# Chapter 1: Atoms, Molecules and Ions

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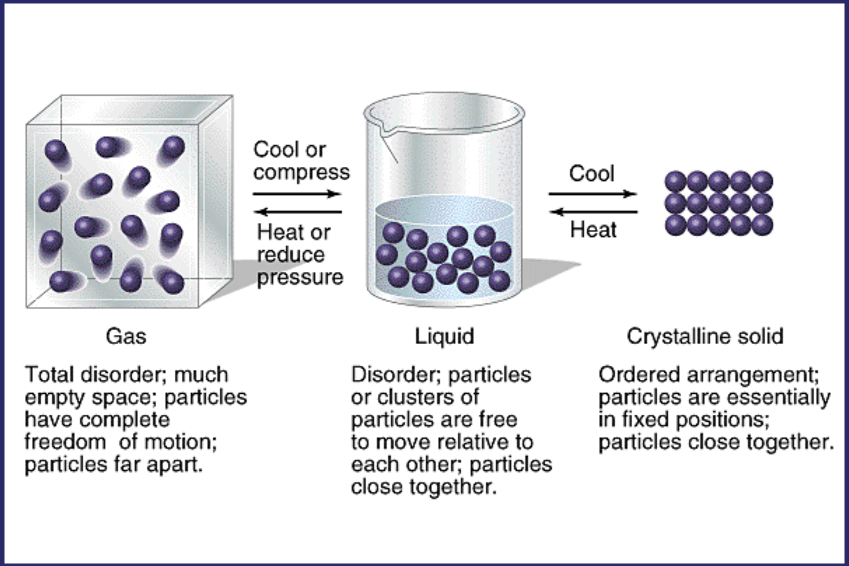
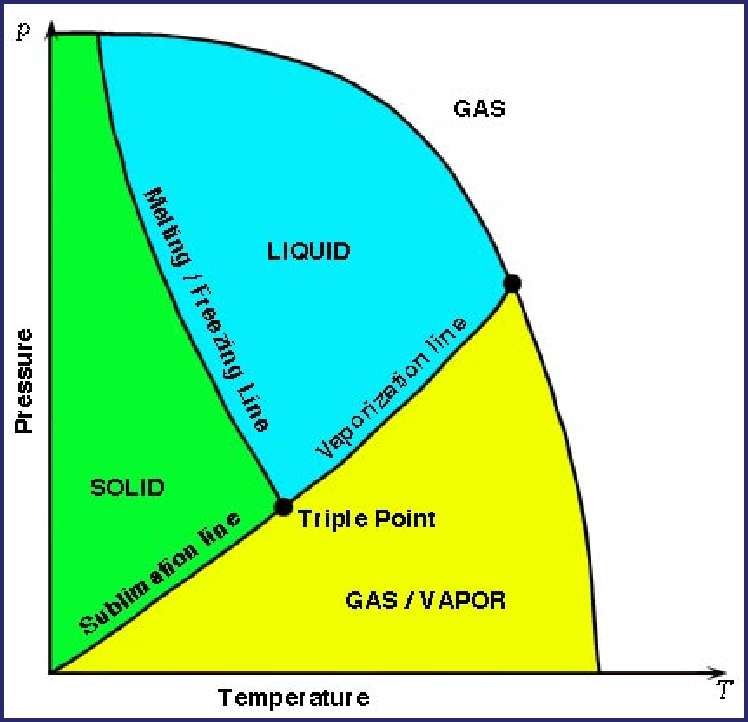
# Section 1.1: Introduction

- **Matter** = the stuff things are made of.
- **Phase =** the physical state a material is in.
  - Depending on the pressure and the temperature, matter can exist in one of three phases (solid, liquid, or gas).
  - The chemical structure of a material determines the range of temperatures and pressures under which this material is a solid, a liquid or a gas.

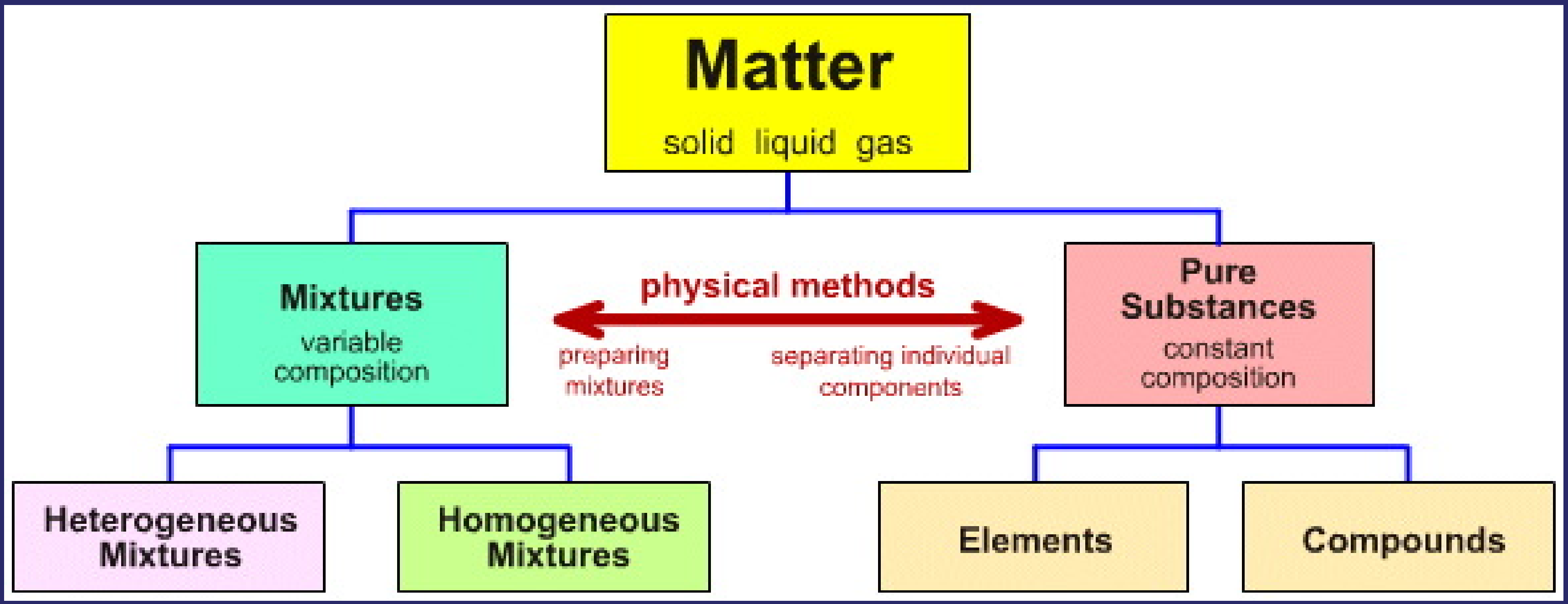


# Section 1.1: Introduction (cont.)

- Example: Water
- The principal differences between water in the solid, liquid and gas states are simply:
  - 1) the average **distance between the water molecules**; small in the solid and the liquid and large in the gas and
  - 2) whether the molecules are **organized** in an orderly three-dimensional array (solid) or not (liquid and gas).



# Section 1.1: Introduction (cont.)



# Section 1.1: Introduction (cont.)

- Another way to classify matter is to consider whether a substance is pure or not.
- A PURE SUBSTANCE has unique composition and properties.
  - Water is a pure substance (each water molecule always contains 2 atoms of hydrogen for 1 atom of oxygen).
  - Under the same atmospheric pressure and at the same ambient temperature, water always has the same density.

# Section 1.1: Introduction (cont.)

- A MIXTURE can be either **homogeneous** or **heterogeneous**.
  - A homogeneous system exhibits a single phase, while a heterogeneous one exhibits one or more phases (different solids, liquids or mixtures of these).

