Unit 10: Periodic Trends

1. Periodic history
   a. Dmitri **Mendeleev** created the first modern periodic table.
      i. He put the elements in order by **mass number**
      ii. He put elements with similar properties in columns (columns are VERTICAL!)
      iii. His table was totally awesome because it could predict properties of unknown elements
      iv. (Lots of other chemists tried to come up with systems. Mendeleev's was the only one fabulous enough to predict properties of elements that hadn't yet been discovered.)
   b. **Moseley** altered Mendeleev's periodic table, changing to elements in order by **atomic numbers**.
      i. This was a major improvement, since there are a few elements that don't go in the right order if you use mass.
      ii. Mendeleev knew there was a problem, and he assumed the masses he was using must have errors.
         (1) (The masses he used were totally valid. He was just wrong– it is PROTON NUMBER (aka atomic number) that controls the properties of an element, not mass number.)
         (2) (In his defense, they hadn't figured out proton numbers yet when he was making a periodic table, so he couldn't really have used them even if he wanted to.)
   c. **The Modern Periodic Law: when elements are in order by atomic number (proton number), their properties repeat *periodically*.**
2. Z_{eff}
   a. This is the symbol for "effective nuclear charge"
      i. "Z" is the symbol for "atomic number" (proton number).
         (1) So a plain Z means "number of protons"
            (a) Since protons are the only *charged* part of the nucleus, Z counts the "nuclear charge"
            (b) (Yeah. They should've just called it "effective protons." I know. Sorry.)
            (c) We're obsessed with *nuclear charge* (number of protons) because that is what gives an element its properties.
      ii. The "eff" stands for *effective*
         (1) Not all protons are considered "effective." (From the point of view of valence electrons.) (Hold on. I'll explain.)
         (2) This brings us to a thing called the "shielding effect"
            (a) Basically, the only electrons we usually worry about are the *outer* electrons (VALENCE electrons).
               (i) Because the valence electrons are the ones that actually do chemical reactions
               (ii) The others electrons just sit there
                  1) These other electrons are called "inner electrons"
                  2) They're necessary, because *they cancel the charge of a bunch of the protons*.
                  3) But they don't get involved in chemical reactions, because they're in the inside layers of the electron cloud.
            (b) So the inner electrons don't react. But because they DO cancel out the charge of protons, they are actually acting as a **shield**. They're protecting the valence electrons from the attractive force of **most** of the protons in the atom.
               (i) Consider sodium. It has 11 electrons and 11 protons.
                  1) Only ONE of the electrons is in the valence.
                  2) The other ten electrons "shield" the valence from the pull of those 11 protons in the nucleus.
                  3) Since one electron can cancel one proton, basically 10 of the 11 protons get cancelled out by the "shielding effect" of the inner electrons.
                  4) Which means there is only ONE proton that is still "effective" after the shield has done its work.
      iii. Wow. Here's a shortcut that saves lives:
         (1) At the beginning of each period (row)
            (a) All the previous electrons become "inner" electrons.
            (b) And they cancel out all the previous protons.
            (c) So the number of valence electrons will always EQUAL the number of "effective nuclear charges" for main group element atoms. (If they are ions, it gets messier.)
         (2) That means Na has one valence electron and one effective nuclear charge because it is in column 1. Same for Li, K, Rb, Cs, Fr.
         (3) That also means that Ne has eight valence electrons and eight effective nuclear charges. Because it is in column 18. Same for Ar, Kr, Xe, and Rn.
         (4) So on a practical level, the "effective nuclear charge" is essentially the number of protons that have been added in the period (row) the element is in.
            (a) It's a little messier in the d-block and f-block, but we're gonna ignore that.
            (b) Also, I'm simplifying the math, because the mathy truth involves some calculus. We're ignoring that too.

b. Why do we even care? Because those "effective nuclear charges" pretty much control the properties of elements. And this whole unit is about the ability to predict element properties from location on the periodic table.
   i. All horizontal (<===>) periodic trends are based on Z_{eff} It is the REASON element properties change horizontally.
   ii. The fact that Z_{eff} starts over at 1 at the beginning of each period (row) is the reason properties start over. It's the reason the properties are PERIODIC (repeating a trend; starting a pattern over at predictable intervals) instead of just one long increase.
   iii. If we weren't building the electron cloud in layers, then none of that would be true.

