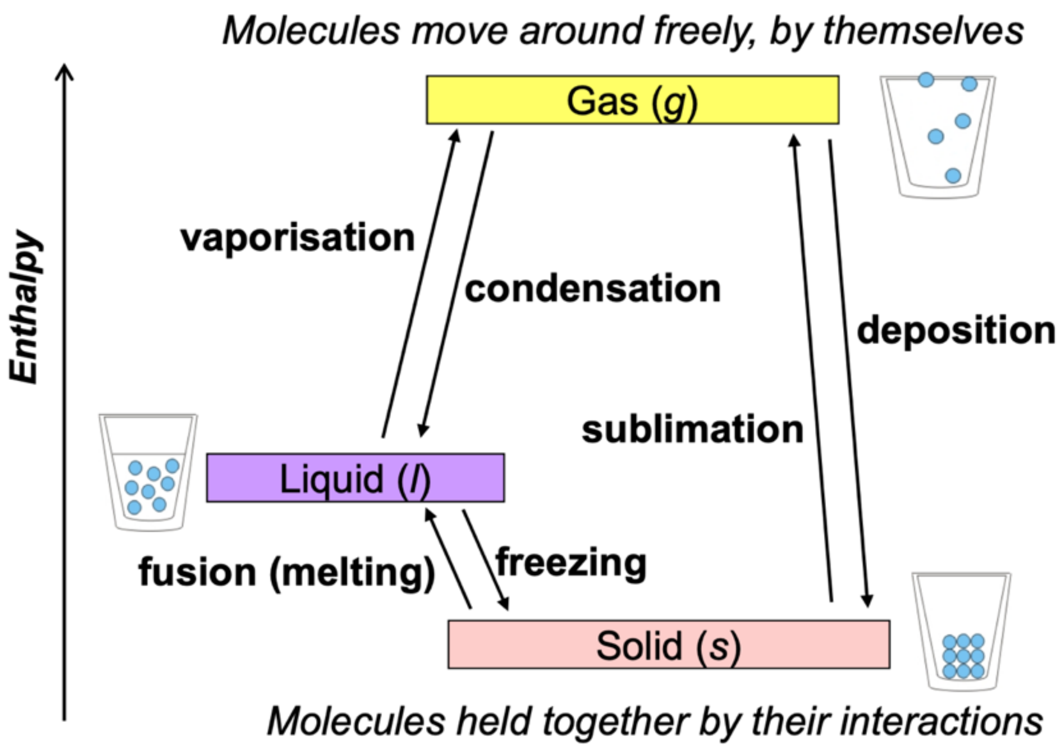
MODULE: Part 1: Matter and Atoms

- **Matter –** Anything that occupies space and mass (gas, liquids, and solids)
- **Plasma –** A fourth state of matter
    - Occurs naturally in interiors of stars
    - **A gaseous state that contains appreciable numbers of electrically charged particles**
    - Also found in other high temperature environments (lighting, televisions, some)
- Mass of object - measure of amount of matter in it
    - Ways of measurement
        - Force it takes to accelerate an object (I.e. takes more force to move a car than a bike)
        - Using a balance to compare mass with a standard mass
- Weight – the force that gravity exerts on object
    - Force – directly proportional to mass
    - Weight of object changes as force of gravity
    - i.e. Astronaut mass does not change but weight does
- Law of conservation
    - Summary of many scientific observations about matter
    - *There is no detectable change in the total quantity of matter present when matter converts from one type to another chemical change) or changes among solid, liquid or gaseous states (physical change)*
        - Brewing beer example

Classifying States of Matter

- **Matter is linked to kinetic molecular model of matter**
    - Solids: particles closely packed usually in regular array vibrate rapidly in fixed positions
    - Liquids and gases are fluid as particles not in fixed position and can move past one another
    - Gas, molecules or atoms are far apart and fly around quickly, colliding with one another and container walls. They go to every part of their container, so the volume and shape are the volume and shape of the container

Phase Changes



- **Pure substance:** Has constant composition/matter that is composed of only one substance
    - **All specimens have the same makeup and properties**
        - **i.e. Sucrose is 42.1% carbon, 6.5% hydrogen, and 51.4 oxygen**

            - Same physical properties in all specimens, same melting point, color, sweetness
    - Pure substances broken down into **elements** and **compounds**
        - **Element** = pure substance that cannot be further purified
            - Iron silver gold oxygen
        - **Compound** = pure substances that can be broken down by chemical changes
            - **Breakdown a produced either elements or other compounds, or both**
                - **I.e. silver chloride can be broken down into silver and chlorine by the absorption of light**
- **Mixture =** Composed of two or more types of matter that can be present in carrying amounts and can be separated by physical changes such as evaporation
    - Heterogeneous mixture: mixture with composition varying from point to point
        - i.e. Italian dressing
        - Chocolate chip cookies
    - Homogenous mixture (solution) exhibits uniform composition and appears the same throughout visually
        - i.e. sports drink, maple syrup, salt solution