# Section 1.1: Introduction (cont.)

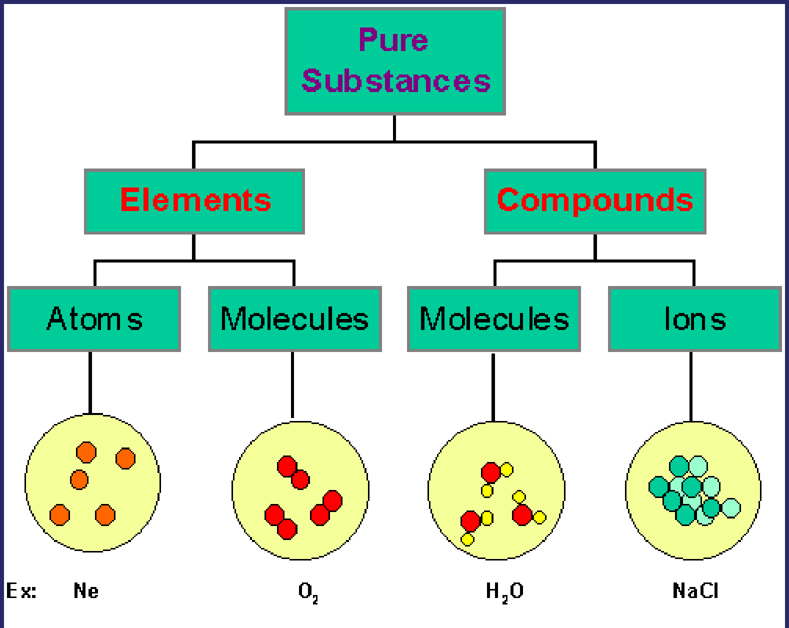
- Homogeneous mixture
  - Example: mixing a teaspoon of salt in a glass of water, the composition of the various components and their properties are the same throughout.
  - Different aliquots of this salt solution would have the same density.

# Section 1.1: Introduction (cont.)

- Heterogeneous mixture
  - Example: dropping gold coins or a teaspoon of oil in a glass of water
  - Different aliquots will contain different amounts of oil or of gold depending on whether these aliquots are taken from the top or the bottom of the mixture.

# Section 1.1: Introduction (cont.)

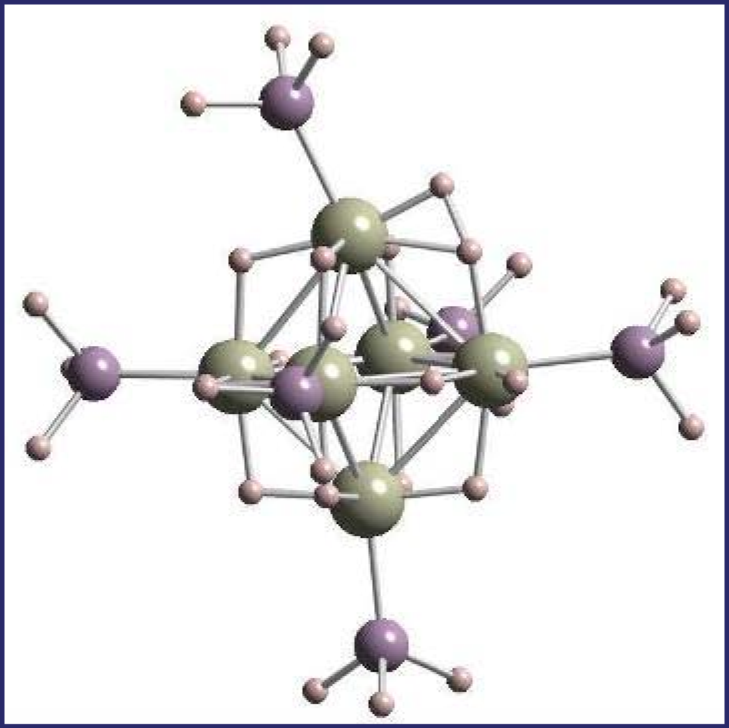
- Elements are the simplest form of matter and cannot be broken down using chemical methods into two or more pure substances.
- Example: iron is a pure substance, you can take a piece of iron and break it down into smaller and smaller pieces, but each of these **smaller pieces has the same properties as the starting material** (hence, it is always the same substance).
- The chart showing all known elements and giving some of their properties is the **Periodic Table of the Elements**.



# Section 1.1: Introduction (cont.)

- Compounds, on the other hand, can be broken down into two or more pure substances.

- Examples:
  - $H_2O$ or water can be broken down into $H_2$ (hydrogen gas) and $O_2$ (oxygen gas).
  - Table salt or NaCl can be broken down into Na (sodium metal) and $Cl_2$ (chlorine gas).



# Section 1.1: Introduction (cont.)

- Compounds are therefore defined as being **made of at least two different elements**.
- A compound is a pure substance with unique composition and properties. Hence, $NO_2$ and $N_2O$ are different compounds since they have different compositions.
- Compounds made with only two elements (such as $H_2O$, $NO_2$, $N_2O$, NaCl) are called **binary compounds**.

# Section 1.2: Basic Periodic Table

- Practice and memorize the first 54 elements (from H: hydrogen to Xe: xenon) of the Periodic Table.

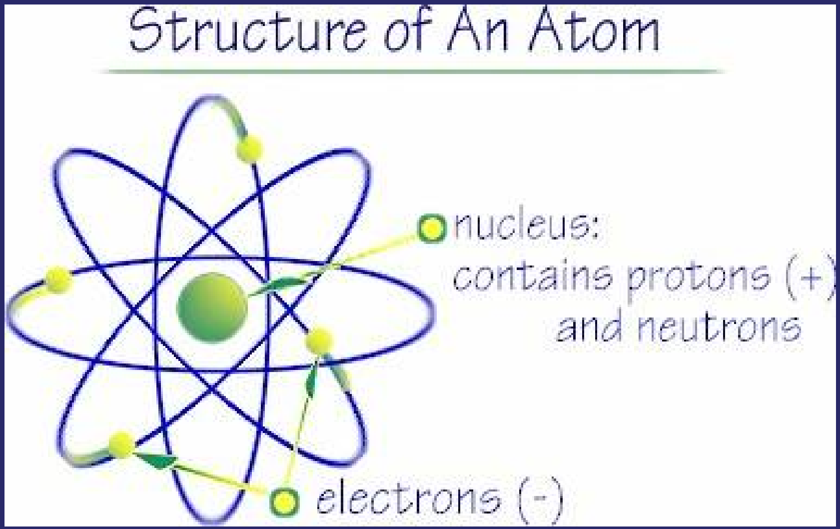
- Also memorize the following elements (Cs, Ba, W, Os, Pt, Au, Hg, Pb, Bi) as they appear frequently in many chemical problems and engineering applications.

# Section 1.3: Concept of Atomic Number

- In 1808, John Dalton developed the **Atomic Model of Matter**. As is the case for every model, Dalton's model is based on a number of assumptions or "postulates".
- An element is composed of particles called **atoms**. All these atoms exhibit the same chemical properties.
  - In a chemical reaction (transformation of pure substances called reactants into other pure substances called products), no atom of any element is destroyed, created or changed into an atom of another element.
  - Compounds are formed when atoms of two or more elements combine in a unique fashion ($CaCl_2$, $H_2O$, etc…).
  - An atom is the smallest "amount" of an element that has the properties of the element. Individual atoms are made of subatomic particles (**electrons**, **protons** and **neutrons**).

