REVIEW: TOPIC 1

Name: ________

ATOMIC CONCEPTS

UNIT 1 – MAJOR UNDERSTANDINGS

- ☆ 3.1a The modern model of the atom has evolved over a long period of time through the work of many scientists.
- ☆ 3.1b Each atom has a nucleus, with an overall positive charge, surrounded by negatively charged electrons.
- ☆ 3.1c Subatomic particles contained in the nucleus include protons and neutrons.
- ☆ 3.1d The proton is positively charged, and the neutron has no charge. The electron is negatively charged.
- ☆ 3.1e Protons and electrons have equal but opposite charges. The number of protons equals the number of electrons in an atom.
- ☆ 3.1f The mass of each proton and each neutron is approximately equal to one atomic mass unit. An electron is much less massive than a proton or a neutron.
- ☆ 3.1g The number of protons in an atom (atomic number) identifies the element. The sum of the protons and neutrons in an atom (mass number) identifies an isotope. Common notations that represent isotopes include: ^{14}C, $^{14}_{6}C$, carbon-14, C-14.
- ☆ 3.1h In the wave-mechanical model (electron cloud model) the electrons are in orbitals which are defined as the regions of the most probable

UNIT 1 – MAJOR UNDERSTANDINGS (CONTINUED)

electron location (ground state)

- ☆ 3.1i Each electron in an atom has its own distinct amount of energy.
- ☆ 3.1j When an electron in an atom gains a specific amount of energy, the electron is at a higher energy state (excited state).
- ☆ 3.1k When an electron returns from a higher energy state to a lower energy state, a specific amount of energy is emitted. This emitted energy can be used to identify an element.
- ☆ 3.1l The outermost electrons in an atom are called the valence electrons. In general, the number of valence electrons affects the chemical properties of an element.
- ☆ 3.1m Atoms of an element that contain the same number of protons but a different number of neutrons are called isotopes of that element.
- ☆ 3.1n The average atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes.

A – NATURE OF THE ATOM

An **atom** is a unit of matter, the smallest unit of an element, having all the characteristics of that element and consisting of a dense, central, positively charged nucleus surrounded by a system of electrons.

Our concept of the nature of the atom has undergone change and will most likely continue to do so into the future. However, one concept remains unchanged – any atomic theory must be based on the structure of the atom and its various fundamental or subatomic particles. Experimentally, it has been shown that the volume of the atom is primarily made up of negatively charged whirling electrons. The mass of the atom is located in a central core particle called the **nucleus**, which contains all the atom's positive charge.

B – SUBATOMIC PARTICLES

An **electron** ($^{0}_{-1}e$) has a mass of approximately $^{1}/_{1836}$ of a proton and a unit negative charge (**-1**). The mass of an electron is considered negligible; however, they do account for the volume of the atom.

Particles that compose the nucleus are called **nucleons** and include:

- **Protons** ($^{1}_{+1}p$) – A proton has a mass of approximately one atomic mass unit and a unit positive charge (**+1**). Although protons and neutrons are the only nuclear particles that have been identified in an intact nucleus, other particles have been identified among the breakdown products of certain nuclear disintegration. The relationship of these particles to the structure and stability of the nucleus is the subject of much current research. In an atom, the number of protons equals the number of electrons, and the number of protons in an atom (atomic number) identifies the element.
- **Neutrons** ($^{1}_{0}n$) – A neutron also has a mass of approximately one atomic mass unit. It has a unit charge of zero (**0**); therefore, the name "neutron." It is found in the nucleus along with the proton.

Since protons (***p+***) and electrons (***e-***) have equal but opposite charges, all atoms are considered neutral (**0**).

C – Atomic Structure

Atoms differ in the amount of protons and neutrons in the nucleus and in the configuration of the electrons surrounding the nucleus. Most of the atom consists of empty space. Ernest Rutherford's gold foil experiments showed an atom to be mostly empty space with the size of the nucleus very small compared to the atom's size.

Nucleus

The mass of the atom is concentrated almost entirely in the nucleus, which is the basis for determining:

- **Atomic Number** – The atomic number indicates the number of protons in the nucleus. It is the number of protons that identifies the element.

