Periodic Table

In the 1860's several scientist realized that by listing the known elements in order of increasing atomic weights, similar elements with similar properties (i.e., melting points, boiling points, density, and chemical activity) appeared at fairly regular intervals.

Dimitri Mendeleev was one of the first to **publish** a table using these properties to arrange the elements. He also placed **similar elements in the same vertical column.** He **left open spaces on his chart** for elements he predicted would someday be found but had not been discovered at that time.

He developed the idea of the **Periodic Law**: "When elements are arranged in order of increasing atomic weight their properties are repeated periodically." This idea of repetition of properties is one of the most important milestones in chemistry.

Today we know that **atomic number** is a better guide to correlating properties. Mendeleev had suspected atomic weights were not the best criteria to use. These early researchers did not know about protons and neutrons and isotopes.

Groups and Periods on the Periodic Table

Group 1 – **Alkali metals**

- are very good conductors of heat and electricity, are solid at room temperatures, are so soft they can be cut with a knife, have low densities and low melting points

- most chemically active of the metals (must be stored under oil as they will react with air, give up one electron to reach a noble gas configuration, are often identified with a flame test

- are used to produce chemicals, metals, soap, glass, ceramics, petroleum products

Group 2 – **Alkaline Earth metals** - are very chemically reactive (give up 2 electrons during reactions) and are never found free in nature (same as alkali metals)

Group 17 – **Halogen family** - named from "salt former", these elements exhibit all 3 physical states at room temp.

Group 18 - **Noble gases** - all are gases which are naturally occurring

Transitition metals - have properties of metals (ductile, malleable, good conductors of heat and electricity)

Lanthanides - many form alloys with magnetic properties

Actinides - all are radioactive

METALS	**NONMETALS**
Good conductors of heat and electricity	poor conductors of heat and electricity
Shiny in appearance (have luster)	dull in appearance (lack luster-exceptions: iodine and carbon are crystalline and have sheen)
Malleable and ductile	shatter or crumble when struck
High tensile strength	cannot withstand stress
High melting and boiling points	low melting and boiling points
High densities	low densities
Sonorous (give a note when struck such as the sound of a blacksmith hammering iron)	not sonorous
All solids except Hg	exist in all three physical states

Extension notes:

a) in chemical reactions, metal atoms tend t donate electrons to nonmetals that accept them (nonmetals can also share electrons with other nonmetals)

b) metalloids act as electron donors with nonmetals and act as electron acceptors with metals – are solid at room temperatures, are brittle, and are poor conductors of heat and electricity

c) hydrogen is sometimes found in groups 1 and 17 because it can sometimes donate electrons and sometimes accept them. Due to its physical characteristics and the type compounds it forms it is usually classed as a nonmetal.

Trends on the Periodic Table

⇨ atom's size decreases from left to right in a given period

Atoms get smaller as you go from left to right in a period. The electron cloud is pulled inward due to an increase in the nuclear charge (the attractive force of the positively charged nucleus for the negatively charged electron cloud.) The more protons, the stronger the nuclear charge.

As you go left to right the added electrons only fill up a shell that already exist and are thus does not the size of the atom.

⇨ atom's size increases from top to bottom in groups

As you go down a group new main shells are added each period with each new main shells larger than the one it surrounds.

Outer shell electrons are shielded from the nuclear charge by interior shells. This **shielding effect** results in a larger cloud.

⇨ as you go down a group the tendency to lose electrons increases

The larger atoms (at bottom of group) have electrons so far from the nucleus (added main shells and shielding effects) that these electrons are more easily ionized. The nuclear charge holding them is weak.

⇨ ionization energies increase across periods and decrease down groups

The first ionization energy is the amount of energy needed to remove the most loosely held electron (and thus forming a positive ion). Metals have low ionization energies and readily form positive ions. Nonmetals (due to small size, stronger nuclear charge, and many electrons in outer shells) have large ionization energies.

⇨ electronegativity increases from left to right in a period and decreases from top to bottom

Electronegativity is a number showing the ability an atom has to attract electrons (usually shared pairs of electrons) to itself when bound with another atom.