# Section 1.3: Concept of Atomic Number (cont.)

- **Electrons:**
  - Are characterized by the symbol **$e^-$**
  - Reside in the outer regions of a given atom
  - Have a very small mass, $m_{e-} = 9.11 \times 10^{-28}$ grams (g)
  - Have a negative electric charge
  - $q_{e-} = -1.6 \times 10^{-19}$ coulombs (C)



# Section 1.3: Concept of Atomic Number (cont.)

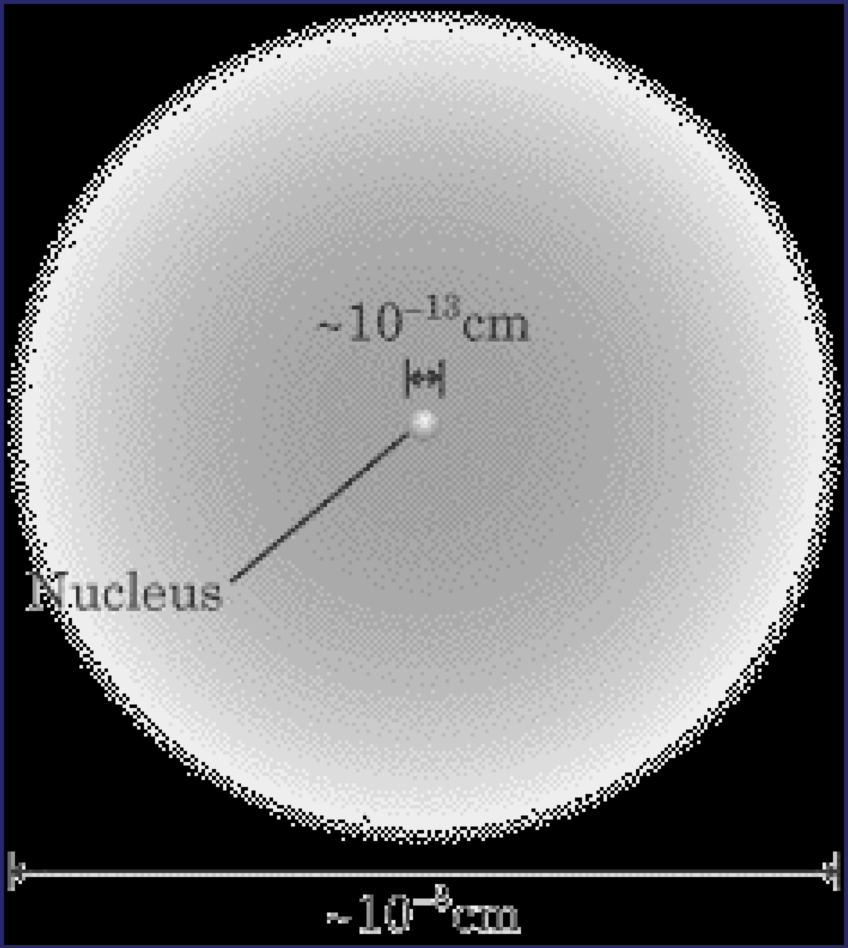
**Protons:**

- Are characterized by the symbol **p**
- Reside in the central part of the atom, known as the **nucleus**
- Have a larger mass than the electron, $m_p = 1.673 \times 10^{-24}$ grams (g)
- Have a positive electric charge $q_p = + 1.6 \times 10^{-19}$ coulombs (C)

# Section 1.3: Concept of Atomic Number (cont.)

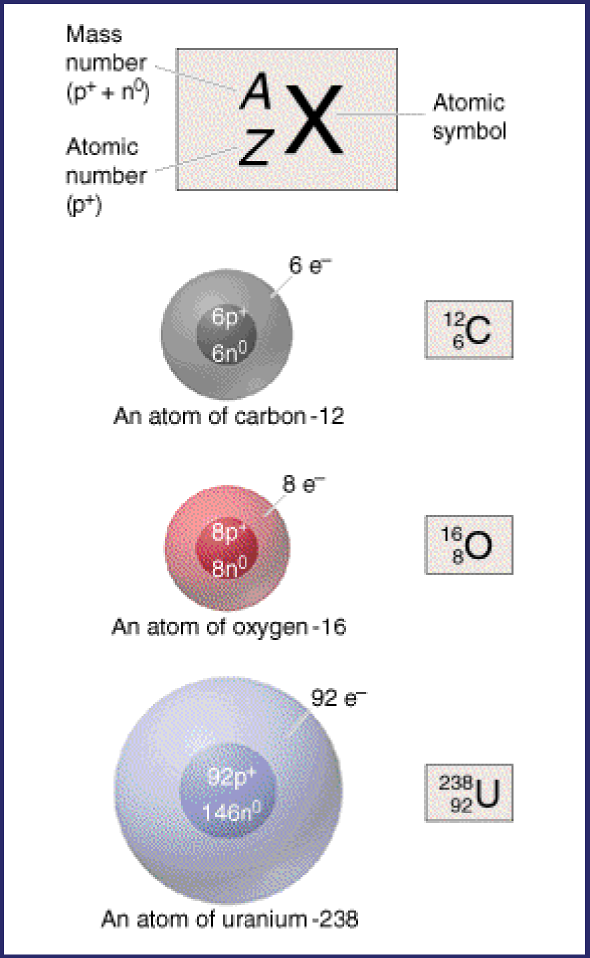
**Neutrons:**

- Are characterized by the symbol **n**
- Reside in the nucleus
- Have about the same mass as the protons, $m_n = 1.675 \times 10^{-24}$ grams (g)
- Have a zero electric charge (i.e. they are neutral)



# Section 1.3: Concept of Atomic Number (cont.)

- **Note:** Protons and neutrons are called **nucleons** as they are all located in the nucleus.
- To make things look simpler, (which sometimes leads to more confusion), we express charges as multiples of the charge of a proton.
  - A proton has a charge of +1
  - An electron has a charge of -1 and in this case we do not specify the units of charge. We understand that the unit of charge is the charge of 1 proton (that is:1.6x $10^{-19}$ C).
  - The key concept to remember is that the charge of the electron is equal in magnitude and opposite in sign to that of the proton.



# Section 1.3: Concept of Atomic Number (cont.)

- The vast majority of the mass of the atom resides in the nucleus.

- All atoms of a given element have the same number of protons in their nucleus. It is the number of protons which characterizes the element. Hence, the number of protons in the nucleus is given the name **atomic number** and is denoted by the symbol **Z**.

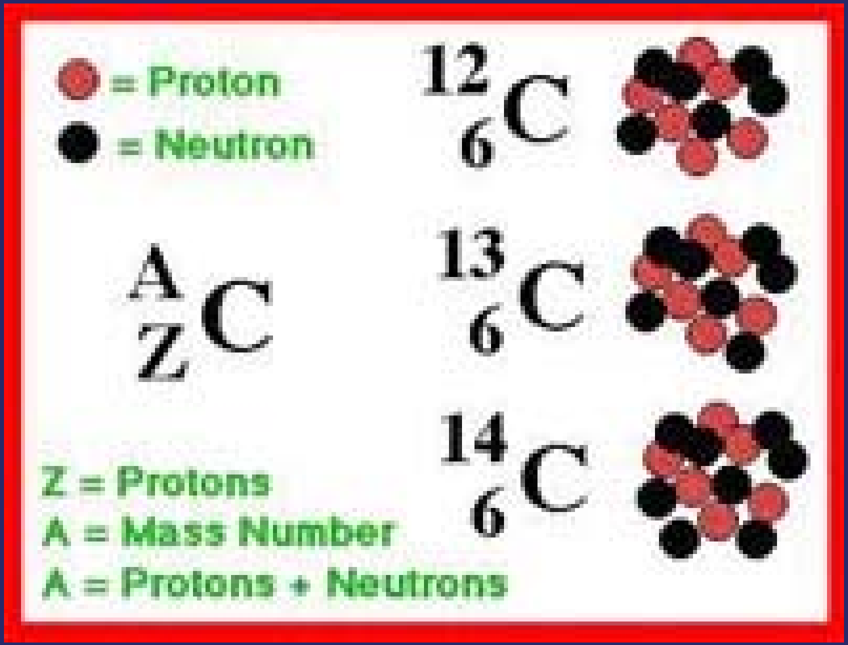
# Section 1.5: Concept of Isotope

- **Mass number** for elements: sum of the number of protons and neutrons , given the symbol **A**.

- Since Z is the number of protons and A is the number of protons + neutrons, then: A – Z is the number of neutrons.