Atoms and Molecules

- **Atom:** Smallest particle of an element that has the properties of that element and can enter into a chemical combination
    - **A lead atom weights 0.**0000001g
- Collections of individual atoms is rare
    - A few such as helium, neon, and argon consist of a collection of individual atoms that move independently of one other
    - Others like hydrogen, oxygen are composed of units that consist of pares of atoms
- Molecule: consists of two more ore atoms joined by a chemical bond
    - Two more atoms like hydrogen or oxygen or two or more different atoms, like molecules found in water (2 H and 1 0)
    - Glucose is 6 carbon, 12 hydrogen, and 6 oxygen

CLASSIFYING MATTER AND THE TRANSFORMATION OF SUBSTANCES

- Characteristics that distinguish one substance from another are called properties
    - **Physical property: C**haracteristic of matter that is not associated with change in chemical composition
        - Include, density, colour hardness, melting and boiling points and electrical conductivity
        - **Physical change** is change in state or properties of matter without accompanying change in chemical identity
            - **I.e. when wax melts or when sugar dissolves in coffee or when steam condenses into liquid**
    - **Chemical property:** Describes substance characteristic behaviour in reaction with other substances and what happens if it does react
        - **Example: flammability, toxicity, acidity**
            - **I.e. iron combines with 02 in presence of water to form rust**
        - **Chemical change** always produces one or more types of matter that differ from matter present
            - Rust is chemical change as rust is a different kind of matter than iron, oxygen and water
- **Properties of matter have two categories**
    - **Extensive Property:** If property depends on amount of matter present
        - **Mass and volume of substance are example of extensive property**
            - **I.e.** gallon of milk has larger mass than cup of milk
                - **Value of extensive property directly proportional to amount of matter in question**

    - **Intensive property:** Not dependent on amount of matter present
        - **I.**e. Temperature. If the gallon and cup of milk are each 20 degrees, when combined, temperature remains 20 degrees
- Whilst many elements differ in chemical and physical properties, some elements have similar properties
    - I.e. many elements conduct heat and electricity, whereas others are poor conductors
        - Metals: conduct well
        - Non-metals: conduct poorly
        - Metalloids: intermediate conductivities

Atoms

The contents of an atom

Properties depend on arrangement of 3 discrete sub-atomic particles that reside within an atom

- **Protons**
- **Neutrons**
- **Electrons**

Protons and neutrons

- Centre of atom lies nucleus consisting of
    - Protons (positively charged particles)
    - Neutrons (Electrically neutral particles) – similar mass to proton
    - Pack very close together
    - Make up 99% of mass

Electrons

- Negatively charged, opposite to protons
- Positively charged nucleus attracts negative particles by electrostatic attraction
- They form a cloud of negative charge around the nucleus which gives the atoms its volume
- Contribute very little mass to atom (~1800 times lighter than proton), but they are in constant motion and take up more space
- The atom is mostly empty space

Electrical neutrality

- **Each electron is a minus, and protons are a one plus**
- **Atoms are electrically neutral**
- **Number of electrons is equal to number of protons**

Size

- The diameter of an atom is on the order of 10^{-10} m, whereas the diameter of the nucleus is roughly 10^{-15} m—about 100,000 times smaller. For a perspective about their relative sizes, consider this: If the nucleus were the size of a blueberry, the atom would be about the size of a football stadium
- Appropriately small units of measure are used to measure atoms
    - **Atomic mass (amu)**
        - Defined with regard to most abundant isotope of carbon, atoms are assigned masses of exactly 12 amu (carbon-12)
        - One amu is exactly 1/12 mass of one carbon-12 atom
            - 1 amu = 1.6605×10^{-24}
        - Dalton (Da) and Unified atomic mass (u) are alternate units that are equivalent to amu

- **Fundamental unit of charge (e or elementary charge)** equals magnitude of charge of an electron (e), e = 1.602×10^{-19}
- **Proton – mass of 1.0073 amu and charge of 1+**
- **Neutron – mass of 1.0087 amu and charge of 0**
- **Electron = mass of 0.00055 amu and charge of 1-**

Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)
electron	outside nucleus	-1.602×10^{-19}	1–	0.00055	0.00091×10^{-24}
proton	nucleus	1.602×10^{-19}	1+	1.00727	1.67262×10^{-24}
neutron	nucleus	0	0	1.00866	1.67493×10^{-24}

Implications

Charged particles within an atom:

- Gives rise to properties of elements
- Gives rise to properties of compounds

Element Identity and Atomic Number

- Number of proteins of an atom is the defining feature of an element
- All atoms of an element have the same number of protons in the nucleus (therefore same number of electrons)
    - Example: any atom that contains six protons is the element carbon and has the atomic number 6 (regardless of how many neutrons or electrons it may have)
    - Example: any atom that contains one proton is Hydrogen
- **Atomic Number (Z):** Total number of protons
- **Mass number (A):** The total number of protons and neutrons
- **Number of Neutrons therefore = A – Z**
- **Atomic Charge =** number of protons – number of electrons
    - Atoms are electrically neutral if they have the same number of positively charged protons and negatively charged electrons
    - **Ion =** When numbers are not equal, the atom is electrically charged
    - **Anion =** When atom gains one or more electrons it will be negatively charged
    - **Cation =** When atom loses one or more electron

