Atoms, Ions & Molecules

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Atomic Theory

- John Dalton (1766-1844) is the scientist credited for proposing the atomic theory.
- This theory explains several concepts that are relevant in the observable world:
  - The composition of a pure gold necklace,
  - what makes the pure gold necklace different than a pure silver necklace, and
  - what occurs when pure gold is mixed with pure copper.
- Before discussing the atomic theory, this article explains the theories that Dalton used as a basis for his theory:
  - the law of conservation of mass and
  - the law of constant composition

John Dalton

John Dalton (6 September 1766 – 27 July 1844) was an English chemist, meteorologist and physicist. He is best known for his pioneering work in the development of modern atomic theory.

He was the first person to explain that everything that exists is made from atoms.

Law of Conservation of Mass

- ***Mass is neither created nor destroyed in ordinary chemical and physical changes***
- **Must start and end with the same amount**
- **<u>Example</u>**

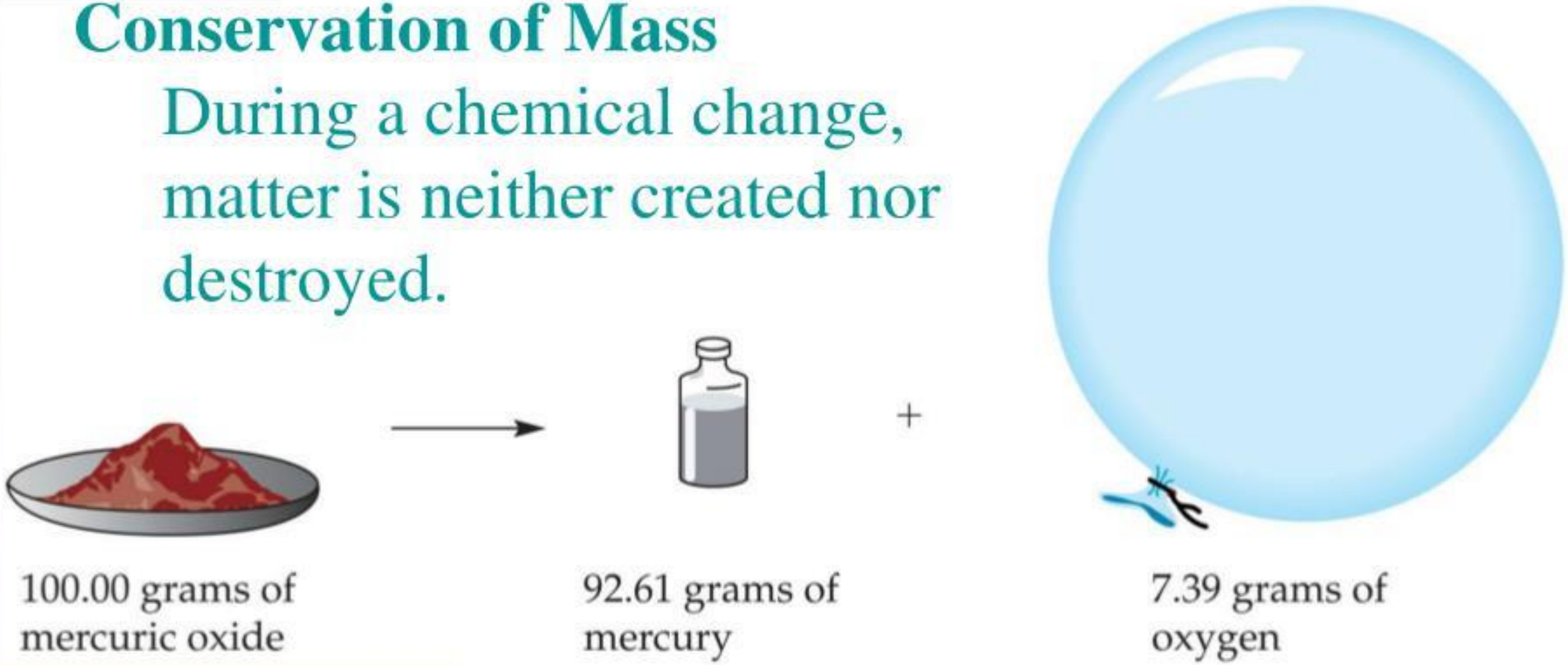



Log	**+**	**Fire**	**→**	**Ashes**	**+**	**Smoke**
30 kg		**1 kg**		**28 kg**		**3 kg**

Lavoisier: The Law of Conservation of Mass

Early 1700's Lavoisier: Law of Conservation of Mass

During a chemical change, matter is neither created nor destroyed.



Law of Definite Proportions

Proust's Law or the Law of Constant Composition

Samples of a given compound always contain the same proportion of elements by mass.

carbon dioxide
CO_2

carbon: 12 g/mol
oxygen: 16 g/mol x 2 = 32 g/mol

12/44 = 0.273 C
32/44 = 0.7272 O

27.3% carbon by mass
72.7% oxygen by mass

1 g CO_2
0.273 g O
0.727 g C

16 g CO_2
4.364 g C
11.64 g O

Law of Multiple Proportions

When two elements form a series of compounds, the ratio of the masses of the 2nd element that combine with 1 gram of the first element can always be reduced to small whole numbers.

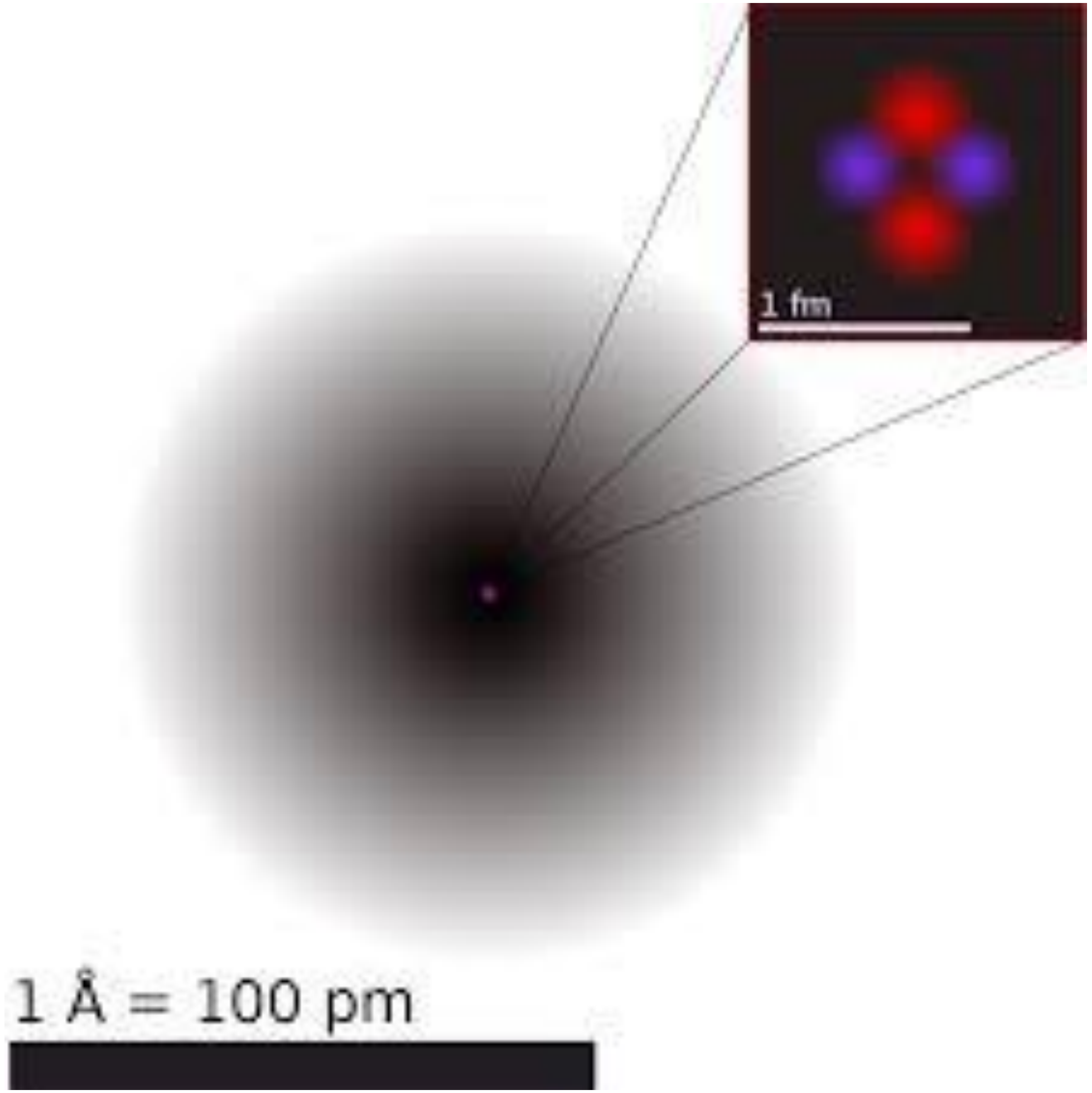
$\boxed{CO}$ $\qquad \frac{2.67}{1.33} = \qquad$ $\boxed{CO_2}$

$12g\ C = 16g\ O \qquad\qquad 12g\ C = 32g\ O$

$1g\ C = \boxed{1.33g\ O} \qquad\qquad 1g\ C = \boxed{2.67g\ O}$

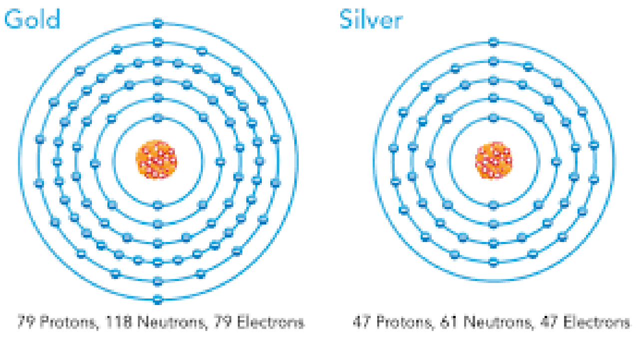
Dalton's Atomic Theory

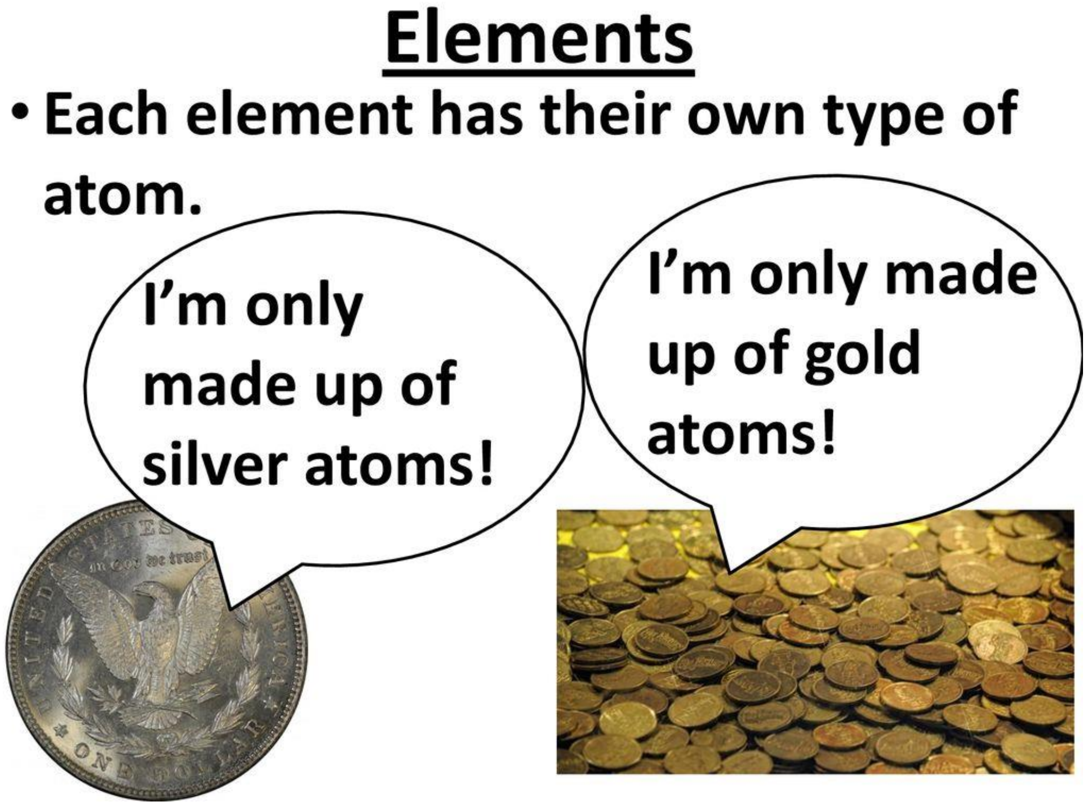
- Each chemical element is composed of tiny indivisible particles that cannot be seen by the unaided eye, called atoms.
- Atoms can neither be created nor destroyed.
- Pictured is a helium atom.
- The purple and red dots represent the neutrons and protons in the nucleus.
- The black area around the nucleus represent the electron cloud.



Dalton's Atomic Theory

- All atoms of an element have the same mass other properties, but the atoms of another element differ from all other elements.
- For example, gold and silver have different atomic masses and different properties.





Dalton's Atomic Theory

- In a compound, different elements combine in a simple numerical ratio.
- The illustration below describes this rule.
- The second equation for the reaction is incorrect because half of an atom does not exist

One Carbon atom + Four Hydrogen atoms → CH_4

~~½ Carbon atom + Two Hydrogen atoms → $C_{1/2}H_2$~~

Dalton's Atomic Theory

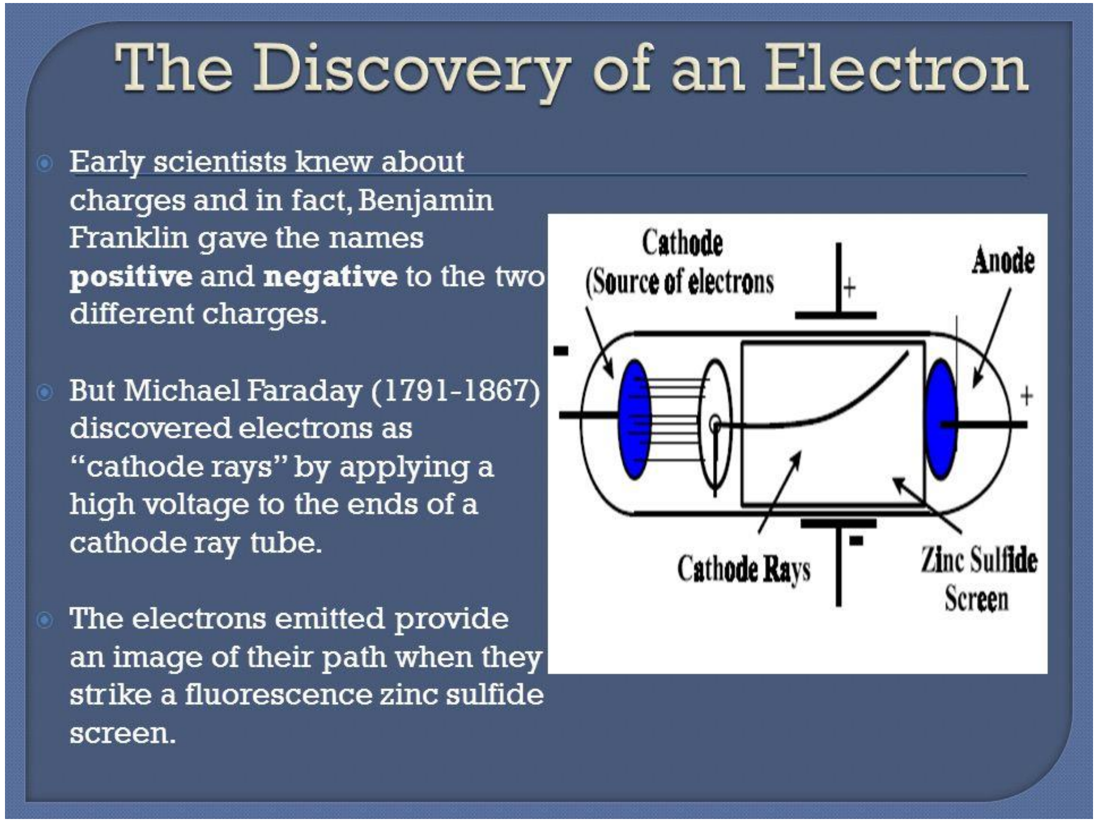
- Atoms can neither be created nor destroyed.
- They just rearrange and/or combine forming a different compound or substance(s).
- When the elements copper (a shiny, red-brown solid) and oxygen (a clear and colorless gas) react, their atoms rearrange to form a compound containing copper and oxygen (a powdery, black solid).

Dalton's Atomic Theory

1. Elements are made of tiny particles called **atoms.**
2. All atoms of a given element are identical.
3. The atoms of a given element are different from those of any other element.
4. Atoms of one element can combine with atoms of other elements to form compounds. A given compound always has the same relative numbers and types of atoms.
5. Atoms are indivisible in chemical processes. That is, atoms are not created or destroyed in chemical reactions. A chemical reaction simply changes the way the atoms are grouped together.

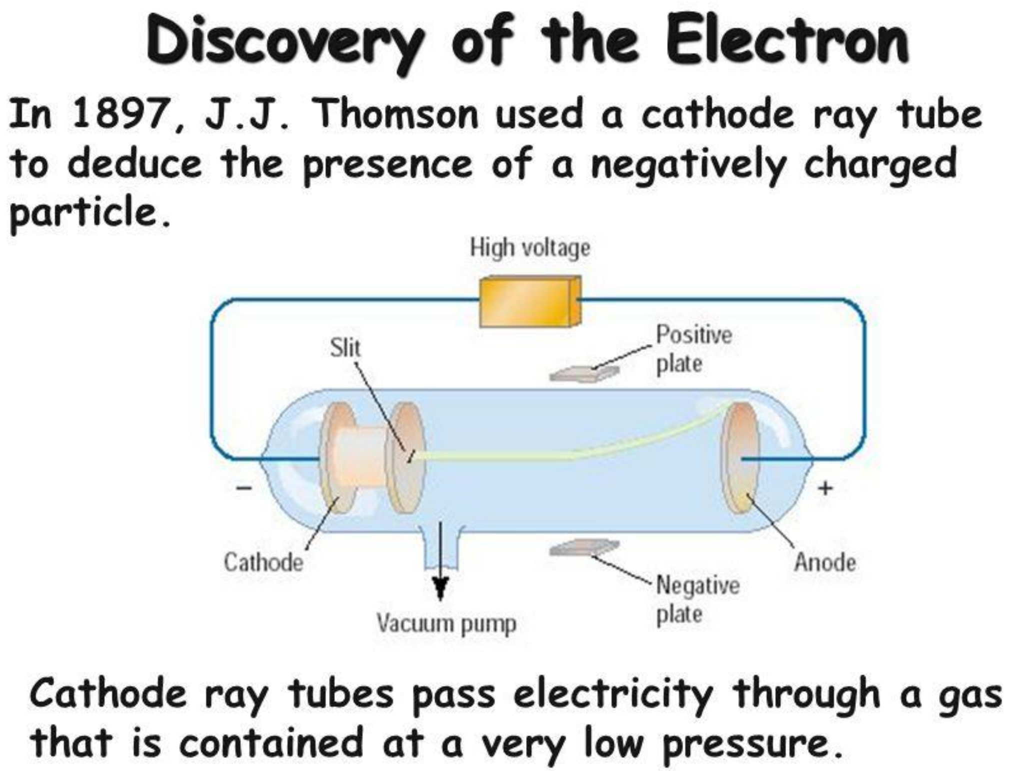
Discovering Electrons

- Michael Faraday's (1791-1867) invention of the first cathode-ray tube (CRT) led to the discovery of the type of radiation emitted by the negative terminal, which is the cathode, upon the passage of electricity.
- The observed radiation crosses the vacuum glass tube to the positive terminal which is the anode.
- The cathode ray tube produced invisible rays that can only be detected by the light emitted by the materials they struck called phosphors.



