within the nucleus to keep protons apart. The neutron is neutral in charge (a charge of zero). The # of neutrons does not have to equal the # of protons in an atom.
- Protons (p^+) and neutrons (n^0) are very close to the same size and mass. Both have a much, much larger mass than the electron (e^-)





The Nucleus: Summary Video



- ➢ But what about ***electrons***?
- ➢ Chemical and physical properties are the result of electrons in the atom.

Atomic Structure – Electron Orbitals

- 1897: Thompson theorized electrons were static in a clump of positive charge.
- 1904: Japanese physicist suggests a central nucleus exists and electrons travel around it like the rings around the planet Saturn.
- 1911: Rutherford's experimental results support the nucleus and he agreed that electrons orbit the nucleus like planets around the Sun. However, it could not explain properties of elements, like why heated metal glowed red/orange.
- 1913: New Zealand physicist Niels Bohr adjusts the model such that electron's have fixed distances from the nucleus, but that electrons can change where they are located in the atom by gaining/losing energy.

Bohr Model of the Atom

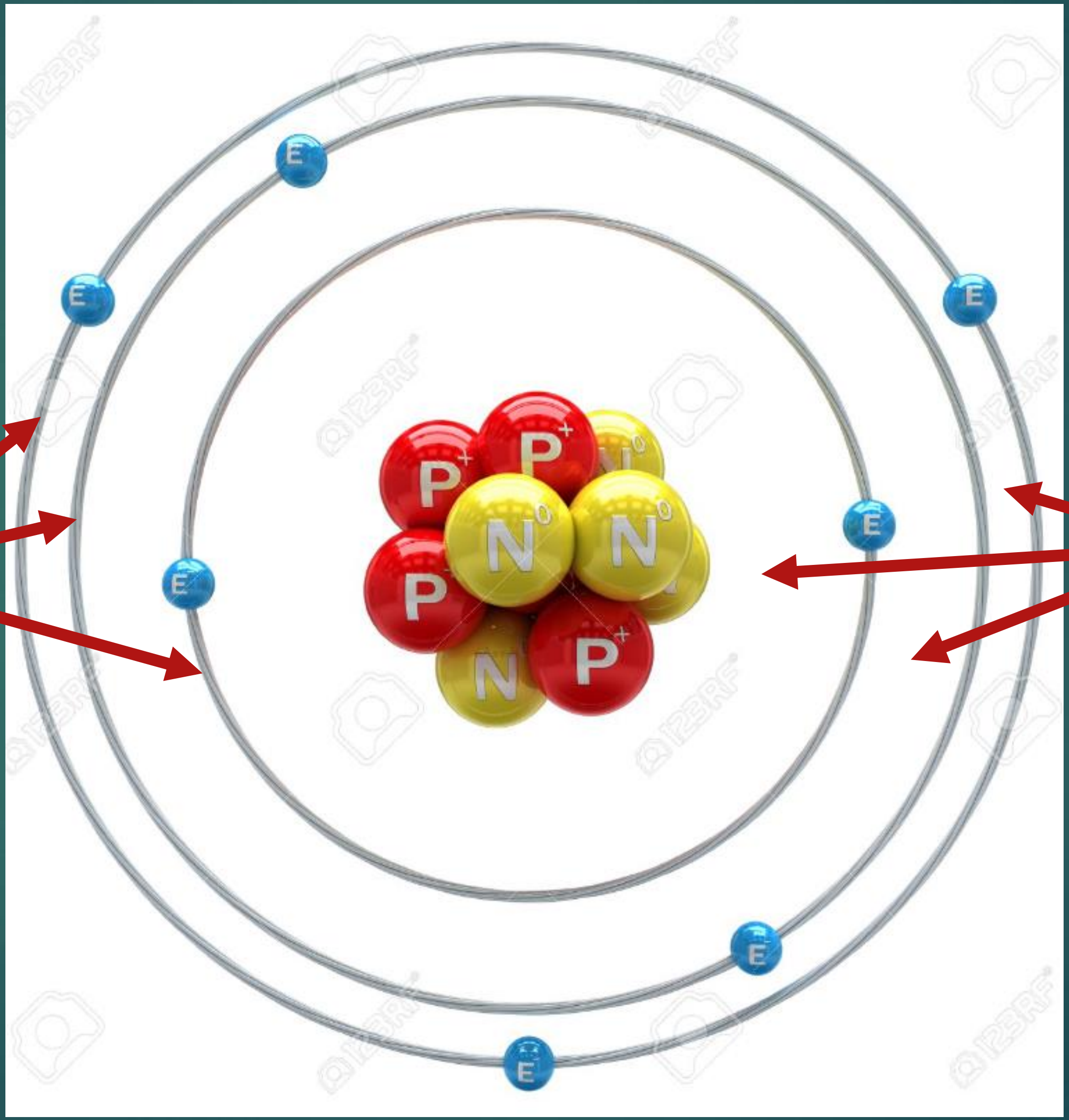
- Explains observations of light coming from the simplest element, hydrogen, but failed for larger atoms, like metallic elements change color when heated.

- The energy electrons have is ***quantized***, they can only have a specific amount of energy at each orbit and they cannot be found at any other orbit.

- Electrons gain or lose a ***quantum*** of energy to change orbital locations around the nucleus.