# Section 1.5: Concept of Isotope (cont.)

- **Isotopes:** atoms that have the same number of protons but different numbers of neutrons
- All isotopes of an element have the same number of electrons and the same number of protons, since atoms are always neutral. All carbon isotopes have 6 electrons and 6 protons (Z = 6). However, carbon-12 has 6 neutrons, carbon-13 has 7 neutrons, and carbon-14 has 8 neutrons.
- Isotopes are generally not present in nature in equal quantities; they have different natural abundances.



# Section 1.7: Metals, Nonmetals and Metalloids (cont.)

- Elements in the Periodic Table can be classified as **metals**, **nonmetals** and **metalloids** or **semimetals**.

- **Metals** are typically on the **left-hand side** of the Periodic Table (exception: H is a nonmetal).

- **Nonmetals** are typically on the **right-hand side** of the Periodic Table

- **Metalloids** are on either side of a **stairway** between metals and nonmetals starting between Boron and Aluminum.

# Section 1.7: Metals, Nonmetals and Metalloids

- Metals are characterized by the following physical properties:
  - Luster, high heat and electrical conductivity
  - Malleability, (ability to make films or sheets)
  - Ductility, (i.e. they can be pulled into wires)
- Nonmetals do not exhibit the above properties.

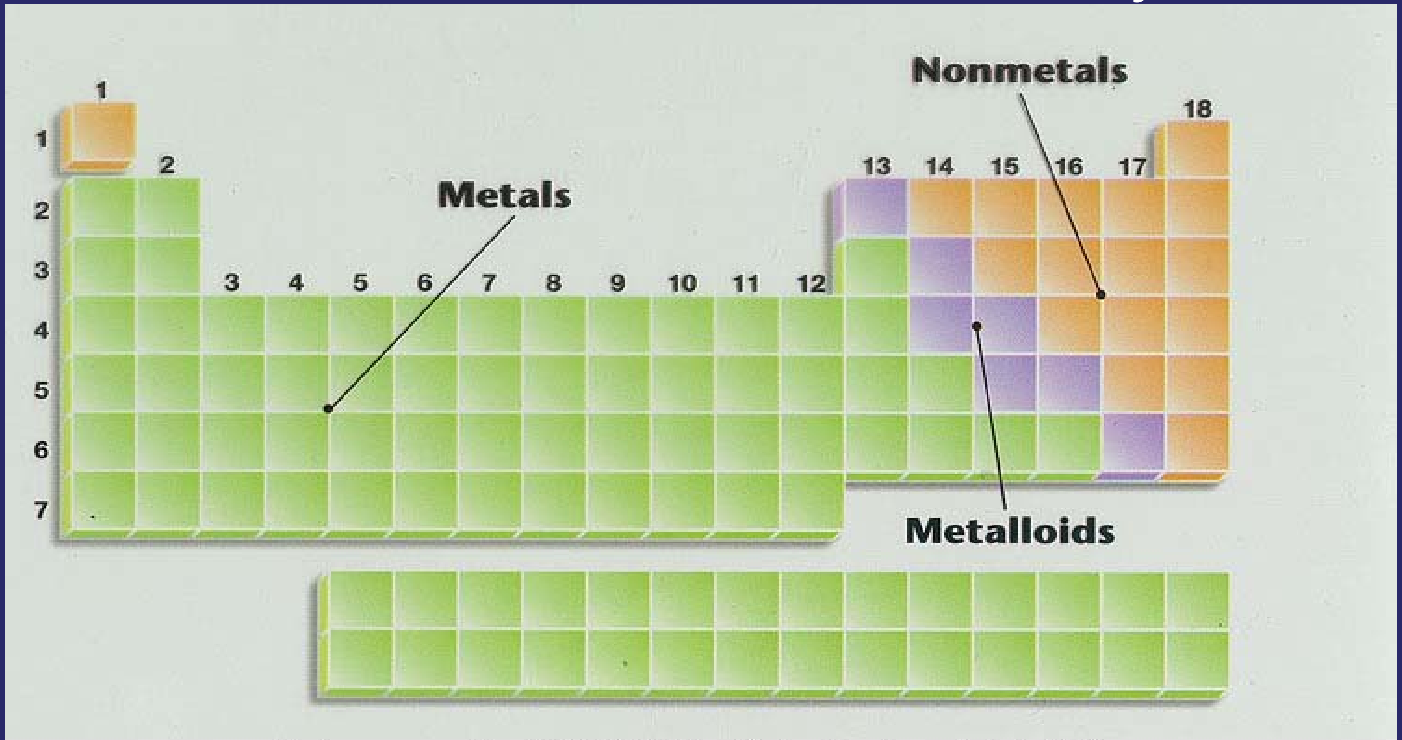
# Section 1.7: Metals, Nonmetals and Metalloids (cont.)

- Metalloids or semi-metals have some properties of metals and some properties of nonmetals. Metalloids include **boron (B), silicon (Si), germanium (Ge), arsenic (As), antimony (Sb), tellurium (Te)** and **polonium (Po)**.

- Different rules apply for the naming of a compound depending on whether the compound includes metallic elements or not.

# Section 1.8: Periodic Table (Metals, Nonmetals and Metalloids)

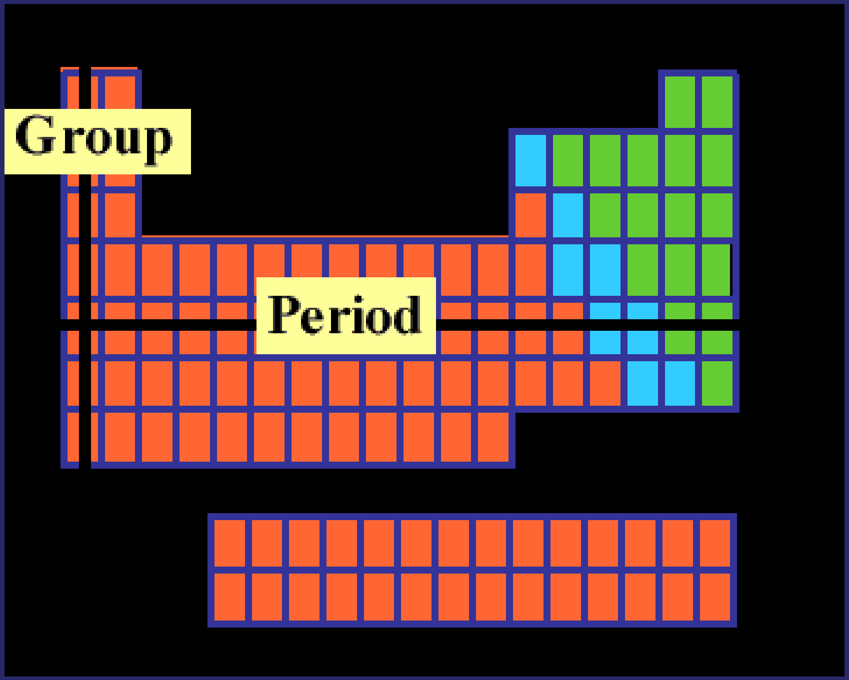
- Review which elements are metals, which are nonmetals and which are metalloids (or semi-metals). Know how to locate the famous “stairway”.




http://www.csun.edu/~psk17793/G%20Chemistry/reading_the_periodic_table.htm

# Section 1.9-1.10: Concept of Group and Period

- The horizontal rows in the Periodic Table are called **periods**.
- The vertical columns in the Periodic Table are called **groups**.
- There are 18 groups and 7 periods in the Periodic Table.