Isotopes

- Every atom of an element has the same number of proteins in the nucleus
- However, most elements have two or more types of atom with different number of neutrons
    - I.e. whilst all carbon atoms have 6 protons and 6 electrons, 98.89% have 6 neutrons in nucleus whilst 1.11% have 7 neutrons and 0.01% have 8 neutrons
- **Isotope =** atoms of the same element with different numbers of neutrons are isotopes of that element

All Mg have 12 protons in nucleus This example has 24 as mass – has 12 neutrons whereas a 25 would have 13 and would be written Mg-25

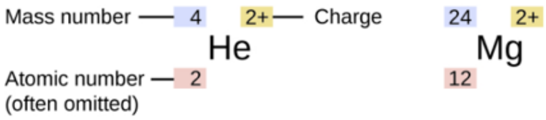


Figure 2.14 The symbol for an atom indicates the element via its usual two-letter symbol, the mass number as a left superscript, the atomic number as a left subscript (sometimes omitted), and the charge as a right superscript.

Atomic Mass

- Each proton and each neutron contribute approx. to one amu to mass of an atom, and electron much less, **atomic mass** is approx. equal to its mass number (whole number)
  - Average masses of atoms are not whole numbers because most elements exist as mixtures of two or more isotopes

$$\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})_i$$

- **Boron Example**
  - **Boron is composed of two isotopes**
    - **19.9% are ^{10}B with mass of 10.0129 amu and 80.1% are ^{11}B with 11.0093 amu**
    - **The average atomic mass is:**
      - **(0.199 x 10.0129 amu) + (.801 x 11.0093 amu)**
      - **= 1.99 + 8.82**
      - **=10.81 amu**

How Ions are formed and charges on them

- Cations are generally formed by chemical reactions by removal of one or more electrons from atoms of metallic elements (i.e. to form K+ ions from K atoms, or MG^{2+} Ions from Mg atoms)
- Anions are formed by gaining one or more electrons by atoms of nonmetallic elements (e.g. Cl^- ions from Cl atoms)
- Noble gases form neither cations nor anions

Topic 2: Part 4: The Electronic Structure of Atoms

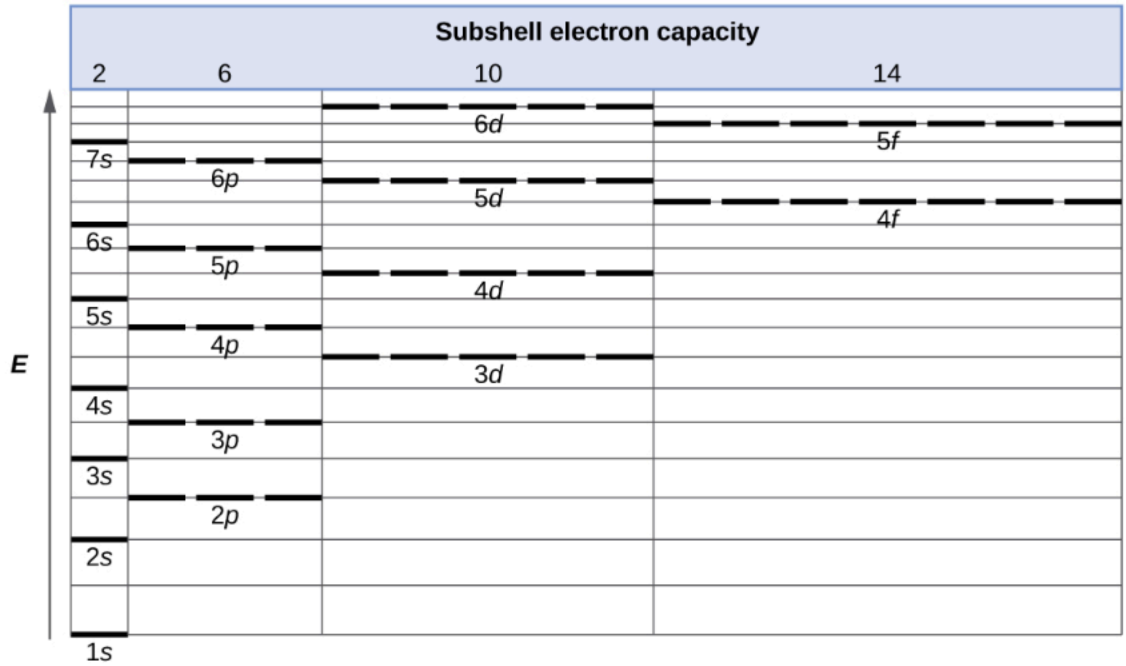
Concepts covered

- Electron behaviour in an atom
- Orbital energies
- Shapes of orbitals
- Filling electron orbitals (Aufbau principle, Hund's rule and Pauli exclusion principle)
- Valence electrons

Gained knowledge/skills to practice

- How many electrons can occupy an orbital
- What the name of an orbital tells us
- How to fill electron orbitals for the first 20 elements
- Which electrons are an atoms valence electron?
- How many electrons occupy the 1st energy shell etc.