Bohr Model of the Atom

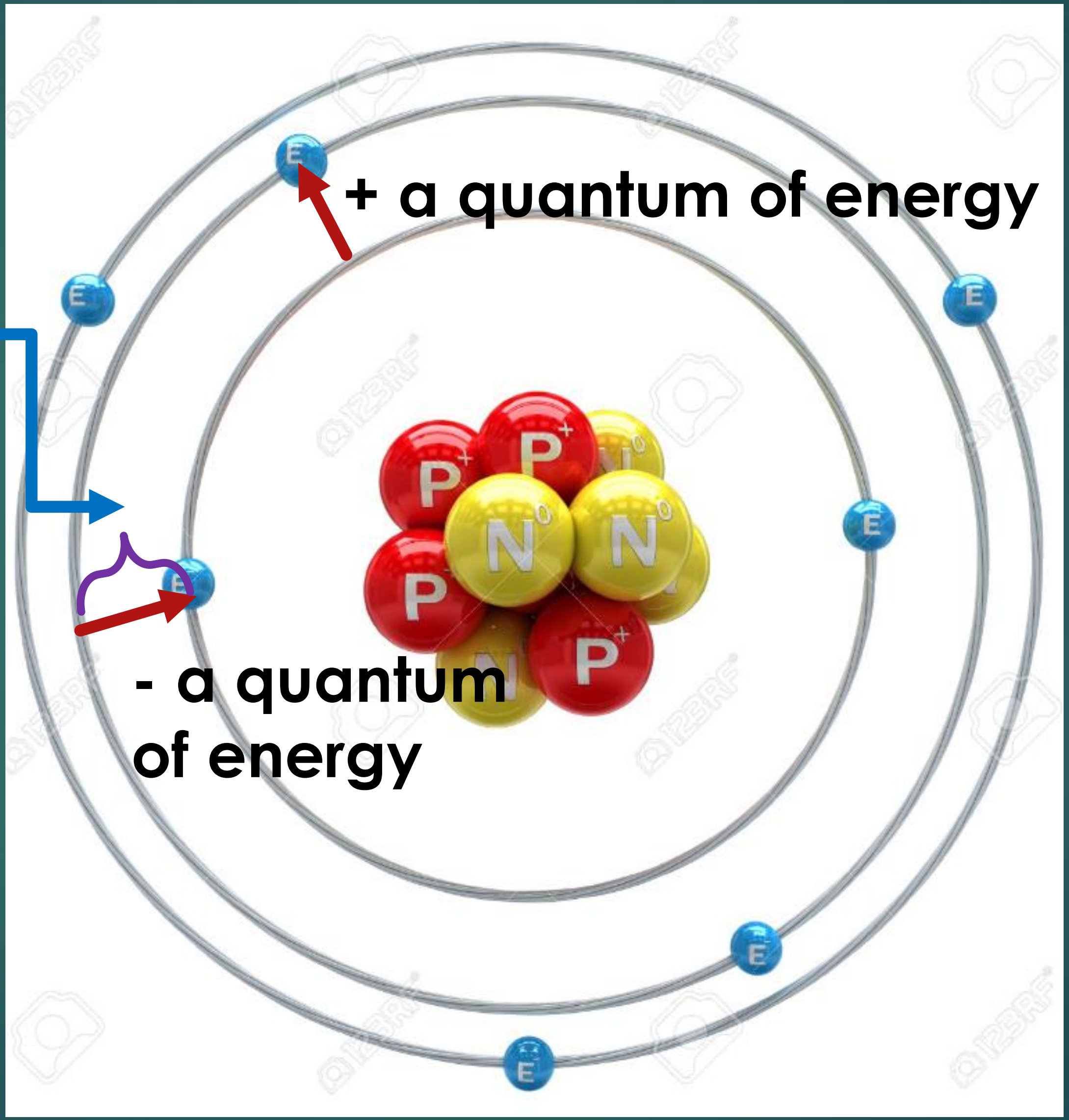
Quantized energy orbitals



Electrons can’t be here!

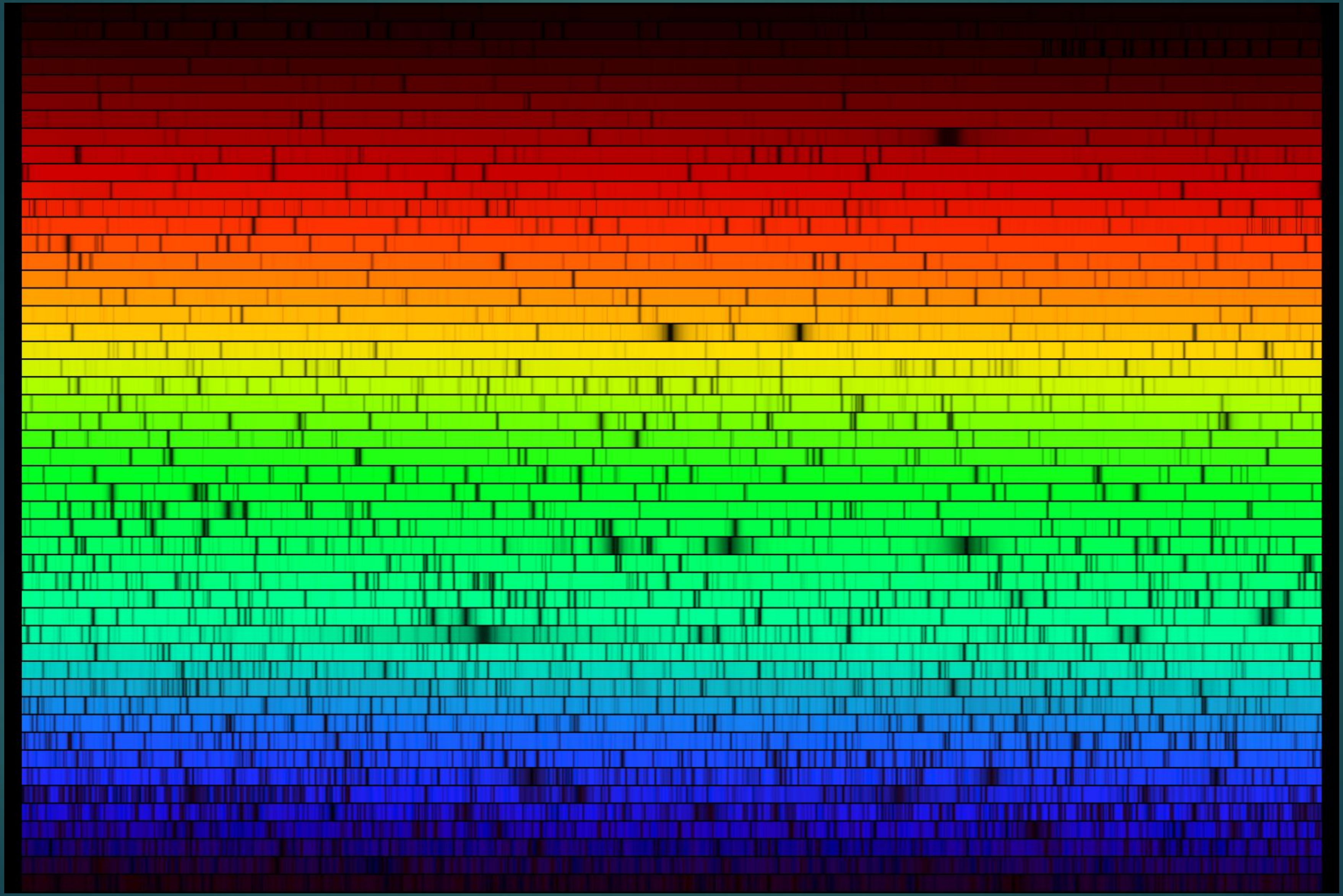
Bohr Model of the Atom

Change in energy is released as radiation, in the case of metals, it could be orange light!



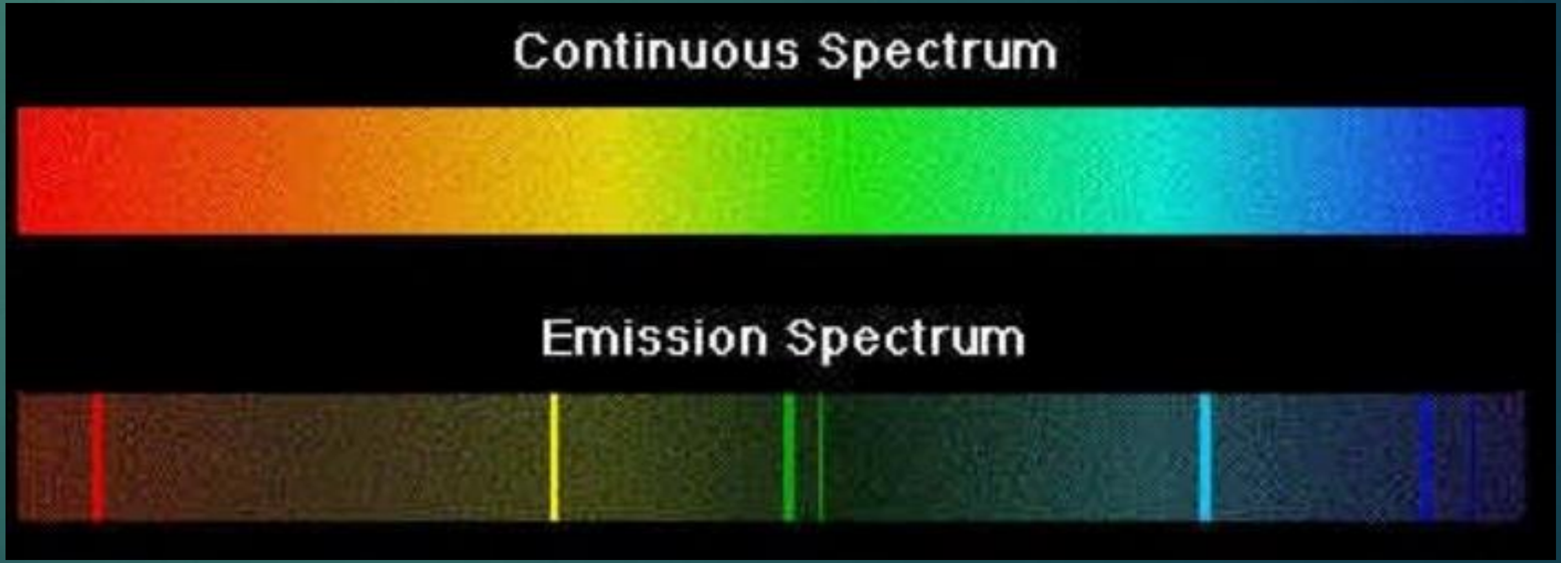
- Again, this model explained hydrogen, but failed for the larger atoms.
- This theory was refined by Erwin Schrodinger in 1926.