- **Isotopes** – Isotopes are atoms with the same atomic number (number of protons) but a different number of neutrons. This difference in the number of neutrons affects the mass of an atom but does not affect its chemical identity.

Note: For an element the number of protons in the nucleus remains constant, but the number of neutrons may vary. The most common isotope of carbon has six protons and six neutrons in its nucleus. It can be written ^{12}C or $^{12}_{6}C$ or Carbon-12 or C-12. However, there are less common isotopes of carbon such as C-14 and C-15. They also contain six protons in the nucleus but C-14 contains eight neutrons in its nucleus and C-15 contains nine neutrons in its nucleus.

- **Mass number** – The mass number indicates the total number of protons and neutrons. Since the masses of the protons and neutrons are approximately one, the mass number approximates the total mass of the isotope. The number of neutrons in an atom can be calculated by subtracting the atomic number from the mass number.

  The sum of the protons and neutrons in an atom (mass number) identifies an isotope. Common notations that represent isotopes include: ^{14}C, $^{14}_{6}C$, carbon-14, C-14.

- **Atomic mass** – The mass of a neutral atom, called its atomic mass, is measured in **atomic mass units (amus)**, which are standardized on the isotope carbon-12 ($^{12}_{6}C$). This isotope of carbon is equal to 12.000 atomic mass units. Therefore, the definition for an atomic mass unit can be stated as one twelfth ($1/12$) of a carbon-12 ($^{12}_{6}C$) atom.

The atomic mass of an element, given in the *Reference Tables for Physical Setting: Chemistry*, is the weighted average mass of the naturally occurring isotopes of that element. Since most elements occur naturally as mixtures of isotopes, this average is weighted according to the proportions in which the isotopes occur. This accounts for fractional atomic masses found in the reference tables. For example, the element hydrogen exists in three different isotopes:

protium ($^{1}_{1}H$) occurs about 99.0% of the time in nature.

deuterium ($^{2}_{1}H$) occurs about 0.6% of the time in nature.

tritium ($^{3}_{1}H$) occurs about 0.4% of the time in nature.

In general, the mass number is determined by rounding off the atomic mass of the element to the nearest whole number. For example, the atomic mass of a single atom (isotope) such as neon-20 is 19.992 amu, while the atomic mass of the element neon (which is the weighted average of all its natural isotopes) is 20.183 amu. The masses of atoms, multiplied by their occurrence, average out to be 1.000797 amu.

Computing the Average Weight (Mass) of an Element

The average atomic weight of an element, as given the *Periodic Table of the Elements*, is the average of the weights of the atoms in a naturally occurring mixture of the isotopes of the element, if any. Computation of some of these average weights is an interesting mathematical extension which leads to an understanding of the process.

Example: Data in the various chemical handbooks indicate that the occurrences of the principal isotopes of magnesium are:

Mg^{24} – 79.3 percent of atomic mass 23.9924
Mg^{25} – 10.1 percent of atomic mass 24.9938
Mg^{26} – 10.6 percent of atomic mass 25.9898

It is obvious that the average atomic mass (chemical atomic weight) is between 24 and 26 and closer to 24. The weighted average is computed as:

Mg^{24} – 23.9924 x .793 = 19.026
Mg^{25} – 24.9938 x .101 = 2.524
Mg^{26} – 25.9898 x .106 = <u>2.755</u>
24.305

This corresponds to the weight given in the table.

The **gram atomic mass** (the mass of one mole of atoms) of an element is the mass in grams of **Avogadro's number** (approx. 6.0225×10^{23}) of atoms of that element as it occurs naturally. It is numerically equal to the atomic mass. Examples:

1 gram atomic mass of carbon-12 has a mass of 12 grams.
1 gram atomic mass of sodium-23 has a mass of 23 grams.

D – Atomic Models

Bohrs' Planetary Model

All atoms possess energy which causes them to vibrate. In 1901, Max Karl Ernst Ludwig Planck proposed that atoms absorb energy only in discrete amounts called **quanta**. A single quanta is called a quantum of energy.