Orbital Energies and Atomic Structure



- First orbital at bottom of diagram is the orbital with electrons of lowest energy
- The energy increases as we move up to the 2s and then 2p and 3s, 3p
- Increasing the n value has more influence on energy than increasing the l value for small atoms
- Energy of orbitals increases within a shell in the order s < p < d < f
- This pattern doesn't hold for larger atoms
    - 3d orbital is higher energy than 4s orbital
    - These overlaps occur more frequently as we move up chart
- Filing order is based on observed experimental results – don't let 5d filing before 6p confuse you
- Within each shell, as value of *l* increases, electrons are less penetrating (less electron density found close to the nucleus) in order s > p > d > f
    - **Shielding: electrons that are closer to the nucleus slightly repel electrons that are further out**
        - These offsets the more dominant electron nucleus attractions slightly
        - Electrons in orbitals that experience more shielding are less stabilized and higher in energy

Electron Configuration

Electronic configuration = the arrangement of electrons in the orbitals of an atom

- Three key pieces for electronic configuration
    1. Number of the principal quantum shell *n*
    2. Letter that designates the orbital type (subshell, *l)*
    3. Superscript number that designates number of electrons in that subshell
        - i.e. 2*p*4
            - 4 electrons

- P subshell
- Principle quantum number of 2

Aufbau Principle

in the ground state of an atom or ion, electrons fill atomic orbitals of the lowest available energy levels before occupying higher levels.

- The filing order begins at hydrogen and includes each subshell as you proceed in increasing Z order

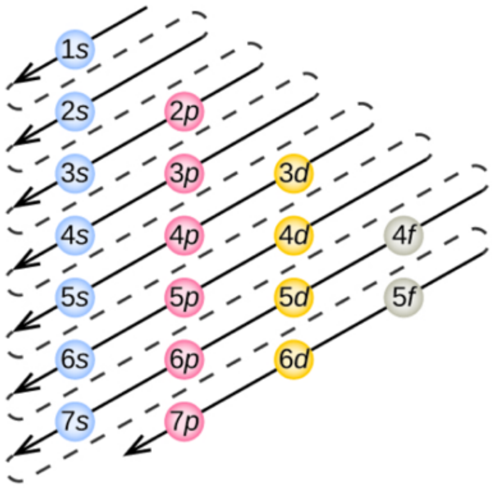


Figure 6.26 This diagram depicts the energy order for atomic orbitals and is useful for deriving ground-state electron configurations.

(see drawn notes for electron configuration)

- When we come to alkali metal potassium (atomic number 10), we expect that we might add electrons to 3d subshell
- All available chemical and physical evidence indicate that potassium is like lithium and sodium and that the next electron is not added to the 3d level but instead added to the 4s level
- 3d orbital with no radial nodes is higher in energy because it is less penetrating and more shielded from nucleus than 4s, which has 3 radial notes
    - Potassium thus has electron configuration of $[Ar]4s^1$
        - **Potassium corresponds to Li and Na in its valence shell configuration**
- Starting with transition metal scandium (21), additional electrons are added successively to the 3d subshell
    - 3d is filled to capacity with 10
    - 4p filled next
- Three series of elements, scandium (sc), through Copper (Cu), yttrium (T) through silver (Ag) and lutetium (Lu), through gold (Au), a total of 10d electrons are added to the (n-1 shell) to bring that (n-1 shell from 8 to 18 electrons) and so on
- Some exceptions to order of fillings exist

