3.1 The Periodic Table (Textbook p.84)

A Table of the Elements
In addition to his Atomic Theory, ***John Dalton*** also devised a system of ***chemical symbols*** and, having ascertained the relative weights of atoms, arranged them into a table.
•Other scientists also tried to arrange the elements in the form of a table.

3.1.1. Describe the arrangement of elements in the periodic table in order of increasing atomic mass.

Dmitri Mendeleev
•A Russian chemist in the mid-1800s
•**Mendeleev** cataloged thousands of facts about the 63 elements known at the time
•He became convinced that groups of elements had similar, ***"periodic"*** properties.

Mendeleev's Table
Elements on **Mendeleev's table** were arranged according to their increasing **atomic mass**, leaving blank spaces where he was sure other, unknown elements would fit.

Valence(Valency)
•The concept that Mendeleev found most helpful in laying out his table was the notion of ***valences***
•Almost all the elements known at the time would combine with either ***hydrogen or oxygen***, so the valence of an element was related to the number of atoms of hydrogen or oxygen that combined with that element.

Valence
•Hydrogen and oxygen form water, H_2O, so hydrogen was given a valence of 1 and oxygen a valence of 2
•For any other element, the valence was defined to be:
–the number of hydrogen atoms, or twice the number of oxygen atoms, that would combine with one atom of that element
•Mendeleev put elements with the same valence into the same group.

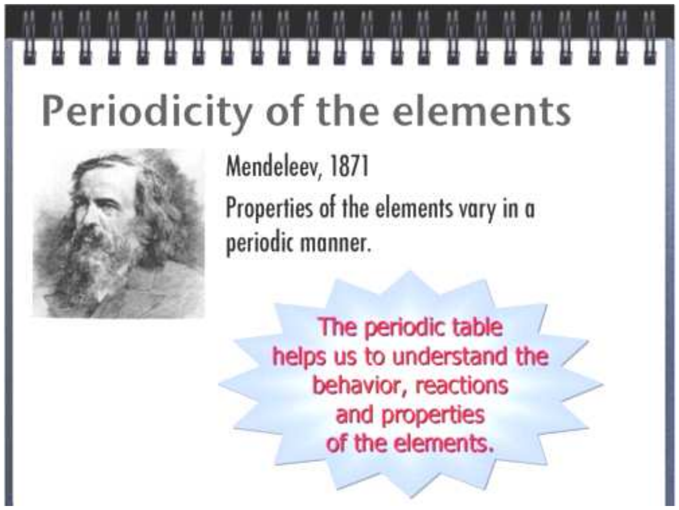
Valence
•Valence is related to the number of electrons that an element has in its ***outermost shell or energy level – the valence electrons***.
•Mendeleev predicted the properties of unknown elements based on the idea of ***periodic properties***
– because of this, Mendeleev is considered to be the "**Father of the Periodic Table"**.

Henry Moseley
•Fifty years after Mendeleev, the British scientist ***Henry Moseley*** discovered that the number of protons in the nucleus of a particular type of atom was always the same
•When atoms were arranged according to increasing **atomic number**, the few problems with Mendeleev's periodic table disappeared.
•Because of Moseley's work, the modern periodic table is based on the ***atomic numbers*** of the elements.

Periodicity of the Elements
•Dimitri Mendeleev and Henry Moseley brought order to the elements:
–by discovering the periodic nature of the elements, they were able to arrange the elements into families or groups and place them on a periodic table
– by organizing the elements, scientists could better study the structure of matter.

The Periodic Law

"The physical and chemical properties of the elements are periodic functions of their atomic numbers".



3.1.2 Distinguish between the terms group and period

The Modern Periodic Table

The modern periodic table (handout) has elements arranged in a series of:

•Vertical columns called ***groups*** or ***families***

•Horizontal rows called ***periods.***

Groups

Groups are numbered from left to right on the table:

– 1 to 8

– or, 1 *to* 18 *(or Roman numerals,) depending upon the particular version of the periodic table*

Groups

•Elements in groups have similar properties

–that's why they are also sometimes called **families**

•Although properties are similar, they change as you go up or down the group.

Groups – chemical activity

•Groups usually contain either metals or nonmetals

–more on this shortly

•Chemical activity generally:

–increases as you go down a metal group

–decreases as you go down a nonmetal group.

Periods

•Periods are numbered from top to bottom on the table: 1 to 7

•The properties of elements in a period are quite different, but there are patterns.