Absorbing Energy: Absorption Spectra

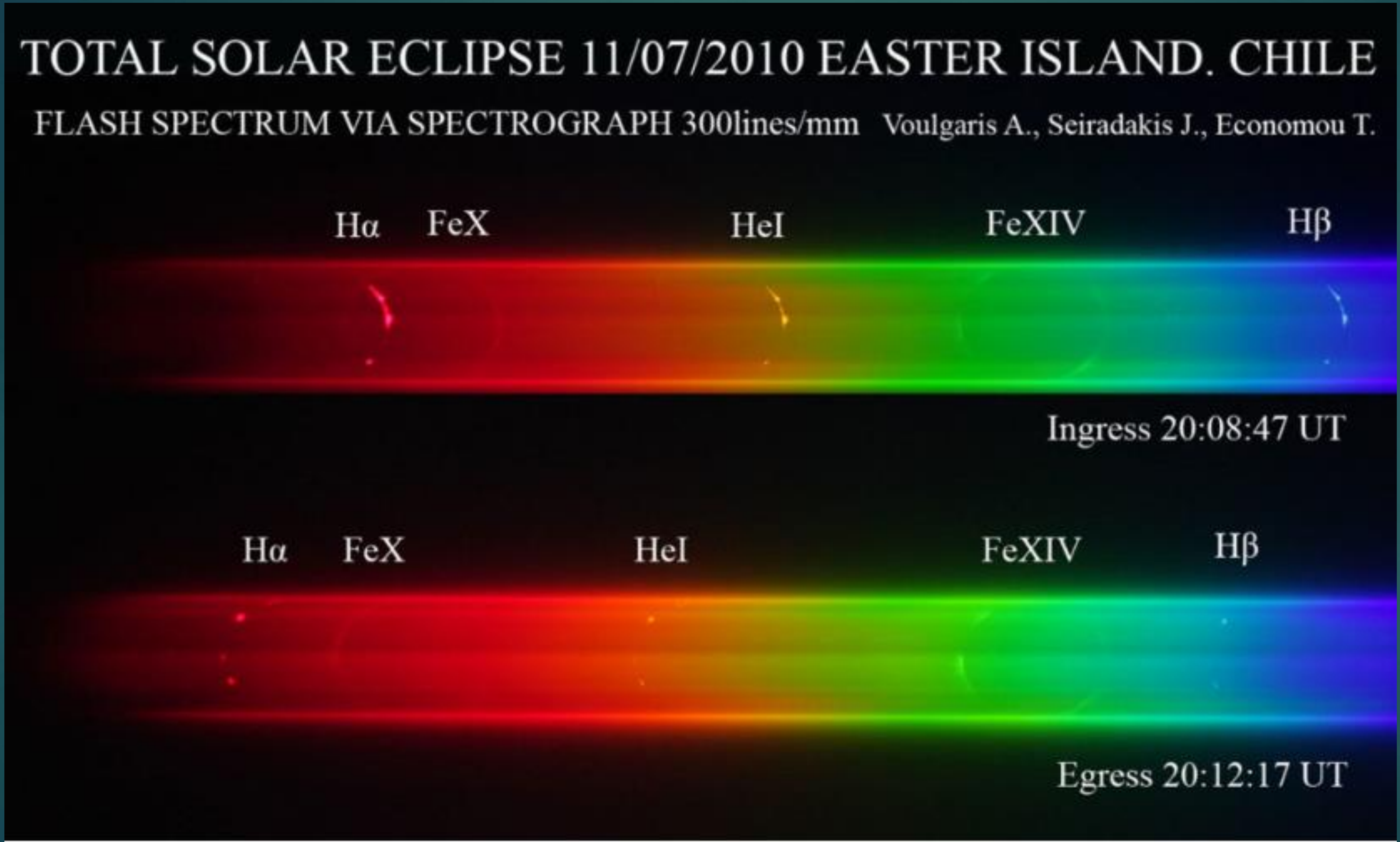


Emitting Energy: Emission Spectrum

- Helium was discovered on the Sun before it was found on Earth.
- During a solar eclipse, extra energies were found being emitted from the hot gas, called the solar corona, surrounding the Sun.



Emitting Energy: Emission Spectrum



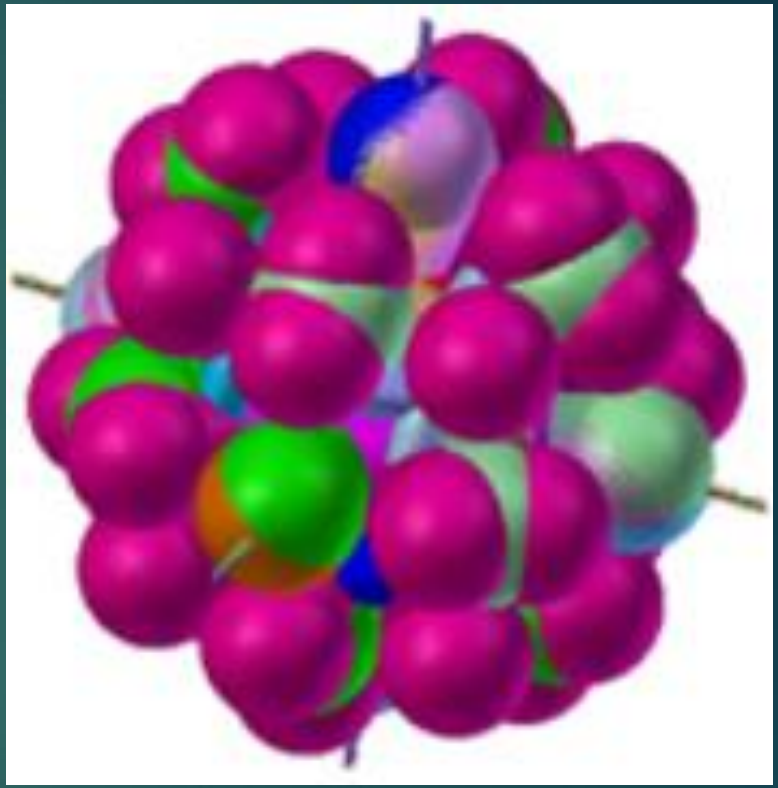
The Flash Spectrum after the 2nd and before 3rd contact from TSE 2010, Easter Island