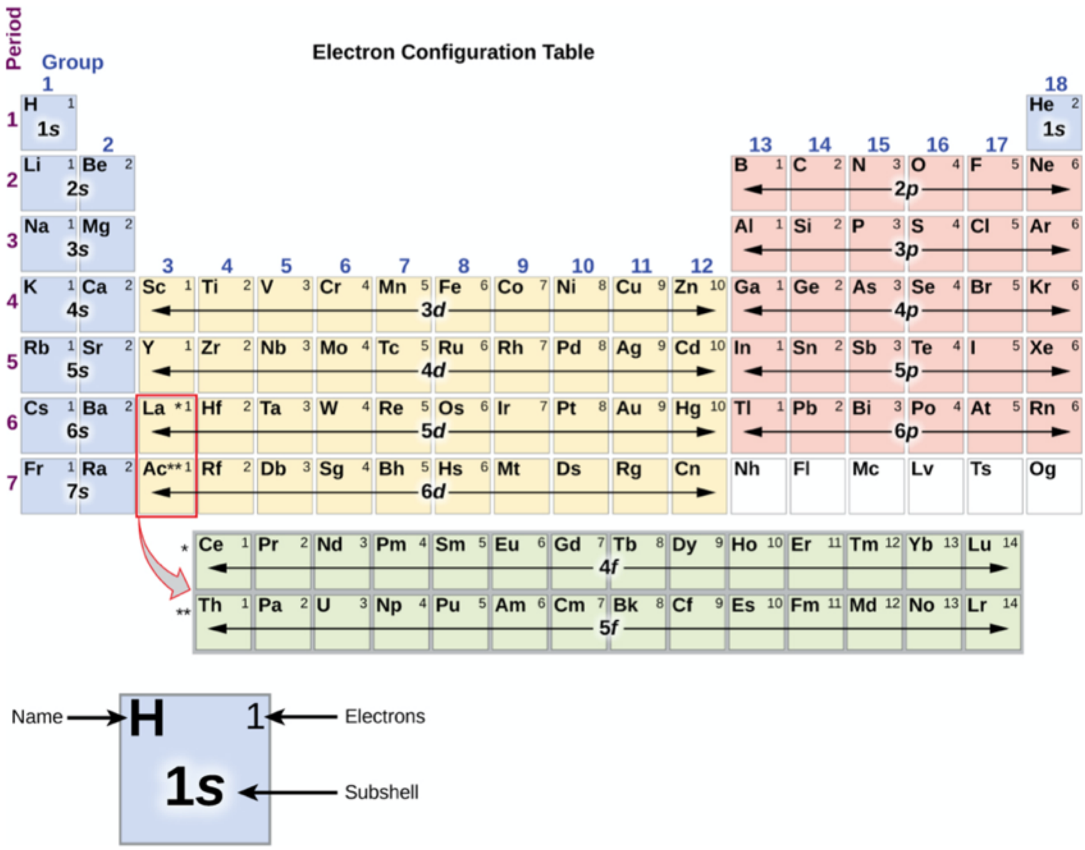


Figure 6.27 This periodic table shows the electron configuration for each subshell. By "building up" from hydrogen, this table can be used to determine the electron configuration for any atom on the periodic table.

- **-i.e. for chromium and copper**
  - Half-filled and completely filled subshells represent conditions of preferred stability
  - Electron shifts from 4s into 3d orbitals to gain extra stability of half-filled 3d subshell in Cr or filled 3d subshell in Cu
  - Other exceptions occur
  - No simple method to predict exceptions

Electron configuration and the periodic table

- Periodic table arranges atoms based on increasing atomic number so that elements with same chemical properties recur periodically
- Valence electrons play most important role in chemical reactions
  - Other electrons have highest energy
    - More easily lost or shared
    - Determining factor of some physical properties
- Elements in same column have same number of valence electrons
- Alkali metals (sodium and lithium) have one valence electron (1a)
- Alkaline earth metals beryllium and magnesium have two valence electrons (2)
- Halogens fluorine and chlorine have seven (7a)

Three main classifications (determine which orbitals are counted in valence shell or highest-level orbitals of an atom)

1. **Main Group Elements (blue and red)**
   - **a.** Last electron added enters an s or p orbital
   - **b.** Nonmetallic elements as well as many metals and metalloids
2. **Transition elements or transition metals**
   - a. Last electron enters the d orbital
   - b. But only partially filled d orbitals, thus, filled orbitals like An, Cd, Hg, Cu, Ag and Au are not technically transition elements (but term used frequently to refer to entire yellow block)
3. **Inner Transition Elements**
   - a. Metallic elements which last electron added occupies an f orbital
     - i. Two inner series

1. Lanthanide series
2. Actinide series

Electron Configurations of Ions

- Ions are formed when atoms gain or lose electrons
- Cation forms when one or more electrons are removed
  - Main group: electrons added last are first electrons removed
  - Transition and inner: electrons in s orbital are easier to remove than d or f and so highest ns electrons are lost and then the d or f electrons are removed
- Anion forms when one or more electrons are added (they fill in order predicted by Aufbau)

Variation in Covalent Radius

- There are several ways to define the radius of atoms and thus to determine their relative sizes
- **Covalent radius (one half the distance between nuclei of two identical atoms when they are joined by a covalent bond)**
- **N increases by one for each element**
- **Electrons being added to region of space that is increasingly distance from nucleus**
- **Size of atom must increase as we increase distance of outermost electrons from nucleus**

Topic 3: Part 1, The Periodic Table, A Chemical Roadmap

Periodic Law: *The properties of the elements are periodic functions of their atomic numbers*

- A modern periodic table arranges elements in increasing order of atomic numbers
- Groups atoms with similar properties in the same vertical column (groups are numbered 1-18)
  - **Main Group Elements:** Groups 1, 2, 13 - 8
  - **Transition Elements/Metals:** Groups 2-12
  - **Inner Transition Metals:** Two rows are bottom of table (top row elements are lanthanides and bottom row elements are actinides)

Metals/Metalloids/non-metals

- **Metals**
  - Shiny, malleable, good conductors of heat and electricity (yellow)
- **Non-Metals**
  - Nonmetals dull, poor conductors of heat and electricity (green)
- **Metalloids**
  - Conduct heat and electricity moderately well, and possess some properties of metals and some properties of non-metals (purple)