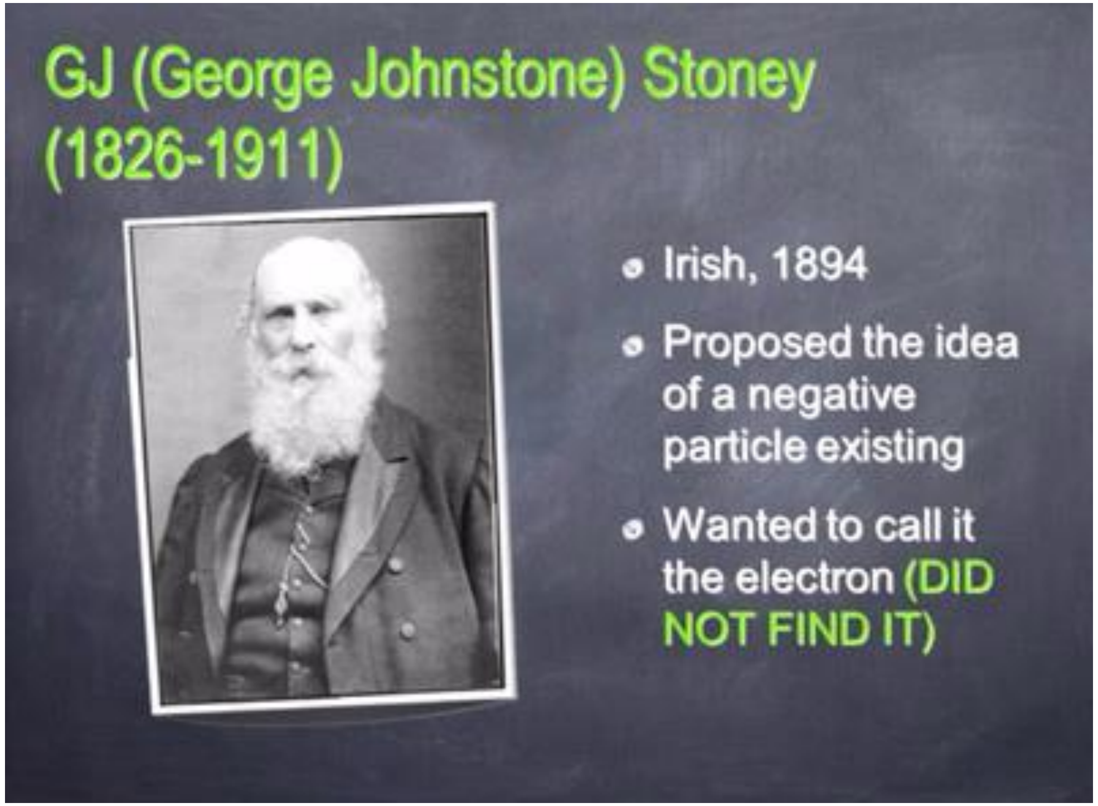
Discovering Electrons

- These cathode rays are influenced by magnetic and electric fields that are deflected similar to the negatively charged materials which prompted J. J. Thompson (1856 – 1940) to conclude cathodes are found in all atoms as the negatively charged particles.



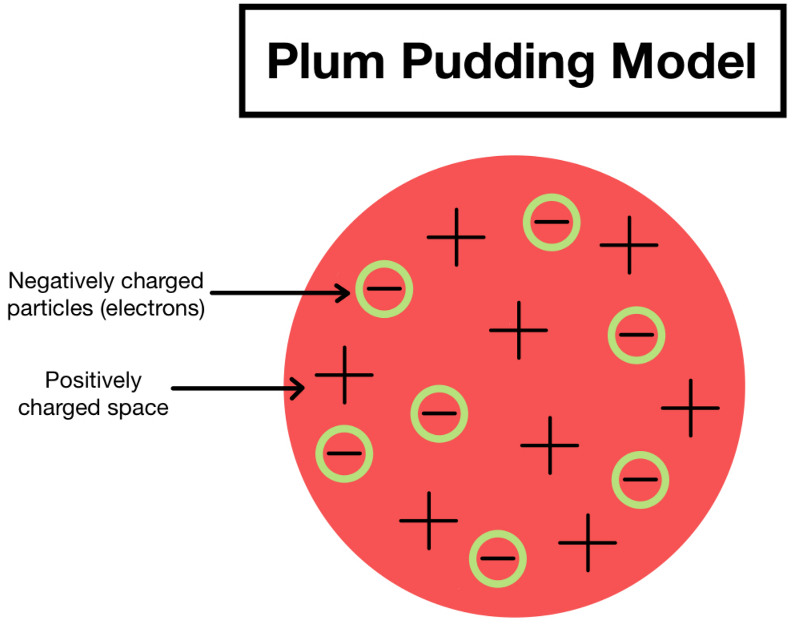
Discovering Electrons

- George Stoney first to give the name electrons referring to the cathode rays



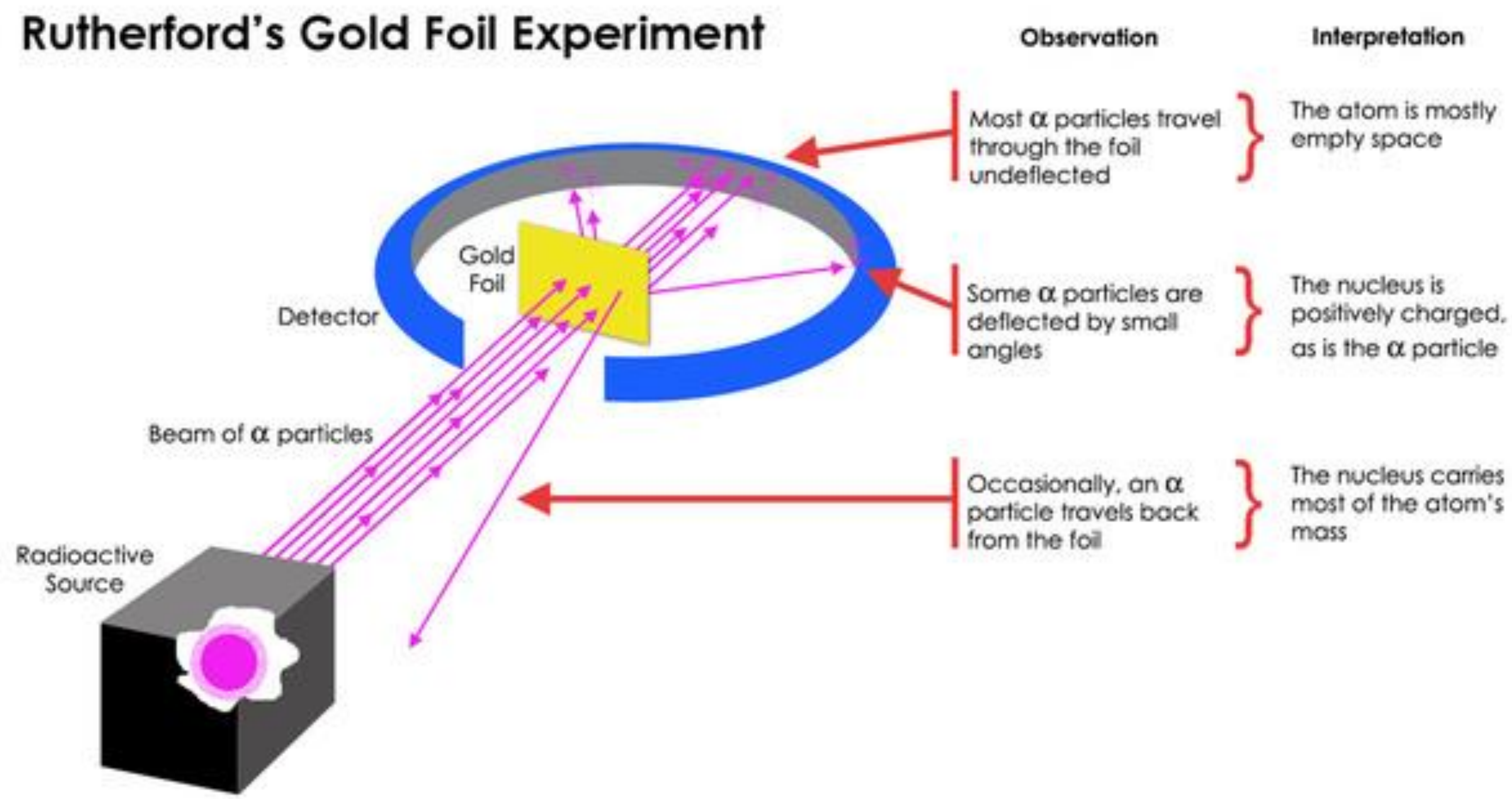
The Plum Pudding Model

- After the discovery of electrons, Thompson, proposed the plum pudding model of an atom, which shows the electrons floating in a positively-charged material.
- This model was named after the plum-pudding dessert.



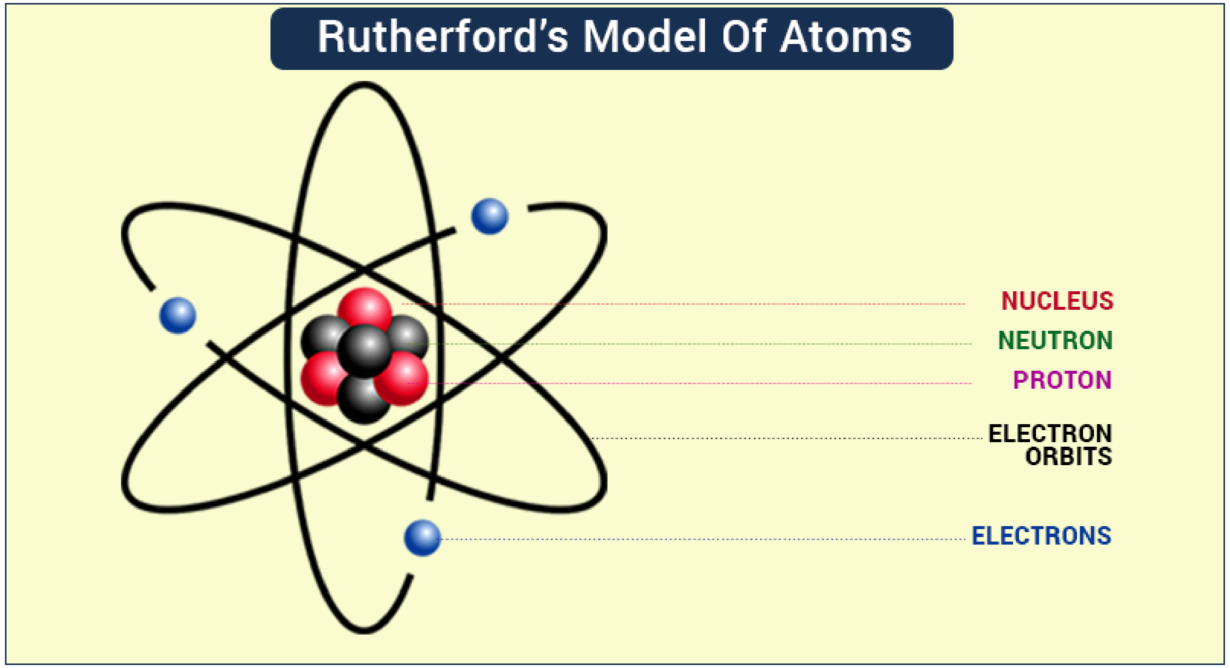
The **Plum Pudding Model** was a model of an atom created by J.J Thompson. The model describes the atom as negatively charged particles swimming in a positively charged sea. It got it's name because the model resembles the look of plum pudding

Discovery of the Proton



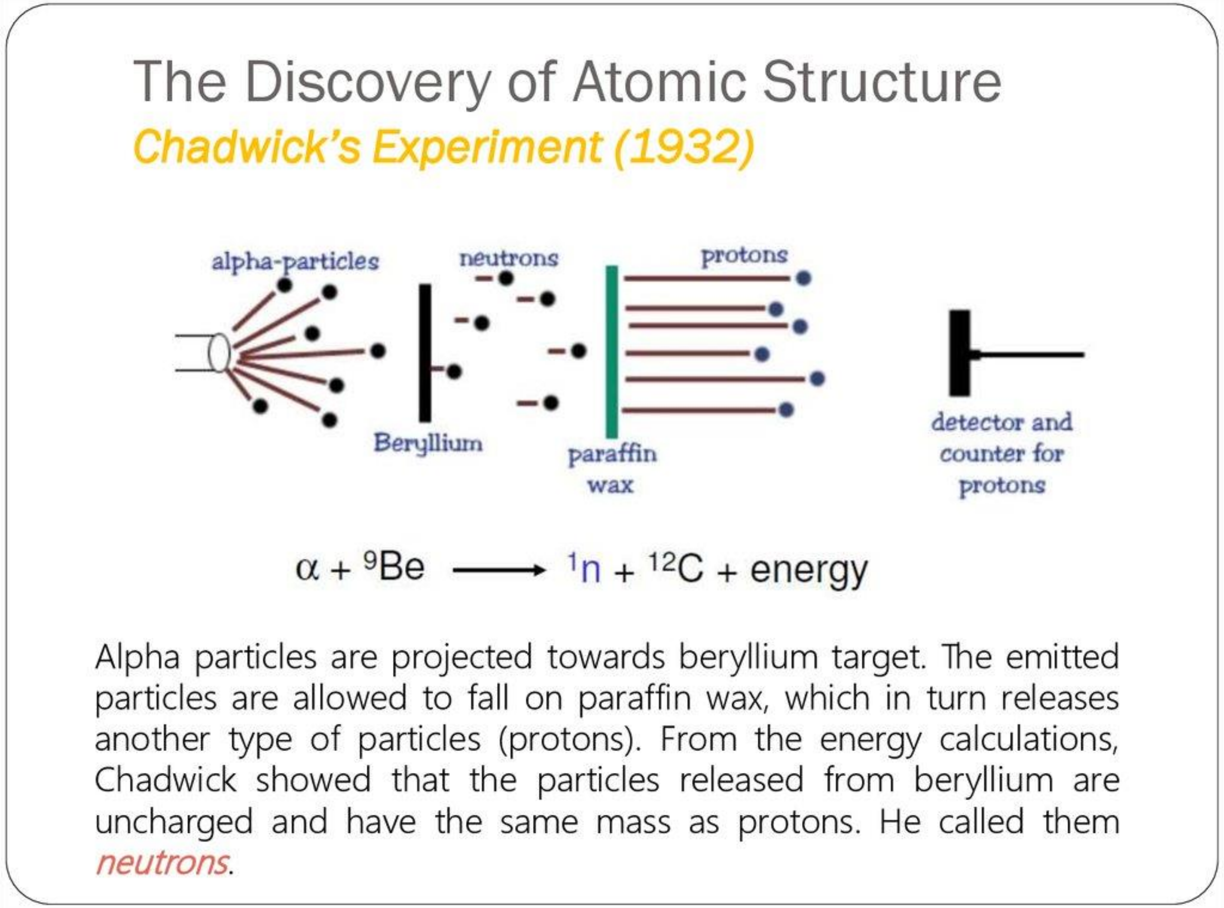
Discovery of the Proton

- To account for these observations, Rutherford devised a model called the nuclear atom.
- In this model, the positive charge is held in an extremely small area called the nucleus, located in the middle of the atom.
- Outside of the nucleus the atom is largely composed of empty space.
- This model states that there were positive particles within the nucleus, but failed to define what these particles are.
- Rutherford discovered these particles in 1919, when he conducted an experiment that scattered alpha particles against nitrogen atoms.
- When the alpha particles and nitrogen atoms collided protons were released.

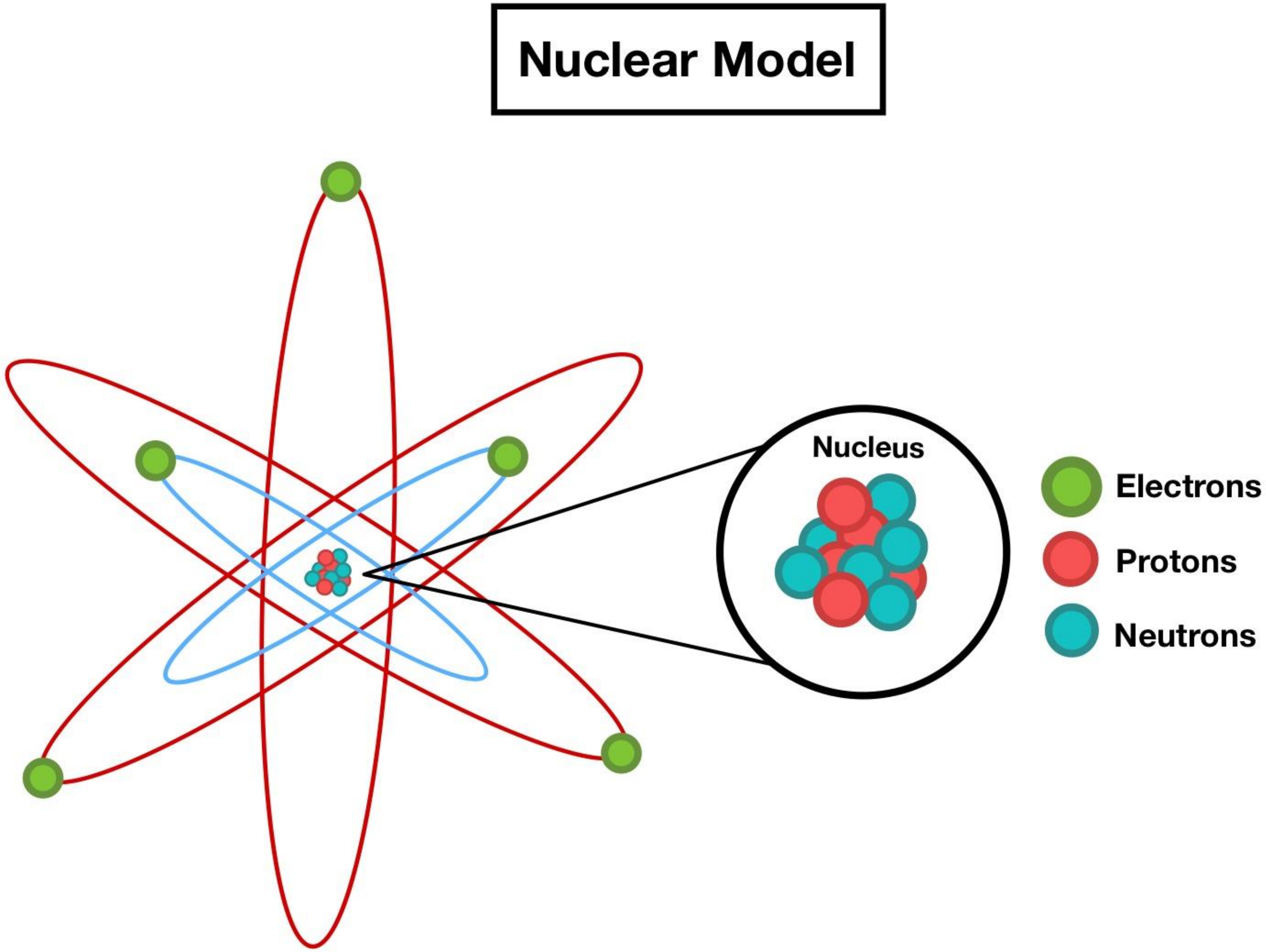


The Discovery of the Neutron

- In 1933, James Chadwick (1891-1974) discovered a new type of radiation that consisted of neutral particles.
- It was discovered that these neutral atoms come from the nucleus of the atom.
- This last discovery completed the atomic model.



The Discovery of the Neutron



Example

- When 32.0 g of methane are burned in 128.0 g of oxygen, 88.0 g of carbon dioxide and 72.0 g of water are produced. Which law is this an example of?
  a) Law of Definite Proportions
  b) Law of Conservation of Mass or
  c) Law of Multiple Proportions.