Quantum Mechanical Model

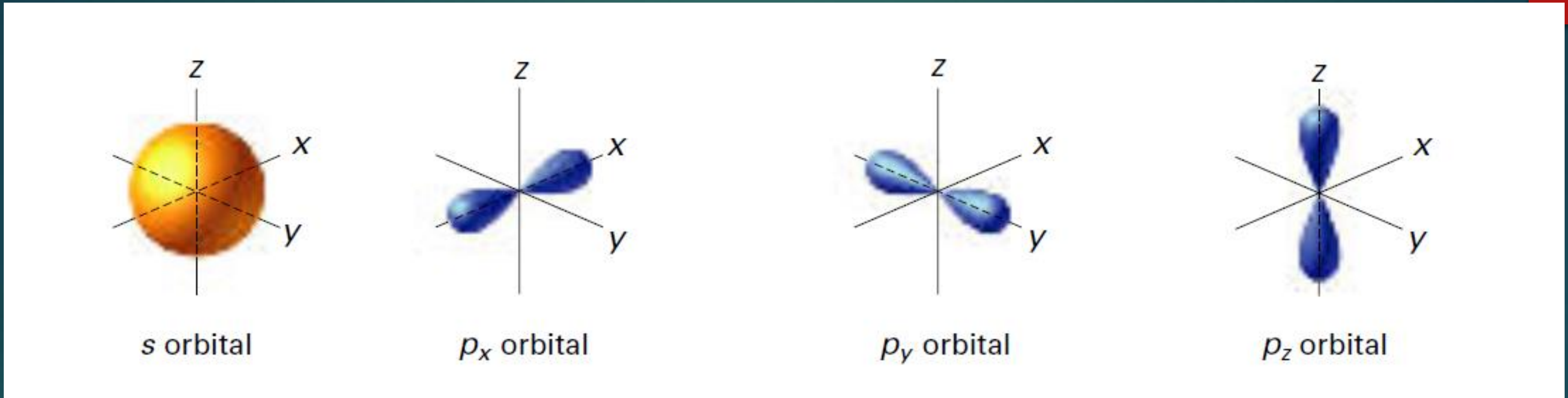
- In 1926 Austrian physicist Erwin Schrodinger spearheaded the development of basis for our current model of the atom, the Electron Cloud Model or Quantum Mechanical Model.
- His model of the atom was mathematical. A mathematical representation of the atom based on all know numerical and scientific information at the time.
- This resulted in a model where electrons have a high probability of being found in a specific region around the nucleus.

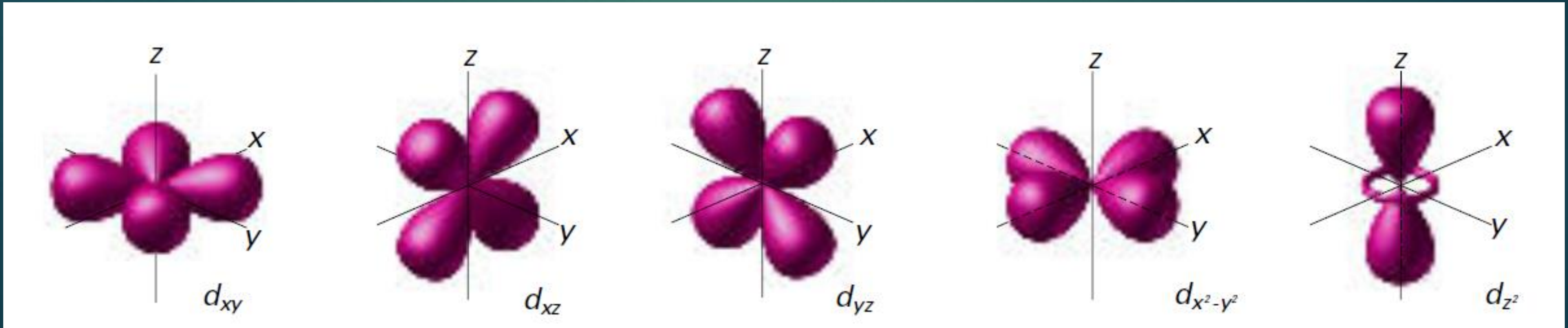
Quantum Mechanical Model

- The regions were not circular orbitals, but blob-shaped 3D spaces.
- Electrons were found to have wave and particle properties.
- Expanded upon Bohr's work and explained the properties of all elements (to a degree, they did not have any electronically powered computers for calculations).