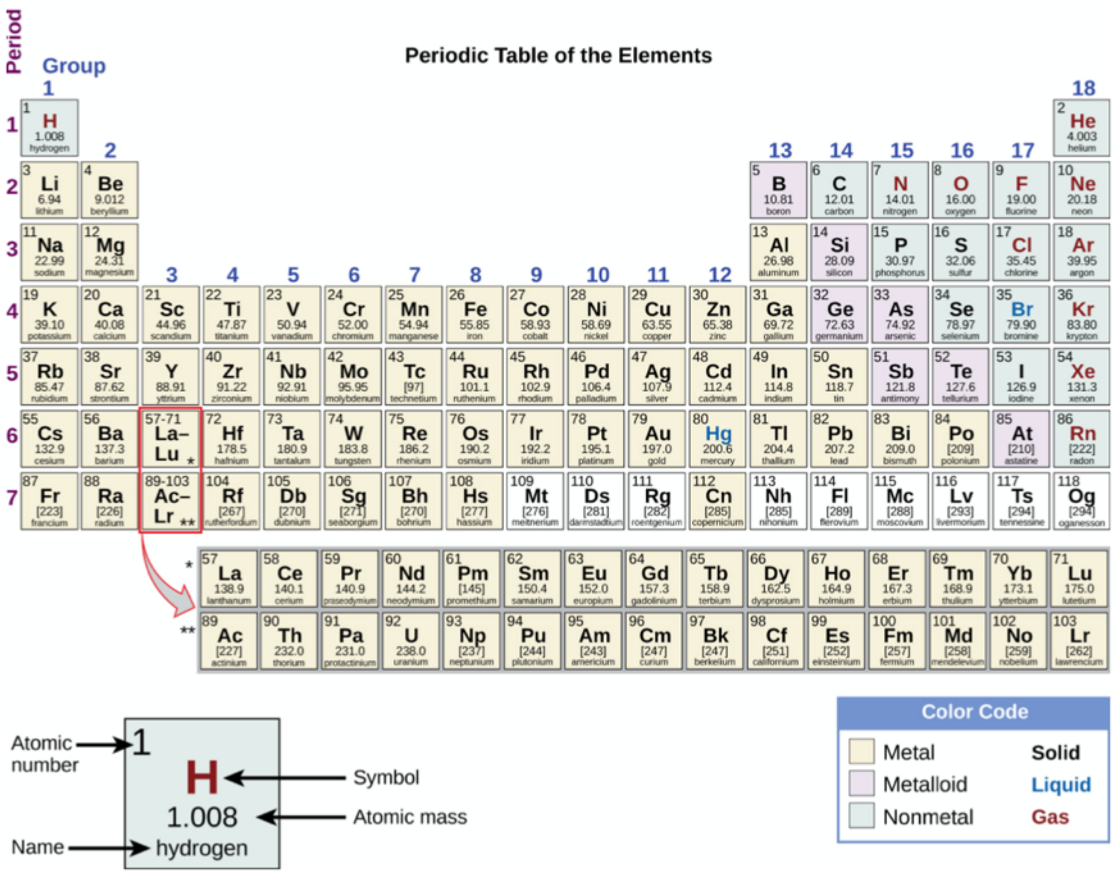


Figure 2.26 Elements in the periodic table are organized according to their properties.

Subdivisions

- **Main Group Elements:** Groups 1, 2, 13 - 8
- **Transition Elements/Metals:** Groups 2-12
- **Inner Transition Metals:** Two rows are bottom of table (top row elements are lanthanides and bottom row elements are actinides)

- **Alkali Metals (Group 1)**
  - Form compounds that consist of one atm of the element and 1 atom of hydrogen
    - All have similar properties
- **Alkaline Earth Metals (Group 2)**
  - Form compounds consisting of one atom of the element and two hydrogen
- **Pnictogens (group 15)**

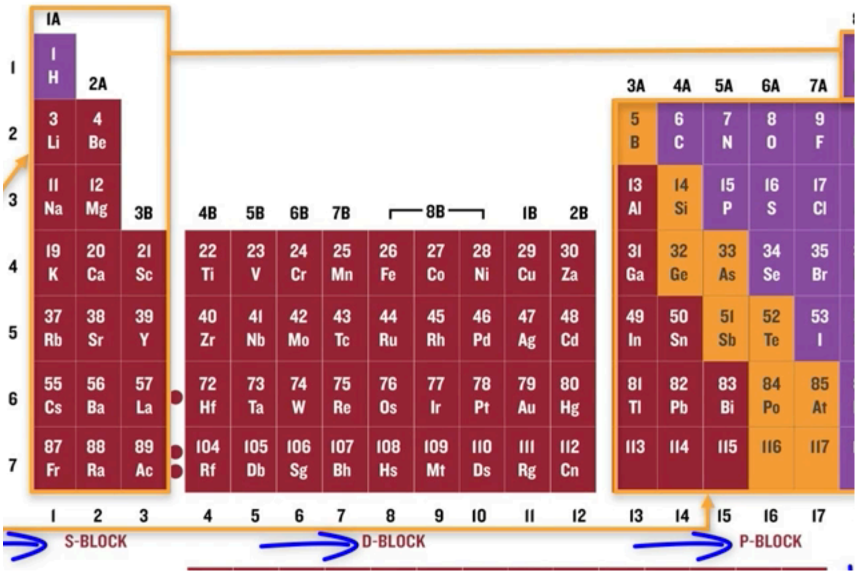
- **Chalcogens (16)**
  - Oxygen family
- **Halogens (17)**
- **Noble or Inert Gases (18)**
- **Hydrogen**
  - Unique, properties similar to group 1 and 17, so shown by itself or at top of both groups