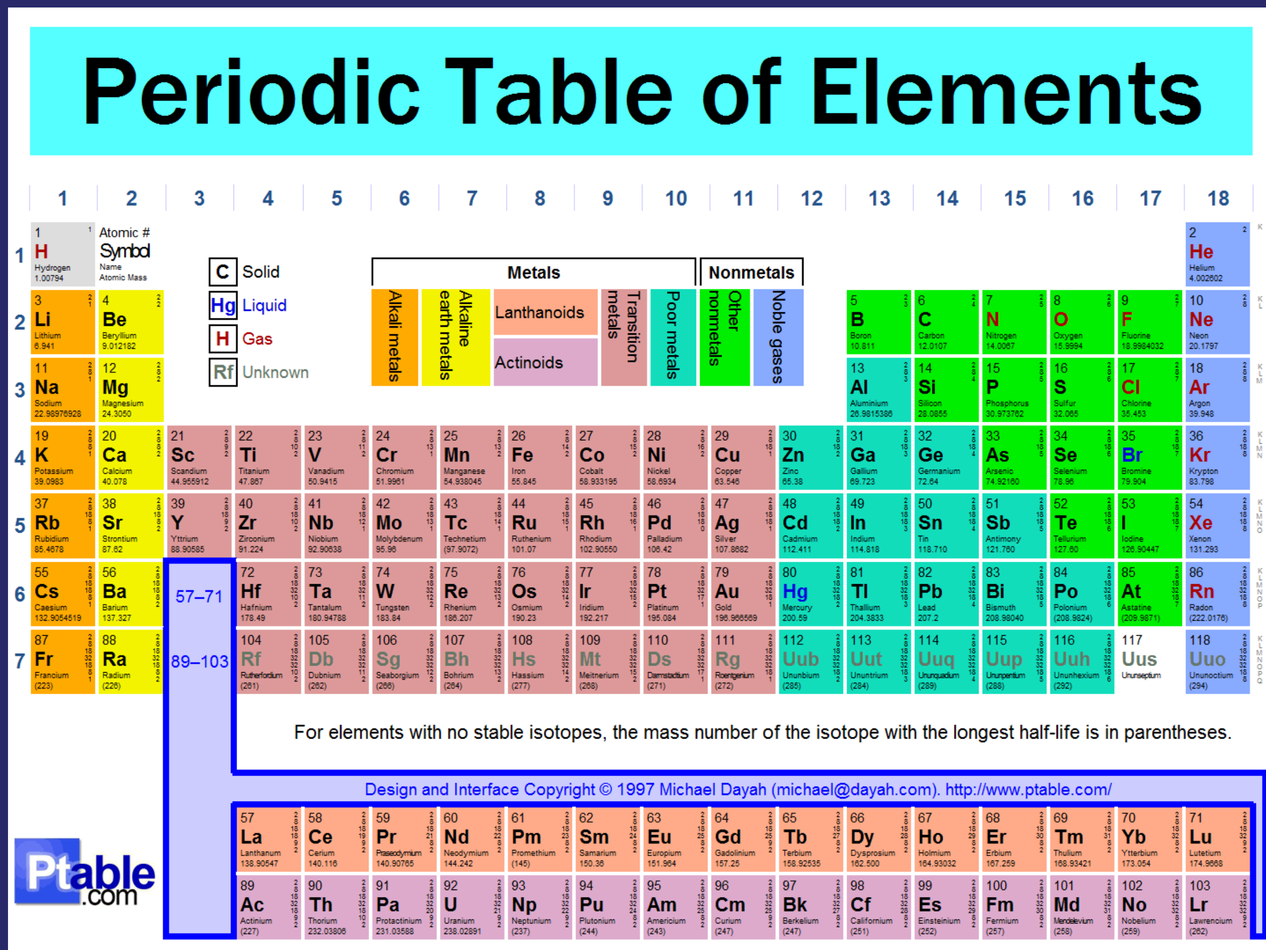


# Section 1.11-1.12: Concept of Family

- Some of the groups and some sets of groups have specific names and constitute **families**.  Here are the families you need to know:
  - Elements in groups 1, 2, 13, 14, 15, 16, 17, 18 are called **main group elements**.
  - Elements in groups 3 through 12 are called **transition elements**.
  - Elements in group 1, except hydrogen, are called **alkali metals**.
  - Elements in group 2 are called **alkaline earth metals**.
  - Elements in group 17 are called **halogens**.
  - Elements in group 18 are called **noble gases**.

- *In Section 1.12, practice locating elements and their respective families using the Interactive Periodic Table.*

# Section 1.11-1.12: Concept of Family (cont.)



# Section 1.13-1.14: Properties of Elements in a Family

- Elements in the same family tend to have similar properties. Properties, while similar, may be of a different magnitude.
- For example, consider C, Si, Ge. We will see in Chapter 8 and Chapter 9 that they exhibit similar bonding with atoms like H or Cl (i.e. form molecules of similar shape).

# Section 1.13-1.14: Properties of Elements in a Family (cont.)

- Consider Bromine, Chlorine, Iodine: They react by a similar mechanism with hydrocarbons (molecules obtained from crude oil that contain only C and H).
- Consider Li, Na and K. They react by a similar mechanism with water.
- *(see video in Section 1.14)*

# Section 1.15: Concept of a Molecule, Part I

- Molecules, like all compounds, are neutral (no net charge). Molecules form when two or more atoms of the same or of different nonmetal elements combine with one another. By "combine" we mean that they form **chemical bonds** between them.

- There are principally two types of chemical bonds:

  - **Ionic bonds** are chemical bonds between a metal and a nonmetal. For example NaCl, CsF, PbCl are ionic bonds.

  - **Covalent bonds** are chemical bonds between two nonmetals. For example CH, NO, HCl, CO, SO, PCl are covalent bonds.

# Section 1.15: Concept of a Molecule, Part I (cont.)

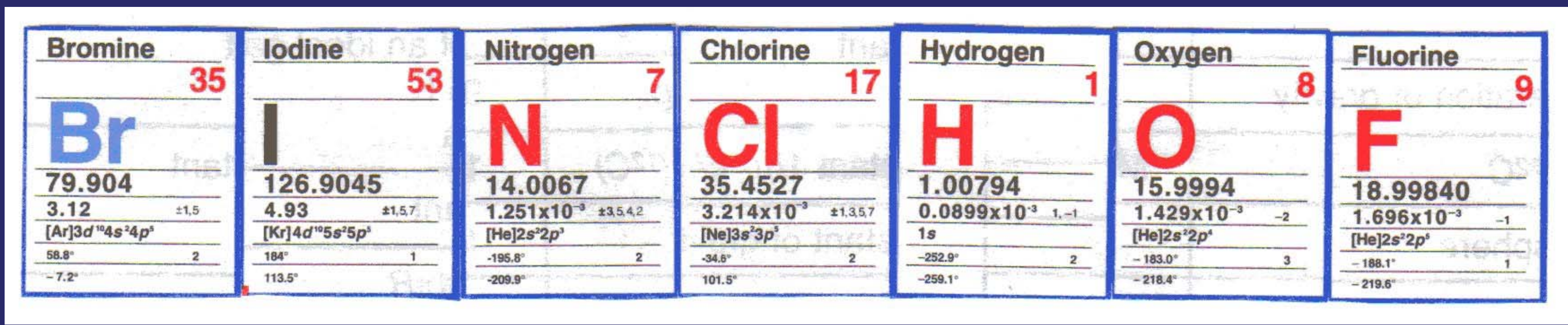
- A significant number of atoms in the Periodic Table exist under normal conditions in the "**elemental form**" as solids.
  - Fe (iron) exists as an element as solid iron
  - C (carbon) exists as solid graphite or solid diamond.
  - There are however a number of notable exceptions that you need to be aware of.

# Section 1.15: Concept of a Molecule, Part I (cont.)

- Under normal conditions (atmospheric pressure and ambient temperature):
  - He, Ne, Ar, Kr, Xe, Rn or noble elements exist as gases.
  - H, N, O, F, Cl are not stable in the elemental form and exist as gases $H_2$, $N_2$, $O_2$, $F_2$, $Cl_2$ in the "**molecular form**".
  - Br and I are not stable in the elemental form and exist as liquid $Br_2$ and solid $I_2$.
  - Phosphorus, P, and sulfur, S, are not stable in the elemental form and exist as $P_4$ and $S_8$ in the molecular form.
  - Hg (a metal) in the elemental form exists as a liquid.