Solution

Answer: (b) Law of Conservation of Mass

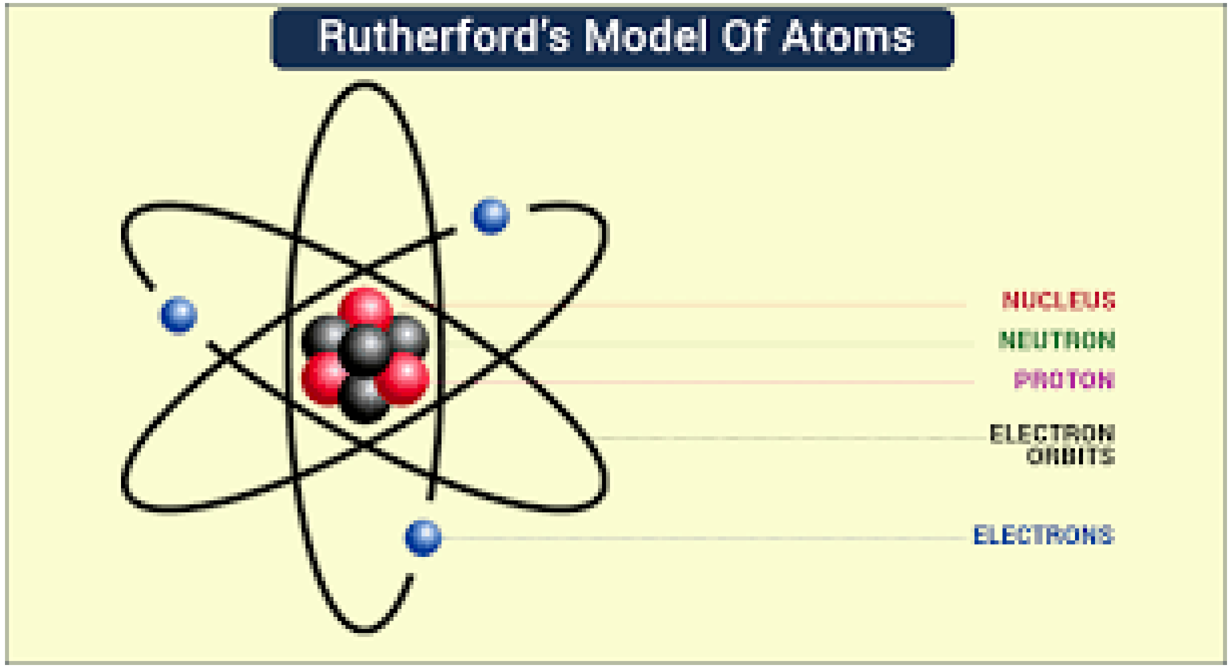
- Reactants: 32.0 g methane + 128.0 g oxygen = 160.0 g
- Products: 88.0 g carbon dioxide + 72.0 g water = 160.0 g

Table 1: Charge and mass of three sub atomic particles

Particle	Charge	Mass (grams)
Electrons	–1	9.1094×10^{-28}
Protons	+1	1.6726×10^{-24}
Neutrons	0	1.6749×10^{-24}

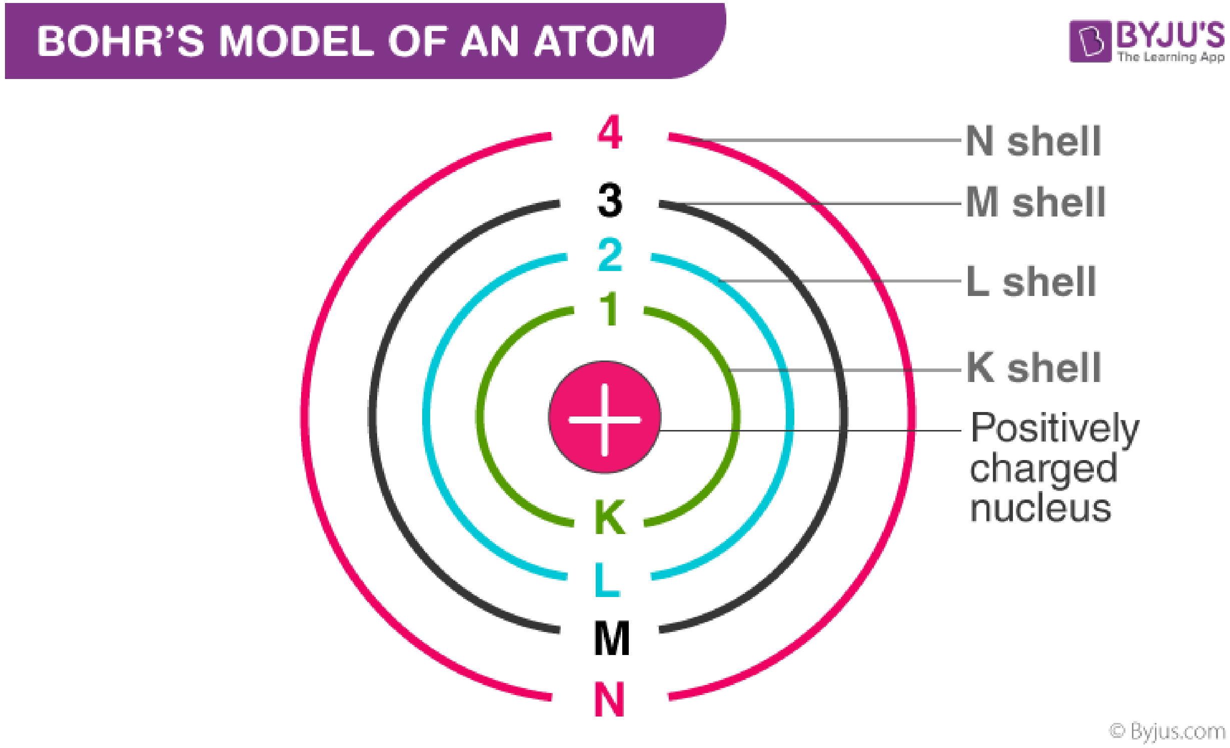
Atom

- The atom is the smallest particle of matter consisting of three sub- atomic particles:
  1. Proton
  2. Neutron
  3. Electron
- Protons with neutron are inside the nucleus which carries most of the mass and is positively charged .
- The electrons are found in an electron cloud surrounding the nucleus.
- Rutherford’s model of the atom was better than earlier models.



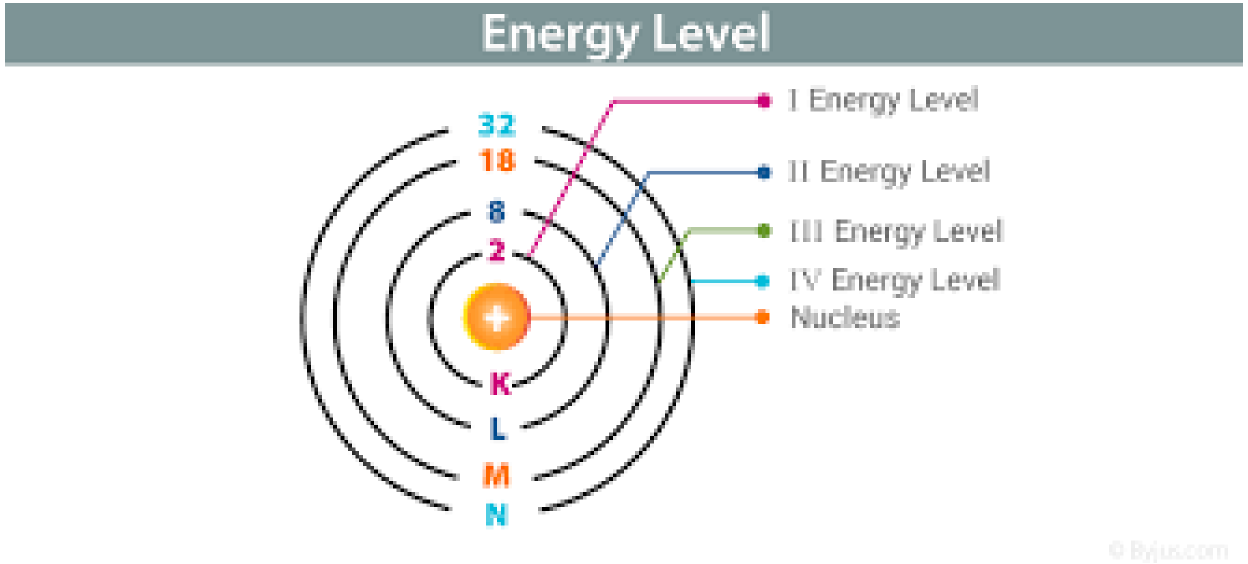
Bohr's Model of the Atom

- Bohr’s research focused on electrons.
- In 1913, he discovered evidence that the orbits of electrons are located at fixed distances from the nucleus.
- Remember, Rutherford thought that electrons orbit the nucleus at random.
- In Bohr’s atomic model, electrons orbit at fixed distances from the nucleus.
- These distances are called energy levels.



Energy Levels

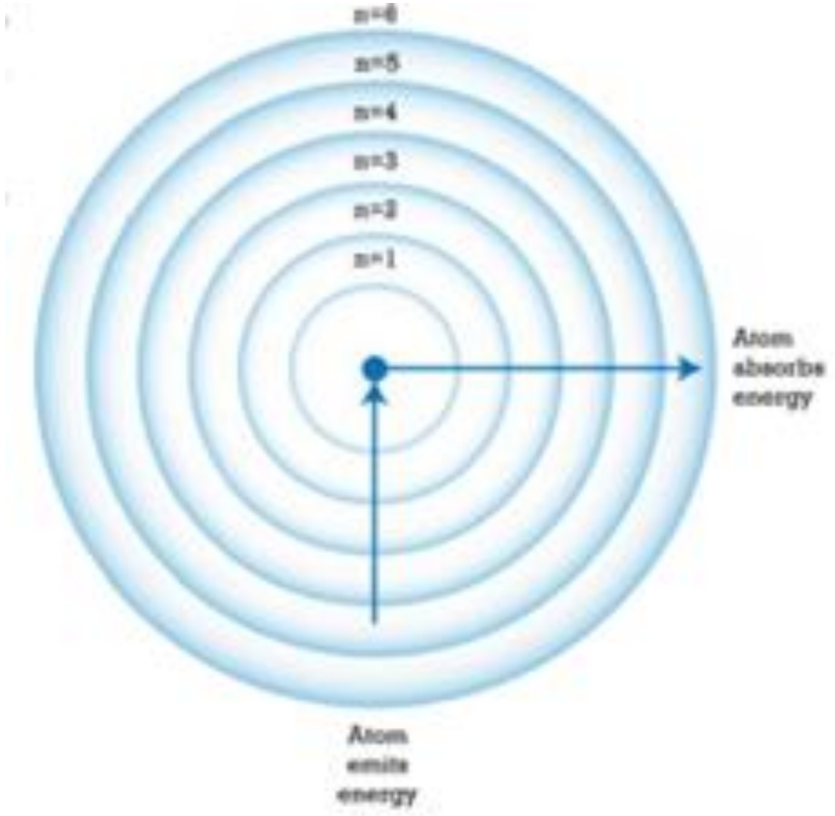
- Founded on the idea on Bohr's model, are the energy levels.
- **Energy levels** are regions around the nucleus at fixed distances where electrons are found.
- They are the only places where electrons can be found.
- Energy levels are a little like rungs on a ladder.
- You can stand on one rung or another but not between the rungs.
- The same goes for electrons.
- They can occupy one energy level or another but not the space between energy levels.



Energy Levels

- Starting from the level with the least energy is the one closest to the nucleus.
- Going farther away from the nucleus, the energy in each level increases.
- Electrons can jump from a lower energy level to a higher energy level and become excited
- As you go farther from the nucleus, the levels have more and more energy.
- Electrons in the lowest energy level are in their **ground state** but they can jump from one energy level to another.
- As electrons jump from a lower energy level, energy is absorbed and thus turned into their **excited state**.
- As the electrons jump back to the lower state, energy is emitted or given off.

- This model of an atom contains six energy levels (n = 1 to 6). Atoms absorb or emit energy when some of their electrons jump to a different energy level.

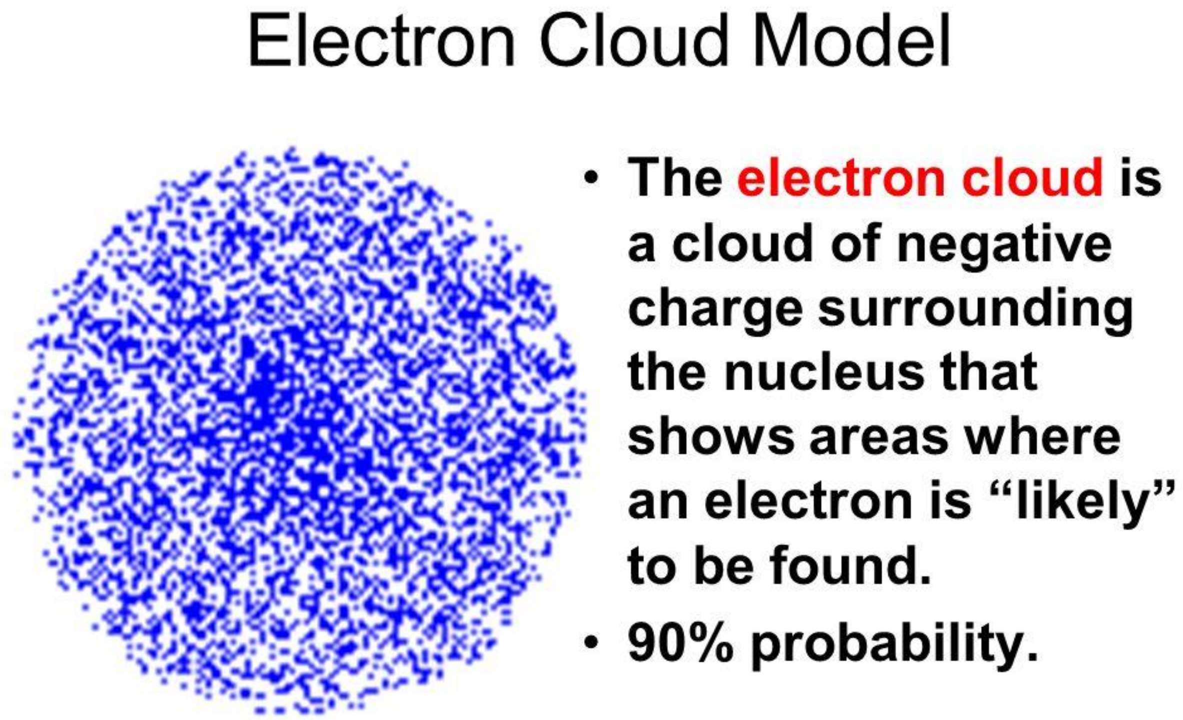


Energy Levels

- Seen a firework?
- When fireworks explode, the atoms in the chemicals composing the fireworks, absorbed energy and so their electrons.
- Some of the electrons jump to a higher energy level as they absorb energy.
- As they go back to their original energy levels, they give off energy in the form of light.
- Different atoms have different arrangements of electrons that accounts for the different colors of light given off.

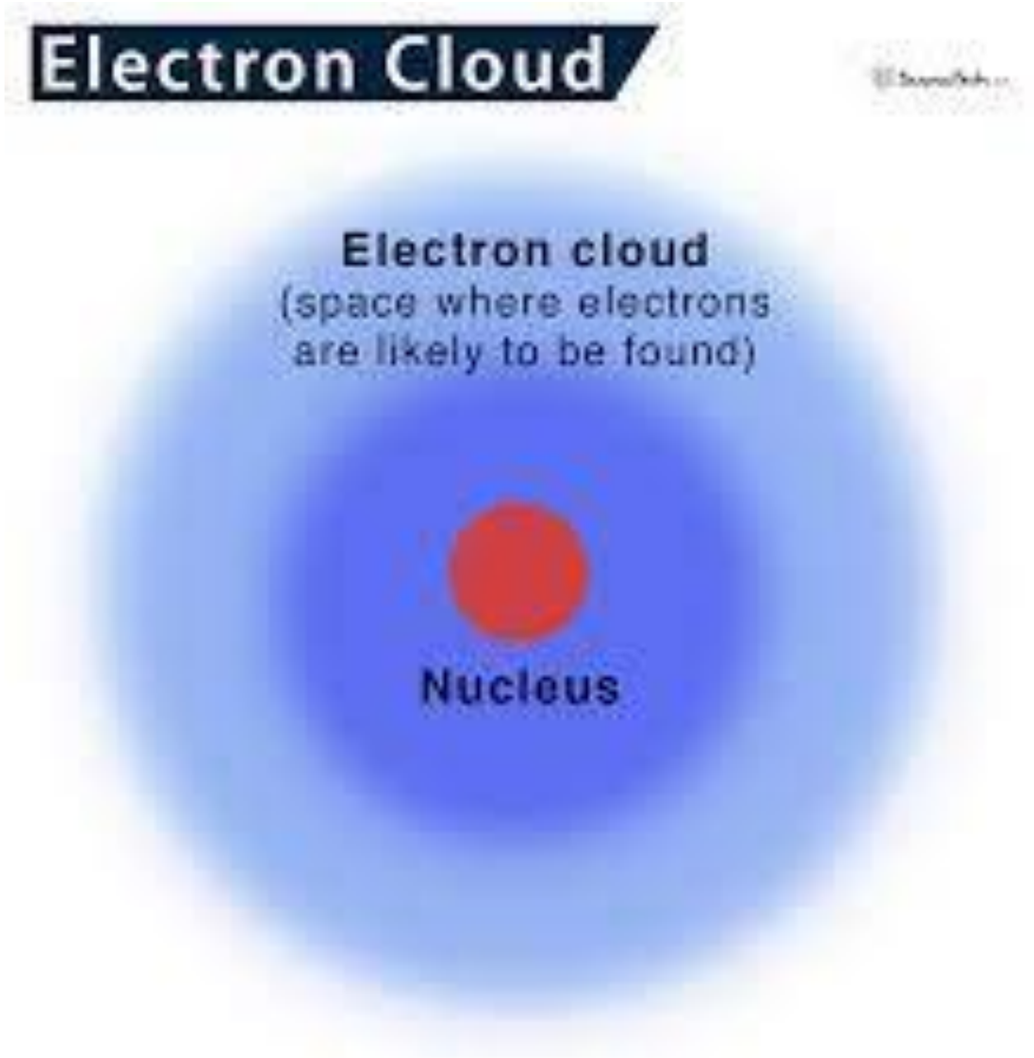
Electron Cloud Model and the Orbitals

- Today, these ideas about electrons are represented by the electron cloud model.
- The electron cloud is an area around the nucleus where electrons are likely to be in an orbital.



Electron Cloud

- The electron cloud visualizes the location of electrons in the nucleus of an atom.
- These electrons and protons are held in place by the **electromagnetic force** in the nucleus.
- Electrons move around the nucleus, creating the cloud; however, they actually form **electron shells**.
- Each electron shell can hold a certain amount of electrons.
- The first shell holds 2 electrons, and subsequent shells hold 8 electrons.
- The number of electrons in a given atom is equal to the number of protons in that atom.
- However, units of the same element can have different numbers of electrons.



Orbitals

- Some regions of the electron cloud are denser than others.
- These are the regions where the probability of finding the electrons is high called **orbitals**.
- The first orbital has a maximum of two electrons.
- Different energy levels in the cloud have different numbers of orbitals.
- Therefore, different energy levels have different maximum numbers of electrons.
- Table 2.1 lists the number of orbitals with their corresponding number of electrons for the first four energy levels.
- The more distance away from the nucleus, the energy levels farther from the nucleus have more orbitals.
- Therefore, these levels can hold more electrons.