3.Atomic Radius

A. **Up and to the right**, (electrons are) held more tight(ly)
   1. In any group (column), the **L**ower an element is, the **L**arger it will be.
      a) Each row has one more **energy level** than the previous one. This makes the atoms in each subsequent row larger than in the previous row.
         (1) (FYI: Energy level, also known as "electron shell" is given by the 1st (n) quantum number)
   2. In any period (row), the farther to the **right** an element is, the **smaller** it will be.
      a) Because: the further to the right on the periodic table the more Z_{eff} (effective nuclear charge)
      b) More Z_{eff} holds the electrons more tightly, making the atom smaller.
         (1) Remember: protons and electrons are attracted to each other. (Opposites attract.)
         (2) The more effective protons, the more power the nucleus has to pull the electron cloud in toward it and make the atom smaller.
   3. **L**ower & **L**efter = **L**ooser (electrons) = **L**arger atom
      a) At the bottom of a group (column) atoms are larger because they have more energy levels
      b) At the left of a period (row) atoms are larger because they have less Z_{eff} and therefore don't pull their electrons toward the nucleus as much.

IV. Ionic Radii

A. Cations (positive ions)
   1. Are much smaller than their parent atoms
   2. When the atom is ionized, it loses its valence electrons.
      a) That means it loses its whole outer energy level, which makes it much smaller.
      b) Also, it holds on even more tightly to the remaining electrons. (It's a tug-of-war with the protons winning.)

B. Anions (negative ions)
   1. Are somewhat larger than their parent atoms
      a) When the atom is ionized, it gains electrons in its valence.
      b) The number of energy levels doesn't change, so the size change isn't as dramatic as for a cation.
      c) Anions are larger than their parent atoms because the electrons repel one another.
         (1) The electrons are winning the tug-of-war, so the electron cloud gets bigger.
         (2) The electrons want the electron cloud to be larger so they don't have to be so close together.
            (a) They are attracted to the nucleus, but they are repulsed by each other (Like charges repel. Electrons hate other electrons.)
            (b) If the electron cloud pulls closer to the nucleus, it gets smaller... which means electrons have to be closer to other electrons. They hate that. Because they hate each other. So if there are more electrons than protons, the electron cloud gets bigger. (So the electrons can be further from each other.)

V. (First) Ionization Energy (Unless specifically stated otherwise, we assume "ionization energy" means *first* ionization energy.)

A. The energy required to remove an electron from a neutral atom
   1. NOTE: 'Ionization energy" is ALWAYS about **removing** electrons from an atom. Never the reverse.

B. **Up and to the right**, held more tight
   1. Smaller atoms have higher ionization energies because their valence electrons are closer to the nucleus and therefore more tightly held.
   2. Elements toward the right end of the periodic table have higher ionization energies because they have greater Z_{eff} (effective nuclear charge)
   3. More Z_{eff} holds the electrons more tightly, making them harder to remove (ionize).

C. The zig-zags in ionization energy graphs:
   1. Some electron configurations are extra stable
      a) Full sublevels are extra stable: 2 s-electrons, 6 p-electrons, 10 d-electrons, 14 f-electrons. Stable!
      b) Half-full sublevels are also stable: 3 (out of 6) p-electrons, 5 (out of 10) d-electrons, 7 (out of 14) f-electrons
      c) These "extra stable" configurations are harder to ionize that you would expect.
   2. Some electron configurations are particularly unstable
      a) The first electron in a new sublevel (1 p-electron, 1 d-electron, 1 f-electron)
      b) The first *paired* electron in a sublevel: 4 (of 6) p-electrons, or 6 (of 10) d-electrons

c) These happen to be the configurations that come right after the "extra stable" ones described above.
d) They are really easy to ionize.

3. The combination of an "extra stable" configuration (say Zn, which has a full d-sublevel) right next to a "less stable" configuration (say Ga, which has only one p-electron) creates a little zig-zag in the ionization energy graph.
   a) There will be a zig-zag pretty much every time you move from one block to another
      (1) Say d-block to p-block
         (a) (the first element in the p block will cause a little "dip" in the Ionization Energy graph)
      (2) And again when you get to the halfway point in the block
         (a) Say between the third p-electron (Half-full! stable!)
         (b) And the fourth p-electron (Has to pair up. Hates life. Unstable!)