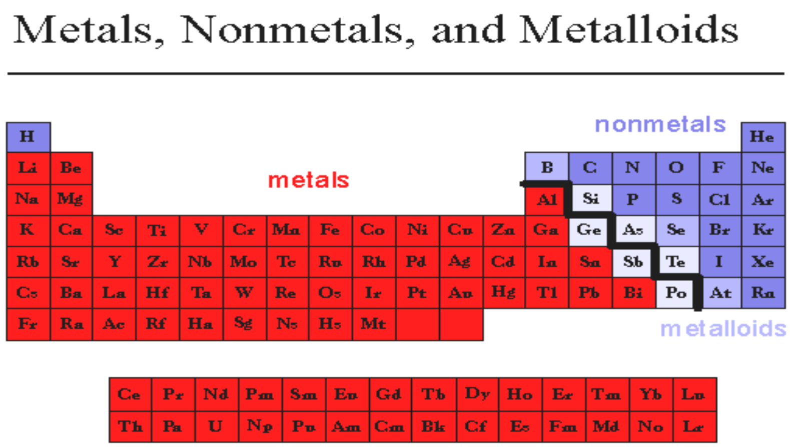
Periods

•The first (far left) element in a period is always an active metal, the last (far right) is always an inactive nonmetal

•Generally, within a period:

–the chemical activity of metals decreases from left to right
–the chemical activity of nonmetals in a period increases from left to right.

Information on The Periodic Table



Sections of the Periodic Table
The periodic table is divided into three main sections:
•metals
•nonmetals
•metalloids
Each one of these groups contains elements with similar physical properties.

Metals
•Metals makeup more than 75% of the elements in the periodic table.

Explanation of Terms
•***Luster*** - metallic shine
•***Malleable*** - can be hammered, pounded, or pressed into different shapes without breaking
•***Ductile*** - can be drawn into thin sheets or wires without breaking.

Nonmetals
•There are 17 nonmetals in the periodic table.

Metalloids
•The seven metalloids are B, Si, Ge, As, Sb, Te and At.

Metalloids
•Elements with properties of both metals and nonmetals
•Elements touching the ***metal-nonmetal line*** on the table -this line is drawn on some tables, but not all.

–**Silicon** and **Germanium** are two metalloids important in the manufacture of computer chips
–Their conducting characteristics allow electric circuits to be "printed" on them.

Summary

•The periodic table is divided into three main sections, metals, nonmetals and metalloids

•Most elements are metals

•Metalloids have properties between metals and nonmetals

•Some groups in the periodic table contain metals only, some nonmetals only and some both.

Periodic Properties

•The periodic table also has certain properties characteristic of certain regions in the periodic table.

Alkali Metals

•These are the metals in the first column of the periodic table

•They are soft shiny metals that usually combine with group 8 (or 17) nonmetals (the halogens) in chemical compounds in a 1:1 ratio e.g. sodium chloride NaCl.

Alkaline Earth Metals

•These are the elements in the second column of the periodic table, and they are very similar to the alkali metals

•They combine with the halogens in a 1:2 ratio e.g. magnesium chloride $MgCl_2$

Halogens

•The halogens are fluorine, chlorine, bromine, and iodine

•Halogens exist as diatomic (two atom) molecules in nature e.g. chlorine $Cl_{2.}$

Noble Gases

•Also called rare gas elements or inert gases - all occur in nature as gases.

•The noble gases make up the last column in the periodic table.

•Noble gases are very ***unreactive***.

Transition Metals

•The transition metals are the metals located between columns 2 and 13 (IIA and IIIA) in the periodic table

•Transition metals are so-called, as they show a gradual transition (change) in properties from one member to the next.

3.1.3 Apply the relationship between the electron arrangement of elements and their position in the periodic table up to Z = 20. (s and p electrons)

3.1.4 Apply the relationship between the number of electrons in the highest occupied energy level for an element and its position in the periodic table. (number of valance electrons is equal to the group number)

Valence, Groups and Reactivity

At the beginning of this topic it was mentioned that Mendeleev used valence to classify his elements into groups. It was also said that valence was related to the number of electrons in the outermost electron shell.

What is the relationship between the number of electrons, valence and groups in the Periodic Table?

Families of metals

Alkali metals - Group 1

•1 electron in the outer energy level

•React with water to release hydrogen gas

•The most reactive metals – stored under oil

Alkaline earth metals - Group 2

•2 electrons in their outer energy level.

Transition metals - Groups 3 to 12
•2 electrons in their outer energy level
•Form compounds that are brightly colored
•Quite often are used as ***catalysts*** - *more later*..

Families of nonmetals
Boron family - Group 13
•3 electrons in their outer energy level
•Aluminum is the most abundant metal and the third most abundant element in the Earth's crust.

Carbon family - Group 14
•4 electrons in their outer energy level
•Carbon's unique characteristic of bonding to itself is responsible for complex molecules composed of long chains of carbon atoms - ***organic chemistry***, which comprises millions of different molecules
•Silicon is the second most abundant element in the Earth's crust.

Nitrogen family - Group 15
•5 electrons in their outer energy level
•Nitrogen is the most abundant element in the Earth's atmosphere
•Phosphorus is used in matches.