The most active metals (lower left on chart) have lowest electronegativities (they tend to lose electrons). Atoms of high electronegativity (upper right nonmetals) tend to gain electrons.

⇨ negative ions are larger than their atoms and positive ions are smaller than their atoms

Nonmetals gain one or more electrons to fill out their outer shell. The cloud tends to swell outward as the nuclear charge holding each electron is now less. Negative ions are larger.

Positive ions (formed usually from metals) grow smaller as they lose their outer shell of electrons. The effective nuclear charge is now stronger on the remaining electrons and they are pulled inward.

Oxidation States

Chemist have found that certain elements are more stable, or less reactive than others. These stable atoms have their outtermost shell (s and p subshell) filled to capacity with 8 electrons (2 electrons for hydrogen and helium). Noble gases show little tendency to form compounds as they already have a filled outer shell.

When 2 or more atoms combine, each atom tends to get a complete outermost shell holding 8 electrons. This outermost s and p shell is called the valence shell. The **Octet Rule or Rule of 8** maintains that atoms try to fill this outer valence shell by losing, gaining, or sharing electrons during reactions.

The oxidation state is used to help identify how many electrons will be transferred or shared. A positive oxidation state indicates that the atom will lose electrons. A negative oxidation state indicates that the atom will gain or share electrons.

Chemical Reactivity

General rules:

a) metal atoms tend to transfer electrons to nonmetals when they react.

b) nonmetal atoms tend to gain or share electrons when they react.

c) to determine which metal or nonmetal is more reactive, compare elements within groups or periods (for diagonal predictions use Reactivity Series)

Chemical Reactivity in a given period:

⇨ For metals in a given period reactivity is determined by the number of electrons that must be transferred to a nonmetal. Those metals with only 1 outer shell electron will be more reactive than metals with two or more electrons. It takes less energy to remove one electron (lower first ionization energy).

⇨ For nonmetals in a given period the greater the number of electrons already in the outer shell the more reactive that element will be. If a nonmetal has seven outer shell electrons it will only need to gain 1 electron to reach its octet. This nonmetal will be more reactive than a nonmetal that must gain two or more electrons.

Summary: In the same period, it is the number of electrons that chiefly determines reactivity. The smaller the number of electrons transferred between reacting atoms, the more vigorous the reaction.

Chemical Reactivity in a given group:

⇨We already know that the negatively charged electrons are held in orbit by their attraction to the positively charged nucleus. This force of attraction decreases considerable as the distance from the nucleus increases. Therefore, electrons that are more distant from the nucleus are held less tightly and are more easily lost that are electrons closer to the nucleus.

⇨Since the more distant electrons are more easily lost, metals having distant electrons (at the bottom of the group) react readily with other nonmetallic elements.

⇨ Nonmetals combine chemically by gaining or sharing electrons (from metals or other nonmetallic elements). Since the attraction for electrons is greater when the atomic radius is small, the closer the outermost orbit (called the valence shell) is to the nucleus of a nonmetal, the more reactive it is. A small radius means that the nucleus pulls strongly on any electrons near it and strongly desires to fill its outer shell (Octet rule). The more reactive nonmetals will be found near the top of the group.

Summary: In the same group, elements have the same number of outershell electrons and it is the atomic radius which largely determines reactivity. The larger metals loose outer shell electrons more easily and smaller nonmetals (whose attraction for electrons by the nucleus is greater) are more likely to take electrons away from other metals (or share with other nonmetals).

The Reactivity Series

By studying replacement reactions we can arrange the metals in decreasing order of reactivity. The reactivity series list the different metals (and hydrogen) in order of their decreasing tendency to lose electrons in water solutions at specified temperatures. Hydrogen is on the list because it behaves as though it were a metal in certain reactions. Metals found above hydrogen will replace hydrogen in acid solutions, while those below will not.

As we know the tendency of metals to lose electrons depends chiefly on the nuclear charge and on the atomic radius of the metal atom. The reactivity series may be used to make reasonable predictions concerning the reactivity of different metals. According to the table, for example, aluminum will replace mercury in an aqueous solution of a mercury compound. Silver will not replace tin.