VI. Successive (subsequent) Ionization Energy

A. It is possible to remove multiple electrons from the same atom.
   1. The more electrons you remove, the harder it gets, so the ionization energy gets bigger each time.
      a) $IE_1 < IE_2 < IE_3$ etc
   2. Because the proton/electron tug-of-war is getting more and more out of balance, so the protons hold on to the remaining electrons more and more tightly.

B. Somewhere in the list of successive ionization energies there will be a big jump in energy.
   1. The number of "easy ionizations" *before the big jump* is the number of valence electrons the atom had.
   2. For example in this list of successive ionizations:
      a) $IE_1 = 577$, $IE_2 = 1815$, $IE_3 = 2740$, $IE_4 = 11600$, $IE_5 = 15000$, $IE_6 = 18310$
      b) The "big jump" was between IE_3 and IE_4
      c) That means the atom being ionized had three valence electrons (because there were three "easy" ionizations before the jump)
      d) That means the element being ionized is in the third column (the B, Al, Ga, In, Tl column)
   3. So the reason we care about "successive ionization energies" is that they can help us determine which column the element is in.

VII. Electronegativity

A. Attraction for electrons shared in a bond

B. This also increases **up and to the right**
   1. Smaller atoms can attract electrons more because their valence is closer to their nucleus, so there is a stronger pull from the protons
   2. Atoms with more Z_{eff} can attract electrons more because... they have more Z_{eff}.
   3. Wow. Like the exact same reasons as ionization energy. Amazing.
   4. Except: **noble gases don't bond**. So they can't have "attraction for electrons shared in a bond." So they don't have electronegativity values.
      a) Fluorine is the most "up and to the right" atom other than the noble gases.
      b) So fluorine has the highest electronegativity of all the elements.

VIII. Electron affinity

A. The amount of energy released when an electron is added to a neutral atom.

B. You're not gonna believe this, but... it increases up and to the right! (What are the odds???)
   1. Why does it increase up and to the right? I bet it has something to do with:
   2. Up: smaller atoms have a stronger pull for electrons because the valence is closer to the nucleus
   3. Right: more Z_{eff} = more attraction for electrons
   4. Wow. Like the exact same reasons as ionization energy AND electronegativity. It's like there's a pattern or something.

C. Actually, those are also the same two reasons as atomic radius–
   1. Increased pull on electrons causes ionization energy, electronegativity, and electron affinity values to be LARGER
   2. Increased pull on valence electrons draws them toward the nucleus and makes the atomic radius SMALLER.
   3. So the atomic radius (size) shows the opposite trend, but for the same reasons. Crazy.

D. **Noble gases don't want any more electrons** (because they are already stable, with full octets and all). So they don't have electron affinity values.

E. Hey! The two trends noble gases don't participate in both start with "electro". Fun fact for remembering!

IX. Melting/Boiling point: increases toward the center of each period. Boring.

X. Fun Vocabulary:

A. Period= row = horizontal!

B. Group = family = column = vertical!

C. Periodic table sections:
   1. Alkali metals (column 1, valence = 1)
   2. Alkaline earth metals (column 2, valence = 2)
   3. Transition metals (the d-block)
   4. Lanthanides (the 4f)

5. Actinides (the 5f)
6. Metalloids (B, Si, Ge, As, Sb, Te, Po, At)
7. Chalcogens (column 6, valence = 6)
8. Halogens (column 7, valence = 7)
9. Noble gases (column 8, valence =8, except He = 2)

D. Zeff = effective nuclear charge
1. The number of protons that are unshielded
2. For all neutral atoms this is approximately equal to the valence number

E. Shielding effect
1. The inner electrons (everything other than the valence) makes a "shield" around the nucleus
   a) Who are we shielding? The valence electrons
   b) Who are we protecting them from? The protons (who want to pull them into the nucleus)
   c) Who is the shield? all of the inner electrons (everything other than the valence)
2. The shield cancels out some of the positive charge so that the valence electrons don't feel as much "pull" from the nucleus.
3. As the shield increases in size, so does the atom.