# Section 1.16: Periodic Table (Molecules)

- Interact with the Periodic Table to learn that H, N, O, F, Cl, Br, I, P and S exist under normal conditions in the molecular form ($H_2$, $N_2$, $O_2$, $F_2$, $Br_2$, $I_2$, $P_4$ and $S_8$).
- Know: BrINClHOF



# Section 1.17: Concept of an Ion, Part I

- When an atom loses or gains electrons, a charged particle is formed. This charged particle is called an **ion**.
- Typically, metal elements tend to lose electrons, forming positively charged ions called **cations**.
- Typically, nonmetal elements tend to gain electrons, forming negatively charged ions called **anions**.

**Table 1: Common Simple Cations and Anions**

| cation | name | anion | name |
|---|---|---|---|
| $H^+$ | hydrogen | $H^-$ | hydride |
| $Li^+$ | lithium | $F^-$ | fluoride |
| $Na^+$ | sodium | $Cl^-$ | chloride |
| $K^+$ | potassium | $Br^-$ | bromide |
| $Cs^+$ | cesium | $I^-$ | iodide |
| $Be^{2+}$ | beryllium | $O^{2-}$ | oxide |
| $Mg^{2+}$ | magnesium | $S^{2-}$ | sulfide |
| $Ca^{2+}$ | calcium | | |
| $Ba^{2+}$ | barium | | |
| $Al^{3+}$ | aluminum | | |
| $Ag^+$ | silver | | |

# Section 1.18: Periodic Table (Ions)

- Interact with the Periodic Table, by clicking on the group numbers highlighted in red.
- Memorize the types of ions that are predicted to form for elements in different groups.
- Remember that it is the group an element belongs to, which matters as far as forming ions is concerned.

# Section 1.19: Concept of an Ion, Part II

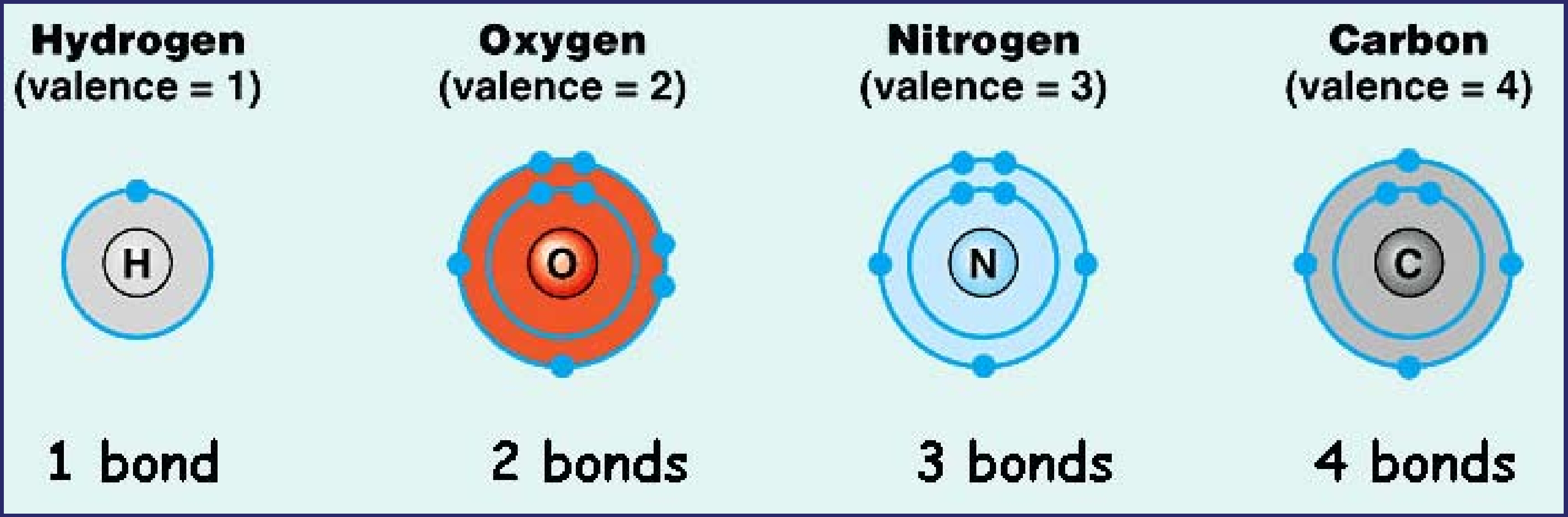
- We learned that:
  - Group 1 elements (Li, Na, K, etc...) tend to lose 1 electron
  - Group 2 elements (Be, Mg, Ca, Ba, etc…) tend to lose 2 electrons.
  - Group 13 metals (Al, Ga, In) tend to lose 3 electrons.
  - Group 15 nonmetals tend to gain 3 electrons.
  - Group 16 nonmetals tend to gain 2 electrons.
  - Group 17 nonmetals (halogens) tend to gain 1 electron, forming **halides.**
  - Group 18 (noble gases) do not form stable ions (they are mostly inert).
- Why is this? Because "**Nobility is Stability**"

# Section 1.19: Concept of an Ion, Part II (cont.)

- What do we mean by Nobility is Stability?
  - Consider Li. Lithium has 3 electrons and 3 protons. Lithium is in group 1. Hence, it forms a stable ion ($Li^+$) by losing 1 electron. Hence, $Li^+$ has 3 protons and 2 electrons. *$Li^+$ has the same number of electrons as Helium (He)*
  - Consider Br. Bromine has 35 electrons and 35 protons. Bromine is in group 17. Hence, it forms a stable ion ($Br^-$) by gaining 1 electron. $Br^-$ has 35 protons and 36 electrons. *$Br^-$ has the same number of electrons as Krypton (Kr).*

# Section 1.19: Concept of an Ion, Part II (cont.)

- Stable ions are often **formed through a gain or a loss of electrons** that is such that the resulting ions have the same number of electrons as the closest noble gas element in the Periodic Table.



# Overview of Section 1.20-1.51: Naming of Compounds

- Recall that there are two main types of compounds: ionic and molecular (or covalent). Different rules apply for ionic and for molecular compounds.
- Don't forget the ion concepts we just learned!
- Ionic compounds: formed by combination of cations with anions in such a way that the resulting substance has a formula that is electrically neutral.

# Section 1.20: Binary Ionic Compounds between Main-group Metals and Nonmetals

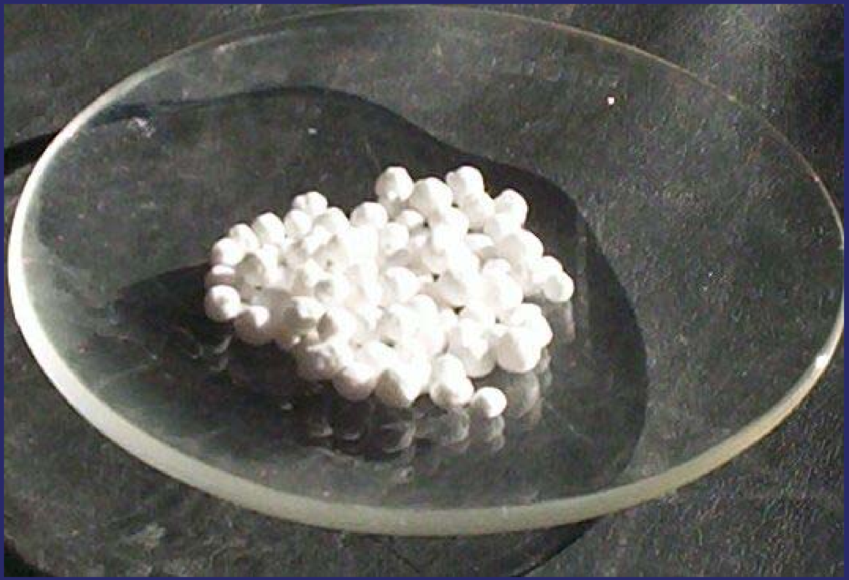
- A **binary ionic compound** is formed using a metal for the cationic species and a non-metal for the anionic species. Here, we consider only main group metal cations (elements in groups 1, 2 and elements Al, Ga of group 13.
- Examples of such compounds are NaCl, $MgF_2$, KBr, $AlCl_3$, etc…

# Section 1.20: Binary Ionic Compounds between Main-group Metals and Nonmetals (cont.)

- The rules for naming these compounds are:
  - The metal cation is always named first and the nonmetal anion second
  - The cation is named exactly as the element
  - The anion is named by keeping the root of the element name and adding “**ide**” to the **root**.
  - Cl is named **chlorine**, hence $Cl^{-}$ is named **chloride**.
  - P is named **phosphorous**, hence $P^{-3}$ is named **phosphide**.
  - O is named **oxygen**, hence $O^{-2}$ is named **oxide**.

