# Unit 8: The Atom Review

1. History of Atomic Theory
   a. Greeks:
      i. Democritus: suggested that there was a "smallest particle" of matter, which couldn't be broken down further. He called it the atom.
      ii. Aristotle: suggested that a few elements could be combined in different ways to make different compounds
   b. John Dalton: Proposed the "Atomic Theory"
      1. All matter is made up of tiny particles called atoms.
         a. Dalton originally said "atoms are indivisible." We no longer believe that—we know atoms can be separated into protons, neutrons, and electrons.
         b. We revised Dalton's postulate to say "atoms are the smallest particle that retain the properties of the element."
         c. The rest of Dalton's atomic theory stands as originally written. When new evidence is gathered, old ideas are *modified*, but not tossed out altogether.
      2. All atoms for a given element are identical
      3. Atoms cannot be created or destroyed in a chemical reaction; they just get rearranged
      4. Atoms can bond together to make compounds with different chemical and physical properties
      5. The Law of Definite Proportions:
         a. For a given compound, the elements are always present in the same ratio.
         b. Water always has twice as many hydrogen atoms as oxygen atoms– that's what makes it water
      6. The Law of Multiple Proportions:
         a. If the ratio between the elements is changed, it will result in a different compound with different properties
         b. Two hydrogens per one oxygen makes water; two hydrogens per two oxygens makes hydrogen peroxide. Definitely not the same. (Don't drink the hydrogen peroxide!)
   c. Research on subatomic particles
      i. J.J. Thompson: the Cathode Ray Tube
         1. Results:
            a. If you change the electric field or put a magnet near a cathode ray tube, the beam shifts.
            b. (The beam is made of electrons.)
            c. You get the same "cathode ray" regardless of what elements are used.
         2. Conclusions:
            a. There is a tiny particle that can be removed from atoms of all elements.
            b. The particle has almost zero mass.
            c. The particle (the electron) has a negative charge.
            d. The technology behind this is how a CRT television/computer screen works.
      ii. Ernest Rutherford: the Gold Foil Experiment
         1. Results:
            a. Shooting alpha particles at a *very* thin sheet of foil.
            b. The particles were expected to pass right through, and most of them did.

c. But some particles veered off course, and one in 8000 particles actually bounced back toward the gun.

2. Conclusions:
    a. Since alpha particles are positive, the ones that came back must have been repulsed by something with a concentrated positive charge (because like charges repel)
    b. Since alpha particles aren't weightless, the concentrated positive charge must also have a concentrated mass.
    c. Since only one in 8000 bounced back, the positive mass that deflected the alpha particles must be very small compared to the size of the atom.
    d. Since most of the alpha particles went straight through the foil, the majority of the atom must be empty space (with a few electrons in it).
    e. The concentrated center is called the nucleus.

3. We say that Rutherford "discovered" the nuclear atom-- because he discovered the existence of the nucleus and its basic properties.

iii. Niels Bohr: "Quantized the atom"

1. Bohr realized that the electrons can't be wherever they want around the nucleus (or they would eventually fall in, since opposites attract).

2. Only certain amounts of energy– and therefore certain radius "orbits"– will work. The other amounts aren't "allowed."
    a. On an orbital diagram, there isn't anything between 1s and 2s– because there isn't an "allowed" energy between them.
    b. An electron is picky! It will only absorb certain amounts of energy– the amounts of energy that can help it get to a different orbital. Since nothing between the orbitals is "allowed", the electron ignores the in-between amounts of energy (because they can't do it any good).

3. This explains why a given atom (hydrogen, for example) always gives off the same colors of light and never any others—the only colors an atom is interested in are the ones with the right amount of energy to move electrons between its *allowed* energy levels.

iv. Thomas Young: the Double Slit Experiment

1. Results:
    a. Electrons shot through a double-slit opening create interference patterns.
    b. *Particles* should not do this.  (Only waves create interference.)

2. Conclusions:
    a. Electrons must be waves—or be capable of acting as waves.

v. Albert Einstein: the Photoelectric Effect

1. Results:
    a. When light of a certain minimum frequency (threshold frequency) is shined on metal, it causes electrons to be ejected from the metal.
    b. Each metal has a different threshold energy.