Li K Ba Sr Ca Na Mg Al Mn Zn Cr Fe Cd Co Ni Sn Pb *H Sb As Bi Cu Ag Pd Hg Pt Au

Most Reactive --→ Least Reactive

Alkali Metals

Common Name	Chemical Name	Formula	Source
baking soda	sodium bicarbonate	$NaHCO_3$	ammonia, sodium chloride, limestone
borax	sodium tetraborate	$Na_2B_4O_7 \bullet 10H_2O$	mineral deposits
caustic potash	potassium hydroxide	KOH	electrolysis of KCl
Glauber's salt	sodium sulfate	$Na_2SO_4 \bullet 10H_2O$	mineral deposits
lye	sodium hydroxide	NaOH	electrolysis of NaCl
muriate of potash	potassium chloride	KCl	mineral deposits
table salt	sodium chloride	NaCl	mineral deposits, brine, salt lakes

Alkaline Earth Metals

Common Name	Chemical Name	Formula	Source	Uses
barite	barium sulfate	$BaSO_4$	mineral deposits	paint pigments, X-Ray diagnosis
Epsom salts	magnesium sulfate	$MgSO_4 \bullet 7H_2O$	mineral deposits	making dyes, laxative
gypsum	calcium sulfate	$CaSO_4 \bullet 2H_2O$	mineral deposits	making plaster of Paris and mortar
lime or quicklime	calcium oxide	CaO	breakdown of limestone in kiln	smelting metals, drying agent
limestone or marble	calcium carbonate	$CaCO_3$	deposits of shells of marine animals	building material, smelting metals
magnesia	magnesium oxide	MgO	decomposition of magnesium carbonate	lining furnaces, making rubber and paint
slaked lime or limewater	calcium hydroxide	$Ca(OH)_2$	adding water to lime	making mortar and bleaching powder, alkalizing soil

Transition Metals

Metal	Symbol	Source	Uses
chromium	Cr	chromite, $FeCrO_4$	stainless steel, nichrome, plating
cobalt	Co	cobaltite, $CoAsS$	alnico magnets
copper	Cu	chalcopyrite, $CuFeS_2$	electric wires, water pipes, coins
gold	Au	elemental state	computer chips, jewelry, plating
iron	Fe	hematite, Fe_2O_3	steels, cast iron
manganese	Mn	pyrolusite, MnO_2	manganese steel, alloys
nickel	Ni	pentlandite, NiS	alnico magnets, steel, catalyst
platinum	Pt	elemental state	surgical tools, catalyst, jewelry
silver	Ag	argentite, Ag_2S	mirrors, coins, tableware, jewelry
titanium	Ti	rutile, TiO_2	aircraft, spacecraft, missiles
tungsten	W	wolframite $FeWO_4$	tungsten steel, light bulb filaments

Halogens

Halogen	Compounds	Properties	Uses
Fluorine	fluorides	protect teeth against decay	drinking water and toothpaste
	Teflon (fluorocarbon)	heat-resistant plastic	non-stick pans and electrical insulation
	Freon (fluorocarbon)	easily liquefied gas with high heat of vaporization	refrigerant
Chlorine	hypochlorous acid and hypochlorites	powerful oxidizing agent, germicide, bleaching agent	purify water for drinking, pools, sewage treatment, bleaching industry, household bleaches and disinfectants
	polyvinyl chloride (PVC)	tough plastic	covering for furniture and floors
Bromine	silver bromide	light sensitive	photographic film, plates, paper
	bromides of sodium and potassium	sedative	headache powders
Iodine	silver iodide, iodides of sodium and potassium	light sensitive, prevent goiter	photography, make ‘iodized’ table salt

Ionization Energy and the Periodic Table

The periodic table organizes the elements according to their increasing atomic number. This is useful because repeating patterns of physical and chemical properties occur as the atomic number of the elements increases. This is called the periodic law.