Orbitals

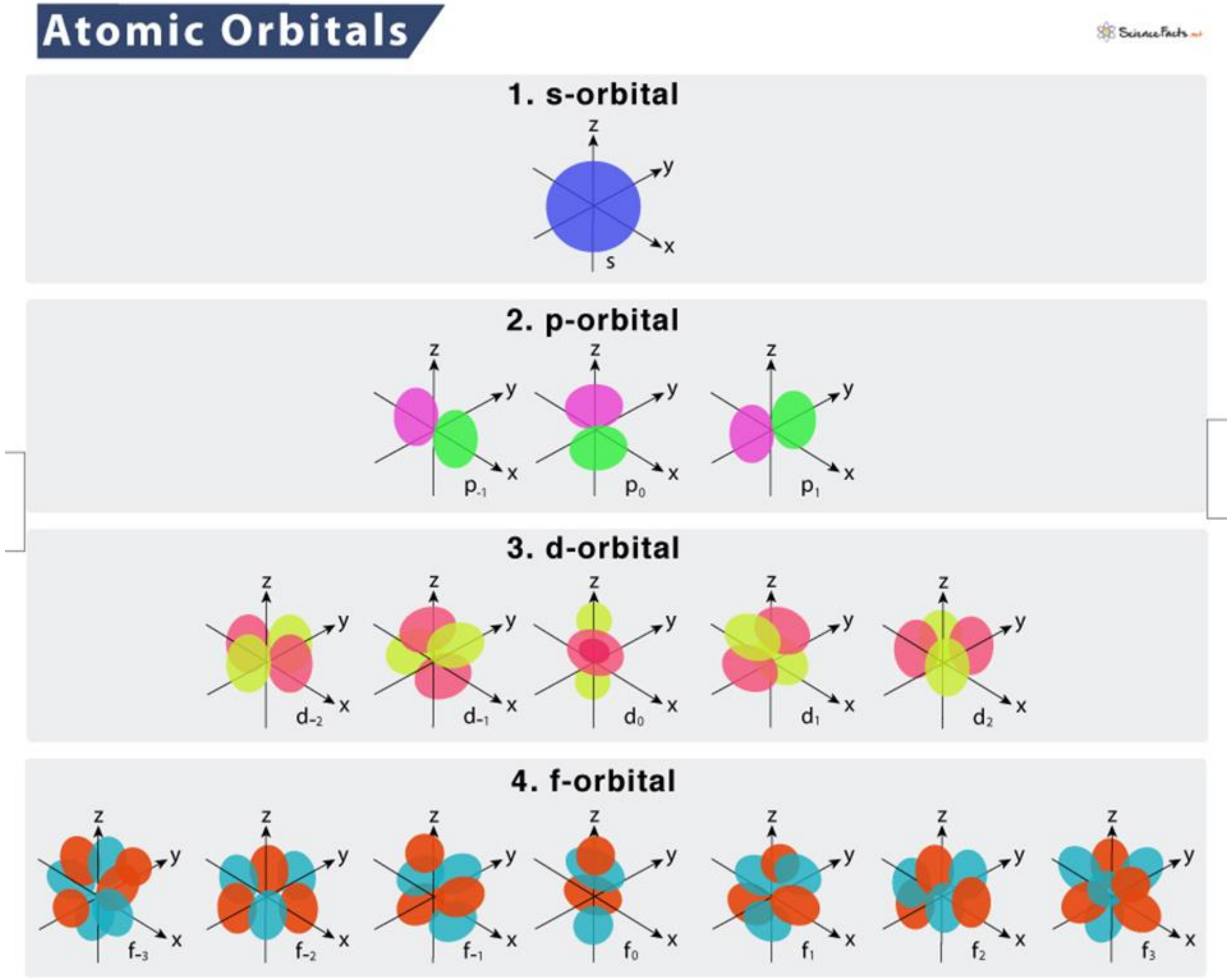
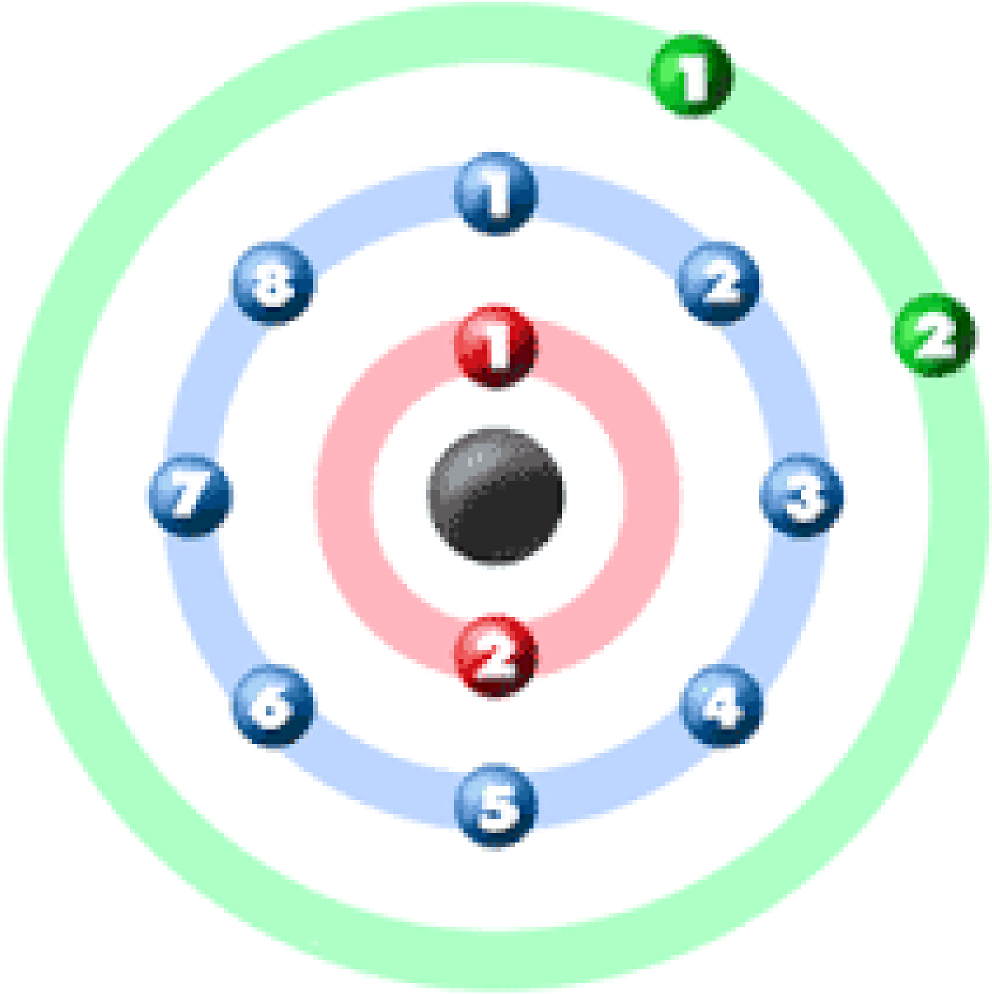


Table 2.2. First Four Energy Levels and Their Orbitals

Energy Level	Number of Orbitals	Maximum No. of Electrons per orbital
1	1	2
2	4	8
3	9	18
4	16	32

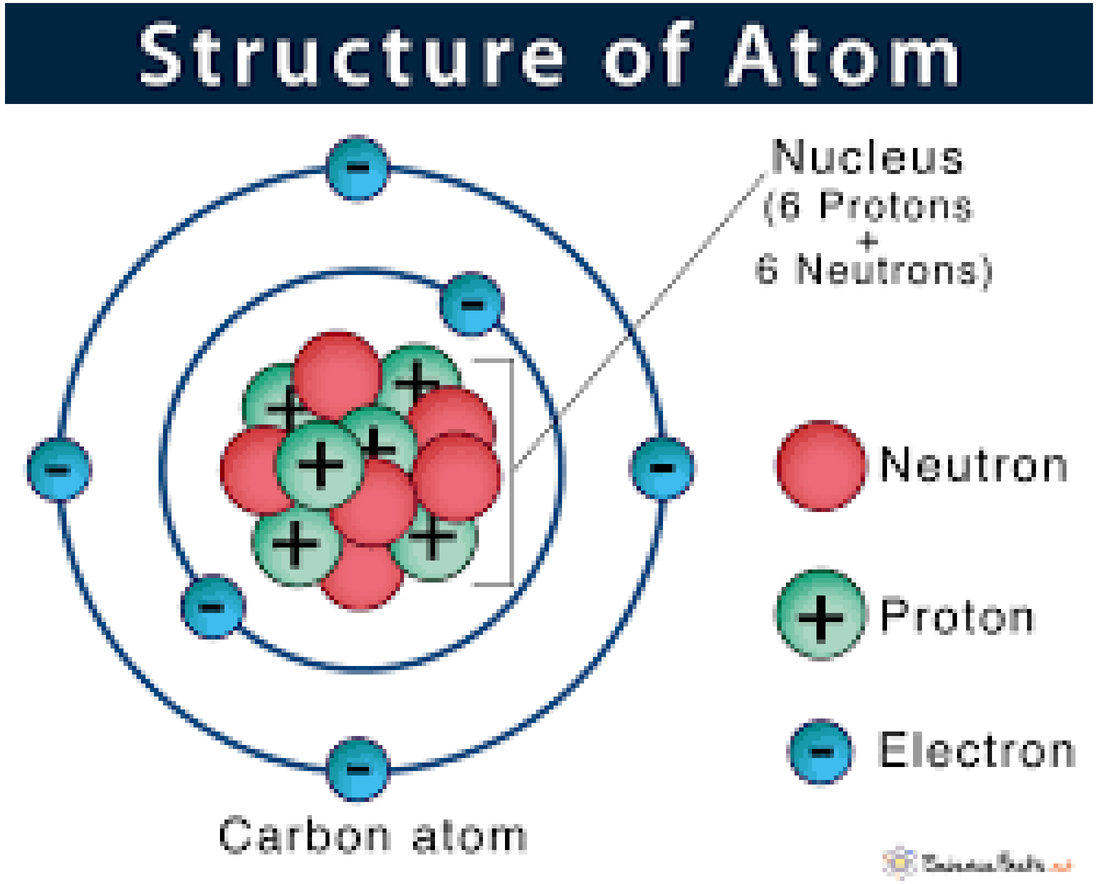
Orbitals

- The figure shows the arrangement of electrons in an atom of magnesium as an example.
- The most stable arrangement of electrons occurs when electrons fill the orbitals at the lowest energy levels first before more are added at higher levels.
- The model represents an atom of the element magnesium (Mg).
- How many electrons does the atom have at each energy level?
- What is the maximum number it could have at each level?



The Structure of Atoms

- An atom is composed of a positively charged nucleus containing the protons and neutrons.
- The nucleus is then surrounded by negatively charged particles - electrons.
- The presence of protons in the nucleus account for the positive charges within the nucleus.
- As earlier discussed, a neutral atom due to the equal number of protons in the nucleus and number of electrons surrounding the nucleus giving a zero charge.
- Most of an atom's mass is in its nucleus; the mass of an electron is only 1/1836 the mass of the lightest nucleus, that of hydrogen.
- Although the nucleus is heavy, it is small compared with the overall size of an atom.



The Structure of Atoms

- The radius of an atom is around 1 to 2.5 angstroms (Å), whereas the radius of a nucleus is about 10^{-5} Å.
- Enlarging an atom to the size of the earth, its nucleus would be only 200 feet in diameter and could fit inside a small football stadium.
- The nucleus of an atom contains protons and neutrons.
- Protons and neutrons have nearly equal masses, but they differ in charge.
- A neutron has no charge, whereas a proton has a positive charge that exactly balances the negative charge on an electron.
- Table 1 lists the charges of these three sub atomic particles, and gives their masses expressed in atomic mass units.

Table 1: Charge and mass of three sub atomic particles

Particle	Charge	Mass (grams)
Electrons	–1	9.1094×10^{-28}
Protons	+1	1.6726×10^{-24}
Neutrons	0	1.6749×10^{-24}

Example

- What is the atomic symbol for bromine, and what is its atomic number?
- Why is it incorrect to use the letter B?
- What other element preempts the symbol B?
- (Refer to the periodic table).

Periodic Table of the Elements

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H Hydrogen 1.01																	2 He Helium 4.00
3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.88	23 V Vanadium 50.94	24 Cr Chromium 51.99	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.38	31 Ga Gallium 69.72	32 Ge Germanium 72.63	33 As Arsenic 74.92	34 Se Selenium 78.97	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.95	43 Tc Technetium 98.91	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.6	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57-71 Lanthanides	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.85	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.20	83 Bi Bismuth 208.98	84 Po Polonium [208.98]	85 At Astatine 209.98	86 Rn Radon 222.02
87 Fr Francium 223.02	88 Ra Radium 226.03	89-103 Actinides	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [269]	109 Mt Meitnerium [278]	110 Ds Darmstadtium [281]	111 Rg Roentgenium [280]	112 Cn Copernicium [285]	113 Nh Nihonium [286]	114 Fl Flerovium [289]	115 Mc Moscovium [289]	116 Lv Livermorium [293]	117 Ts Tennessine [294]	118 Og Oganesson [294]

57 La Lanthanum 138.91	58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium 144.91	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.06	71 Lu Lutetium 174.97
89 Ac Actinium 227.03	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium 237.05	94 Pu Plutonium 244.06	95 Am Americium 243.06	96 Cm Curium 247.07	97 Bk Berkelium 247.07	98 Cf Californium 251.08	99 Es Einsteinium [254]	100 Fm Fermium 257.10	101 Md Mendelevium 258.10	102 No Nobelium 259.10	103 Lr Lawrencium [262]

Alkali Metal | Alkaline Earth | Transition Metal | Basic Metal | Metalloid | Nonmetal | Halogen | Noble Gas | Lanthanide | Actinide

Mass Number and Atomic Number

- The **atomic number** is equal to the number of protons within the nucleus which is also equal to the number of electrons surround the nucleus.
- The **mass number** is the sum of the protons and neutron.
- It is the whole number of an **atomic mass.**
- Atomic mass a relative mass assigned to an atom which is exactly 1/12th the mass of C-12.
- The SI unit is expressed in **amu** (atomic mass unit) represented by u.

- **1.0 amu = 1.6605 x 10^{-24} g**
- **1.0 g = 6.0221 amu**

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19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.88	23 V Vanadium 50.94	24 Cr Chromium 51.99	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.38	31 Ga Gallium 69.72	32 Ge Germanium 72.63	33 As Arsenic 74.92	34 Se Selenium 78.97	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.95	43 Tc Technetium 98.91	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.6	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57-71 Lanthanides	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.85	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.20	83 Bi Bismuth 208.98	84 Po Polonium [208.98]	85 At Astatine 209.98	86 Rn Radon 222.02
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89 Ac Actinium 227.03	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium 237.05	94 Pu Plutonium 244.06	95 Am Americium 243.06	96 Cm Curium 247.07	97 Bk Berkelium 247.07	98 Cf Californium 251.08	99 Es Einsteinium [254]	100 Fm Fermium 257.10	101 Md Mendelevium 258.10	102 No Nobelium 259.10	103 Lr Lawrencium [262]

Alkali Metal	Alkaline Earth	Transition Metal	Basic Metal	Metalloid	Nonmetal	Halogen	Noble Gas	Lanthanide	Actinide

Mass Number and Atomic Number

- Oxygen-16 atom is 16.00 amu
- Carbon-12 atom is 12.00 amu
- Hydrogen-1 atom is 1.00 amu

- **Mass Number = Number of Protons + Number of Neutrons**
- **Atomic Number = No. of Protons = No. of Electrons**
- **Mass Number = Whole Number of Atomic Mass**
- **Charge = Number of protons (+) + Number of electrons (-)**

Mass Number and Atomic Number

- The number of neutrons in an atom is determined by subtracting the atomic number (number of protons = number of electrons) from the mass number.
- Using the format notation as

$${}^{A}_{Z}X^{C}$$

- X represents the atomic symbol of the specific element
- A is the mass number
- Z is the atomic number
- C is the charge

Examples

- How many protons, neutrons and electrons are there in an atom of

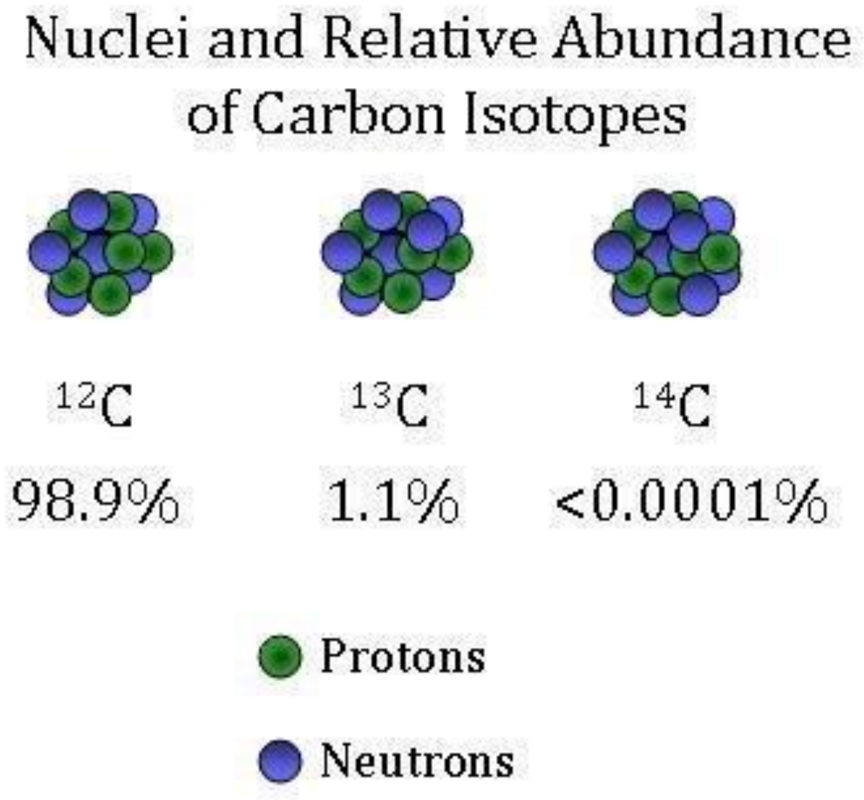
$$^{197}Au_{79}$$

What is the charge, C?

- Give the complete chemical symbol for the atom that has 18 protons, 18 electrons and 22 neutrons.
- Give the complete chemical symbol for the element that has 10 electrons, 11 protons and 12 neutrons.

Isotopes

- Isotopes are atoms of the same element having a different number of neutrons with the same atomic number.
- The atomic mass of a particular element is equal to the average of the relative abundance of all its isotopes found in nature.
- For example, there are three naturally occurring isotopes of carbon: carbon-12, carbon-13, and carbon-14.
- Carbon-12 is the most common of these three, making up about 98.89% of all carbon, whereas carbon-13 has 1.11% natural abundance.
- Carbon-14 occurs rarely in nature.
- Atomic masses for other elements use the carbon-12 scale as a reference.



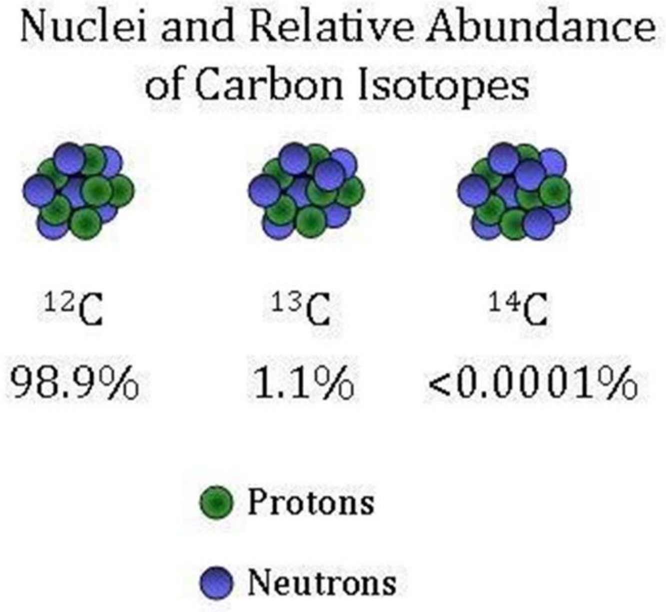
Isotopes of Carbon

- average atomic mass

= [mass of carbon-12 x (% natural abundance/100) + mass of carbon-13 x (% natural abundance/100)]

= (12 x 0.9889) + (13 x 0.0111)

= 12.0111



Examples

- How many protons, neutrons, and electrons are there in an atom of uranium- 238?

Write the symbol for this isotope.

- Magnesium (Mg) has three significant natural isotopes:
  - 78.70% of all magnesium atoms have an atomic weight of 23.985 u,
  - 10.13% have an atomic weight of 24.986 u, and
  - 11.17% have an atomic weight of 25.983 u.

How many protons and neutrons are present in each of these three isotopes?

What are the symbols for each isotope?

What is the weighted average of the atomic weights?

Natural and Artificial Isotopes

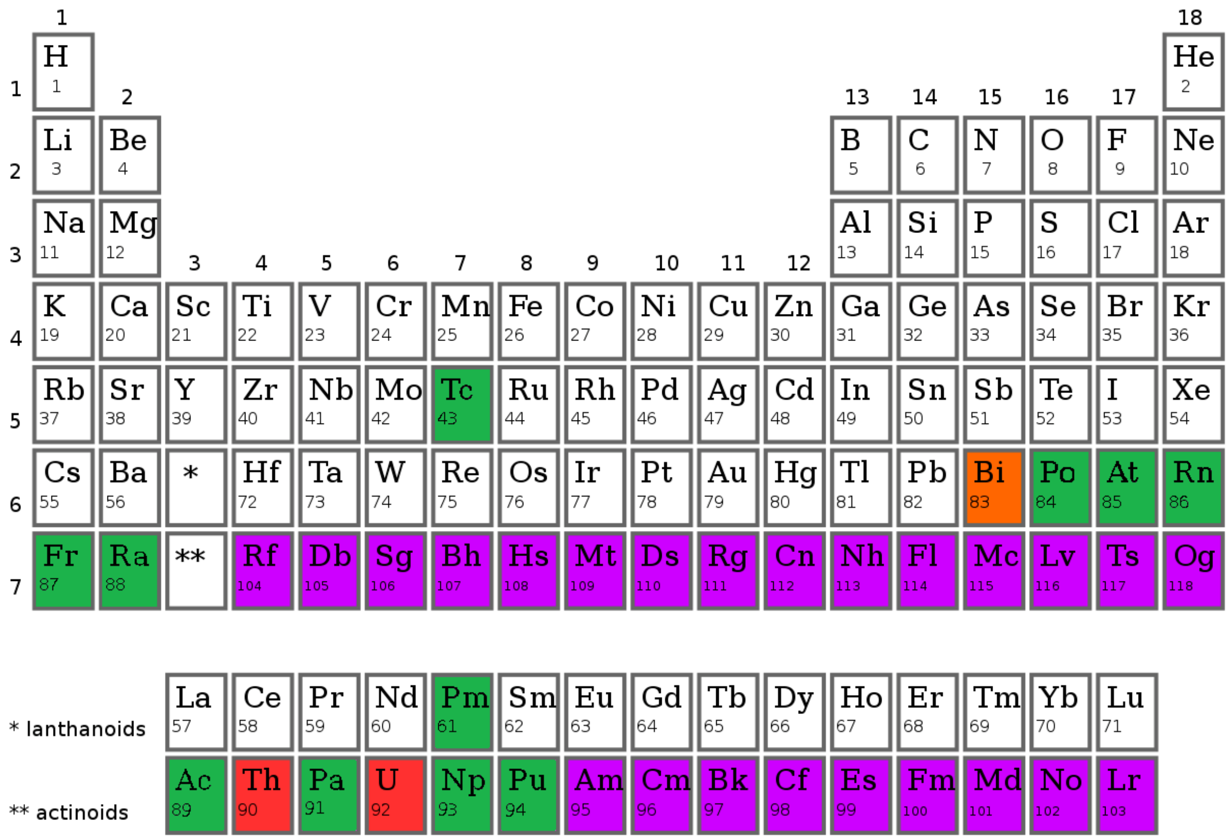
- Most of the elements are having naturally occurring isotopes.
- There are those occurring more predominantly with greater abundance and there are those occurring only in traces.