2. Conclusions:
    a. Different frequencies (colors) of light have different energies.
    b. Intensity (brightness) of light is NOT a measure of energy. (A bright light gives off more photons per second, but the brightness is not an indicator of energy, frequency, or wavelength.)
    c. Light has momentum, therefore it must be a particle—or be capable of acting as a particle.  (Particles of light are called photons.)

vi. Louis de Broglie

1. Wave-particle duality

2. Building on the conclusions of Young and Einstein, de Broglie suggested that both electrons and light (and other things) have a dual nature.
   a. They act like particles in some situations
   b. They act like waves in other situations
   **c.** The double-slit experiment which showed that electrons can experience **interference** and **diffraction** despite being considered particles most of the time.
   d. The photoelectric effect showed that light can act as a stream of particles, having momentum and a specific frequency for each photon, regardless of intensity.
vii. Werner Heisenberg: the *Heisenburg Uncertainty Principle*
   1. Observing an electron requires a light wave to bounce off of the electron (and into the eye of the observer).
   2. The lightwave transfers energy/momentum to the electron (as in the photoelectric effect.
   3. The electron's position/velocity change due to its interaction with a lightwave. The act of observing the electron *changes* the electron—so measurements of electrons will ALWAYS be inherently flawed.
      a. Knowing the exact velocity means you cannot know the position.
      b. Knowing the exact position means you cannot know the velocity.
      c. Measuring one of these quantities changes the value of the other, because observing an electron causes it to move.
viii. Erwin Schrödinger: the Schrödinger Equation(s)
   1. An equation to define the 'wave state' or 'wave shape' of an electron cloud.
      a. Electrons don't really follow little circular orbits around the nucleus– they occupy "wave-states"
      b. This explains why an electron would occupy a space shaped like a p-orbital (∞). The funky shape comes from a wave pattern.
      c. The electron moves at such a high speed that it effectively occupies *all* the volume defined by its path. This volume is called the *electron cloud*.
         i. Bohr's proposed electron *orbits* were 2-dimensional (like the Jimmy Neutron symbol). *We don't believe in **orbits** anymore.*
         ii. Electron cloud **orbitals** predicted by the Schrödinger equation are 3-dimensional. This is the modern theory.
   2. The Schrödinger wave equation can be used to determine the probability of finding an electron at certain point. (But we don't bother to solve that equation in high school chemistry because it is seriously ridiculous.)

2. Subatomic particles:
    a. Proton
        i. Charge = +1
        ii. Mass = 1 amu
        iii. Location = nucleus
        iv. Nuclear symbol: $^{1}_{1}\text{p}$
    b. Neutron
        i. Charge = 0
        ii. Mass = 1 amu
        iii. Location = nucleus
        iv. Nuclear symbol: $^{1}_{0}\text{n}$
    c. Electron
        i. Charge = -1
        ii. Mass .0
        iii. Location = electron cloud
        iv. Nuclear symbol: $^{0}_{-1}\text{e}^{-}$
3. Isotopes and subatomic math
    a. Atomic Number (Z)
        i. Is the number of *protons* in the atom.
            1. The protons determine the identity of the atom.
                a. If you change the number of protons, you've changed to a different element
                b. (It is impossible to change the number of protons in a normal chemical reaction. That is why chemical reactions have to have the same elements on both sides of the arrow.)
        ii. Z can be the same for two atoms that have *different* atomic masses.
            1. These are called isotopes.
                a. A pair of isotopes will have the same number of protons in their nuclei
                b. But they will have different numbers of neutrons in the nuclei (and therefore different masses)
            2. If you want to talk about an isotope you have to identify it by its mass
                a. If I say "carbon-12" everyone knows I am talking specifically about the isotope with 6 neutrons (and a mass of 12)
                b. If I say "carbon-14" everyone knows I am talking specifically about the isotope with 8 neutrons (and a mass of 14)
                c. If I just say "carbon" everyone will assume I am talking about a *mixture* of all the naturally occurring isotopes of carbon– some with 6 neutrons, some with 8 neutrons– and I will use the *average mass* of carbon (12.011) that is on the periodic table. (The average mass is a *weighted* average.)
    b. Mass number (A)
        i. Is based on how many protons + neutrons.
            1. Mass number *minus* protons = neutrons
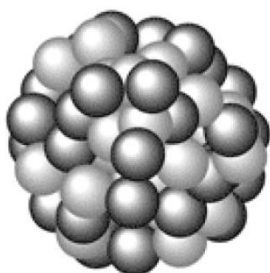            2. A - Z = neutrons
    c. Electrons!
        i. If there is no charge written next to the symbol of the element, it has the same number of electrons as protons (ie, the atomic number (Z)).
        ii. If there is a charge:
            1. The number of electrons = the atomic number *minus* the charge.
                a. This works for positives: $Ca^{2+}$

i.   Atomic number (Z) = 20
ii.   Charge = +2
iii.   20 - 2 = 18 electrons
b.   And for negatives: $N^{3-}$
i.   Atomic number (Z) = 7
ii.   Charge = -3
iii.   7 - ⁻3 = 10 electrons
2.   When in doubt, remember, compared to the number of protons
a.   (Protons ALWAYS equal the atomic # (Z))
b.   Negatively charged things should have EXTRA electrons
c.   Positively charged things should have FEWER electrons
d.   Nuclear symbols (these explanations are based on the graphic above)
i.   Protons

## The Nucleus

chemical symbol, $X$, for the element

$$^A_Z X$$

atomic number, $Z$, equals the number of protons in the nucleus.

mass number, $A$, equals the number of nucleons (protons plus neutrons) in the nucleus.

neutron number, $N = A - Z$

*Isotopes* of an element have the same number of protons but a different number of neutrons in the nucleus, in other words, the same atomic number, $Z$, but different neutron number, $N$, and, therefore, different mass number, $A$.

