The History of the Atom

1) Atom – smallest particle of an element that retains its identity in a chemical reaction
2) Size of an atom
   - 100, 000, 000 Cu atoms side by side would make a cm.
3) **Activity – Black Box inference activity.**
4) Democritus (460 BC – 370 BC) – Greek philosopher, believed atoms were indivisible and indestructible.
5) John Dalton (1766-1844) – beginning of modern atomic discovery. He transformed Democritus's ideas on atoms into a scientific theory.
   - Dalton's Atomic Theory
     - All elements are composed of tiny, indivisible particles called atoms.
     - Atoms of the same element are identical. Atoms of one element are different from those of another.
     - Atoms of two different elements can physically mix together or can chemically combine in simple whole-number ratios to form compounds.
     - Chemical reactions occur when atoms are separated, joined, or rearranged. Atoms of one element can never be changed into atoms of another element via a chemical reaction.
6) Atomic structure
   - Atoms are now considered divisible.
   - Electrons – discovered in 1897 by J.J. Thompson (1856-1940) an English physicist. Negatively charged subatomic particle.
   - Draw experiment from pg 104-105.
   - Robert A. Millikan (1868-1953) – US physicist that determined the quantity of a charge of an electron.
   - Eugen Goldstein (1850-1930) – observed rays traveling opposite of cathode rays and assumed they had a positive charge. Later called protons.
   - James Chadwick (1891-1974) – English physicist in 1932 confirmed the existence of a neutral subatomic particle the neutron.
   - Table 4.1 pg 106 summary (1/2000)
   - Ernest Rutherford (1871-1937) – former student of Thomson, gold foil experiment said the atom was mostly empty space with a small dense middle coined the nucleus.

- Nucleus contains protons and neutrons.

Rd: pg. 101-108

Differences Between Atoms

1) Elements are difference because of their subatomic particles.
2) Atomic number (Z) – number of protons.
3) Protons must equal Electrons.
4) Mass number (A)– number of protons and neutrons.
5) Neutrons = Mass number (p & n) – Atomic number (p)
6) Shorthand symbol writing. Mass # on the top Atomic # on the bottom or name-mass number.
7) **Group practice sheet.**
8) Isotopes – same number of protons, but a different number of neutrons. Also different mass numbers.
   - This makes them chemically alike, but leads to some more interesting properties.
9) Atomic mass unit (amu) – one-twelfth the mass of carbon-12 atom.
10) Atomic mass - weighted average of all of the naturally occurring isotopes.
11) **Use Table 4.3 pg 114 to calculate atomic mass on periodic table.**
12) **Do a few examples using tables. (Honors Only!)**

Rd: pgs 110-116
HW: #1-4

Models of the Atom

1) Time line on pg 128-129
    - Dalton
    - Thompson (plum-pudding)
    - Rutherford (couldn't explain properties of elements)
    - Niels Bohr (1885-1962) – proposed that electrons are found only in specific circular paths (orbits) around the nucleus.
        - Energy levels to classify electrons.
        - Quantum – amount of energy required to move between levels.
        - Ex. Ladder rungs not uniformly spaced.
        - Ground State – normal state of an atom. Lowest energy.
        - Excited State – addition of energy to put it above the ground state.
    - Erwin Schrödinger (1887-1961) – quantum mechanical model (accepted today) Based on probability. Electron cloud model.
    - NOTE after all of this Chadwick confirms the neutron.
2) Schrödinger lead to atomic orbitals or the region of space where there is the highest probability of finding an electron. Equation plots out a shape.
    - Energy levels labeled by principal quantum number (n) 1, 2, …
    - Sublevels correspond to different shapes.
    - Table 5.1 pg 131 with explanation.
    - Pictures of shapes.
    - Maximum # of electrons Table 5.2 pg 132

Rd: pg 127-132

Electron Configurations

1) **Activity – Electron configuration sheet with magnetic boards.**
2) Electron config – arrangement of electrons into appropriate orbitals.
3) Rules of assigning electrons
    - Fill the lowest energy levels first.
    - Each orbital can only have 2 electrons and they must have opposite spins.
    - If there is more than one level in an orbital each level must get one electron before a gets two. Sharing.
4) Draw the way to remember energy level order.
5) Include noble gas configurations.
6) Identify elements based on electron configs.
7) Do many examples. **Do examples over Element 20 for honors as well as exceptions.**

Rd. pg 133-136
HW: #5-6

Physics Review

1) Brief review of a wave: amplitude, wavelength, frequency, crest, and trough. Include units and symbols.
2) Inverse relationship between wavelength and frequency and a direct relationship between energy and frequency.
3) Review of electromagnetic spectrum.
4) When atoms absorb energy it makes electrons jump to a higher energy level it will lose energy and will emit light when it falls to a lower level.
5) **Lab – Flame test for atoms.**
6) Explanation of Bohr Model for the H atom.
    - Review reference tables Bohr model for the H atom.
    - Relate identification of color frequency and wavelength of the light emitted to the energy of a photon.
8) Louis de Broglie (1892-1987) – French deduced since light behaved as a particle and a wave (duality principle) could matter.
9) Classical mechanics good for motion larger than atoms while quantum mechanics describes motion of subatomic particles.
10) Heisenburg uncertainty principle – impossible to know the position and velocity of a particle at the same time.

Rd: pg 138-145
HW: #7