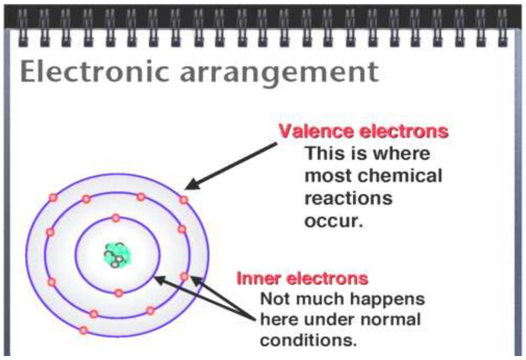
Oxygen family - Group 16
•6 electrons in their outer energy level
•Oxygen is the most abundant element in the Earth's crust
•Oxygen supports ***combustion***.

Halogens - Group 17
•7 electrons in their outer energy level
•Halogens easily combine with metals to form ***salts***
•Most reactive of all the nonmetals.

Noble gases - Group 18
•8 electrons in their outer energy level
•Because of their electron arrangement Noble Gases are almost complete inactive, ***"inert"***
•All members of the family are colorless gases
•Argon is the most abundant Noble Gas, making up almost one percent of Earth's atmosphere.

Valence electrons and group number
•The number of valence electrons is generally equal to the group number of the element, or, the ***group number minus 10*** *(using the 1 - 18 group numbering system)*
–There are exceptions:
•Helium - actually has two electrons, but is included in group 8 (18)
•Transition elements.

•Valence (*combining power*) is related to the number of ***valence electrons.***



Valence and valence electrons

•The ***valence*** of an element is generally equal to the group number, or, 8 minus the group number using the old group numbering system (roman numerals A block)
•For example: Oxygen in Group 6A
–Valence = 8 - 6 = **2**

Valence and chemical formulae

•A knowledge of valence is useful in determining the formula of a compound
What is the formula for magnesium chloride?
•Mg = group 2, valence = 2
•Cl = group 17, valence = 8 - 7 = 1
•Formula of magnesium chloride = $MgCl_2$

Group	No. of valence electrons	Valence	Example	Name
Alkali metals	1	1	NaCl	sodium chloride
Alkaline earths	2	2	$MgCl_2$	magnesium chloride
Transition metals	2(variable)	2	$FeCl_2$	iron (II) chloride
Boron group	3	3	BCl_3	boron trichloride
Carbon group	4	4	CH_4	methane
Nitrogen group	5	3	NH_3	ammonia
Oxygen group	6	2	H_2O	water
Halogens	7	1	HCl or NaCl	hydrogen chloride
Noble gases	8 (He=2)	0	none	none

Now complete exercise 3.1 on page 85

Periodic Trends

•The Periodic Table is arranged according to the Periodic Law

–the Periodic Law states that ***"when elements are arranged in order of increasing atomic number, their physical and chemical properties show a periodic pattern"***

•Certain properties of the elements exhibit a gradual change in properties as we go across a period or down a group

–knowing these trends can help in our understanding of chemical properties.

3.2.2 Describe and explain the trends in atomic radii, ionic radii, first ionization energies, electronegativities and melting points for the alkali metals (Li to Cs) and the halogens (F to I).

3.2.3 Describe and explain the trends in atomic radii, ionic radii, first ionization energies and electronegativities for elements across period 3.

3.2 Periodic Trends – Physical properties (Textbook p. 86)

•Because properties of elements are based on their electron configurations, many of their properties are predictable and repeat in periodic patterns.

The properties that will be examined are:

•atomic size (diameter and radii)

•ionic radii

•first ionization energy

•electronegativity

•melting points

Atomic size

Atoms get larger down a group

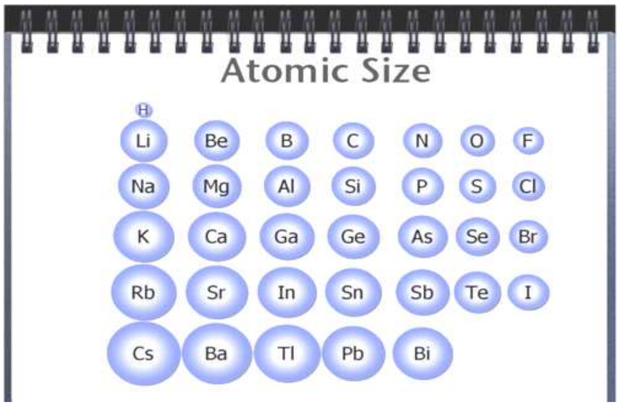
•WHY?

–the number of electron shells increases

–each additional shell is further from the nucleus

–atomic size increases.

–



Atoms get smaller across a period

•WHY?

–electrons are added to the same energy level (shell)

–more protons in the nucleus creates a "**higher effective nuclear charge**"

–a stronger force of attraction pulls the electrons closer to the nucleus.

Ionic radii

Ions aren't the same size as the atoms they come from. Compare the sizes of sodium and chloride ions with the sizes of sodium and chlorine atoms:

Positive ions

Positive ions are smaller than the atoms they come from. Sodium is 2,8,1; Na^+ is 2,8.
A whole layer of electrons, has been lost and the remaining 10 electrons are being pulled in by the full force of 11 protons.