### Isotopes of Hydrogen

$^1_1H$     $^2_1H$     $^3_1H$

normal hydrogen          deuterium          tritium
nucleus                  nucleus            nucleus

1.   All three isotopes of hydrogen have 1 proton
2.   That's what makes them hydrogen– and why they have the symbol H
ii.   Neutrons
1.   the normal hydrogen isotope has 1-1 = 0 neutrons
2.   the deuterium isotope has 2-1 = 1 neutron
3.   the tritium isotope has 3-1 = 2 neutrons
iii.   Electrons
1.   All three isotopes of hydrogen have 1 electron
2.   Because they all had 1 proton
3.   There is no charge on any of the nuclear symbols in the graphic

4. So the # protons and the # electrons must be the same

e. Isotope Names

    i. Have the format: Name-A

        1. Like "Carbon-12" or "Carbon-14"

        2. The name is followed by a hyphen and the mass number of the isotope

        3. The isotope name gives the same information as the nuclear symbol

    ii. Since the name is given:

        1. You can look it up on a periodic table

            a. The atomic number (Z) will be given in the periodic table square

            b. The proton number equals the atomic number.

        2. The mass number is the number after the hyphen

            a. mass number - atomic number = neutrons (same as before)

            b. A - Z = neutrons

        3. The electron number still equals the proton number unless a charge is given.

4. Average Mass

a. Is the mass given on the periodic table, which has decimals at the end.

    i. It isn't a nice round number

    ii. Because it comes from *averaging* the masses of all the naturally occurring isotopes.

    iii. To find average mass:

        1. You need to do a *weighted* average.

        2. For each isotope:

            a. Multiply each % abundance by its isotope mass

            b. Add all those the answers together.

            c. Divide by 100.

                i. Psst. The percents better add up to 100% total, because that's what percent means.

        3. If there is 90% of isotope A, and only 10% of isotope B, you would expect the average to be closer to the mass of A, since it is more abundant.

        4. We measure atomic mass in terms of "atomic mass units"

            a. (You can abbreviate it: amu)

5. Electron Emissions & Spectra

a. $E = h \cdot f$

    i. E: the energy absorbed/released by the electron that changes energies.

        1. E is measured in joules (J)

    ii. h: Planck's constant.

        1. The value of h is $6.626 \times 10^{-34}$ J·s

    iii. f: the frequency of the light absorbed/emitted

        1. "f" has units of 1/s

        2. 1/s is also called Hertz (Hz)

b. Different frequencies of light are different COLORS.

    i. $c = \lambda \cdot f$

        1. c: the speed of light

            a. this is a constant

            b. its value is $3 \times 10^8$ m/s

        2. $\lambda$: the wavelength of the light

            a. This needs to be in meters (m)

            b. Problems often give this value in nanometers– then you have to convert it.

                i. 1 nm = 1x10-9 m

            c. the name of that Greek letter $\lambda$ is: lambda

        3. f is the same as it was in the previous equation, mmkay?

    c. The specific wavelengths of light absorbed by the electrons in an atom are different for each element.
        i. Altogether, these are called the "spectrum" of the element.
        ii. When the energy is being GIVEN OFF by the electrons, it is called the:
            1. Emission spectrum
                a. or the bright line spectrum
                b. This is what you see in the gas discharge tubes and neon lights.
                c. Fluorescent light bulbs work because of electrons emitting light.
        iii. When the energy is being ABSORBED by the electrons, it is called the:
            1. Absorption spectrum
                a. or the dark line spectrum
                b. Motion detectors in automatic doors at the grocery stores work because of electrons absorbing light.
        iv. The "spectra" of various elements can be used to identify them.
            1. Every different element
                a. has different sized energy gaps between its energy sublevels
                    i. and therefore absorbs and emits different wavelengths of light
                b. Because the wavelengths of the light are different and distinct for each element, they can be used as "fingerprints"
                    i. This is how we can determine what elements are in a distant star by analyzing the light it gives off– specific colors indicate specific elements
                    ii. The same idea can be used to determine which elements are in a sample of an unknown substance at a crime lab.
    d. The "electromagnetic spectrum" is the range of all possible wavelengths of light.
6. "Electromagnetic radiation" is just a fancy word for light.
    a. The lowest energy light is radio waves, followed by microwave, infrared, visible, ultraviolet, x-ray, and finally gamma ray.
        i. Visible light (the colors human eyes can detect) is a VERY narrow part of the spectrum.
            1. Red is the lowest energy (longest wavelength) color of visible light
            2. Violet is the highest energy (shortest wavelength) color of visible light
    b. Gamma ray is the kind of "light" with the most energy. It also can do the most damage. (Causing cancer etc.)