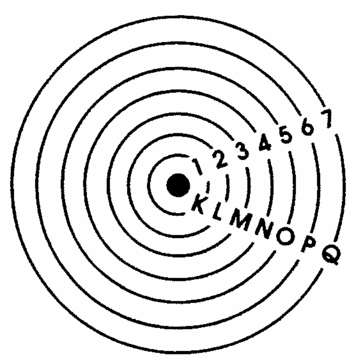
Planck also proposed an equation that relates the energy of a vibrating atom to its frequency: $E = h\nu$. Where E is energy in ergs, ν (Greek letter "nu") is frequency in reciprocal seconds sec^{-1}, and h is a fundamental constant of nature, called **Planck's constant**. Its value is **6.63×10^{34} Joules/second**. This equation proposed that high frequency violet light waves have more energy than low frequency red light waves. He also proposed that energy is not given off or absorbed in a continuous flow – but in small packets of quanta.

The model for atomic structure of the elements has passed through many stages of development. Less than a century ago, Danish physicist Niels Bohr made one proposal for the model.

Although not currently used by chemists to describe atomic structure, the significance of the Bohr model of the atom is that Bohr's model is concerned with the first applications of the quantum mechanical concepts to atomic structure.

In the Bohr model (at the right), electrons were considered to revolve around the nucleus in one of several concentric circular orbits, similar to the solar system.

In the Bohr model of an atom, the principal energy level approximates how far the electron is from the nucleus and can be denoted by the letters K, L, M, N, O, P, Q, or by the numbers 1, 2, 3, 4, 5, 6, 7.

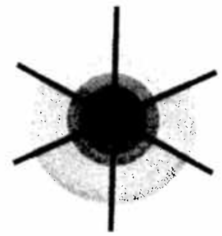


Bohr Model

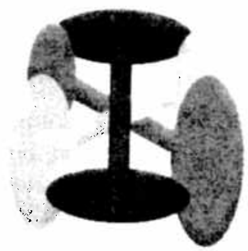
Wave-Mechanical Model (Electron Cloud Model)

Although Bohr's model accounted for the lines of the hydrogen spectrum, it did not account for the spectra of heavier and more complicated atoms.

Electrons occupy orbitals that may differ in size, shape, or orientation in space. The term **orbital** refers to the average region of the most probable electron location. The orbital model differs from the Bohr model in that it does not represent electrons as moving in planetary orbits around the nucleus. Instead, it is defined so that no two electrons will have the same four (4) quantum numbers.



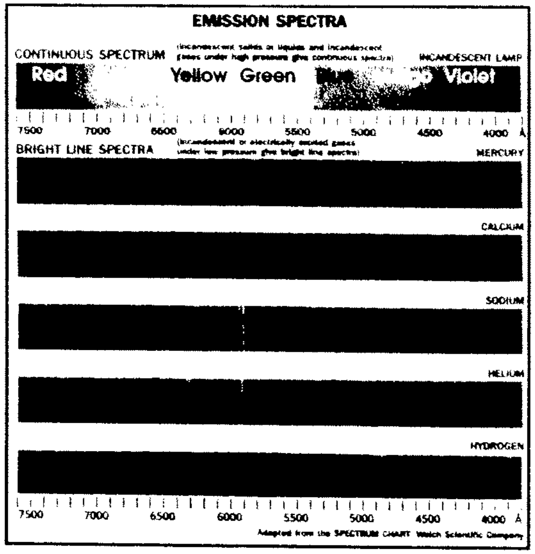
Orbital Models of Atom



Electrons in orbits near the nucleus are at lower energy levels than those in orbits more distant from the nucleus. When the electrons are in the lowest available energy levels, the atom is said to be in the "**ground state**."