# Section 1.20: Binary Ionic Compounds between Main-group Metals and Nonmetals (cont.)

- **Example 1: Name KBr**
- K is a metal, Br is a nonmetal.
- K is a group 1 metal, so the cation formed from K is $K^+$.
- Br is a group 17 nonmetal, hence the anion formed from Br is $Br^-$.
- Combining the two ions forms a neutral compound (1 positive and 1 negative charge).
- KBr is potassium bromide.

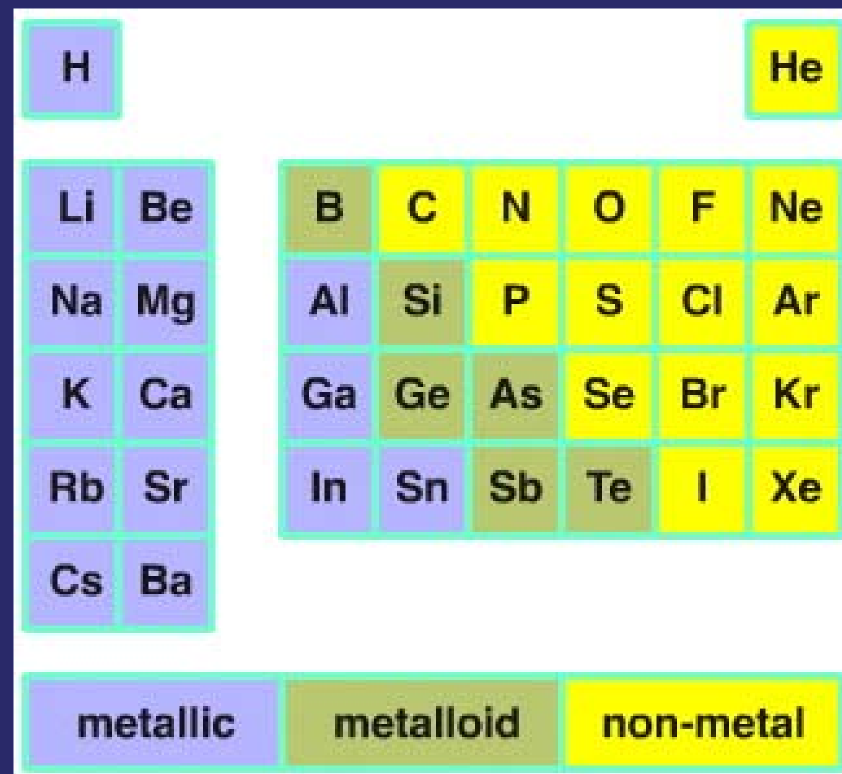


# Section 1.20: Binary Ionic Compounds between Main-group Metals and Nonmetals (cont.)

- **Example 2: Name $MgF_2$**
- Mg is a metal, F is a nonmetal.
- Mg is a group 2 metal, so the cation formed from Mg is $Mg^{+2}$.
- F is a group 17 nonmetal; the anion formed from F is $F^{-}$.
- Combining the two ions to form a neutral compound requires 1 $Mg^{+2}$ and 2 $F^{-}$ (2 positive charges and 2 negative charges).
- $MgF_2$ is magnesium fluoride.

# Section 1.21: Periodic Table: Binary Ionic Compounds between Main-group Metals and Nonmetals (cont.)

- Practice making compounds using group 1, 2 or 13 metals and group 15, 16 or 17 nonmetals. Remember the names and chemical formulas of these compounds.



# Section 1.22: Binary Ionic Compounds between Transition Metals and Nonmetals

- Many transition metals are capable of forming cations bearing different charges.
  - Fe can form $Fe^{+2}$ and $Fe^{+3}$ cations
  - Fe can in general form two compounds with a given non metal
  - compounds made from $Fe^{+2}$ or $Fe^{+3}$ and $Cl^{-}$ are $FeCl_2$ and $FeCl_3$

- The systematic name is derived by naming the cation as the transition metal itself followed (without space) by a Roman numeral in parentheses indicating the charge of the metal cation.

- $Fe^{+2}$ and $Fe^{+3}$ are named <u>iron(II)</u> and <u>iron(III)</u>, respectively.

# Section 1.22: Binary Ionic Compounds between Transition Metals and Nonmetals (cont.)

- To complicate matters, there is a different way to name such cations.

- This method is grounded in the history of chemistry and consists in using different names for each of the possible ions.

- For instance, in cases where only two common ions are observed, the ion with the lowest charge is named by adding “ous” to the latin root of the element.

- Consider the following non-exhaustive list of cations that are commonly used in the chemical literature.

# Section 1.22: Binary Ionic Compounds between Transition Metals and Nonmetals (cont.)

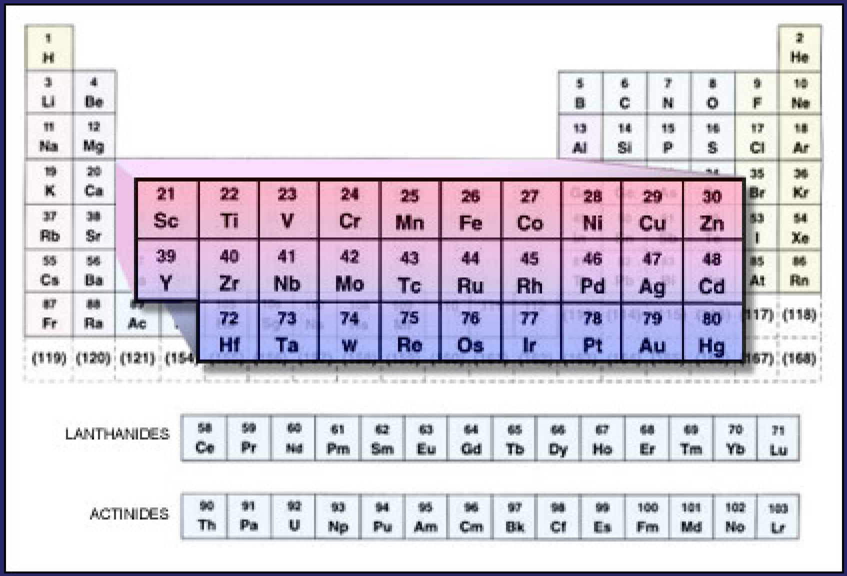
- You should use the systematic nomenclature but recognize the old names.

  The systematic nomenclature for naming binary ionic compounds having a transition metal consists in:

  - Naming the transition metal using a Roman numeral in parentheses to characterize the charge of the cation,
  - Naming the anion using the root of the element followed by "ide".
  - For example $FeCl_2$ and $FeCl_3$ are called iron(II) chloride and iron(III) chloride, respectively.

# Section 1.23: Periodic Table: Binary Ionic Compounds between Transition Metals and Nonmentals

- Practice naming these compounds and note the various cations exhibited by some of the most common transition metals.
- In particular note that most of these cations, in contrast with cations from Main Group Metals, do not have the same number of electrons as noble gases.
- You do not need to memorize all these ions, just know how to name them following the systematic nomenclature.