Negative ions

Negative ions are bigger than the atoms they come from. Chlorine is 2,8,7; Cl^- is 2,8,8.
Although the electrons are still all in the 3-level, the extra repulsion produced by the incoming electron causes the atom to expand. There are still only 17 protons, but they are now having to ‘hold’ 18 electrons.

Anions (negatively charged) are almost invariable larger than cations (positively charged).
In general, ionic radius decreases with increasing positive charge and increases with increasing negative charge.

As with atomic radius, ionic radii increase on descending a group and decrease across a period:. For groups 1 and 7:

Ion	Radius (nm)	Ion	Radius (nm)
Li^+	0.068	F^-	0.133
Na^+	0.098	Cl^-	0.181
K^+	0.133	Br^-	0.196
Rb^+	0.148	I^-	0.219

For Period 3
Na^+ = 0.098 nm
Mg^{2+} = 0.065 nm
Al^{3+} = 0.045 nm
N^{3-} = 0.171 nm
O^{2-} = 0.146 nm
F^- = 0.133 nm

Note: 1 nm = One-billionth of a meter (10^{-9} m).

The ionic radius, r_{ion}, is a measure of the size of an ion in a crystal lattice. The concept of ionic radius was developed independently by Goldschmidt and Pauling in the 1920s to summarize the data being generated by the (then) new technique of X-ray crystallography.
The ionic radius is not a fixed property of a given ion, but varies with coordination number. Nevertheless, ionic radius values allow periodic trends to be recognized.
An "anomalous" ionic radius in a crystal is often a sign of significant covalent character in the bonding.

IB Chemistry

3.2.1 Define the terms first ionization energy and electronegativity.

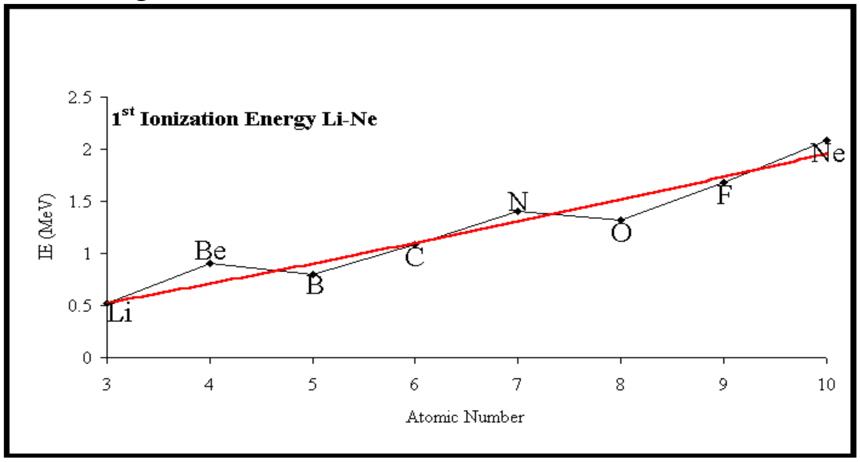
First Ionization Energy (1st IE)

•What is ***first ionization energy***?

–" the energy needed to remove the outermost, or highest energy, electron from a neutral atom in the gas phase"

•very high ionization energy values indicate a stable electron arrangement (like full shells).

•note high values for noble gases.

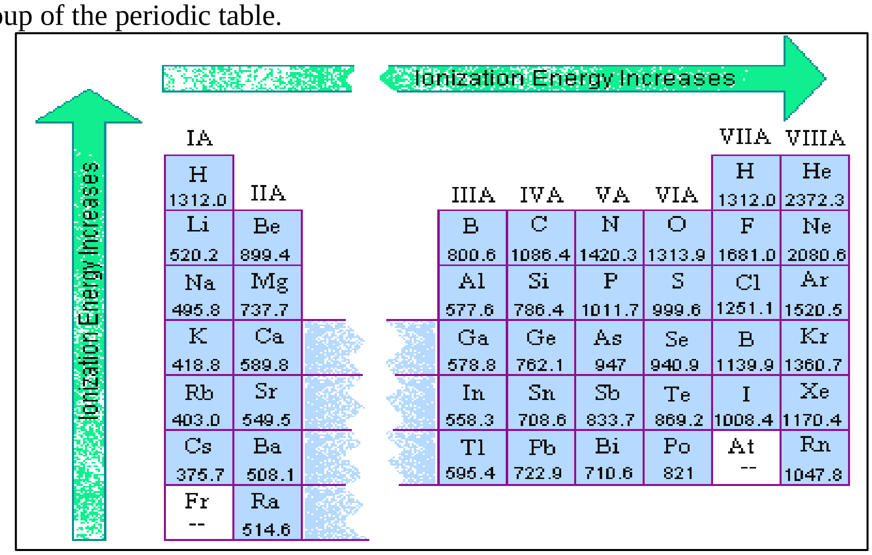


First Ionization Energy

•Two trends are apparent:

–an increase from left to right across a period

–a decrease down a group of the periodic table.