Ionization Energy is the amount of attraction an atom has for its own electrons. An element with a high ionization energy attracts its own electrons very strongly. It is not likely to release electrons to participate in chemical reactions. An element with a low ionization energy gives up electrons easily. It is more likely to participate in chemical reactions. Thus, the ionization energy of each element gives us one measure of its reactivity compared to other elements.

This information is particularly useful in predicting how each metallic element will behave chemically. Ionization energy is not as useful for predicting the chemical activity of nonmetals and metalloids. They are more likely to accept or share electrons in chemical reactions rather than to give up their electrons. In this activity you will examine how the ionization energy varies as the atomic number increases.

Procedure:

1. Study the chart of ionization energies and electronegativies found below. The units for ionization energy on this chart are kilojoules/mole.

2. If you are plotting this graph by hand turn the paper horizontally. The x-axis is the **atomic number** and should go to 92 (uranium). The y-axis is **ionization energy (kJ/mol)** and should go from 0 to 1400.

3. Write out in words your conclusions from the graph.

1314 **H** 2.1			Ionization Energies and Electronegativites														2368 **He** --
519 **Li** 1.0	900 **Be** 1.5											799 **B** 2.0	1088 **C** 2.5	1401 **N** 3.0	1036 **O** 3.5	1682 **F** 4.0	2076 **Ne** --
498 **Na** 0.9	736 **Mg** 1.2											578 **Al** 1.5	787 **Si** 1.8	1063 **P** 2.1	1000 **S** 2.5	1256 **Cl** 3.0	1519 **Ar** --
418 **K** 0.8	590 **Ca** 1.0	632 **Sc** 1.3	661 **Ti** 1.5	649 **V** 1.6	653 **Cr** 1.6	716 **Mn** 1.5	762 **Fe** 1.8	757 **Co** 1.8	736 **Ni** 1.8	745 **Cu** 1.9	904 **Zn** 1.6	578 **Ga** 1.6	782 **Ge** 1.8	1013 **As** 2.0	942 **Se** 2.4	1142 **Br** 2.8	1352 **Kr** --
402 **Rb** 0.8	548 **Sr** 1.0	636 **Y** 1.2	670 **Zr** 1.4	653 **Nb** 1.5	695 **Mo** 1.8	720 **Tc** 1.9	724 **Ru** 2.2	745 **Rh** 2.2	804 **Pd** 2.2	728 **Ag** 1.9	866 **Cd** 1.7	557 **In** 1.7	707 **Sn** 1.8	833 **Sb** 1.9	870 **Te** 2.1	1008 **I** 2.5	1172 **Xe** --
377 **Cs** 0.7	502 **Ba** 0.9	540 **La** 1.1	531 **Hf** 1.3	586 **Ta** 1.5	770 **W** 1.7	757 **Re** 1.9	841 **Os** 2.2	887 **Ir** 2.2	862 **Pt** 2.2	887 **Au** 2.4	1008 **Hg** 1.9	590 **Tl** 1.8	716 **Pb** 1.8	770 **Bi** 1.9	820 **Po** 2.0	-- **At** 2.2	1038 **Rn** --
-- **Fr** 0.7	510 **Ra** 0.9	678 **Ac** 1.1	-- **Ku** --														
1008 - Ionization energy C - symbol 2.5 - electro- negativity			665 **Ce** 1.1	557 **Pr** 1.1	607 **Nd** 1.1	-- **Pm** 1.1	540 **Sm** 1.1	548 **Eu** 1.1	594 **Gd** 1.1	649 **Tb** 1.1	657 **Dy** 1.1	-- **Ho** 1.1	-- **Er** 1.1	-- **Tm** 1.1	598 **Yb** 1.1	481 **Lu** 1.2	
			-- **Th** 1.3	-- **Pa** 1.5	385 **U** 1.7	-- **Np** 1.3	-- **Pu** 1.3	-- **Am** 1.3	-- **Cm** 1.3	-- **Bk** 1.3	-- **Cf** 1.3	-- **Es** 1.3	-- **Fm** 1.3	-- **Md** 1.3	-- **No** 1.3	-- **Lr** --	