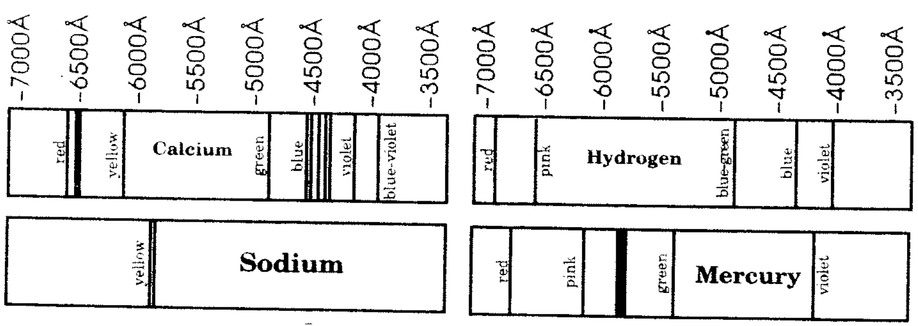
When atoms absorb energy, electrons may shift to a higher energy level. At this higher energy level, the atom is said to be in an "**excited state**." The excited state is unstable, and the electrons fall back to lower ground state energy levels. In the process, they release energy equal to the energy difference of the two energy levels involved. This emitted energy can be used to identify the element.

Spectral lines. When electrons in an atom in the excited state return to lower energy levels, the quanta of energy emitted is called **radiant energy**.

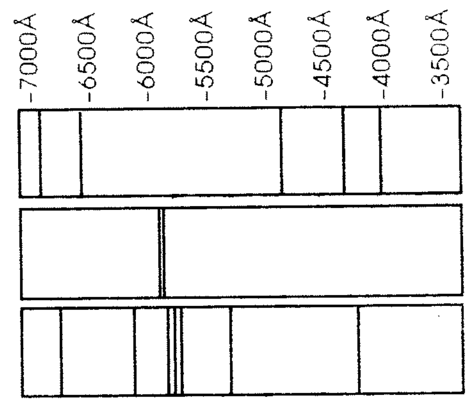


This radiant energy, made up of **quanta** (also called **photons**), has a wave characteristic with a specific frequency. Elements are identified by measuring the wavelength of the radiant energy being emitted. These waves are called spectral lines.

The study of spectral lines has provided much of the evidence regarding energy levels within the atom. This study is carried on with an instrument called the **spectroscope**, which investigates the wavelength of one type of radiant energy called **light**. A spectroscope separates light into specific bright line bands of color called a **spectrum**. Each element gives off a characteristic bright line spectra when investigated by a spectroscope and can therefore be identified by this method. Some examples of bright line spectral lines (measured in **Å** = **angstrom** is a unit of length equal to one hundred-millionth (10^{-8}) of a centimeter):



Directions: Identify the element for each of the bright line spectral line graphs in questions 20 through 22.



20 Question: ____________________

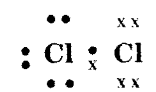
21 Question: ____________________

22 Question: ____________________

F – VALENCE ELECTRON

The electrons in the outermost principal energy level of an atom are called the **valence electrons**. The chemical properties of an atom are related to the valence electrons. The term **"kernel"** is sometimes used to refer to the atom exclusive of the valence electrons. The kernel consists of the nucleus and all electrons except the valence electrons.

The valence electrons may be represented by Lewis **electron dot symbols** in which the kernel of the atom is represented by the letter symbol for the element and the valence electrons are represented by dots. For example, to the right is the electron dot diagram for a chlorine molecule (Cl_2).



The chemical properties of an atom are dependent on the number of valence electrons available to react with other atoms and the strength of the **electrostatic force** which holds them to the nucleus.

As atoms increase in size, the amount of protons and electrons increase causing the nucleus to have greater mass and also a greater positive charge. With an increase in electrons the kernel holds a greater number of electron shells and the electrostatic attraction of the valence electrons to the nucleus diminishes. This allows the valence electrons to wander and be more chemically active.