Explanation of 1st IE

•Increase from left to right across a period

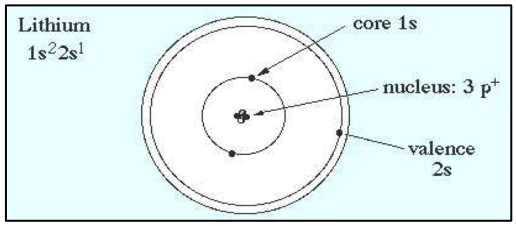
–the force of attraction between the nucleus and an electron becomes larger as the number of protons in the nucleus increases

•Decrease down a group of the periodic table

–additional shells ***shield*** *the outermost electrons from the attractive force of the nucleus.*

Shielding?

Example: the valence electron of lithium is ***"shielded"*** to some extent, from feeling the entire 3+ charge of the nucleus by the ***"core"*** electrons.



Electronegativity

In the 1930's, Linus Pauling (1901 - 1994), an American chemist who won the 1954 Nobel Prize, recognized that atoms in a molecule differ in their ability to attract electrons.

Electronegativity

•Pauling defined electronegativity as:

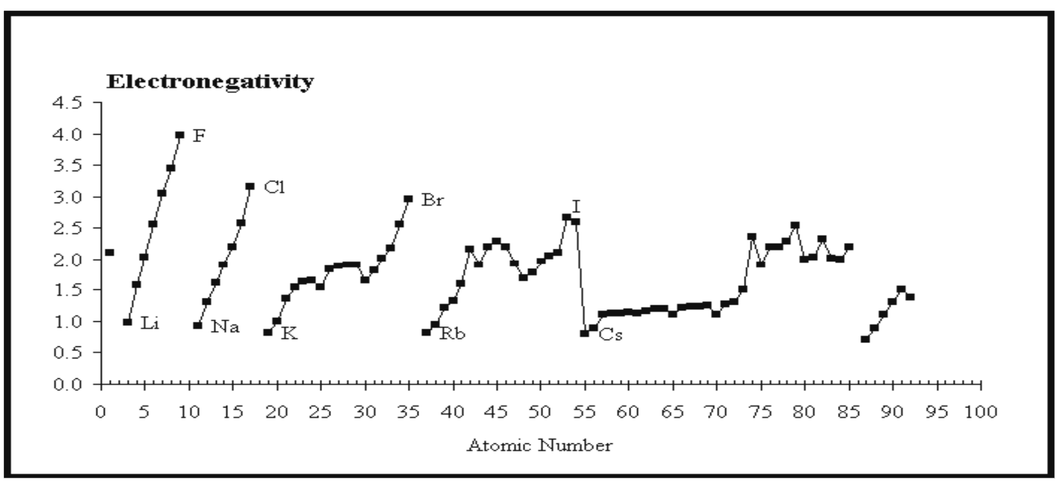
"The power of an atom in a molecule to attract electrons to itself"

•Pauling assigned electronegativity values to the elements.

3.2.4 Compare the relative electronegativity values of two or more elements based on their positions in the periodic table.

Electronegativity values – Pauling scale

H=2.1	X	X	X	X	X	X
Li=1.0	Be=1.5	B=2.0	C=2.5	N=3.0	O=3.5	F=4.0
Na=0.9	Mg=1.2	Al=1.5	Si=1.8	P=2.1	S=2.5	Cl=3.0
K=0.8	Ca=1.0	Ga=1.6	Ge=1.8	As=2.0	Se=2.4	Br=2.8
Rb=0.8	Sr=1.0	In=1.7	Sn=1.8	Sb=1.9	Te=2.1	I=2.5
Cs=0.7	Ba=0.9	Tl=1.8	Pb=1.9	Bi=1.9	Po=2.0	At=2.2



Note: - fluorine is the most electronegative, cesium the least.
- noble gases not included due to exceptional stability.

Trends in electronegativity
Increases up a group and from left to right in a period.

Explanation of electronegativity
•Along a period

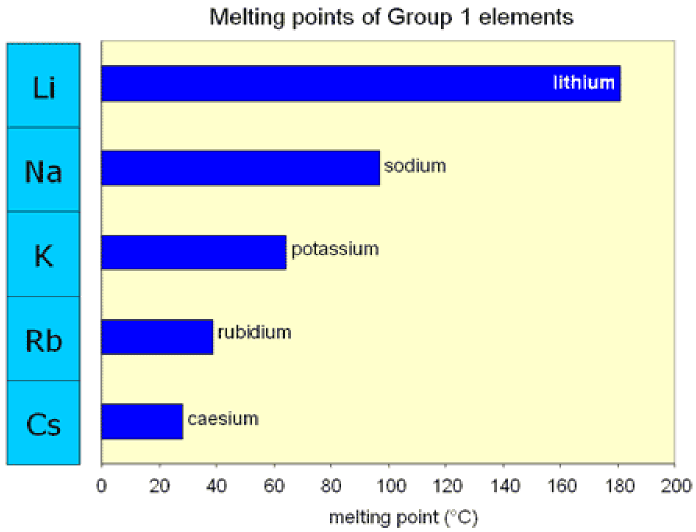
–*number of protons increases so electrons are attracted more strongly*

