Chemistry Study Sheet

Antoine Lavoiser – Compiled the first list of elements that contained 23 elements categorized into 4 subcategories.

1800’s – New Elements, plus advent of electricity that is used to break down compounds. Spectrometer developed = Used to identify newly isolated elements

petrochmicals: soaps, dyes, fertilizers

1870 – Approx. 70 elements known (almost three times Lavoisers list)

1860 – Chemists agreed upon method for determining atomic mass of elements.

John Newlands – Organized first **14 elements.** Noticed elements that when elements order by atomic mass, properties repeated every eigth time. (1864) (element 1 and element 8 have same), (element 2 and element 9 have same) 🡨 Called this the law of octaves.

{ Law of octaves was eventually rejected because it didn’t work for all elements }

Lothar Meyer – saw connection between atomic mass and elemental properties.

Dmitri Mendeleev – Saw the same connection, but is given more credit because he **published organization scheme first**, and went on to better demonstrate.

Later Mendeleev noticed that when order with increasing atomic mass into columns with similar properties, there was a periodic pattern of their props.

Reason it was accepted: Predicted the existence and properties of unknown elements.

Henry Mosely : Discovered that atoms of each element have an equal number of protons in their nuclei that is also equal to the atomic number. 🡪 By ordering in increasing atomic number, Moseley eliminated problems found in previous tables by Mendeleev (placing according to atomic mass resulted in certain elements being placed in incorrect groups of elements with differing properties.

🡪

Resulted in a periodic repetition of chemical and physical properties, or

**Periodic Law :** The statement that there is a periodic repetition of chemical and physical properties when elements are arranged by increasing atomic number.

Periodic Table (Basic Structure)

**Groups** (or Families): columns on the periodic table (There are 18 groups on the PT)

**Periods**: Rows on the periodic table (There are 7 periods in PT)

**Representative Elements**: “Group A Elements” (1A – 8A) or (Groups 1, 2, 13 – 18)

**Transition Elements** : “Group B Elements” (1B – 8B)

Element Classification

**Metals :** Elements that are shiny when smooth and clean, solid at room temp, good conductors of heat and electricity.

Metals are also **malleable** (can be pounded to thin sheets) and **ductile** ( can be drawn to wires).

Most Group A and B elements are metals. (Except for Hydrogen all elements on the left side of the PT are metals.

**Alkali metals :** Group 1A elements (except for hydrogen)

**Alkaline Earth Metals** : Group 2A elements

🡪 Both are chemically reactive, but **alkali metals are more reactive of the two**

The group B elements or transition elements can be divided into sections : **transition metals** and **inner** **transition metals**.

Inner Transition Metals (two bottom rows) (elements 58-103) can be further divided into the **lanthanide series** (elements 58-71) ( normally substances that emit light when struck by electrons ) and the **actinide series** (elements 90 – 103)

**Nonmetals** : Elements that are generally gases or brittle, dull looking solids. These are poor conductors of electricity and heat. (The only nonmetal liquid at room temperature is bromine (Br))

**Halogens** : Group 7A elements ( **extremely reactive** )

**Noble Gases** : Group 8A elements (**extremely unreactive** )

**Metalloids :** (semimetals) Elements with the chemical and physical properties of both metals and nonmetals.

Classifying Elements By Electron Configuration

**Valence Electrons and Valence Electrons in the Period :** Electrons of an atom within the outermost principal energy level. Groups have similar chemical properties because the group number corresponds with the number of valence electrons in every element within that group.) (Ex: 1A elements all have 1 valence electron.)

**Valence Electrons in the Period** : The energy level that the valence electrons of an atom of element are found indicates which period they are place in. (Example: An element with 4 principle energy levels will be in period 4.)

Periodic Trends:

Atomic Radius: Half the distance between adjacent nuclei in a crystal of the element. (for metals) (This is because the electron cloud does not exist in a physical state, and does not have a defined edge such as a golf ball does.)

For nonmetals the atomic radius is defined as half the distance between nuclei of **identical atoms** that are **chemically bonded together**. ( atomic radius of a nonmetal is determined from the diatomic molecule of element)

(Note Atomic **Size** is defined as how closely an atom lies to a neighboring atom. Because the nature of the neighboeing atom can vary from one substance to another, the size of the atom itself also tends to vary somewhat from substance to substance.)