# Section 1.24-1.27: Introduction to Polyatomic Ions

- Besides the group of ions formed from metals and nonmetals, there exists a large group of ions called polyatomic ions.
- Polyatomic ions contain different atoms linked together by covalent bonds.
- The names of these ions, their chemical formula and net charge *must be memorized*.

# Section 1.24-1.27: Introduction to Polyatomic Ions (cont.)

- Note that the common features shown by the last three sets. The ions, whose names end with "ate" contain one oxygen more than those, whose names end with "ite". Note also that the charge is the same for the "ite" and "ate" anions.

| | | | |
|---|---|---|---|
| $NH_4^+$ | ammonium ion | $OH^-$ | hydroxide ion |
| $C_2H_3O_2^-$ | acetate ion | $MnO_4^-$ | permanganate ion |
| $CrO_4^{-2}$ | chromate ion | $Cr_2O_7^{-2}$ | dichromate ion |
| $CN^-$ | cyanide ion | $CO_3^{-2}$ | carbonate ion |
| $NO_3^-$ | nitrate ion | $NO_2^-$ | nitrite ion |
| $SO_4^{-2}$ | sulfate ion | $SO_3^{-2}$ | sulfite ion |
| $PO_4^{-3}$ | phosphate ion | $PO_3^{-3}$ | phosphite ion |

# Section 1.24-1.27: Introduction to Polyatomic Ions (cont.)

Anions made using halogens and oxygen

- “per” indicates that these compounds have the most oxygen content (4 oxygen atoms per halogen atom).
- “hypo” indicates that these compounds have the least oxygen content (1 oxygen atom per halogen atom).
- Note again that chlorate, bromate and iodate have one more oxygen atom that chlorite, bromite and iodite.

| | | | | | |
|---|---|---|---|---|---|
| $ClO_4^-$ | perchlorate ion | $BrO_4^-$ | perbromate ion | $IO_4^-$ | periodate ion |
| $ClO_3^-$ | chlorate ion | $BrO_3^-$ | bromate ion | $IO_3^-$ | iodate ion |
| $ClO_2^-$ | chlorite ion | $BrO_2^-$ | bromite ion | $IO_2^-$ | iodite ion |
| $ClO^-$ | hypochlorite ion | $BrO^-$ | hypobromite ion | $IO^-$ | hypoiodite ion |

# Section 1.24-1.27: Introduction to Polyatomic Ions

- The final series of polyatomic anions to consider are the series of anions such as: $CO_3^{-2}$, $SO_4^{-2}$, $SO_3^{-2}$, $PO_4^{-3}$, $PO_3^{-3}$.
- **Consider the phosphate ion: $PO_4^{-3}$**
  - Adding $H^+$ to $PO_4^{-3}$ results in the formation of $HPO_4^{-2}$ or hydrogen phosphate.
  - Adding $H^+$ to $HPO_4^{-2}$ leads to the formation of $H_2PO_4^-$ or dihydrogen phosphate.
- Similarly, the following compounds are:
  - $HCO_3^-$ hydrogen carbonate
  - $HSO_4^-$ hydrogen sulfate
  - $HSO_3^-$ hydrogen sulfite
  - $HPO_3^{-2}$ hydrogen phosphite
  - $H_2PO_3^-$ dihydrogen phosphite

# Section 1.28-1.43: Ionic Compounds Made with Polyatomic Ions

## COMPOUNDS/FORMULAE TABLE

You need to know the names and formulae of the following chemical compounds:

| Compound | Formula | Compound | Formula |
|---|---|---|---|
| Ammonia | $NH_3$ | Barium chloride | $BaCl_2$ |
| Carbon dioxide | $CO_2$ | Sodium chloride | $NaCl$ |
| Methane | $CH_4$ | Calcium carbonate | $CaCO_3$ |
| Water | $H_2O$ | Copper carbonate | $CuCO_3$ |
| Hydrochloric acid | $HCl$ | Sodium carbonate | $Na_2CO_3$ |
| Sulphuric acid | $H_2SO_4$ | Potassium nitrate | $KNO_3$ |
| Calcium oxide | $CaO$ | Silver nitrate | $AgNO_3$ |
| Iron oxide | $Fe_2O_3$ | Barium sulphate | $BaSO_4$ |
| Lead oxide | $PbO$ | Copper sulphate | $CuSO_4$ |
| Sodium hydroxide | $NaOH$ | Sodium sulphate | $Na_2SO_4$ |

# Section 1.44-1.45: Introduction to Binary Molecular Compounds

- A binary molecular compound is made by **combination of two nonmetals**.
- Bound together by covalent bonds.
- Binary molecular compounds containing hydrogen have the following formulae and names.

# Section 1.44-1.45: Introduction to Binary Molecular Compounds

- Note that in binary molecular compounds made with hydrogen and nonmetals, **hydrogen is usually named first.**

| Nonmetal | Compound Formula | Compound Name |
|---|---|---|
| N | $NH_3$ | Ammonia |
| O | $H_2O$ | Water |
| S | $H_2S$ | Hydrogen sulfide |
| F | HF | Hydrogen fluoride |
| | | Hydrofluoric acid (in solution) |
| Cl | HCl | Hydrogen chloride |
| | | Hydrochloric acid (in solution) |
| Br | HBr | Hydrogen bromide |
| | | Hydrobromic acid (in solution) |
| I | HI | Hydrogen iodide |
| | | Hydroiodic acid (in solution) |

# Section 1.46-1.47: Introduction to Molecular Compounds Containing Polyatomic Anions and Hydrogen

- **Note:**
  - Acids whose names finish with “ic” come from anions whose names finish in “ate”.
  - Acids whose names finish with “ous” come from anions whose names finish in “ite”.
- **Examples:** Sulfuric acid ($H_2SO_4$) comes from the sulfate ion ($SO_4^{-2}$)
- Sulfurous acid ($H_2SO_3$) comes from the sulfite ion ($SO_3^{-2}$)

# Section 1.50-1.51: Concept of a Molecule, Part II: Naming Binary Molecular Compounds

The following rules are used when naming binary molecular compounds.

- The first element is named using its full name (as if the element was a Main Group metal).
- The second non-metal is named as if it were the anion of a non metal.
- Greek prefixes are used to indicate the number of atoms present in the compound chemical formula.
- The prefix “mono” for 1 is never used in front of the name of the first element.
- [http://www2.pvc.maricopa.edu/tutor/chem/chem130/nomenclature/ncrules.html](http://www2.pvc.maricopa.edu/tutor/chem/chem130/nomenclature/ncrules.html)

| NUMBER | PREFIX | EXAMPLE |
|---|---|---|
| 1 | NONE | CHLORIDE |
| 2 | DI- | DICHLORIDE |
| 3 | TRI- | TRICHLORIDE |
| 4 | TETRA- | TETRACHLORIDE |
| 5 | PENTA- | PENTACHLORIDE |
| 6 | HEXA- | HEXACHLORIDE |
| 7 | HEPTA- | HEPTACHLORIDE |
| 8 | OCTA- | OCTACHLORIDE |
| 9 | NONA- | NONACHLORIDE |
| 10 | DECA- | DECACHLORIDE |

# Section 1.50-1.51: Concept of a Molecule, Part II: Naming Binary Molecular Compounds (cont.)

**Exceptions to the Rules for Binary Molecular Compounds:**

- The oxides of phosphorus are often named assuming phosphorus was a metal. Hence, $P_4O_{10}$ would be phosphorus(V) oxide.
- Molecular formula of binary hydrogen compounds with group 15 non-metals are written with the group 15 non-metal first (i.e. $NH_3$ for ammonia, $PH_3$ for phosphine).
- No Greek prefix is used with the binary hydrogen compounds
  (for example, $H_2S$ is hydrogen sulfide and NOT dihydrogen monosulfide).