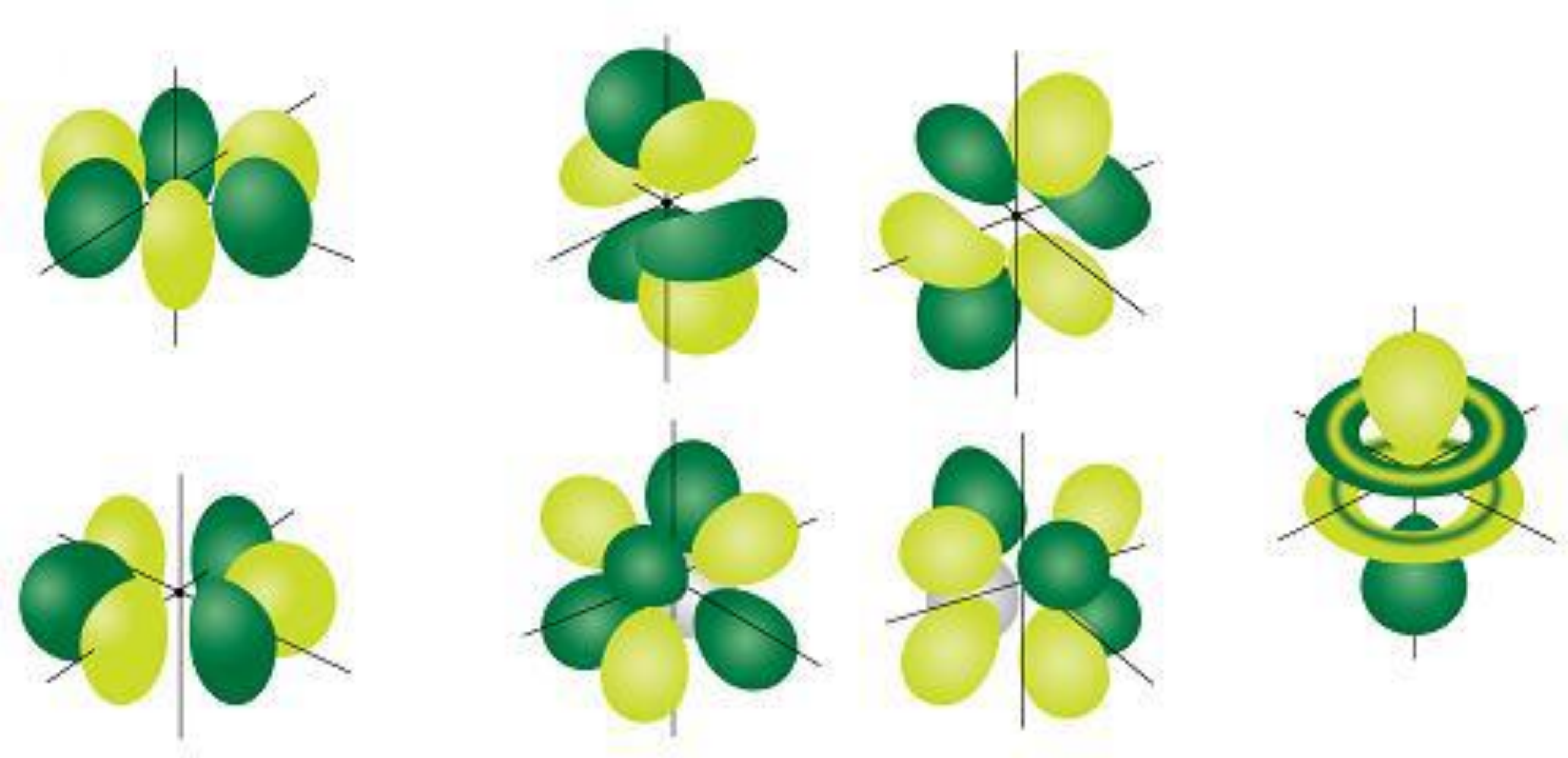


Quantum Mechanical Model

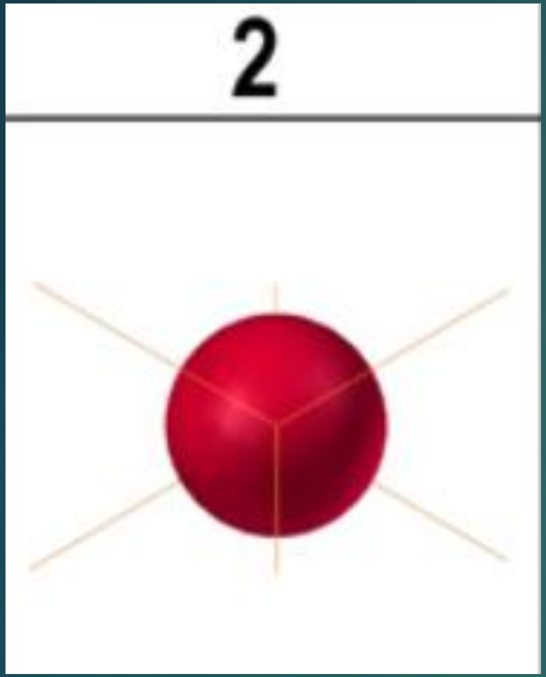


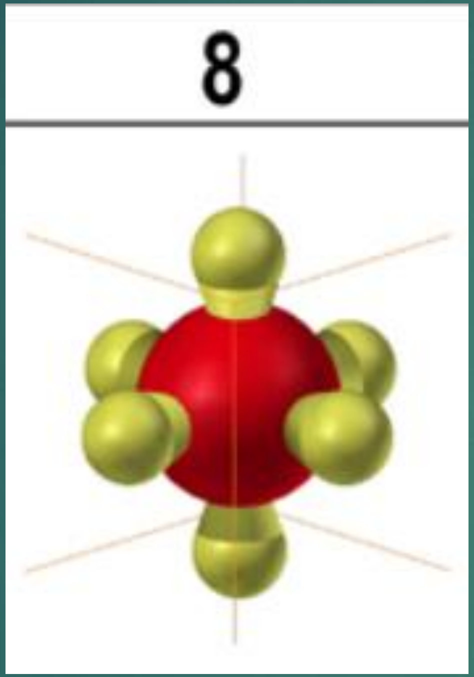


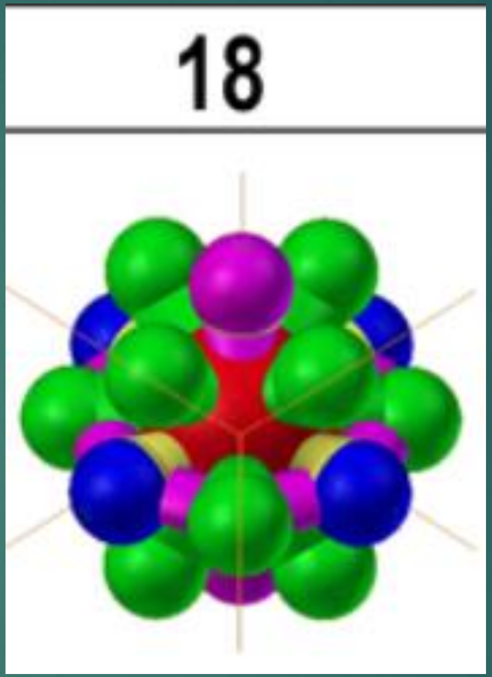
Quantum Mechanical Model

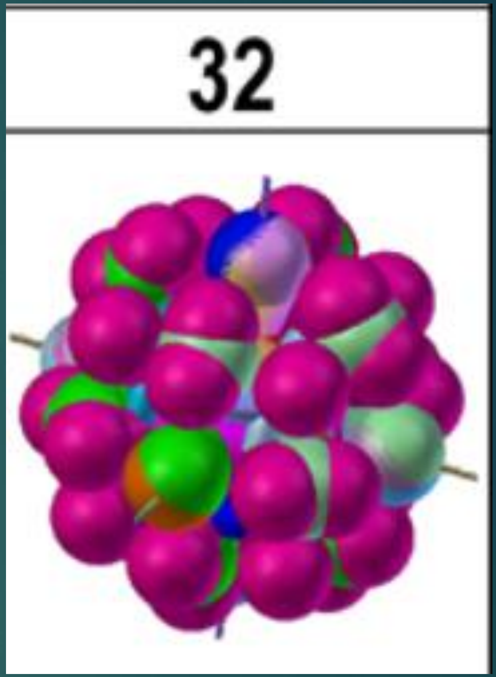


Overlapping Orbitals = Electron Cloud

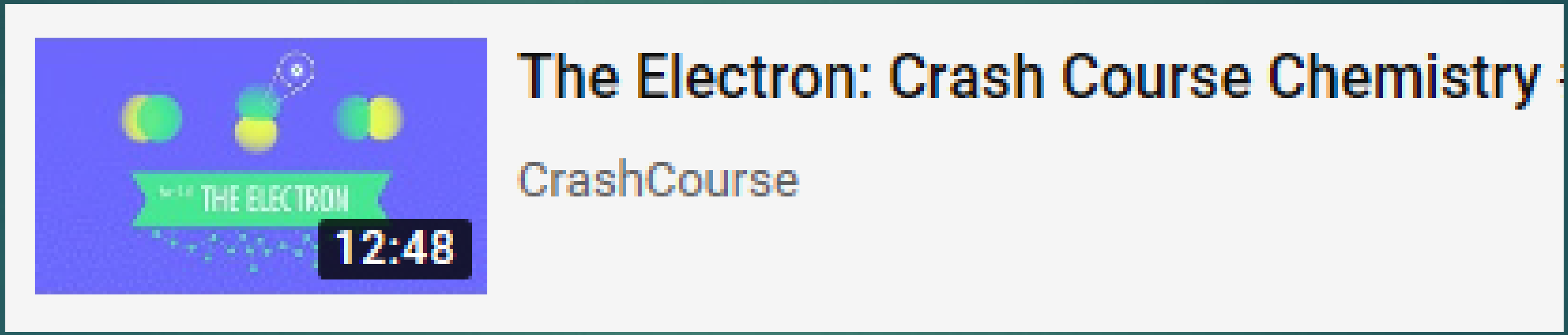








The Electron: Summary Video

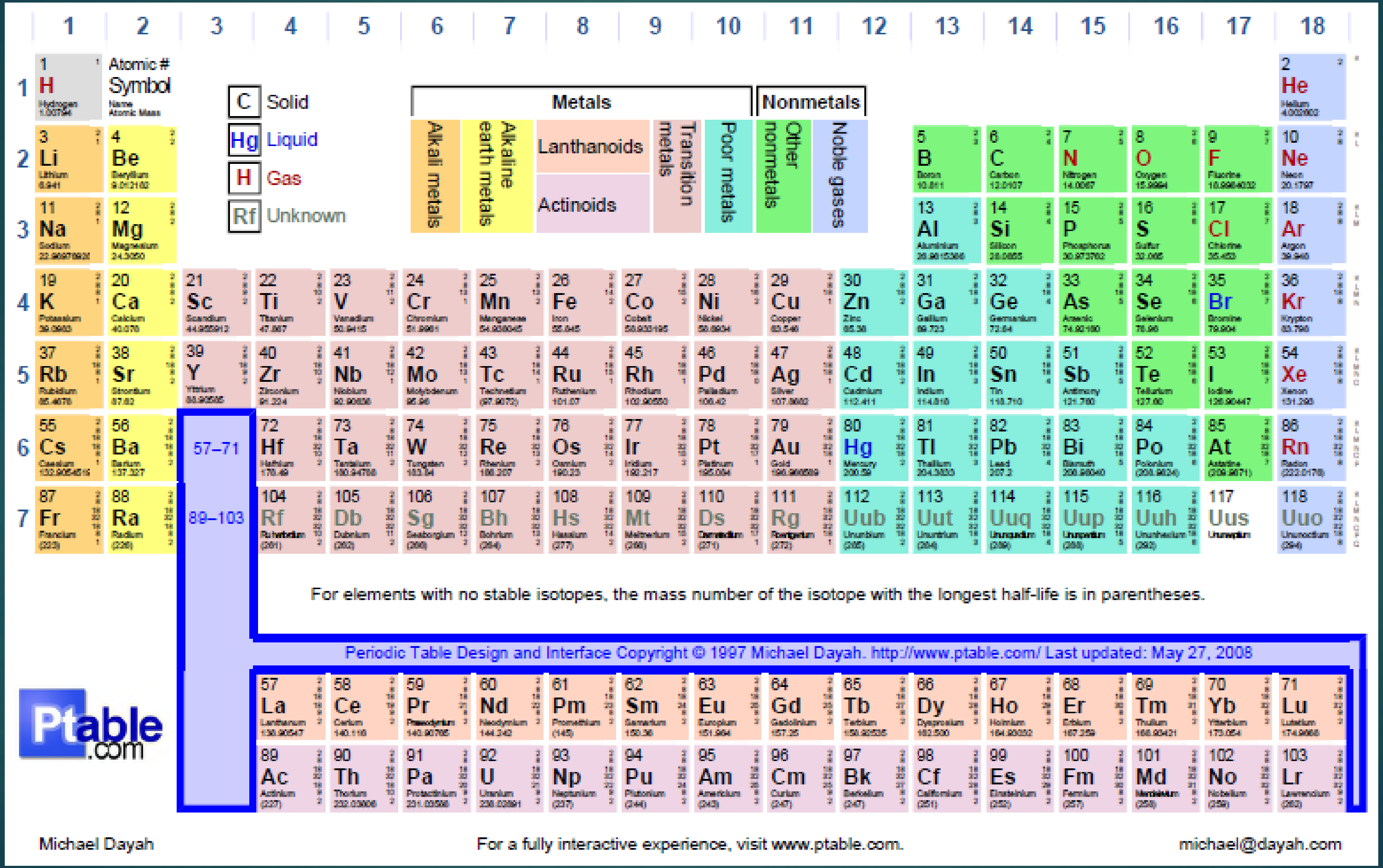


Organizing Elements: The Periodic Table

The Periodic Table: Crash Course Chemistry

CrashCourse

The Periodic Table of the Elements

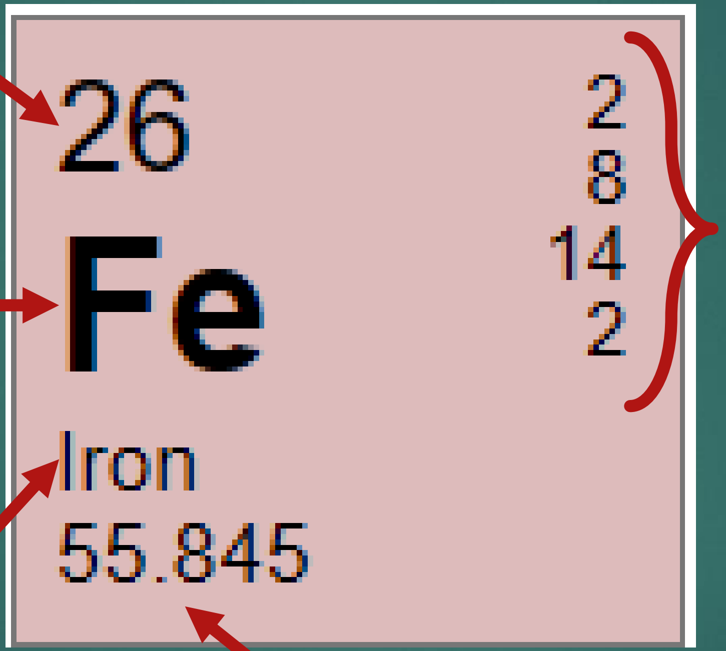


Our Periodic Table

Atomic Number

Element Symbol

Element Name



Electrons In Each Energy Level

Average Atomic Mass

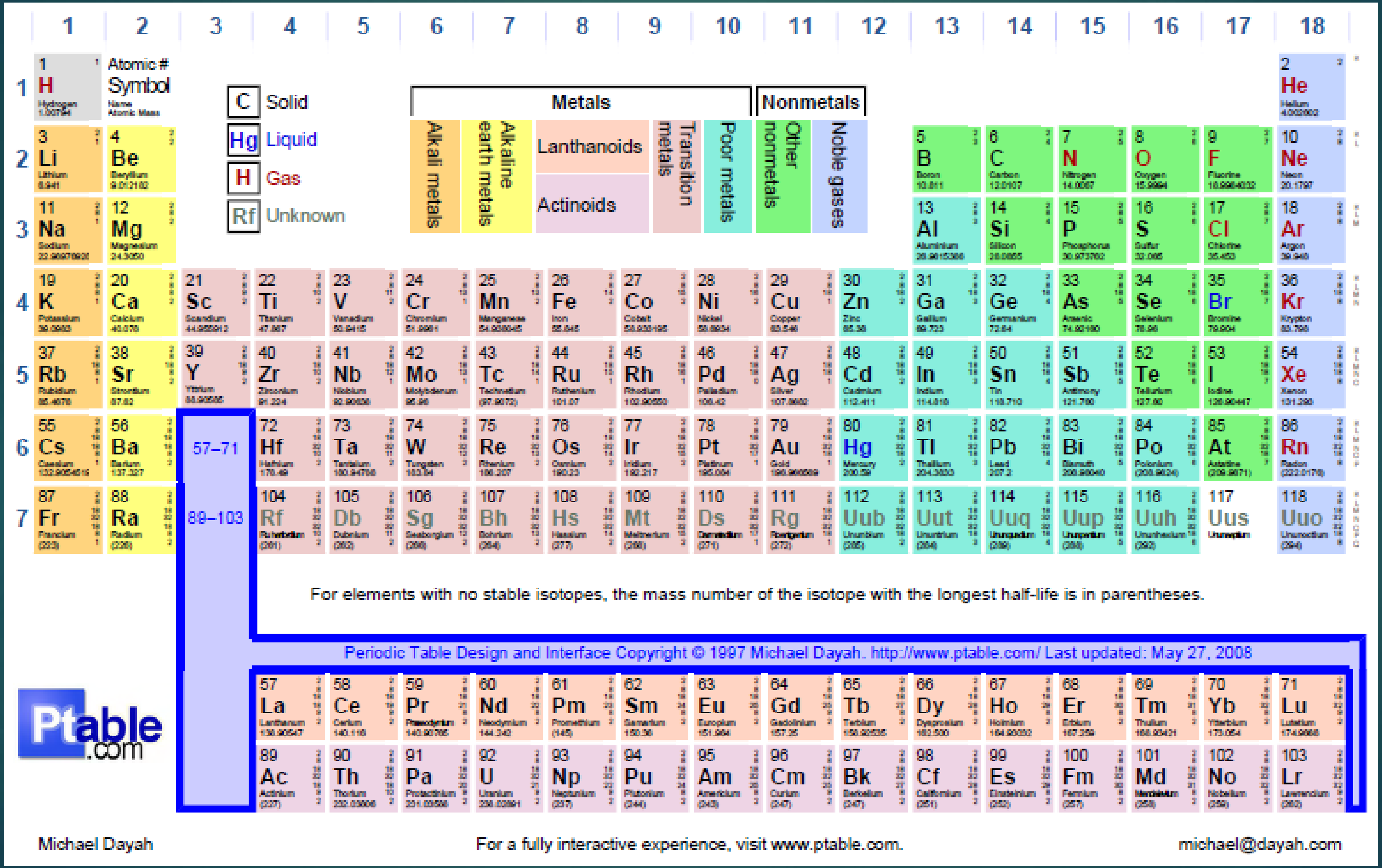