Topic 3, Part 2: The Periodic Table and Electronic Configuration

Electron configuration and the periodic table

- Periodic table arranges atoms based on increasing atomic number so that elements with same chemical properties recur periodically
- Valence electrons play most important role in chemical reactions
  - Other electrons have highest energy
    - More easily lost or shared
    - Determining factor of some physical properties
- Elements in same column have same number of valence electrons
- Alkali metals (sodium and lithium) have one valence electron (1a)
- Alkaline earth metals beryllium and magnesium have two valence electrons (2)
- Halogens fluorine and chlorine have seven (7a)

Blocks and Electronic Structure

- Elements in each of the four blocks of the periodic table: s, p d or f have the same type of subshell as their highest energy shell
- The highest energy subshell of an element in the s-block is an s-subshell
- The highest energy subshell of an element in the. Block is a p-subshell
- 

USING THE PERIODIC TABLE

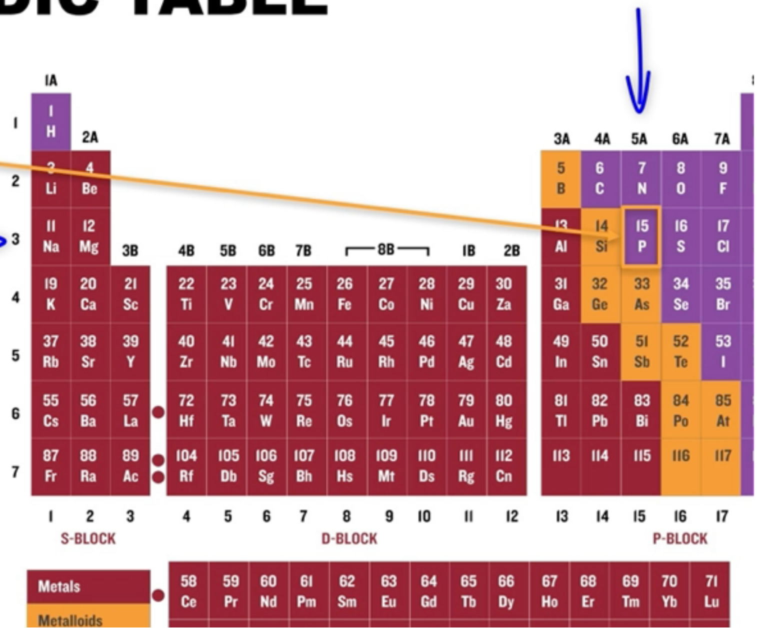
The element **phosphorus**

Atomic number 15: 15 protons

Group 5A: has 5 valence electrons

Period 3: has electrons in its 1^{st}, 2^{nd} and 3^{rd} energy shells

***p*-block:** highest energy orbital that contains electrons is *p*-orbital



Topic 4: Part 1 The Nature of the Chemical Bond

Ions/Cations/Anions

- Ions are atoms that have an electrical charge
- A cation is a positive ion
    - Forms when neutral atom loses one or more electrons from valence shell
- Anion is a negative ion
    - Forms when neutral atom gains electron

The Chemical Bond

- Most substances consist of more than one atom
- Atoms are combined by chemical bonds
- Pair of atoms bond to form diatomic molecules
- Two or more atoms combined are polyatomic molecules
- Many atoms are lattice
- Chemical bond = electrostatic forces – negatively charged electrons of one species attracted to positively charged protons of another
- TYPE of bonding has effect on property of chemical
- Why do bonds occur?
    - Atoms like being in low energy state
    - Low energy = more stability
    - Form bonds if energy of newly formed aggregate is lower than separate atoms

Metallic Bonding

- Many metal atoms pack together to form aggregate that is lower energy than single atom
- Bonding model explains
    - Good conductors of heat
    - Malleable and ductile
    - High boiling point
- Metal cations surrounded by sea of mobile electrons
- Bonds form through attraction of electrons and positive nuclei
- Electrons free to move and conduct heat and electricity
- Metals can be shaped as local bonds can be easily broken and reformed

Octet Rule

There is stability in having a full outer shell of Electrons
Octet Rule: Elements tend to react in such a way that results in atoms having a FULL OUTER SHELL, often meaning 8.

-Atoms can obtain a full shell via:

1) Transfer of electrons (donate or accept electrons) or IONIC BONDING

2) Sharing of electrons COVALENT BONDING

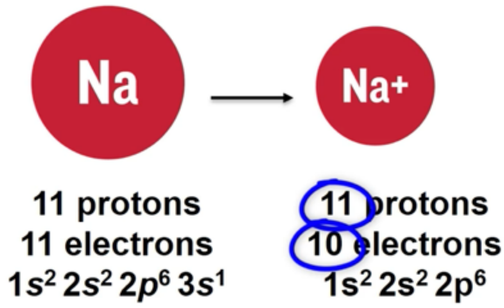
Topic 4: Part 2: Ionic bonding

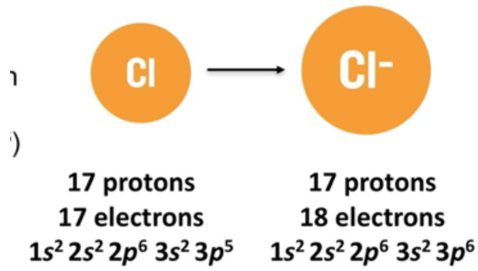
Ionic Compounds

- Exhibit crystalline structure
- Rigid and brittle
- High melting and boiling point
  - Ionic bonds are very strong
- Poor conductors of electricity
- Most dissolve readily in water