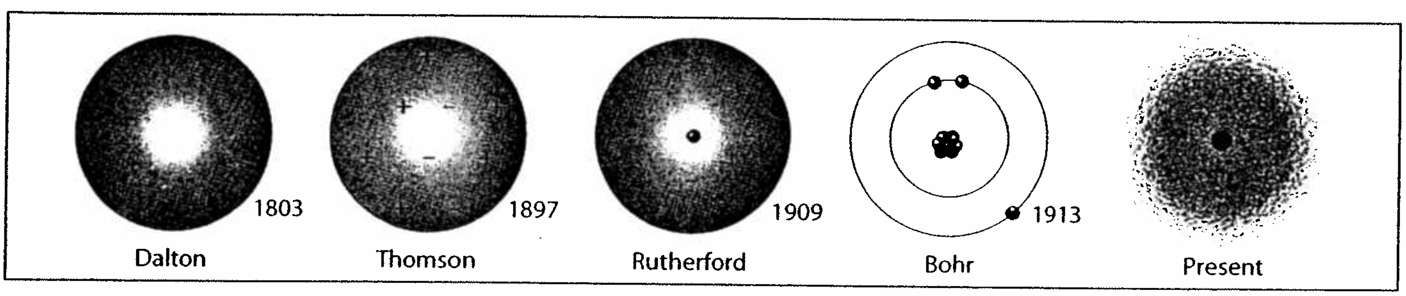
Unit One Atomic Concepts

Key Concepts

- Subatomic particles; proton, neutron, electron
- Atomic Mass
- Isotopes
- Allotropes
- Ion
- Atom
- Bohr Model
- Lewis Dot Diagrams
- Valence Electrons
- Electron Configuration; (2-8-4) only
- Ground State
- Excited Sate

I Atomic Theory

The modern atomic theory has developed over time. Originally atoms were thought to be solid, indivisible and invisible particles. Additionally it was believed that all atoms of the same element were identical. We now know that atoms are made up of ***subatomic particles***. ***The atom*** is made up of a positively charged, dense nucleus with negatively charged electron located in orbitals going around it. Most of the volume of an atom is empty space.



II Subatomic Particles

The three main particles are the ***proton***, the ***neutron*** and the ***electron***. Protons are found in the nucleus, have a mass of one AMU and a charge of +1. The number of protons in an atom or ion is equal to the atomic number. The neutrons are also found in the nucleus and have a mass of 1 AMU but they have no charge. ***Isotopes*** of atoms have the same number of protons but a different number of neutrons. To determine the amount of neutrons in an atom subtract the atomic number form the mass number. Electrons are found in orbits going around the nucleus. Electrons have a mass so small that it is considered to be zero and a charge of -1. Atoms are electrically neutral which means that there are equal numbers of protons and electrons. ***Ions*** have either lost or gained electrons. Positive ions have lost electrons and the number of electrons is equal to the protons minus the charge. Negative ions have gained electrons and the number of electrons is the protons minus the charge. ***Allotropes*** of atoms have the same number of protons, neutrons and electrons but have different properties. An example is carbon; it can be diamond, graphite or coal.

II Nucleus

- contains protons and neutrons, all the mass of the atom = nucleons
- nuclear charge=protons=atomic number
- very, very, very small
- very dense
- isotopes of the same element have a different number of neutrons in the nucleus, the number of protons never change
- ***atomic mass*** is the weighted average of all naturally occurring isotopes; the atomic mass is closest to the most common isotope.

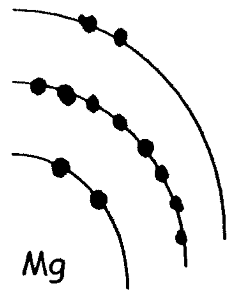
III Electrons

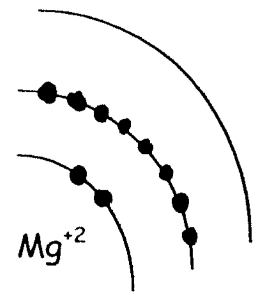
- in energy levels around the nucleus
- can never predict exact location of an electron, orbitals are the regions of highest probability of find the electron
- ***electron Configuration*** tells which energy levels the electrons can be found in, these are written out on the periodic table
- ***ground state***- lowest possible energy level, matches the periodic table.
- ***excited state***- electrons have absorbed energy and moved to a higher energy level, unstable, when electrons drop back down to lower energy levels they give off energy in the form of light
- first energy level can hold 2 electrons, second energy level can hold 8 electrons, third energy level can hold 16 electrons, in general the number of electrons $=2n^2$ where n is the energy level
- ***valence electrons*** are the ones in the outer most energy level, or the last number in the electron configuration

- ions are formed when the valence electrons are added to, negative ion, or taken away from, positive ion.

IV Models

A. ***Bohr Model:*** Shows the number of electron in each energy level.





B. ***Lewis Dot Diagram***: Shows only the electrons in the valence shell

Mg Mg^{+2}

F F^-