•Up a group

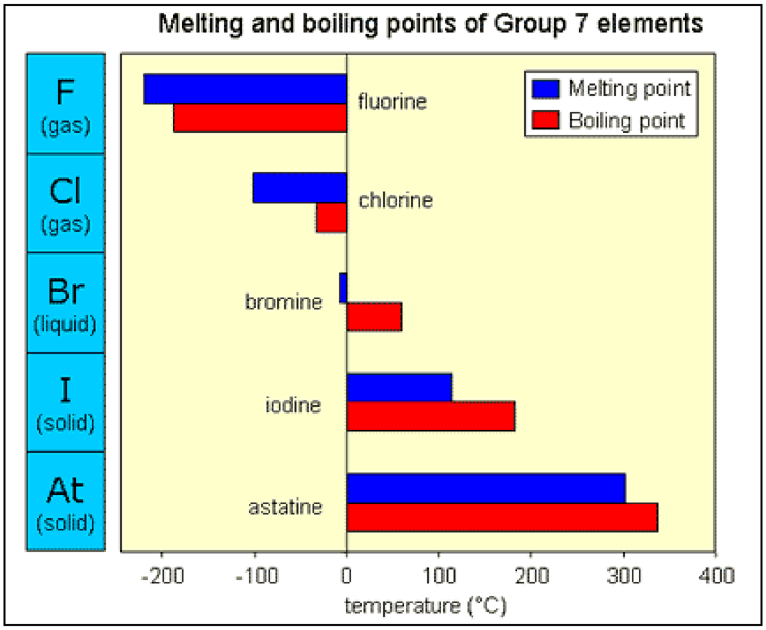
–*fewer inner shells to **shield** attractive forces from the nucleus.*

Note that the relative difference in electronegativity between two atoms will give an indication of the type of bonding that will occur between them in a compound.
The larger the difference then the greater the ionic nature of the bond.
If there is little difference in electronegativity values then a compound formed from them is likely to show covalent bonding.

Melting points
Physical properties such as melting point and boiling point depend on the nature of bonding between particles of the element. In the alkali metals, as the strength of metallic bonding increases so will the melting point.
The alkali metals have low melting and boiling points compared to most other metals. Apart from the other alkali metals, only three metals (indium, gallium and mercury) have lower melting points than lithium. You can see from the graph that lithium, at the top of Group 1, has the highest melting point in the group. The melting points then decrease as you go down the group.



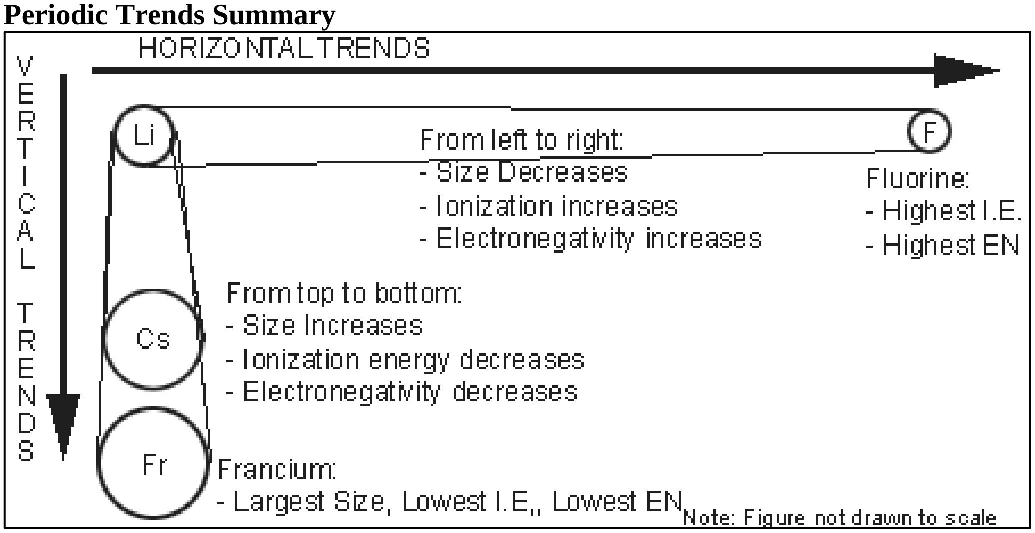
The halogens have low melting points and boiling points. This is a typical property of non-metals. The halogens have molecular covalent structures and there are only weak forces (van der Waals) between the molecules. You can see from the graph below that fluorine, at the top of Group 7, has the lowest melting point and boiling point in the Group. The melting points and boiling points then increase as you go down the Group.