Atoms can achieve full outer shell through **transfer of electrons meaning atoms can donate or accept electrons, the ions that are formed have a full outer shell – this is IONIC BONDING**

- **Oppositely charged ion are attracted**
- **This occurs between a metal and a non-metal ion**
- **NaCl**
  - Metal atom loses electron
  - Non-metal electron gain electron
  - Na
    - $1s^{2,}\ 2s^{2,}\ 2p^{6,}\ 3s^1$
      - Loses ONE electron so that outer shell is full!
      - This creates Na^+
  - Cl
    - $1s^{2,}\ 2s^{2,}\ 2p^{6,}\ 3s^{1,}\ 3p^5$
    - $1s^{2,}\ 2s^{2,}\ 2p^{6,}\ 3s^{1,}\ 3p^6$
      - **Only needs to gain ONE electron to fill electron shell**
      - This creates Cl^-
  - Positively charged Na^+ Cation is electrostatically attracted to the oppositely charged Cl^- anion that this creates the IONIC BOND





Model

- When cations and anions come together a large number of each ion combine to form a 3-d lattice
  - Cations surrounded by anions
  - Anions are surrounded by cations
- It would be incorrect to refer to NaCl as a molecule because there is not a single ionic bond between any specific pair of sodium and chloride ions
- The attractive forces are isotropic meaning they are the same in all directions – any particular ion is equally attracted to all nearby ions
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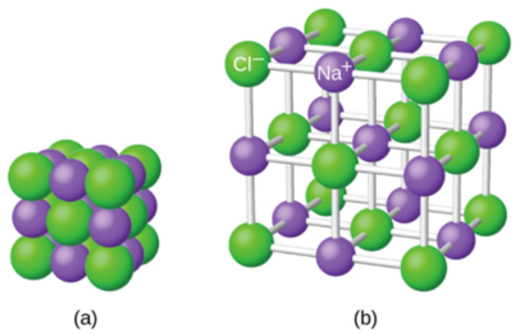

Figure 7.3 The atoms in sodium chloride (common table salt) are arranged to (a) maximize opposite charges interacting. The smaller spheres represent sodium ions, the larger ones represent chloride ions. In the expanded view (b), the geometry can be seen more clearly. Note that each ion is "bonded" to all of the surrounding ions—six in this case.

Electrical neutrality

- Ionic compounds are electrically neutral
  - Total number of positive charge equals negative charge
- Example sodium oxide
  - Sodium ion has +1 charge ($Na^{+)}$
  - Oxygen ion has -2 charge ($O^{2-)}$
    - To balance, two Na_ ions are required, and the ratio is now 2:1

Electronic Structure of Cations

- When forming cation, atom of main group element tends to lose all its valence electrons *thus assuming electronic structure of noble gas that precedes it in periodic table*
- Group 1 (alkali metals) and group 2 (alkaline earth metals) group numbers equal to number of valence shell electrons and to charges of cations formed from atoms of these elements when all valence shell electrons removed
    - Ca (group 2)
        - Neutral atoms have 20 electrons
            - $1s\ 2s\ 2p\ 3s\ 3p\ 4s2$
            - Lose both valence electrons it has a 2+ charge
                - $1s\ 2s\ 2p\ 3s\ 3p$.
                    - Isoelectronic with noble gas Ar
- Group 13-17 – group numbers exceed number of valence electrons by 10
    - Cation formed by loss of all valence electrons is equal to the group number minus 10
        - Aluminium (13) forms 3+ ions Al3+
- Some exceptions exist whereby there is a partial loss of atoms valence shell (inert pair effect)
- Transition and inner transition metals behave differently
    - Most have 2+ or 3+ charges that result from loss of outermost s electrons first sometimes followed by one or two d electrons from next to outer shell
        - I.e. Iron $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$ forms Fe2+ by loss of 4s electrons $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$
        - Iron also forms Fe3+ by loss of 4s electrons and one 3d electrons $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$
    - Although d orbitals according to Aufbau are the last to fill, the outermost s electrons are the first to be lost when ionization occurs
    - When inner transition metals form ions they usually have a 3+ charge from loss of outermost s electrons and a d or f electron

Electronic Structure of Anions

- Most monatomic anion form when neutral non-metal atom gains enough electrons to fill its outer s and p orbitals reaching the electron configuration of the next noble gas
- Charge is equal to number of electrons that must be gained to fill the s and p orbitals of the parent atom
    - i.e. Oxygen has 1s22s22p4 whereas oxygen anion has electron configuration of neon (ne) 1s22s22p6
    - This is 02-

Topic 4: Part 3: Covalent Bonding

Sharing of Electrons is covalent bonding!

- Covalent bonds typically occur between non-metal atoms only
- Two atoms share:
    - One pair of electrons – single bond
    - Two pairs of electrons – double bond
    - Three pairs of electrons – triple bond
- Energy is needed to break covalent bonds; conversely the same amount of energy is realised when one covalent bond molecule is formed