1. The element mercury has seven naturally occurring isotopes. Based on the isotopic masses and abundances shown below, calculate the average atomic mass of mercury.

Isotope	Mass (amu)	Abundance (%)
^{196}Hg	195.9658	0.146
^{198}Hg	197.9668	10.02
^{199}Hg	198.9683	16.84
^{200}Hg	199.9683	23.13
^{201}Hg	200.9703	13.22
^{202}Hg	201.9706	29.80
^{204}Hg	203.9735	6.85

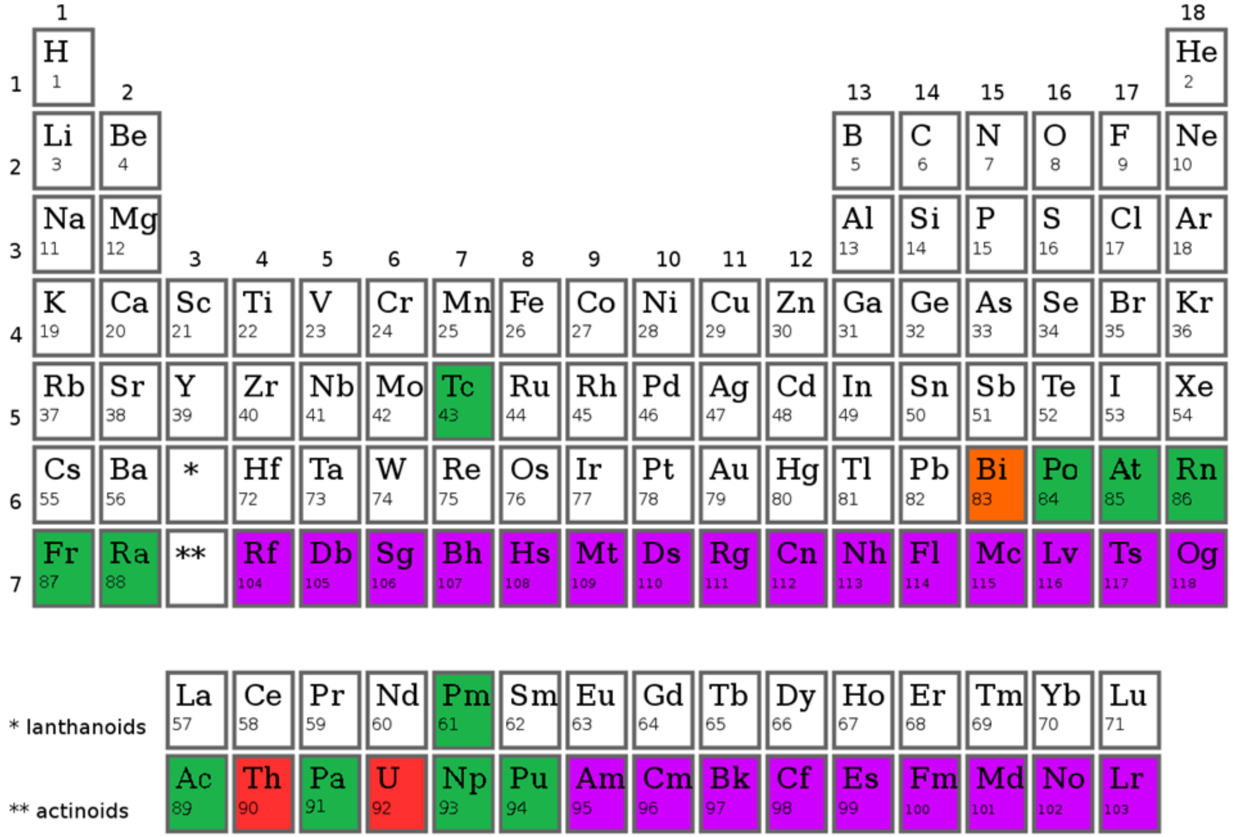
Natural and Artificial Isotopes

- There are 20 elements with only artificially produced isotopes.
- The majority of these are heavier elements; the lightest elements with artificial isotopes are ^{43}Tc and ^{61}Pm.
- The other elements that only have artificial isotopes are those with atomic numbers of 84-88 and 89-103, otherwise known as the **actinoids or actinides**, but excluding ^{90}Th and ^{92}U.



Natural and Artificial Isotopes

- Some naturally occurring and artificially produced isotopes are radioactive.
- The nucleus of a radioactive isotope is unstable; radioactive isotopes spontaneously decay, emitting alpha, beta, and gamma rays until they reach a stability, usually in the state of a different element.
- Bismuth $^{209}Bi_{83}$ has the highest atomic and mass number of all the stable nuclides.
- All nuclides with atomic number and mass number greater than 83 and 209, respectively, are radioactive.
- However, there are some lighter nuclides that are radioactive

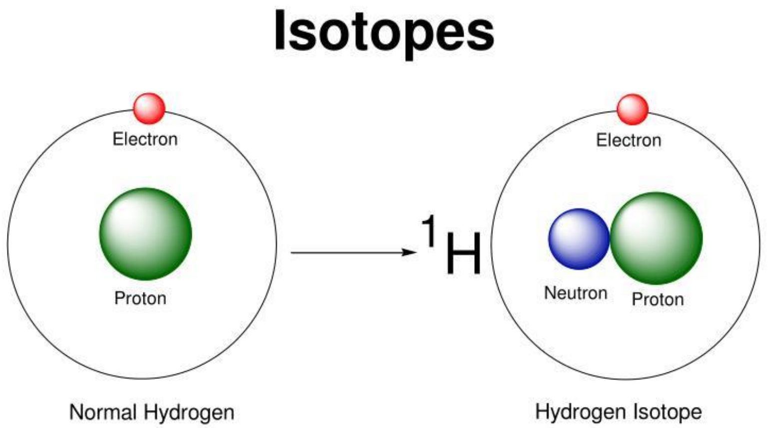


Use of Isotopes

- Structural determination using isotopes is often performed using
  - nuclear magnetic resonance spectroscopy and
  - mass spectrometry

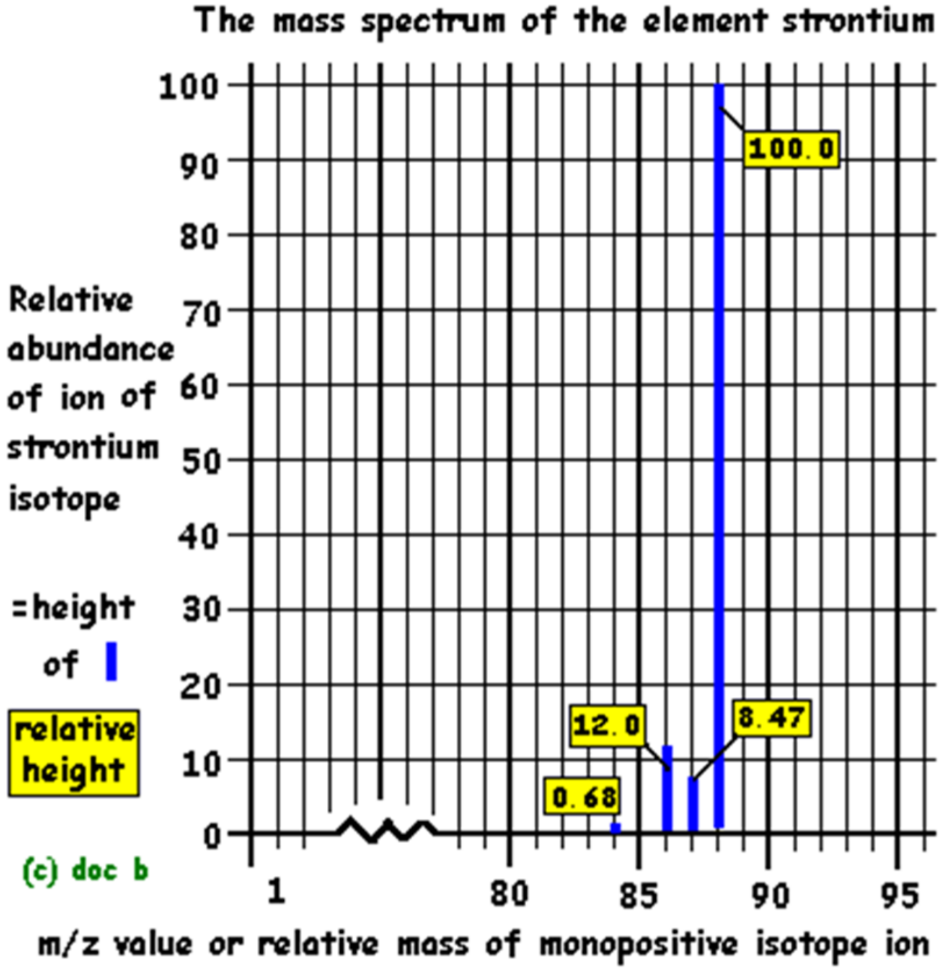
Nuclear Magnetic Resonance Spectroscopy

- NMR analysis is isotope-dependent, and it often relies on trace isotopes of a molecule for detection.
- For example, the most abundant isotope of carbon, C-12, is invisible to NMR, whereas the minor isotope C-13 is NMR active, but only comprises 1.1 percent of a given sample of carbon.
- By replacing C-12 in a molecule with C-13, NMR analysis of that position is greatly enhanced.
- Similarly, H-1 is an NMR active nucleus, whereas H-2 is NMR invisible, so it is possible to determine where a specific hydrogen atom is by replacing it with H-2 and watching for the disappearance of the corresponding signal.



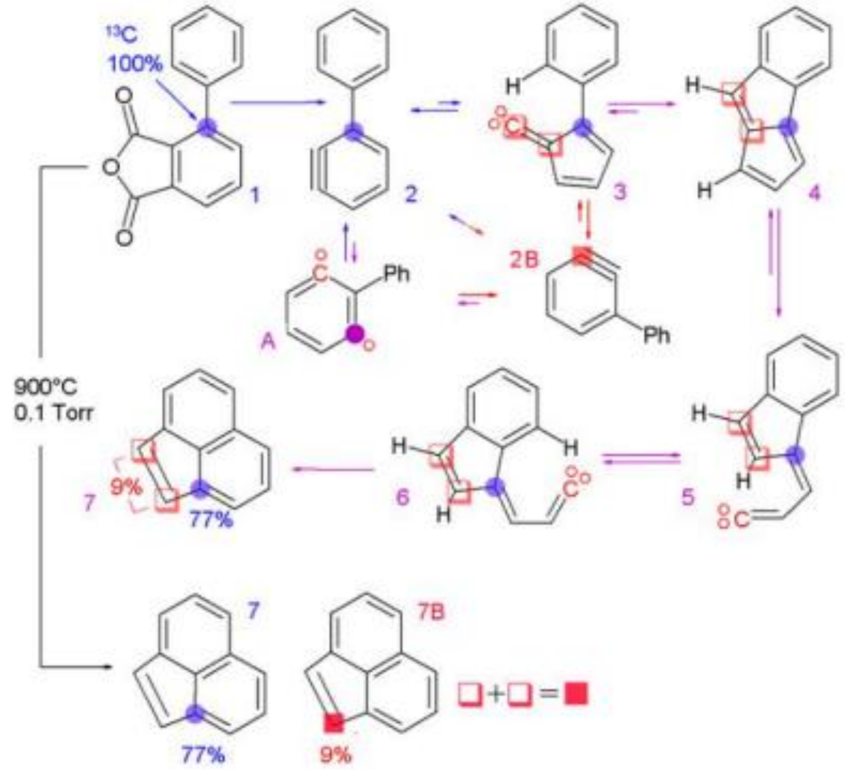
- **NMR works with the isotopes of common organic atoms (hydrogen, carbon, nitrogen, phosphorus) . If we look above we see that the hydrogen on the right is an isotope because of the additional neutron.**

Mass Spectrometry



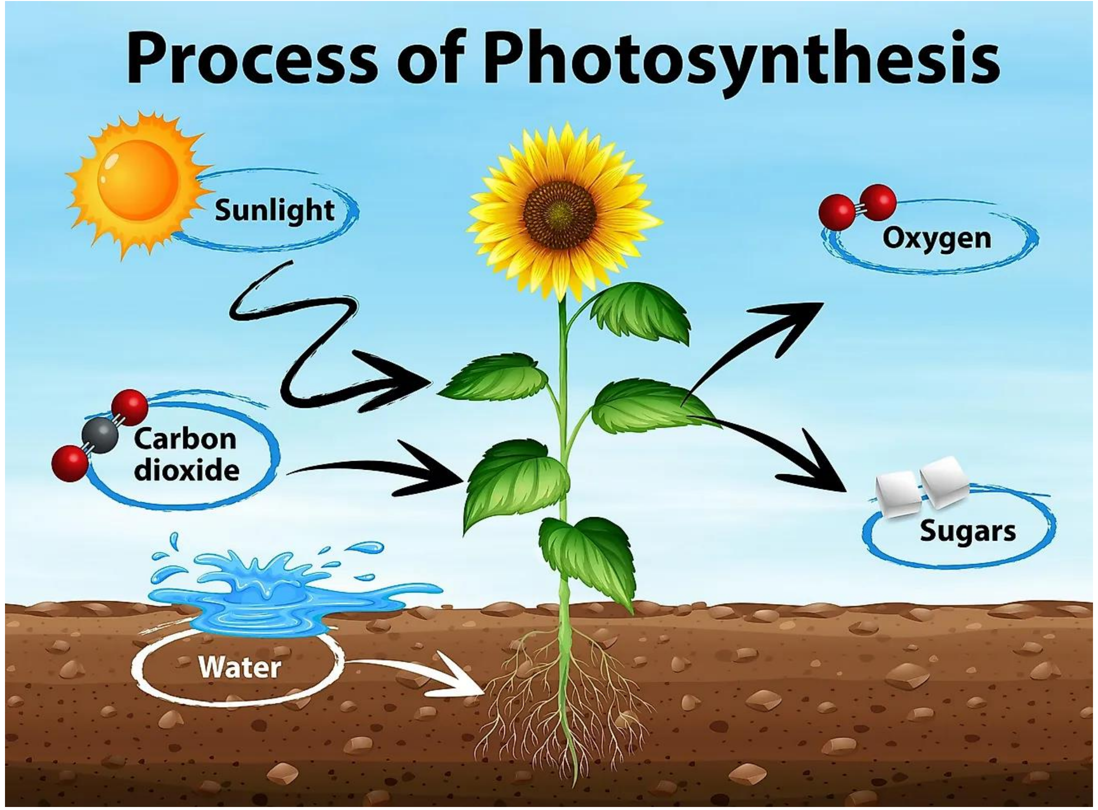
Isotopic Labelling

- A carbon-13 labeled precursor was used to track the position of an atom through a reaction.



Study of Photosynthesis

- Mass spectrometry has been used to study the ratio of carbon isotopes in various plants to understand the mechanisms of photosynthesis.
- The study of the ratio of carbon isotopes in various plants uses spectrometry to understand the mechanism of photosynthesis.



Isotopes in Photosynthesis

Expt 1: Photosynthesis in the presence of isotopic water

$$\underset{\text{Carbon dioxide}}{CO_2} + \underset{\text{Heavy Water}}{2H_2O^{18}} \xrightarrow[\text{Chlorophyll}]{\text{Light}} \underset{\text{Glucose}}{C_6H_{12}O_6} + \underset{\text{Water}}{H_2O} + \underset{\text{Isotopic oxygen}}{O_2^{18}}$$

Expt 2: Photosynthesis in the presence of isotopic carbon dioxide

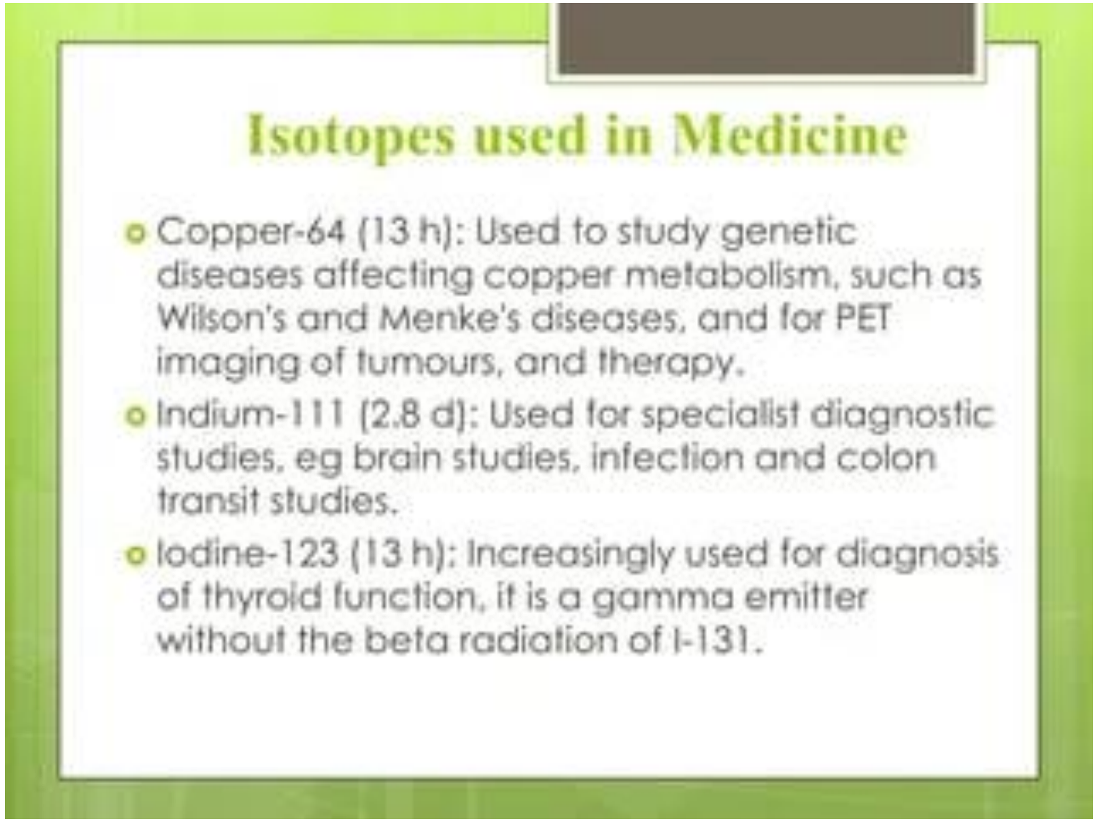
$$\underset{\text{Carbon dioxide}}{CO_2^{18}} + \underset{\text{Water}}{2H_2O} \xrightarrow[\text{Chlorophyll}]{\text{Light}} \underset{\text{Glucose}}{C_6H_{12}O_6^{18}} + \underset{\text{Heavy Water}}{H_2O^{18}} + \underset{\text{Oxygen}}{O_2}$$

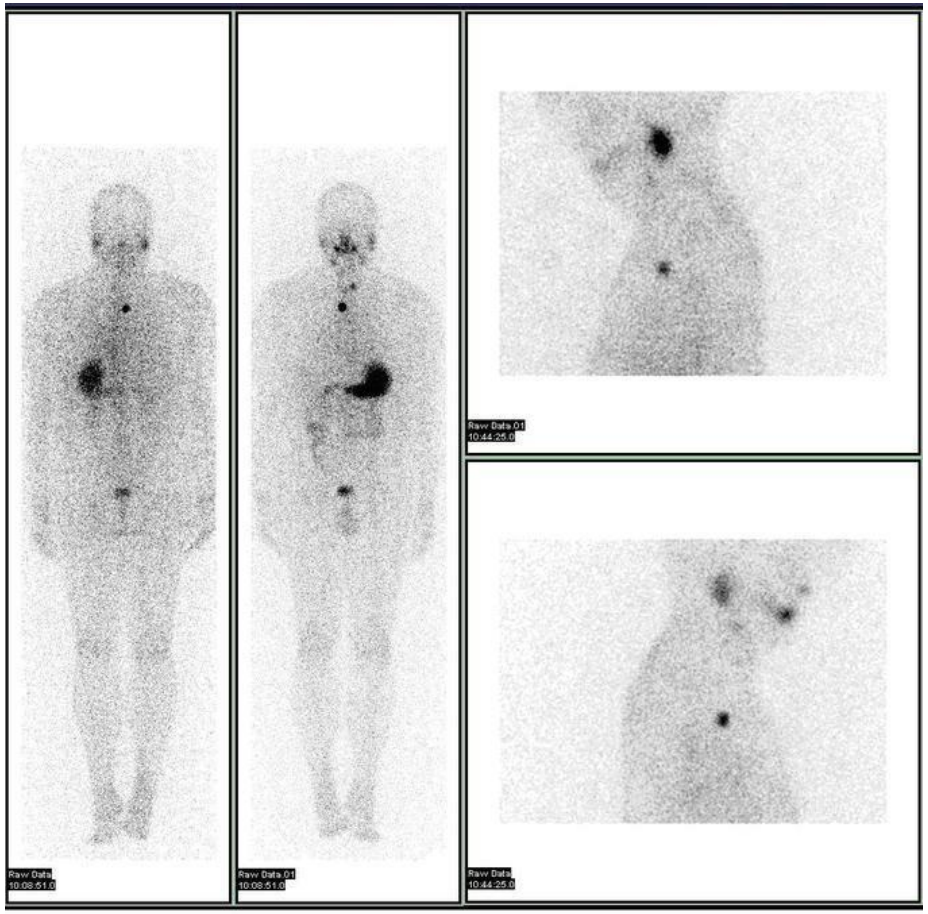
Ruben, Randall, kamen and Hyde (1940): proved that the O_2 evolved during photosynthesis comes from H_2O.