State of halogens at room temperature
Room temperature is usually about 20°C. At this temperature, fluorine and chlorine are gases, bromine is a liquid, and iodine and astatine are solids. You should remember this trend in state - the top two elements are gases, the bottom two are solids and the middle element is liquid.

Periodic Trends - Summary
•Periodic trends are caused by the interactions of three factors:
–nuclear charge (the number of protons in the nucleus)
–electron shell(s)
–***shielding*** (the effect of the electrons between the outer electrons and the nucleus).

Periodic Trends Summary



Now complete exercise 3.2 on page 89

3.3.1 Discuss the similarities and differences in the chemical properties of elements in the same group.

3.3 The Periodic Table and Chemical Behavior (Textbook p. 91)

•Groups react chemically in a similar fashion

•They react the way they do because of ***valence electrons***

–valence electrons - electrons in the outer energy level (shell) of an atom, regardless of which shell

•***Periodic trends*** are related to chemical properties.

•What kind of reactions are of interest?

–reaction with water/acids

–reaction with air/oxygen

–reaction with halogens.

Chemical Activity & Valency Electrons

•We have already mentioned that the noble gases are ***unreactive***

–this is because they have a ***full outer shell*** of electrons

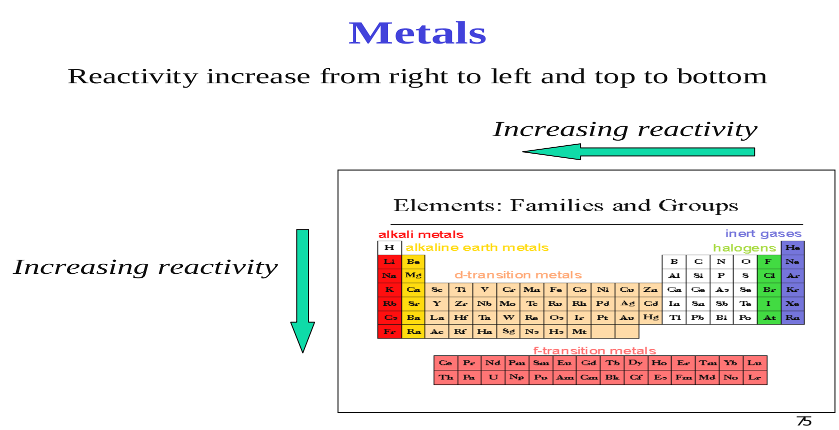
•Elements combine together (react) in order to also obtain a full outer shell - this is ***very stable***

–explains the valences of elements (bonding topic)

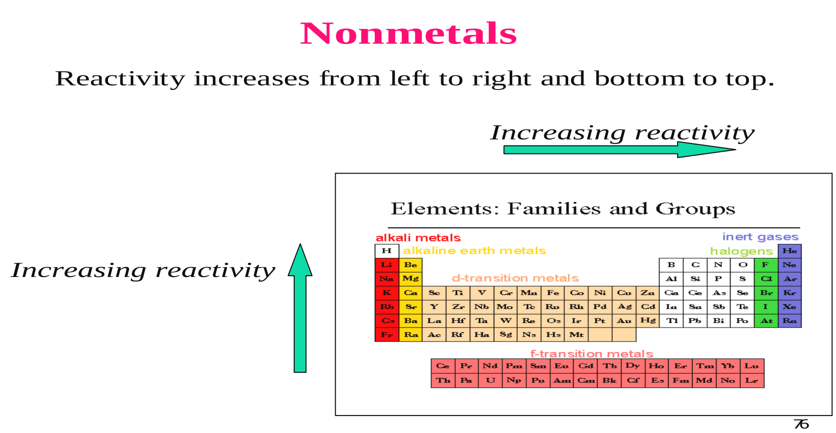
–known as the ***Octet rule***

•*more on this later..*

Groups and chemical activity: Metals



Groups and chemical activity: Nonmetals



Group 1 - Alkali Metals

This is a group of very reactive metals. The most common members of the family are lithium, sodium and potassium.

Element	*Symbol*	*Appearance*	*Melting Point (^{0}C)*	*Density (g/cm^3)*
Lithium Sodium Potassium	Li Na K	Soft grey metal Soft light grey metal Very soft blue/grey metal	181 98 63	0.54 0.97 0.86

The metals have to be stored under oil to exclude air and water. They do not look much like metals, at first sight, but when freshly cut, they all have a typical shiny metallic surface.