Periodic Table: Element Information

- **Element Symbol**: First letter is always a capital letter. If there is a second letter it is lower case. Two letters maximum.
- **Atomic Number**: The number of ***protons*** in the nucleus; how the table is organized. The number of protons defines the element.
- **Average Atomic Mass**: The mass of the element taking into account its various ***isotopes*** (atoms of the same element but a different number of neutrons). The unit is the ***atomic mass unit***, ***amu***.
- 1 amu is defined as exactly one-12^{th} (1/12) the mass of the Carbon-12 atom (6 protons and 6 neutrons).

The Periodic Law

- When elements are tabled in order of increasing atomic number, there is a periodic repetition of chemical and physical properties.
- **Groups**: The vertical columns; elements in the same column have similar chemical and physical properties (boiling points, luster, conductivity, reactivity, etc.).
- **Periods**: The horizontal rows. As you read left to right, elements get properties of non-metals.

The Periodic Table of the Elements

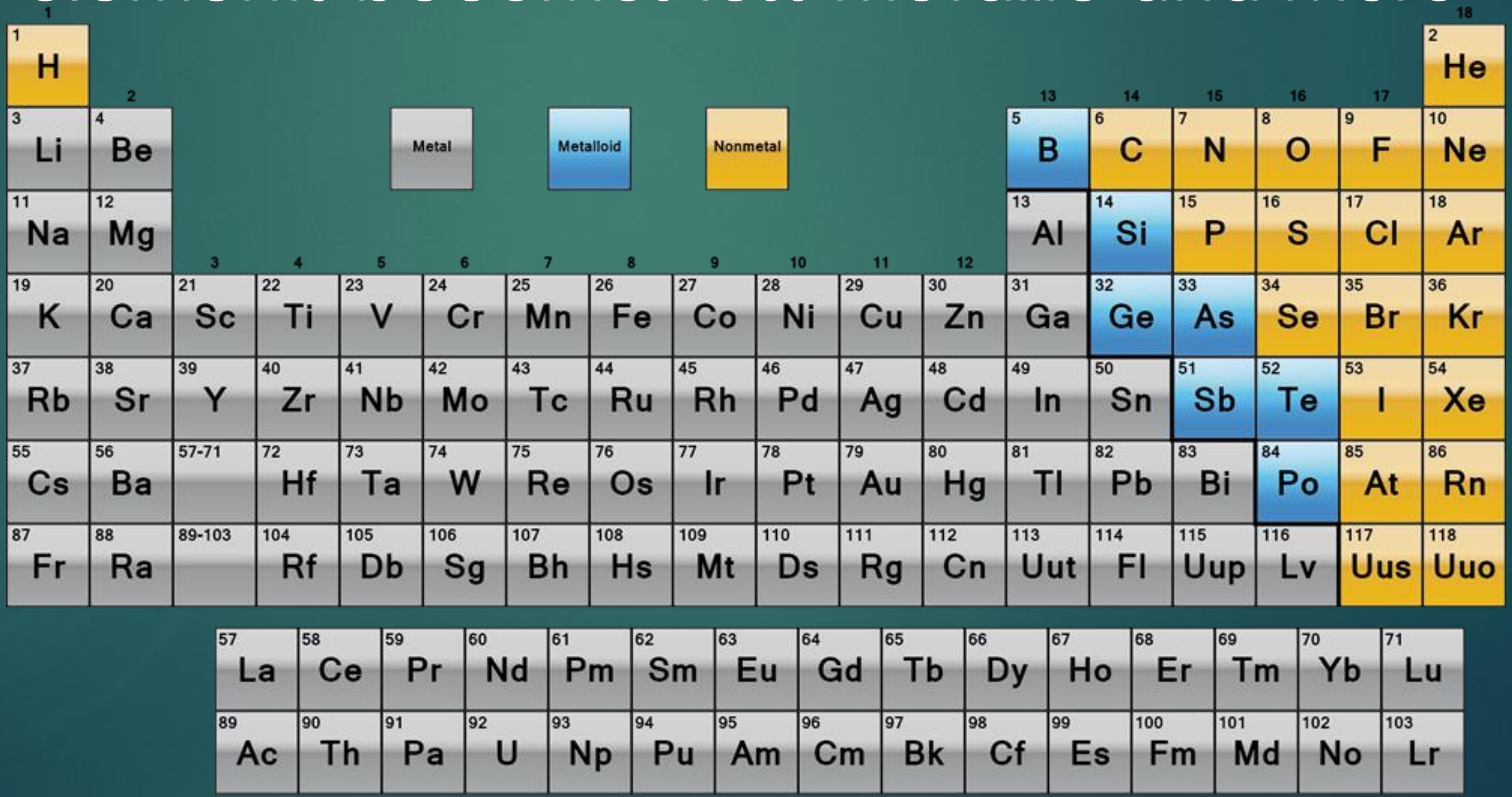


Using the Periodic Table

- What element is atomic number 74?
- What is the symbol for the element with atomic number 82?
- What is the atomic mass of cesium (located in group 1)?
- How many electrons in the 3rd energy level of cadmium?
- What element has the most electrons in group 2?
- What element is located in period 4 group 11?
- How many protons in an atom of carbon?
- How many electrons in the highest energy level of group 1? 17? 18?