Isotopes in Medicine

- Nuclear medicine is a medical specialty that involves the application of radioactive substances to diagnose or treat disease which employs procedures where radionuclides are combined with other elements forming a chemical compound.
- These radiopharmaceuticals, once administered to the patient, can localize to specific organs or cellular receptors.
- This property of radiopharmaceuticals allows nuclear medicine the ability to image the extent of a disease process in the body.
- These images are based on cellular function and physiology, rather than on physical changes in the tissue anatomy.
- Therefore, with some diseases, nuclear medicine studies can identify medical problems at an earlier stage than other d

Diagnosis

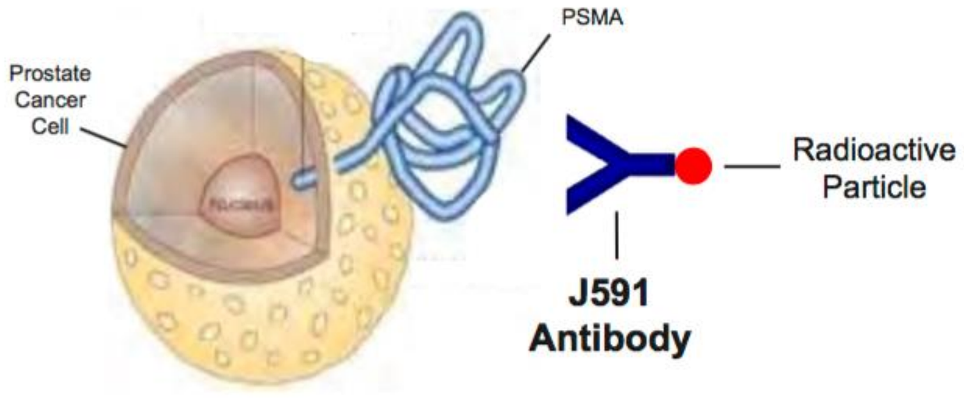

Isotopes used in Medicine
Copper-64 (13 h): Used to study genetic diseases affecting copper metabolism, such as Wilson's and Menke's diseases, and for PET imaging of tumours, and therapy.
Indium-111 (2.8 d): Used for specialist diagnostic studies, eg brain studies, infection and colon transit studies.
Iodine-123 (13 h): Increasingly used for diagnosis of thyroid function, it is a gamma emitter without the beta radiation of I-131.



Treatment

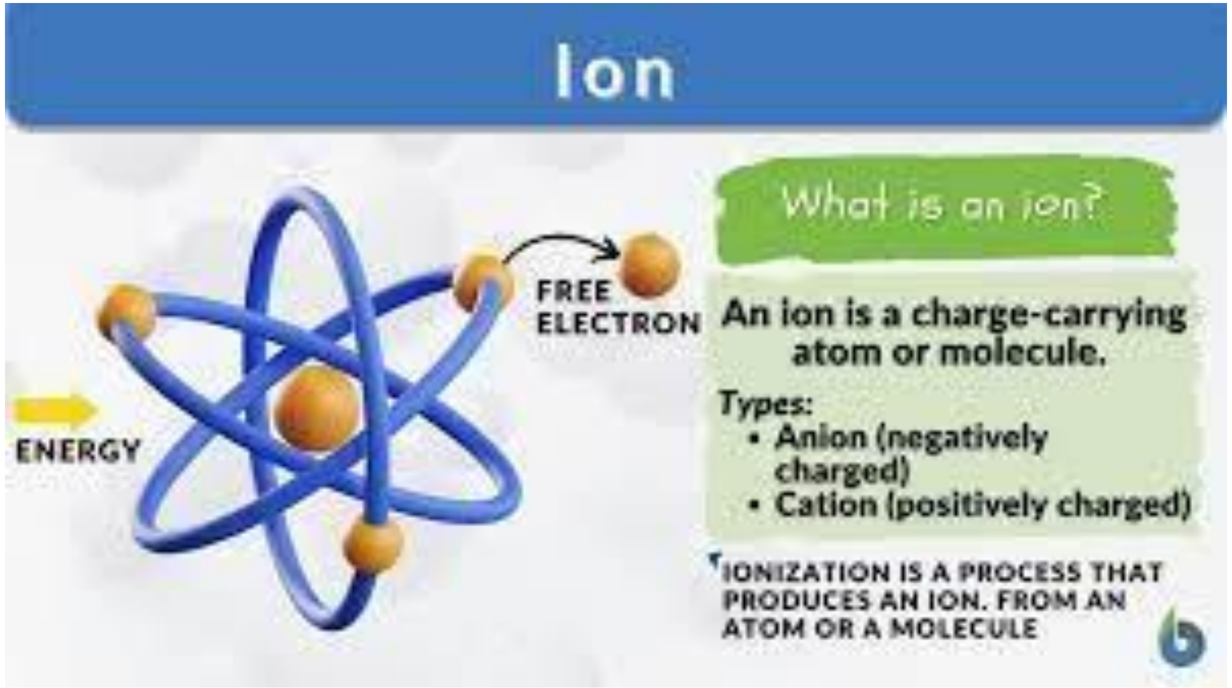
Medical uses of Radio active isotopes: Radiation from radio active isotopes is used in medicine to treat cancer, some of radio active isotopes are:

1. Cobalt – 60 – its symbol $^{60}{}_{27}C$, its β&γ emission, its widely used as external source of radiation for cancer therapy with $t_{1/2}$ 5.3 years.
2. Iodine – 131 – $^{131}{}_{53}I$, its β & γ emission, its used in treatment of thyroid gland cancer with ty_2:8 days.
3. Iodine – 123 – $^{123}{}_{53}I$: its symbol, its β – emission, its used in treatment of thyroid cancer with ty_2 13.3hr.



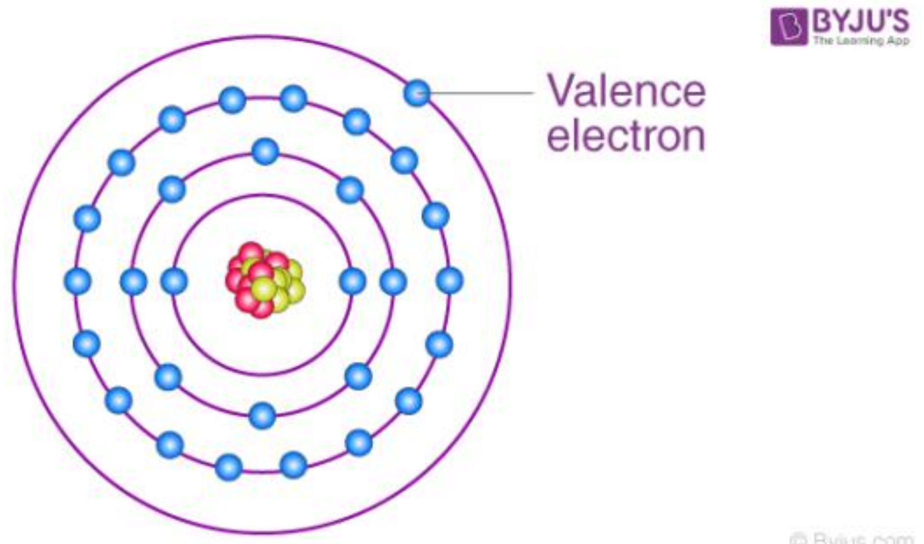
Ion

- Once an atom has either more than protons than electrons or more than electrons than protons, the atom is said to carry a charge either positive or negative.
- Atoms carrying either a positive or negative charge are said to be **ion.**



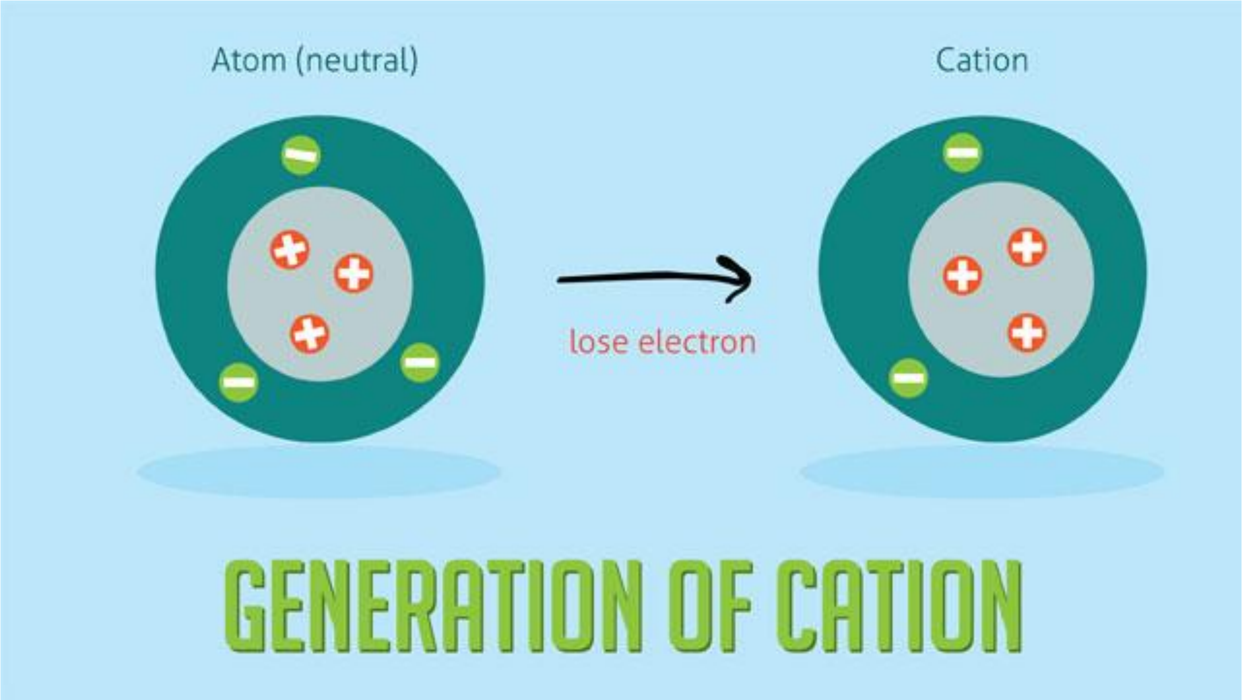
Ion

- Ions are formed when an atom gains or losses electrons during chemical reaction.
- A **valence** electron is the electron or electrons located at the outermost shell of an atom.
- These valence electrons are the ones involved in the reactions.



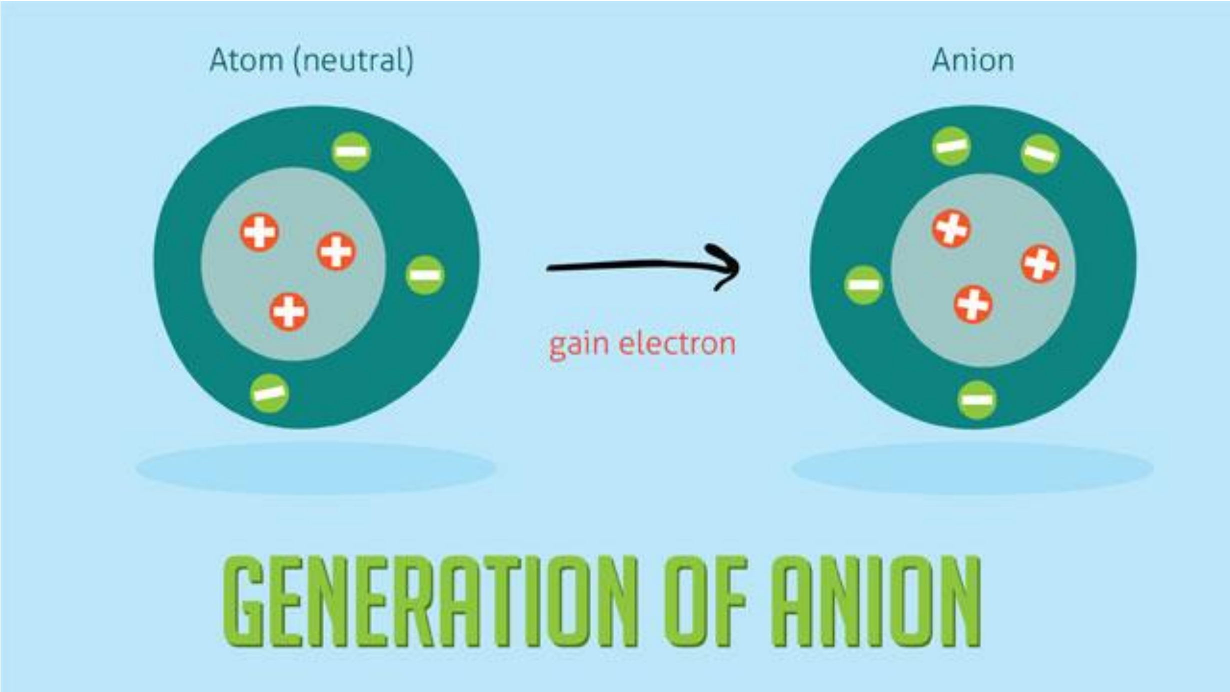
Ion

- An atom with more than four valence electrons tends to give up these electrons during chemical reactions resulting to more protons and thus become a positive ion or a **cation**.
- The charge of a cation depends upon the number of electrons an atom has given up.



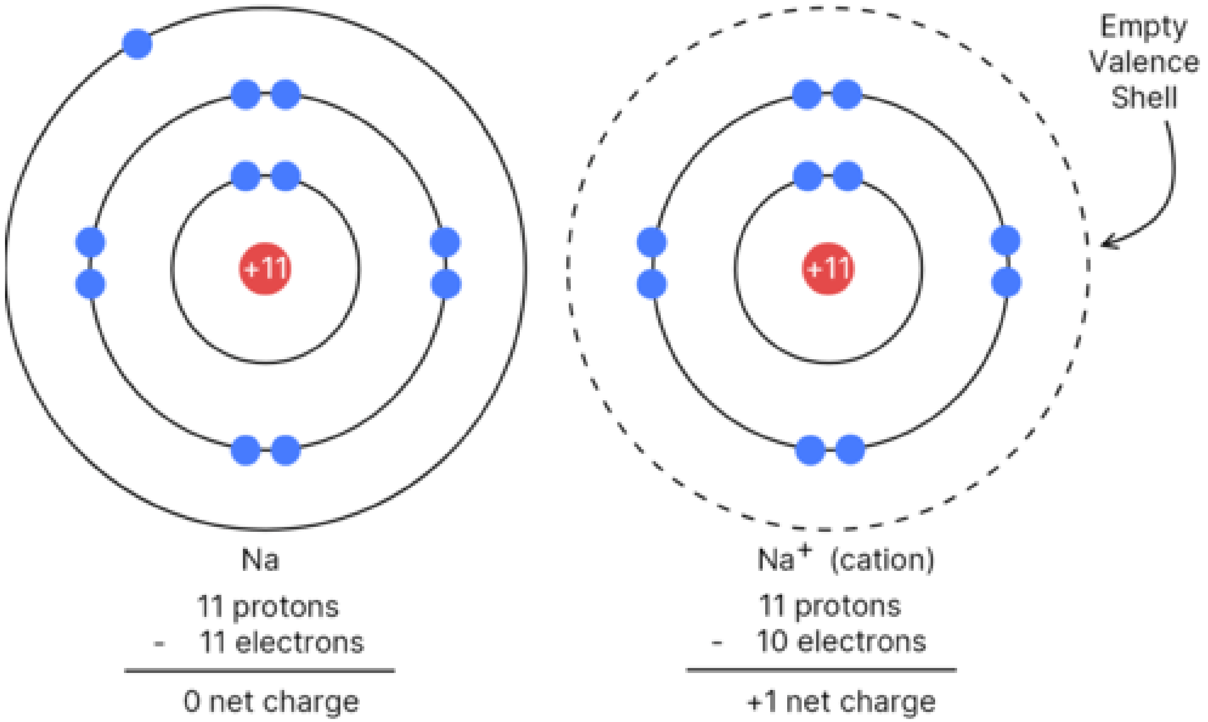
Ion

- On the other hand, atoms with valence electron or electrons less than four tends to accept or gain electrons during chemical reactions.
- This results to more number of electrons than proton and thus becomes a negative charged ion – the **anion**.
- The charge of the anion depends upon the number of electrons the atom has gained.



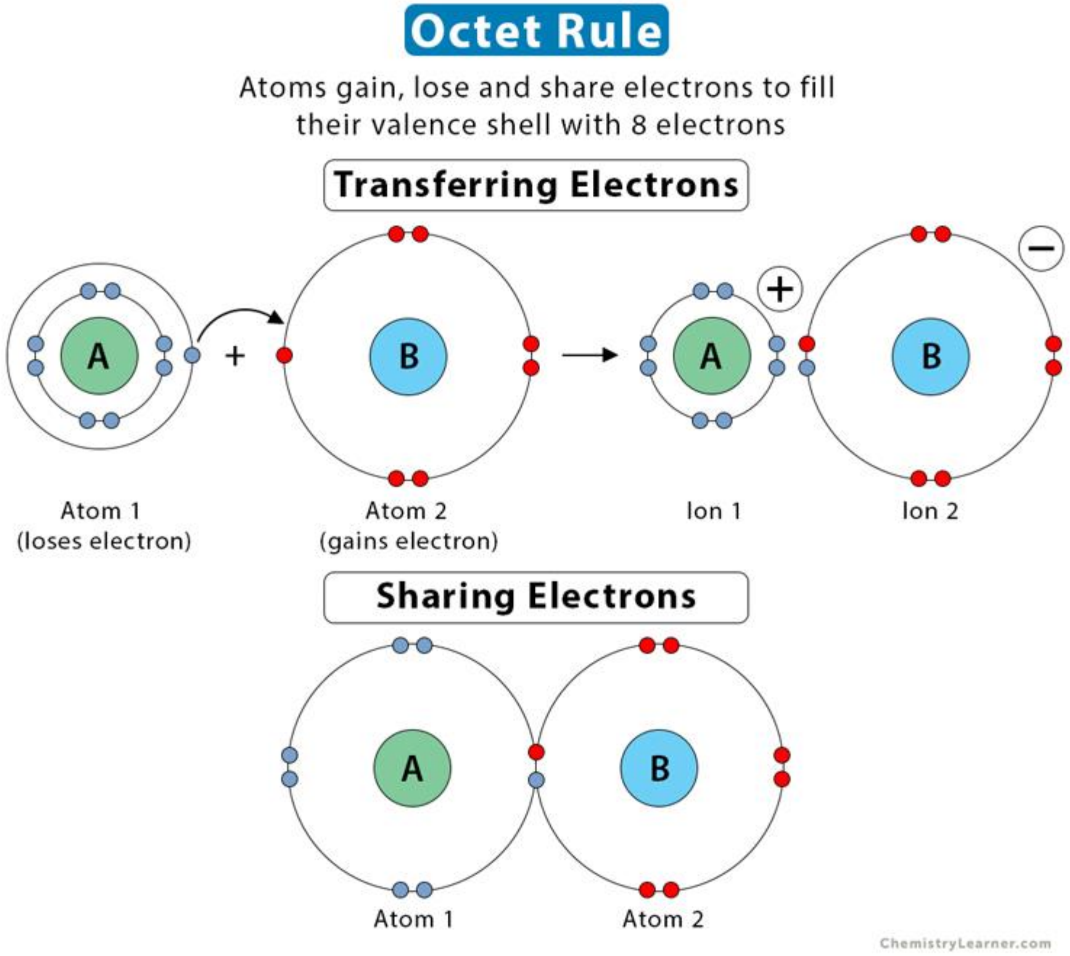
Ion

- A common example is sodium.
- Sodium has eleven electrons and has 1 (less than four) valence electrons.
- During chemical reaction sodium has to give up its 1 valence electron, thus sodium is a cation and since it has given up 1 electron, it has a +1 charge.



