Chapter 08: Atoms and Periodic Properties

The development of the modern atomic model illustrates how modern scientific understanding comes from many different fields of study. Understanding the atom and all the changes it undergoes not only touches your life directly but also shapes and affects all parts of civilization.

Section 8.1: Atomic Structure Discovered

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/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

Electrons...Finally!

- Existence of charge well understood, but e^- not actually discovered until 1897
- If you understand gravity (thanks, Newton!) and you understand 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

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.

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 good choice). 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

- If most particles pass through, then atoms must be mostly open space
- If some alphas get slammed straight back, then there must be something really massive to run into
- Rutherford calculated that electrons were moving outside the nucleus at a distance of about 100k × the nuclear radius

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 pf protons

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 12C (not quite exactly the mass of one proton)
- Individual atoms have specific mass: Every ^{1}H masses exactly 1.00 amu, but not every hydrogen atom is a ^{1}H 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

Section 8.2: The Bohr Model

Tiny Baby Solar Systems: Not

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

Everything Is Connected To Everything Else

- JJ Balmer (about 1885) tries to make sense of hydrogen emission spectra: Why so many lines????
- Max Planck (about 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 e^-s 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?

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

- 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

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

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 and 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

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

- 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: $1s^2 2s^2 2p^6$ means the first orbit is full ($2e^-$), and the second orbit has two sub-levels, also full (total $8e^-$)

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 Mendeleyev 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 e– 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

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

- Unfilled valence shell: Yes, yes, we know
- Might gain or lose e^-; just as easy to empty the valence shell as to fill it
- Metalloid: Can conduct electricity (metallic property), but typically brittle/non-malleable (non-metallic property)

Noble Gases

- Right-most column on periodic table: Completely full valence shell
- Colorless, odorless, inert gases
- Non-reactive: Do not form compounds with other elements