Trend in Period : **Generally decreases** as you move across the period. 🡪 Caused by increasing positive charge within the nucleus **and** the fact that the **principal energy level within a period remains the same**. Each successive elements gains one electron and proton, and **no additional electrons are added between the outermost level and the nucleus**, thus **valence electrons are not shielded from increasing nuclear charge**. (ultimately electrons are pulled closer to the nucleus)

Trend within Group: **Generally increases** (Nuclear charge increases but electrons are added to successively higher principle energy levels.) (**outermost orbital size increases** along with **increasing principle energy levels** 🡪 making the atom **larger.**)

Atoms can gain / lose electrons thus forming **ions.**

Because electrons are negatively charged 🡪 atoms gain or lose electrons acquiring **net charge**.

**Ion** – atom or a bonded group of atoms that has a positive or negative charge.

When atoms **lose electrons** the form **positively charged ions** (and become **smaller**)**.** The electron lost will always be a valence electron, the **loss of a valence electron may leave a completely empty out orbital**, or **electrostatic repulsion between now fewer number of remaining electrons decreases**, allowing them to be pulled into the nucleus.

When atoms **gain electrons** they from **negatively charged ions** (and become **larger**).

The addition of an electron to the atom **increases the electrostatic repulsion** between the atoms valence electrons **forcing them to move father apart**. The increased distance between outer electrons results in a **larger radius**.

Trend within period: Ionic radii generally decreases with positive ions on the left side, and starting at 5A / 6A negative ions decrease as well. (Generally Decreases across a period) (Note: Ions on the left side of periodic table are mostly +, while ions on the right side of the periodic table are mostly -,)

Trends within Group: Ionic Radii generally increase. The reason for this is that the ions outer electrons are in higher principal energy levels. Resulting in gradual increase in ionic size.

Ionization Energy: the energy needed to remove an electron from a gaseous atom.

First Ionization Energy: energy required to remove the first electron from an atom..

A high ionization energy indicated that an atoms has a strong grip on its electrons and is less likely to form positive ions, while a low ionization energy suggest that an atom has a weak grip and is more likely to form positive ions.

After removing first electron, it is possible to remove second (second ionization energy) to form a 1+ ion, and a third (third ionization energy) to form a 2+ ion.

\*\* On table of first ionization energies, a sudden jump after the first ionization, signifies that the atom is likely to lose its first electron, but unlikely to lose its second. The area at which the jump occurs is related to the atoms number of valence electrons. For example, a lithium atom has one valence electron, so its jump occurs after the first ionization energy. (this means it is likely for the atom to for a lithium 1+ ion, but unlikely for it to for a lithium 2+ ion.

Trend in **period**: First ionizations **generally increase** as you move left to right. This is due to the increased nuclear charge having a stronger and strong grip on the atoms electrons. (More energy needed to remove them)

Trend in-**group**: First ionization energies **generally decrease** moving down the group. This occurs because atomic size increases as you move down the group, thus valence electrons are farther from the nucleus making them easier to pull aways thus requiring successively less energy moving down the group.

Octet Rule: atoms tend to gain lose, or share electrons in order to acquire a **full set of eight valence electrons.**

Electronegativity: The relative ability of the element to attract electrons in a chemical bond. These are shown in table and are expressed in numbers that are 3.98 (F) or less.

Flourine (3.98) is the most electronegative element in the periodic table.

Cesium and Francium are the least electronegative elements with values 0.79, and 0.7 respectively.

Trend within period: increases

Trend within group: decreases

Thus the **most** electronegative elements are found in the **upper right corner** of the PT, while the **least** electronegative are found in the **lower left corner**.

In regards to the questions in the packet…

When a metal atom combines with a nonmetal atom, the nonmetal atom will gain electrons and **increase in size.**

Nonmetals will always have a high electronegativity, when asked a question: “Which of the following could be the electronegativity of a nonmetal?” **Always pick on of the extremes.**

Keep in mind that **nonmetals** generally have **high ionization energies**, and **high electronegativity.**

Elements found in the **lower left corner have the strongest metallic properties,** thus the **most active metallic element is found at group one on the bottom**.

The **most active nonmetals** are located in the **upper right corner**.

Metals generally have a low electronegativity, and low ionization energy.

Compared to atoms of metals, atoms of nonmetals have higher electronegativity values.