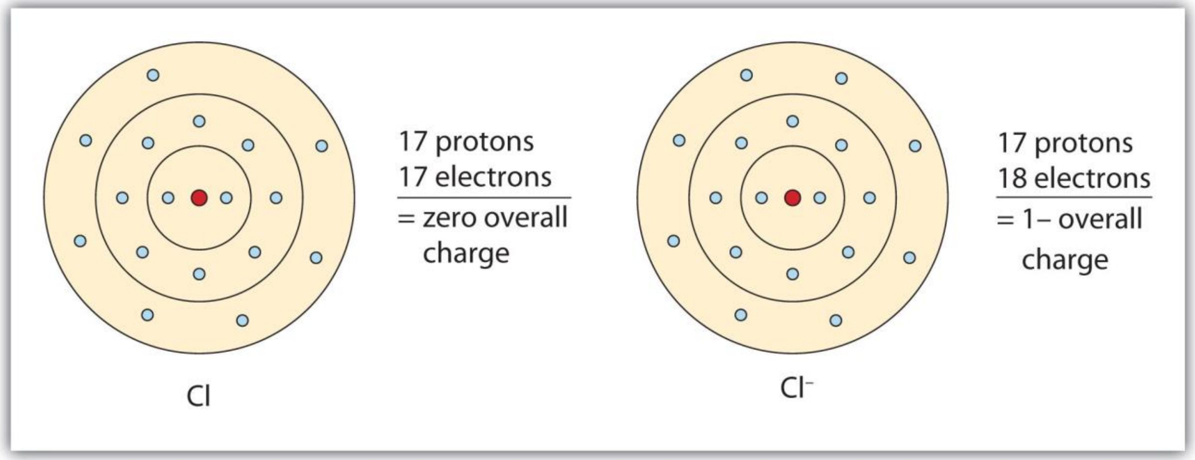
Ion

- In terms of the octet rule an energy level with 2 electrons in the first level is stable and from second level onwards, an eight electron is required to become stable.



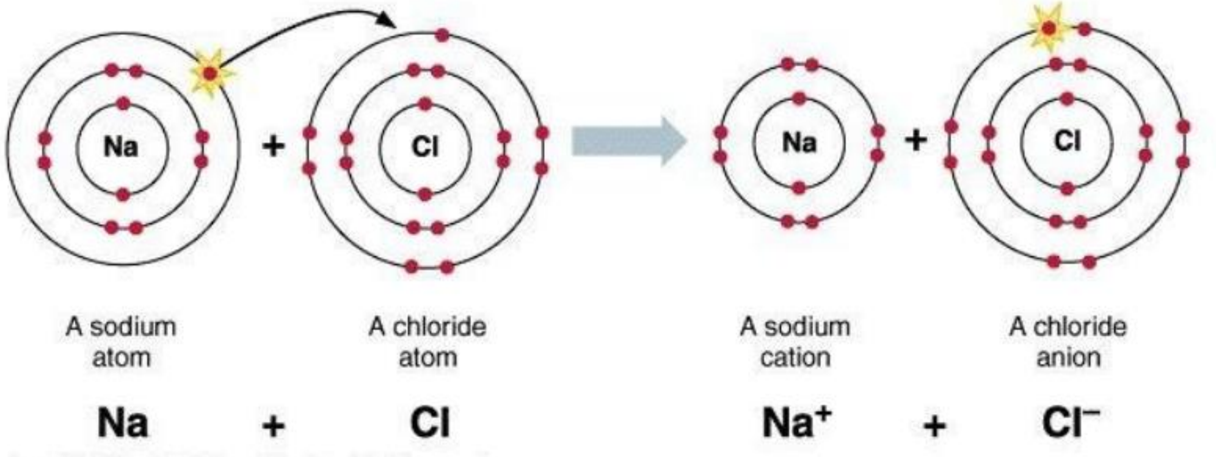
Ion

- On the other hand, chlorine has seven valence electrons (more than four) and tends to gain 1 electron to complete the 8 e- (octet rule).
- Since chlorine atom gained 1 e-, it becomes an anion.
- The charge is -1 corresponding to the 1 electron gained.



Ion

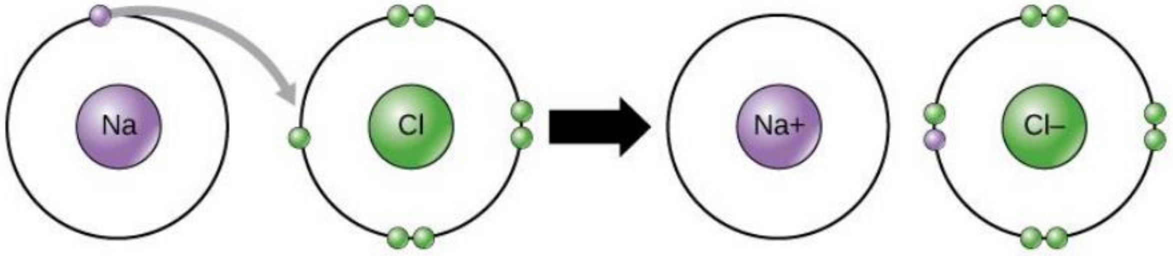
- Therefore sodium needs to give up 1 electron to attain a stable configuration (i.e. 2 e- in the first shell and 8 e- in the second shell).
- On the other hand, chlorine has seven valence electrons (more than four) and tends to gain 1 electron to complete the 8 e- (octet rule).
- Since chlorine atom gained 1 e-, it becomes an anion.
- The charge is -1 corresponding to the 1 electron gained.



Ionic Bonds

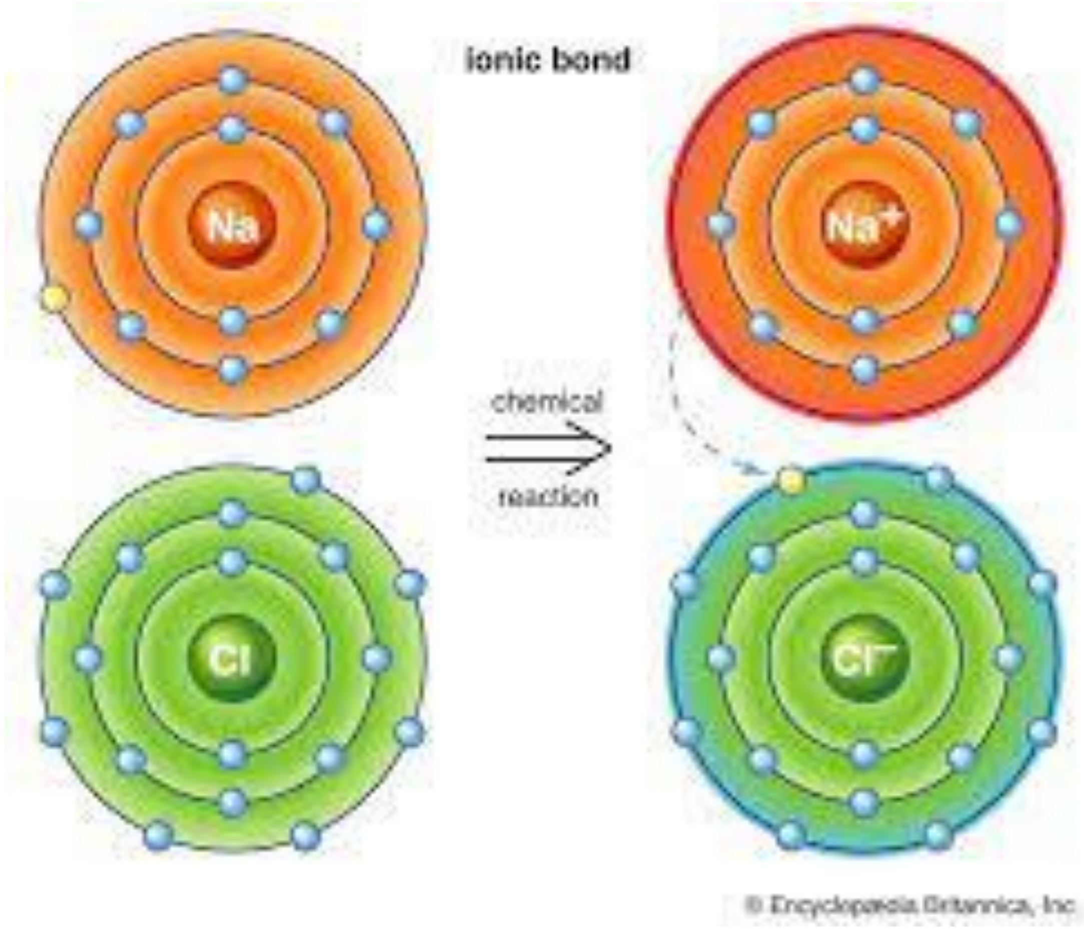
- Sodium atom can lose an electron in forming cation Na^+ and at how chlorine atom can gain an electron in forming the anion Cl^-.
- In reality, however, this process can occur all in one step when sodium gives its electron away to chlorine!
- We can illustrate this as follows:

- Oxidation and reduction of sodium and chlorine. Sodium donates its electron to chlorine to form Na+ and Cl-.



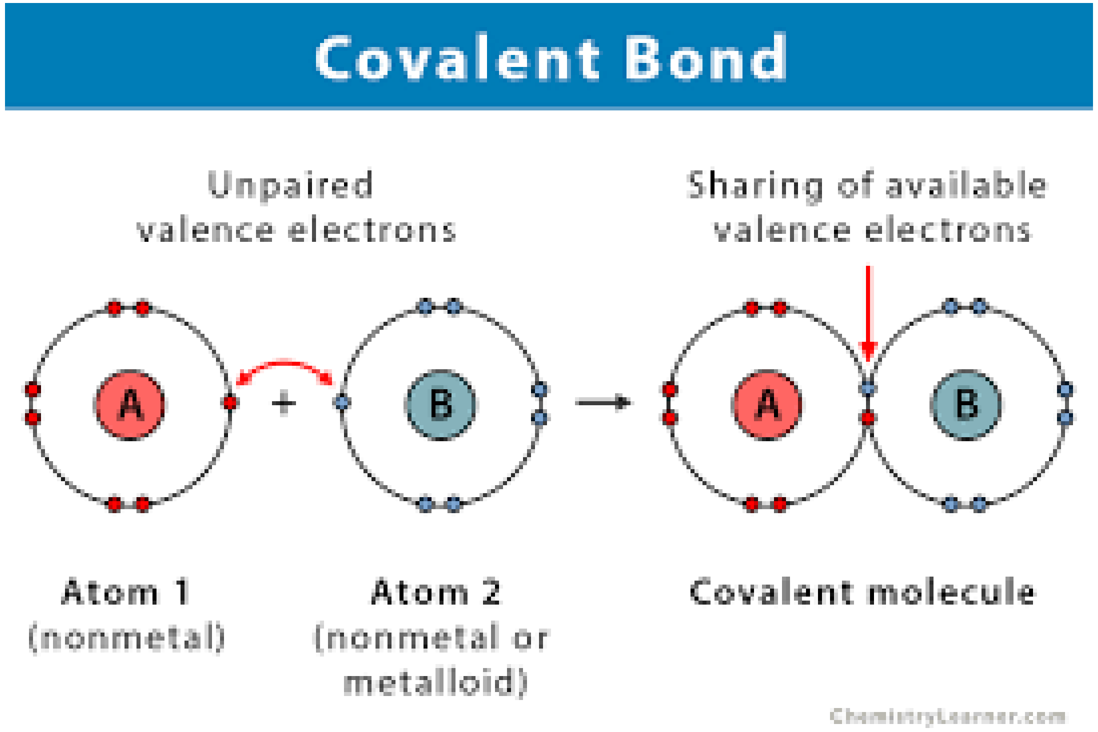
Ionic Bonds

- An illustration on how an electron is transferred from sodium to chlorine in order to form the ions Na+ and Cl-.
- Once these ions are formed, there is a strong electrostatic attraction between them, which leads to the formation of an ionic bond.
- Ionic bonds are formed by electron transfer and that there is unequal sharing of electrons.



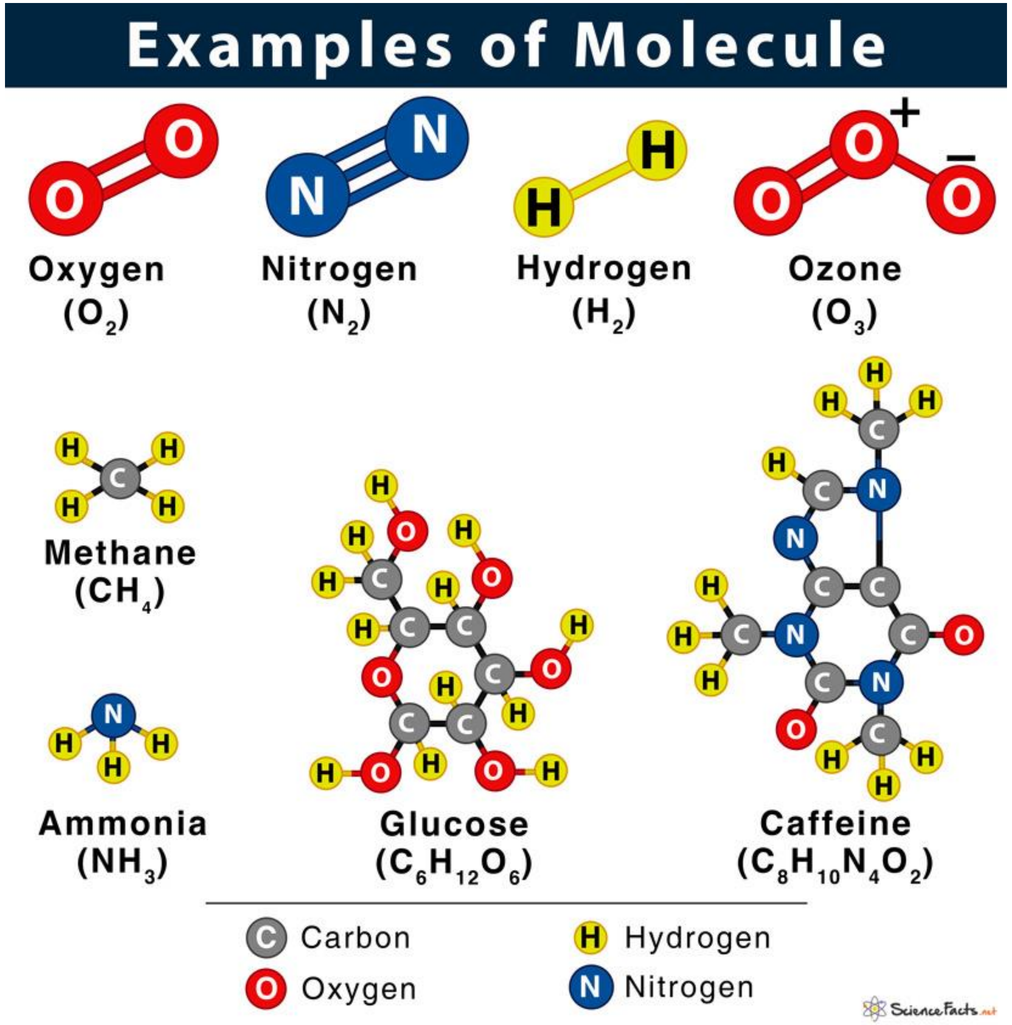
Covalent bonds and molecules

- Sharing of electrons pairs in the formation of compounds results in a **covalent bond**.
- In a covalent bond, the stability of the bond comes from the shared electrostatic attraction between the two positively charged atomic nuclei and the shared, negatively charged electrons between them.



Covalent bonds and molecules

- When atoms combine by forming covalent bonds, the resulting collection of atoms is called a **molecule**.
- We can therefore say that a molecule is the simplest unit of a covalent compound.
- As we will now see, there are a variety of different ways to represent and draw molecules.



Representing molecules: chemical formulas

- **Chemical formulas, or molecular formulas**, are the shorthand way of representing molecules.
- In a chemical formula, we use the elemental symbols from the periodic table to indicate which elements are present, and we use subscripts to indicate how many atoms of each element exist within the molecule.

Compound	Molecular Formula
Glucose	$C_6H_{12}O_6$
Butane	C_4H_{10}
Octane	C_8H_{18}

Representing molecules: structural formulas

- Chemical formulas only tell us how many atoms of each element are present in a molecule, but structural formulas also give information about how the atoms are connected in space.
- In structural formulas, we actually draw the covalent bonds connecting atoms.
- The chemical formula for ammonia is NH_3.
- Now, let's consider its structural formula.
- The figure illustrates the structural formula.
- In this chapter, we will not be discussing the structural formula.
- We will deal with the molecular or chemical formula.

Two structural formulas for ammonia, NH_3. The formula on the left gives only a two-dimensional approximation of molecular structure, whereas the formula on the right shows the orientation of atoms in space using dashes going into the plane of the page and wedges coming out of the plane of the page. The two dots on nitrogen in the right formula indicate a lone pair of electrons.

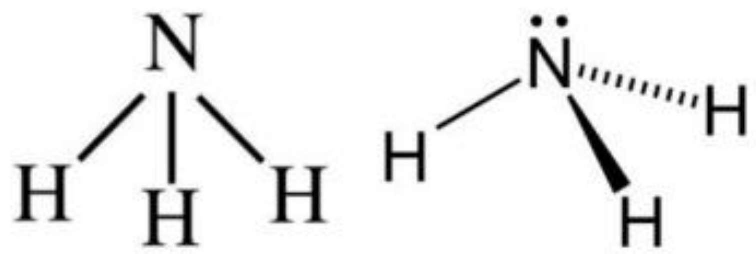
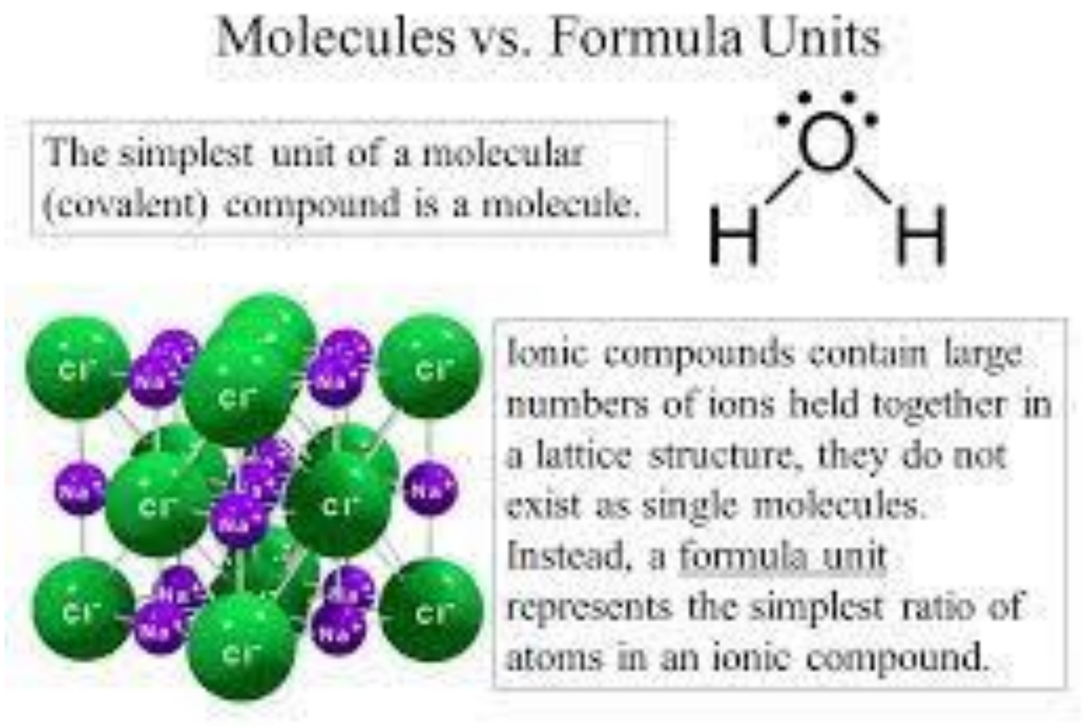


Figure 2.18: Two structural formulas for ammonia.

Covalent vs. ionic compounds: molecules vs. formula unit

- A group of atoms joined by only covalent bonds is known as a molecule.
- The word molecule should only be used in reference to covalent compounds.
- In an ionic compound, such as sodium chloride, there is no such thing as a single molecule of sodium chloride since, in reality, sodium chloride is actually made up of multiple sodium and chloride ions joined together in a large crystal lattice.
- Refer to one piece of NaCl not as a molecule but as a formula unit.
- Single formula units, unlike single molecules, do not exist in nature—formula units used for ease of reference and convenience.