Reaction of alkali metals with water

When a small piece of the alkali metal is put into a trough of water, the metal reacts immediately, floating on the surface of the water and evolving hydrogen. With sodium and potassium, the heat evolved from the reaction is sufficient to melt the metal. The hydrogen evolved by the reaction of potassium with cold water is usually ignited and burns with a pink flame. Sodium reacts quicker than lithium and potassium reacts quicker than sodium.
In each case the solution remaining at the end of the reaction is an alkali.

$$2M(s) + 2H_2O(l) \text{-------}> 2MOH(aq) + H_2(g) \qquad [M = \text{Alkali Metal}]$$

Group 7 - The Halogens

This is a family of non-metals. In the halogen family, the different members have different appearances but they are put in the same family on the basis of their similar chemical reactions.

$$2M(s) + X_2\,(g \text{ or } l \text{ or } s)\text{-------}> 2MX(s) \quad [M = \text{Alkali Metal},\ X = \text{halogen}]$$

A Table showing Trends down Group 7

Element	*Formula*	*Appearance at room temperature*	*State at room temperature*	*Order of reactivity*
Fluorine Chlorine Bromine Iodine	F_2 Cl_2 Br_2 I_2	Pale yellow gas Yellow green gas Red/Brown volatile Liquid Dark grey crystalline solid	Gas Gas Liquid Solid	most reactive least reactive

Displacement reaction of the halogens

A more reactive halogen will displace a less reactive halogen from one of its compounds. For example when chlorine is bubbled into a solution of potassium bromide, the chlorine displaces the less reactive bromine. This means the colourless solution turns orange as the free bromine is formed.

$$2KBr(aq) + Cl_2(g) \text{ -------> } 2KCl\,(aq) + Br_2(aq)$$

potassium bromide + chlorine potassium chloride + bromine

Reactivity of the halogens decreases going down the group and a more reactive halogen will displace a less reactive halogen from a solution of its ions. This is also a redox reaction. No reaction would take place if iodine solution was added to potassium bromide solution because iodine is less reactive than bromine.

	Cl^{-} (aq)	**Br^{-}(aq)**	**I^{-}(aq)**
Cl_2	Colorless no reaction	turns red due to formation of bromine	turns brown due to formation of iodine
Br_2	no reaction	no reaction	turns brown due to formation of iodine
I_2	no reaction	no reaction	no reaction

The reactions of Halide ions with Silver ions

	Cl^{-} (aq)	**Br^{-}(aq)**	**I^{-}(aq)**
Ag+	white ppt	cream ppt	yellow ppt
reason	insoluble AgCl formed	insoluble AgBr formed	insoluble AgI formed
equation	$Ag^{+} + Cl^{-} \longrightarrow AgCl$	$Ag^{+} + Br^{-} \longrightarrow AgBr$	$Ag^{+} + I^{-} \longrightarrow AgI$

3.3.2 Discuss the changes in nature, from ionic to covalent and from basic to acidic, of the oxides across period 3.

Period 3 Elements

Elements change from metals to non-metals.

Character of the oxides

Oxides change from basic to acidic. Chlorides react with water forming acidic HCl and hydroxides or oxy-acids (except sodium chloride).

- Metal oxides are basic
- Aluminium oxide is amphoteric (reacts with both acids and bases)
- Non-metal oxides are acidic

	Na_2O(s)	MgO(s)	Al_2O_3(s)	SiO_2(s)	P_4O_{10}(s) (or P_4O_6(s))	SO_3(g) (or SO_2(g))	Cl_2O_7(l)
Adding H_2O	**Na_2O + H_2O -> 2NaOH**	**MgO + H_2O -> $Mg(OH)_2$**	Insoluble	Insoluble	**P_4O_{10} + $6H_2O$ -> $4H_3PO_4$**	**SO_3 + H_2O -> H_2SO_4**	Cl_2O_7 + H_2O -> $HClO_4$
Adding HCl	Na_2O + H^+ -> $2Na^+$ + H_2O	MgO + $2H^+$ -> Mg^{2+} + H_2O	Al_2O_3 + $6H^+$ -> $2Al^{3+}$ + $3H_2O$	No reaction	No reaction	No reaction	No reaction
Adding NaOH	No reaction	No reaction	Al_2O_3 + $2OH^-$ + $3H_2O$ -> $2Al(OH)_4$	SiO_2 + $2OH^-$ -> SiO_3^{2-} + H_2O	P_4O_{10} + $12OH^-$ -> $4PO_4^{3-}$ + $6H_2O$	SO_3 + OH^- -> SO_4^{2-} + H_2O	Cl_2O_7 + OH^- -> $2ClO_4^-$ + H_2O
Nature	Basic Oxide	Basic Oxide	Amphoteric Oxide	Acidic Oxide	Acidic Oxide	Acidic Oxide	Acidic Oxide

Notes

Dichlorine heptoxide (Chlorine(VII) oxide): Cl_2O_7 (b.p.82ºC)

Phosphorus has two common oxides, phosphorus(III) oxide, P_4O_6, and phosphorus(V) oxide, P_4O_{10}.

Phosphorus(III) oxide is a white solid, melting at 24°C and boiling at 173°C.

Phosphorus(V) oxide is also a white solid, subliming (turning straight from solid to vapour) at 300°C.

Now complete exercise 3.3 on page 95