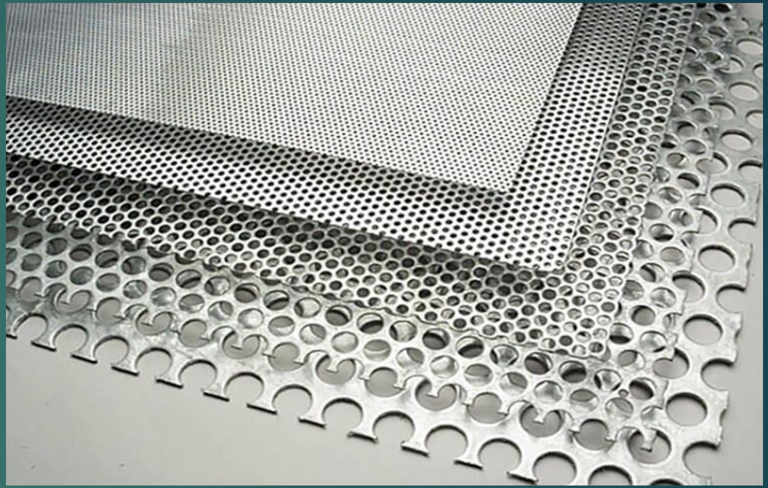
Metals, Nonmetals & Metalloids

- The periodic table classifies/divides elements into one of three groups: **metals**, **non-metals** and **metalloids**.
- Scanning across the periodic table (from left-to-right), the properties of elements becomes less metallic and more nonmetallic.



Physical Properties of Metals

- Most elements are metals.
- Good conductors of heat and electricity.



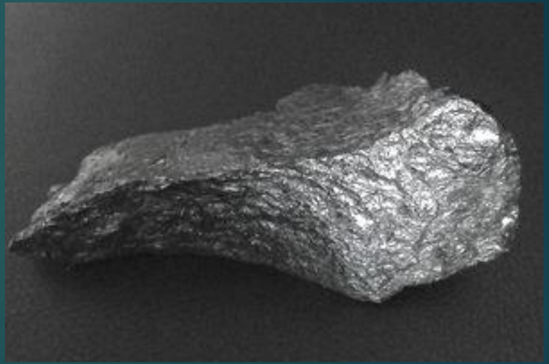
- High luster and sheen; shiny.
- Malleable – hammered into thin sheets.
- Ductile – drawn into wires.
- Solids at room temperature (except for mercury).

Physical Properties of Nonmetals

- State at room temperature varies as many are gases but some are liquids and a few are solids.
- Properties opposite of metals.
- Not good conductors of electricity and heat (carbon is an exception to this)
- Not shiny, so they are dull.
- Not malleable.
- Not ductile.

Physical Properties of Metalloids

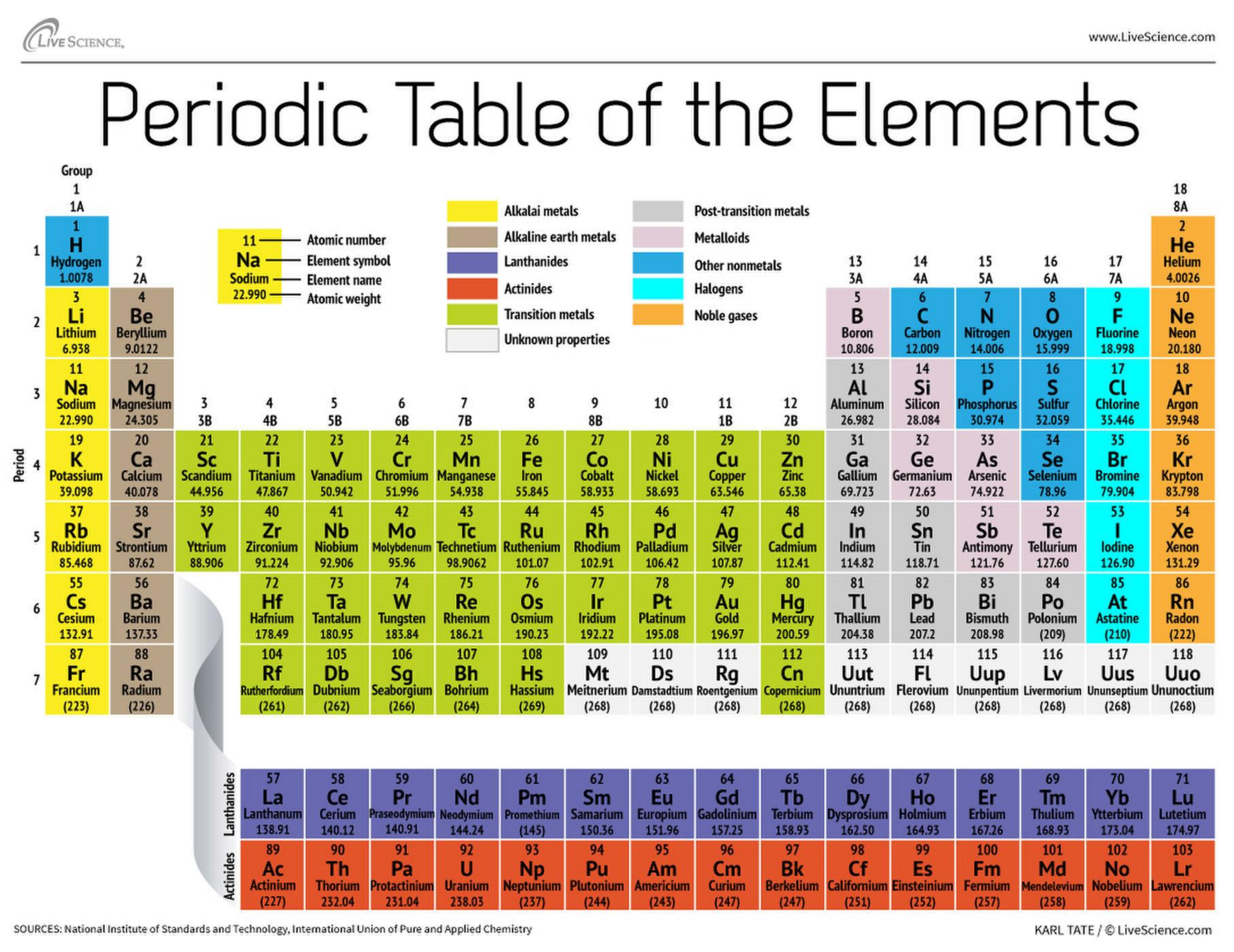
- There is a heavy staircase, bolded line that separates metals and nonmetals.
- The metals that border the line are the metalloids.
- Metalloids tend to have properties of metals and nonmetals; this depends on the conditions the element is under.
- For example, silicon is a poor conductor of electric current, but mix in a small amount of boron and the mixture is a good conductor of electricity (used in electronics).



Groups on the Periodic Table

- Many groups on the periodic table are given a unique name, based on the properties of the elements in that group.

Periodic Table: Group Names



Group 1: Alkali Metals

- All shiny, silvery metals.
- Very violently reactive with water.
- Forms basic compounds with water (baking soda, soaps are bases)
- Form compounds that are mostly white solids and those compounds are very soluble in water (table salt – NaCl).