Symbols

- As defined earlier, a chemical symbol is a letter or two-letter word representing the name of the element.
- The first letter is always written on capital letter and followed by a small letter for a two-word chemical symbol.
- It should not bear a period.
- The first scientist to use symbols for chemical elements was Dalton.
- He was the one who used the symbols of the elements in a specified quantity, i.e one atom of the element.
- Later on, Berzilius suggested that we could use one or two letters of the element to form its symbol.
- Initially, symbols for the name of elements were derived from the place they originated from.
- So **copper** was taken from **Cyprus**, while **gold** was meant **yellow** and so it was denoted accordingly.
- However, today the **International Union of Pure and Applied Chemistry** approves the element names.
- For example, the chemical symbol for **Sodium** is **Na** .
- This ‘Na’ is derived from **Natrium**.
- **Iron** is denoted by **‘Fe’** which comes from the Latin word **“Ferrum**”.
- **H** for **Hydrogen**, **Hg** for **mercury**.

Symbols

Name of element	Symbol	Latin name
1. Gold	Au	Aurum
2. Silver	Ag	Argentum
3. Mercury	Hg	Hydrargyrum
4. Copper	Cu	Cuprum
5. Lead	Pb	Plumbum
6. Iron	Fe	Ferrum
7. Sodium	Na	Natrium
8. Potassium	K	Kalium
9. Tin	Sn	Stannum

Name in English	Symbol	Name in English	Symbol
Magnesium	Mg	Hydrogen	H
Aluminium	Al	Nitrogen	N
Calcium	Ca	Oxygen	O
Chromium	Cr	Fluorine	F
Manganese	Mn	Chlorine	Cl
Cobalt	Co	Bromine	Br
Nickel	Ni	Iodine	I
Zinc	Zn	Carbon	C
Silver	Ag	Sulphur	S
Barium	Ba	Phosphorus	P
Tungsten	W	Boron	B
Platinum	Pt	Silicon	Si
Radium	Ra	Germanium	Ge
Uranium	U	Arsenic	As
Helium	He	Antimony	Sb
Argon	Ar	Neon	Ne

Oxidation Number

- Ions are charged atoms.
- Cations are positively charged while anions are negatively charged.
- **Monoatomic ion** is either cation or anion of a single atom.
- The charge of an atom is written as the superscript and this is known as the **oxidation number.**
- The **polyatomic ion**, is a group of atoms carrying a charge either positive or negative.
- Table 2.3 and Table 2.4 list the oxidation numbers of monoatomic ions and polyatomic ions, respectively.

Element	Symbol	Charge	As written
sodium	Na	+1	Na^{+1}
chlorine	Cl	-1	Cl^{-1}
carbonate	C, O	-2	CO_3^{-2}

Table 2.3 Oxidation Numbers of Monoatomic Ions

I	II											III	IV	V	VI	VII
H +1																
Li +1	Be +2											B +3	C +4 +2 -4	N +5 +4 +3 +2 +1 -3	O -2	F -1
Na +1	Mg +2											Al +3	Si +4	P +5 +3 -3	S +6 +4 -2	Cl +7 +5 +3 +1 -1
K +1	Ca +2	Sc +3	Ti +4 +3	V +5 +4 +3 +2	Cr +6 +3 +2	Mn +7 +4 +3 +2	Fe +3 +2	Co +3 +2	Ni +2	Cu +2 +1	Zn +2	Ga +3 +1	Ge +4 +2	As +5 +3 -3	Se +6 +4 -2	Br +7 +5 +3 +1 -1
Rb +1	Sr +2									Ag +1	Cd +2	In +3 +1	Sn +4 +2	Sb +5 +3	Te +6 +4 -2	I +7 +5 +3 +1 -1

Table 2.4 Oxidation Numbers/Charges of Polyatomic Ions

+1 CHARGE		−1 CHARGE		−2 CHARGE		−3 CHARGE	
ion	**Name**	**ion**	**Name**	**ion**	**name**	**ion**	**name**
NH_4^+	ammonium	$H_2PO_3^-$	dihydrogen phosphite	HPO_3^{2-}	hydrogen phosphite*	PO_3^{3-}	phosphite
H_3O^+	hydronium	$H_2PO_4^-$	dihydrogen phosphate	HPO_4^{2-}	hydrogen phosphate*	PO_4^{3-}	phosphate
Hg_2^{2+}	mercury(I)	HCO_3^-	hydrogen carbonate*	CO_3^{2-}	carbonate	PO_2^{3-}	hypophosphite
		HSO_3^-	hydrogen sulfite*	SO_3^{2-}	sulfite	AsO_3^{3-}	arsenite
		HSO_4^-	hydrogen sulfate*	SO_4^{2-}	sulfate	AsO_4^{3-}	arsenate
		NO_2^-	nitrite	$S_2O_3^{2-}$	thiosulfate	$Fe(CN)_6^{3-}$	ferricyanide
		NO_3^-	nitrate	SiO_3^{2-}	silicate		
		OH^-	hydroxide	C_2^{2-}	carbide		
		CH_3COO^-	acetate	$C_2O_4^{2-}$	oxalate		
		CrO_2^-	chromite	CrO_4^{2-}	chromate		
		CN^-	cyanide	$Cr_2O_7^{2-}$	dichromate		
		CNO^-	cyanate	$C_4H_4O_6^{2-}$	tartrate		
		CNS^-	thiocyanate	MoO_4^{2-}	molybdate		
		O_2^-	superoxide	O_2^{2-}	peroxide		
		MnO_4^-	permanganate	S_2^{2-}	disulfide		
		ClO^-	hypochlorite				
		ClO_2^-	chlorite	−4 CHARGE			
		ClO_3^-	chlorate	$P_2O_7^{4-}$		pyrophosphate	
		ClO_4^-	perchlorate	$Fe(CN)_6^{4-}$		ferrocyanide	
		BrO^-	hypobromite				
		BrO_2^-	bromite				
		BrO_3^-	bromate				
		BrO_4^-	perbromate				
		IO^-	hypoiodite				
		IO_2^-	iodite				
		IO_3^-	iodate				
		IO_4^-	periodate				
		AlO_2^-	aluminate				
		N_3^-	azide				

*Note:

The following hydrogenated oxyanions maybe name also with prefix "**bi**" instead of "hydrogen".

hydrogen carbonate or bicarbonate
hydrogen sulfite or bisulfite
hydrogen sulfate or bisulfate
hydrogen phosphite or biphosphate
hydrogen phosphate or biphospahte

Determination of the Oxidation Numbers of an Atom in a Polyatomic Ion

- The oxidation number of atoms are listed in the given table above.
- There are atoms with variable oxidation numbers or having more than one oxidation number.
- Take the case for example: $\mathbf{Cr_2O_7^{-2}}$
- It is a dichromate polyatomic ion carrying a -2 charge.
- How do we determine the oxidation number of Cr?

Steps in the Determination of Oxidation Number

1. Free elements or neutral elements are assigned an oxidation number of zero (0).

    A neutral compound also carries a charge of zero.

    Cl in its free state is Cl^-, and NaCl is a neutral compound is assigned a charge of zero: $NaCl^0$
2. Hydrogen in covalent compounds, the oxidation number is +1. In hydrides, hydrogen is -1.
3. Oxygen in covalent compounds, oxidation number is -2. The peroxide, it has -1.
4. The oxidation numbers of a monoatomic ion and polyatomic ion are equal to the charge they carry.
5. In covalent compounds, since it contains two nonmetal atoms, the atom which is more electronegative carries the negative charge.

    Example: OCl is a covalent compound between two elements in the Group VIIA.

    Both are considered negative monoatomic ions.

    However, Cl is more electronegative than O, therefore Cl is considered a negative ion
6. The sum of the oxidation numbers of the atoms in a polyatomic ion is equal to its charge

Determination of the Oxidation Numbers of an Atom in a Polyatomic Ion

Example:

- What is the oxidation number of **Cr** in **$K_2Cr_2O_7$?**

Solution:

Referring to Table 2.3:

The oxidation number of K is +1.
The oxidation number of O is -2.

Solving for the oxidation number of Cr:

(No. of K x oxidation number of K) + (No. of Cr x oxidation number of Cr) = charge

- No. of K = 2
- No. of Cr = 2
- No. of O = 7

Determination of the Oxidation Numbers of an Atom in a Polyatomic Ion

Example:

- What is the oxidation number of **Cr** in $\mathbf{K_2Cr_2O_7}$ **?**

Solution: … cont.

Charge = 0 ($K_2Cr_2O_7$ is a neutral compound)

$[(2)x(+1)] + [(2)x(Cr)] + [(7)x(-2)] = 0$

$+2 + 2Cr – 14 = 0$

$2Cr – 12 = 0$

$2Cr = 0 + 12$

$Cr = +12/2$

$= +6$

Determination of the Oxidation Numbers of an Atom in a Polyatomic Ion

Example

- What is the oxidation number of **Cr** in $\mathbf{CrO_4^{-2}}$?

Solution

The oxidation of O is -2.

No. of Cr = 1

No. of O = 4

Charge = -2

Solving for the oxidation number of Cr:

$$[(1)x(Cr)] + [(4)x(-2)] = -2$$

$$Cr - 8 = -2$$

$$Cr = -2 + 8$$

$$= +6$$

Chemical Formula

- A chemical symbol represents the name of the chemical element.
- **Chemical formula** serves as the shortest way of writing the name of the chemical compound.
- It shows the number of atoms of each kind.

Compounds	Formula
Carbon dioxide	CO_2
Magnesium Oxide	MgO
Sodium Chloride	$NaCl$
Copper Sulphate	$CuSO_4$
Sulphuric Acid	H_2SO_4
Nitric Acid	HNO_3
Sodium Hydroxide	$NaOH$
Sulphur dioxide	SO_2
Water	H_2O
Ammonia Gas	NH_3
Sodium Chloride	$NaCl$
Potassium Chloride	KCl
Carbon dioxide	CO_2
Magnesium Chloride	$MgCl_2$
Hydrogen Sulphide	H_2S
Methane	CH_4
Hydrochloric acid	HCl

Steps for writing a chemical formula

Step 1: First, you have to decide the type of the bond.

- If the prefixes are used, then it is a covalent bond.
- In case there are no prefixes, it is an ionic bond.
- After deciding move to Step number 2.

Step 2: Write the correct symbol element and the correct formula for the polyatomic ions.

Step 3: Identify the cation and the anion.

- Always remember to write the cation first, followed by the anion.

Step 4: Determine the oxidation numbers of the cation and the anion by writing as superscripts.

Step 5. Remove the signs and cross-multiply the oxidation numbers – the superscript of the cation becomes the subscript of the anion and the superscript of the anion becomes the subscripts of the cation.

Steps for writing a chemical formula

Step 6. Write the correct chemical formula and examine the subscripts.

a. For ions with subscripts 1, no need to write 1 (understood to have one).

NaCl – Sodium has a subscript of 1 from the oxidation number of chlorine which is -1 and chlorine has a subscript of 1 from the oxidation number of sodium which is +1.

$BaCl_2$ – Ba has a subscript of 1 from chlorine which is -1, and Cl has a subscript of 2 from the oxidation number of Ba which is +2.

b. Reduce the subscript to the lowest term possible.

BaS – Ba has +2 and S has -2, thus, Ba_2S_2. Reducing the subscripts – BaS.

SnO_2 – Sn has +4 and O has -2. Sn_2O_4. Reducing the subscripts – SnO_2

Steps for writing a chemical formula

c. Polyatomic ions taken more than once, should be enclosed in a parenthesis.

$Fe(OH)_3$ – Fe is +3 while a polyatomic OH is -1. Since OH has a subscript of 3, it should be enclosed in a parenthesis.

$(NH_4)_2CO_3$ - NH_4 a polyatomic ion with +1 and CO_3 is a polyatomic ion with -2. NH4 has a subscript of 2 and is enclosed in a parenthesis.

d. Polyatomic ions taken more than once should be enclosed in parenthesis, then in a bracket.

$Al_4[Fe(CN)_6]_3$ – Al is +3 and $Fe(CN)_6$ is -4. A bracket is used to enclose the polyatomic $Fe(CN)_6$ as it already contains a parenthesis.

Chemical Formula

Example:

Write the chemical formula.

- Calcium and sulfate
- Magnesium and phosphate
- Aluminum and sulfur

Naming of Inorganic Compounds

- Naming of inorganic and organic compounds follow the system of nomenclature or system of naming set by International Union of Pure and Applied Chemistry (IUPAC).
- Organic compounds are discussed separately.

Common Examples of Inorganic Compounds Significant in Physiology

- Sodium chloride:
  - $NaCl$
- Calcium phosphate
  - $(Ca)_3(PO_4)_2$
- Potassium iodide
  - KI
- Sodium bicarbonate
  - $Na(HCO_3)$
- Hydrochloric acid
  - HCl
- Sodium hydroxide
  - $NaOH$
- Magnesium hydroxide
  - $Mg(OH)_2$
- Calcium chloride
  - $CaCl_2$

Binary Compounds

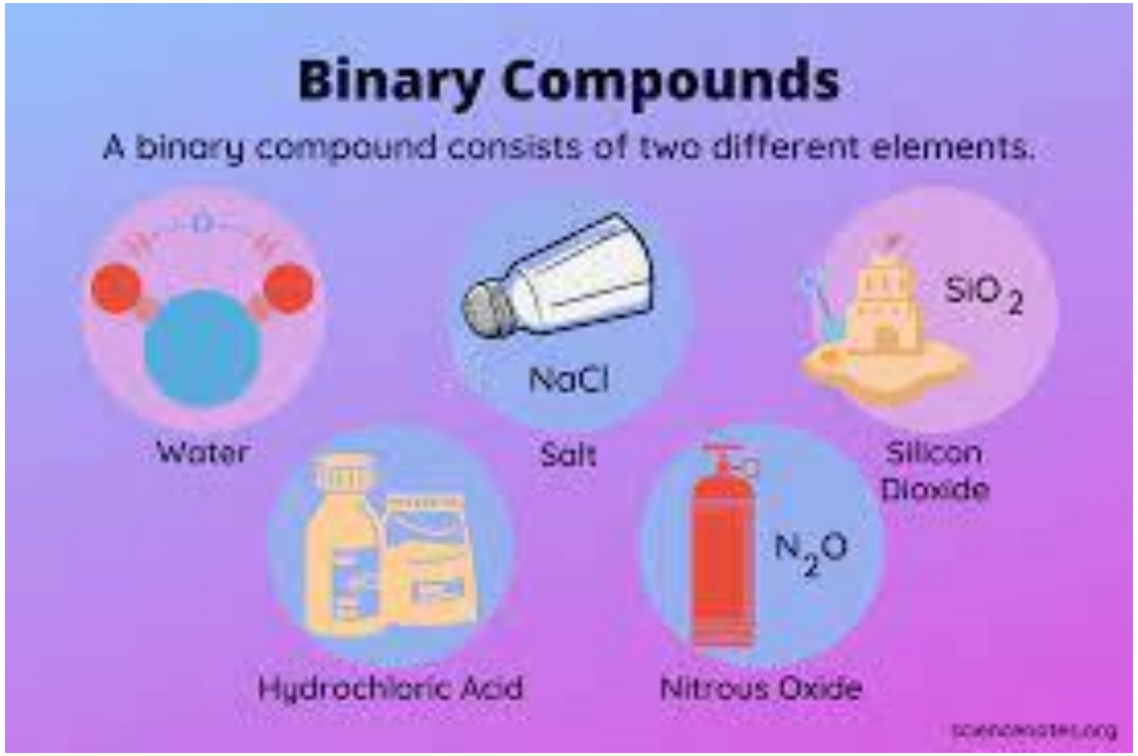
- Binary from the prefix “bi” meaning 2.
- So binary compounds are compounds containing only two elements.
- There are three types of binary compounds:

Type I – an ionic compound formed between a metal that forms only one cation and an anion.

Type II – an ionic compound formed between a metal with more than one type of cations and an anion.

Type III – a covalent compound formed between two non-metals.

- All binary compounds are named ending in “ide.”



Type I Binary Compounds

For Type I binary compounds, they are usually from cations of Group 1 and Group 2.

- A cation takes its name from name of the element.
- An anion is named by taking the root of the name of the element and changing the ending to "ide".

Examples: Give the name of the compound

- Na_2S
- MgI_2

Type II Binary Compound

Type II binary compounds are named in two ways:

- Systematic or Stock System
  - The name of the cation takes its name from the name of the element.
  - A Roman numeral is added enclosed in parenthesis after the name of the cation.
  - The Roman numeral indicates the oxidation number of the metal cation.
  - An anion is named by taking the root of the name of the element and changing the ending to “ide”.
  - Examples:
    - Tin (II) for Tin +2
    - Tin (IV) for Tin +4
- Old System
  - The cation is named from the root of the Latin name of the cation (metal) changing the ending “um” to ”ous” for lowest oxidation state, and “ic” for higher/highest oxidation state.
  - Examples:
    - Tin +2 – from stannum to stannous
    - Tin +4 – from stannum to stannic

Examples:

Give the name of the compound using the Systematic and the Old Systems.

- CuS
- $PbCl_2$

Type III Binary Compounds – The Binary Covalent Compounds

This type of compounds are composed of two non-metals. They are named based on the following:

- The first element is named first taking the name of the first element.
- The second element is named taken from the name of the second element ending in “ide”.
- Greek prefixes are added to denote the number of atoms for both elements.
- However a prefix “mono” is never used for the first element with only one atom.

Table 2.4 Greek Prefixes

Number	Prefix	Number	Prefix
1	Mono	6	Hexa
2	Di	7	Hepta
3	Tri	8	Octa
4	Tetra	9	Nona
5	Penta	10	Deca

Type III Binary Compounds – The Binary Covalent Compounds

If the ending of the name of the first element is a vowel (o or a), and the first letter of the second elements is also a vowel (a or o), the ending vowel (a or o) is dropped at the end of the prefix.

- Nonaoxide – the vowel a in the prefix Nona is dropped giving nonoxide.
- Monooxide is monoxide
- Pantaoxide – the vowel a at the end of the prefix Penta is dropped giving pentoxide.
- The vowel "i" at the end of tri and di are never dropped

Examples:

Give the name of the compound:

- SO_3
- N_2O

Hydrates

- Some ionic compounds occur as hydrates.
- Hydrates are compounds containing number of loosely bound water molecules.
- The water molecule is called the water of hydration.
- To name a hydrate, the name of the main chemical compound is named first followed followed by a Greek prefix added to the word hydrate.

Example:

- Give the name of the hydrate, $Ba(OH)_2 \cdot 5H_2O$

Naming Compounds of Polyatomic Ions

- Polyatomic ions are molecules containing two or more atoms carrying a charge.
- Oxyanions (or oxoanions) are polyatomic ions containing oxygen.

a. With four oxyanions and a halogen bonded to O:

Prefix	Suffix	Number of O Atoms	Superlative/Comparative	Example	Name
Per (greater than common)	-ate	4	Largest number	ClO_4^-	perchlorate
Common	-ate	3	Large number	ClO_3^-	chlorate
Less than one than common	-ite	2	Smaller number	ClO_2^-	chlorite
Hypo (lesser than one) – lowest (only one)	-ite	1	Smallest number	ClO^-	hypochlorite

Naming Compounds of Polyatomic Ions

b. With two oxyanions

- ions with most number of O, the name of the non-metal ending in suffix “ate”.
- ions with least number of O, the name of the non-metal ending in suffix “ite”.
  - SO_4^{-2} and SO_3^{-2}
  - SO_4^{-2} has more O atoms than SO_3^{-2}
  - SO_4^{-2} is named sulfate and SO_3^{-2} is named sulfite

Example:

- Give the name of the compound, $Ca(ClO_3)_2$
- The formula $Mg(NO_2)_2$. Is it magnesium nitrate or magnesium nitrite?

Naming of Inorganic Acid

Naming Binary Acid

- Binary acids are compounds containing a non-metal covalently bonded to hydrogen.
- They are named starting with the prefix “hydro” followed by the root of the acid forming element (usually the non-metal) and the suffix “ic” and word acid.

Name of Acid
$HF(aq)$, hydrofluoric acid
$HCl(aq)$, hydrochloric acid
$HBr(aq)$, hydrobromic acid
$HI(aq)$, hydroiodic acid
$H_2S(aq)$, hydrosulfuric acid

Naming of Oxyacids

- Oxyacids are acids containing oxygen atom bonded to hydrogen and another element.
- Also known as oxoacids or ternary acids.
- To name oxyacids:
  - Remove the word hydrogen, then begin with the root name of the anion, replacing the suffix – ate with –ic and the suffix –ite with –ous and add the word acid

Example:

Give the name of the name of the oxyacid, $H_2SO_{3(aq)}$

Note:

Acids have always the subscript (aq) for aqueous in their formula

Atoms, Ions & Molecules

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end