ADAMA SCIENCE AND TECHNOLOGY UNIVERSITY

General Chemistry (Chem1101)

Introduction to the study of modern chemistry

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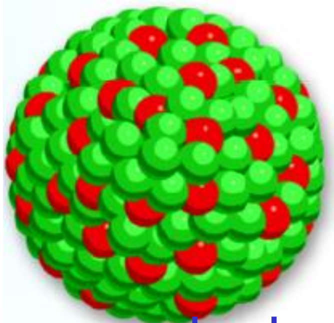
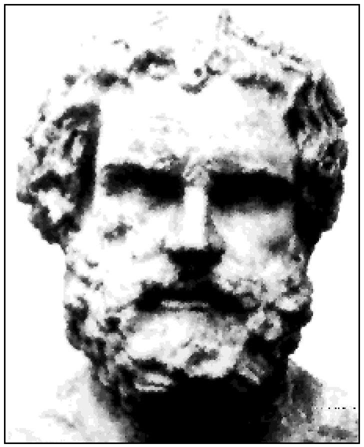
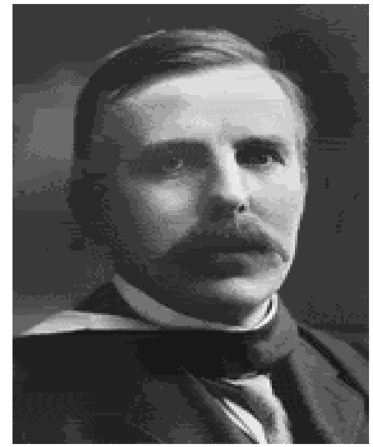
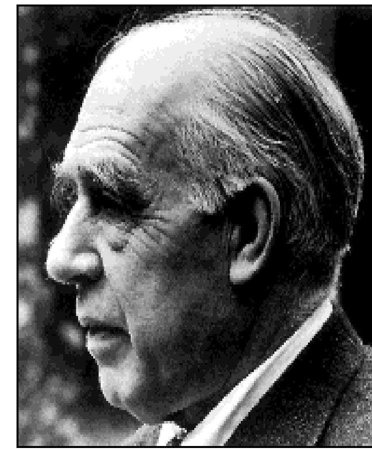


How and why has the atomic theory changed over time?







How engineers use their knowledge of atoms to create new technologies?

Modern Atomic Theory

Dalton's Atomic Theory (1808)

1. Elements are composed of extremely small particles called **atoms**.

2. All **atoms** of a given element are identical, having the same size, mass and chemical properties. The atoms of one element are different from the atoms of all other elements.

3. **Compounds** are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.

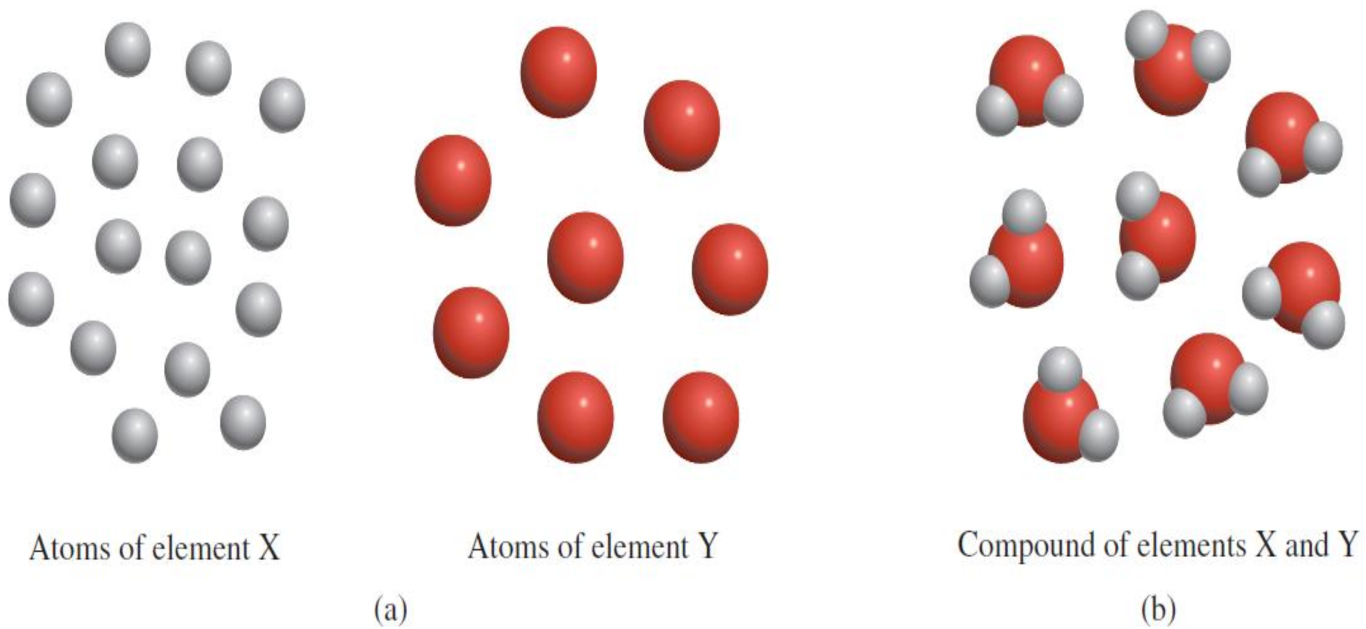
4. A **chemical reaction** involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.

✓Dalton’s atomic theory hypothesis supports (is in agreement with) the following laws:

❖**The Law of Definite Proportion**: States that different samples of the same compound always contain its constituent elements in the same proportion by mass.

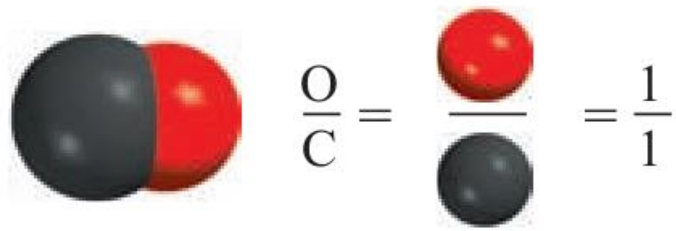
❖Irrespective of sources of the compound!

(a) According to Dalton’s atomic theory, atoms of the same element are identical, but atoms of one element are different from atoms of other elements. (b) Compounds formed from atoms of elements X and Y. In this case, the ratio of the atoms of element X to the atoms of element Y is 2:1.

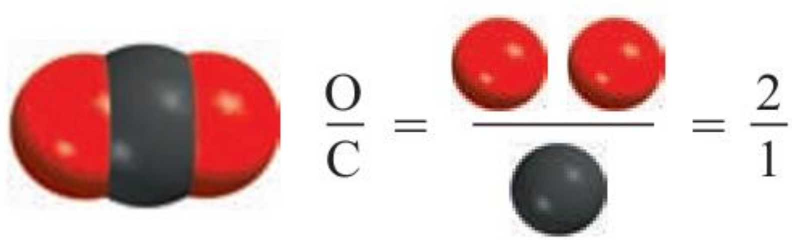


❖**The Law of Multiple Proportions**: States that if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.

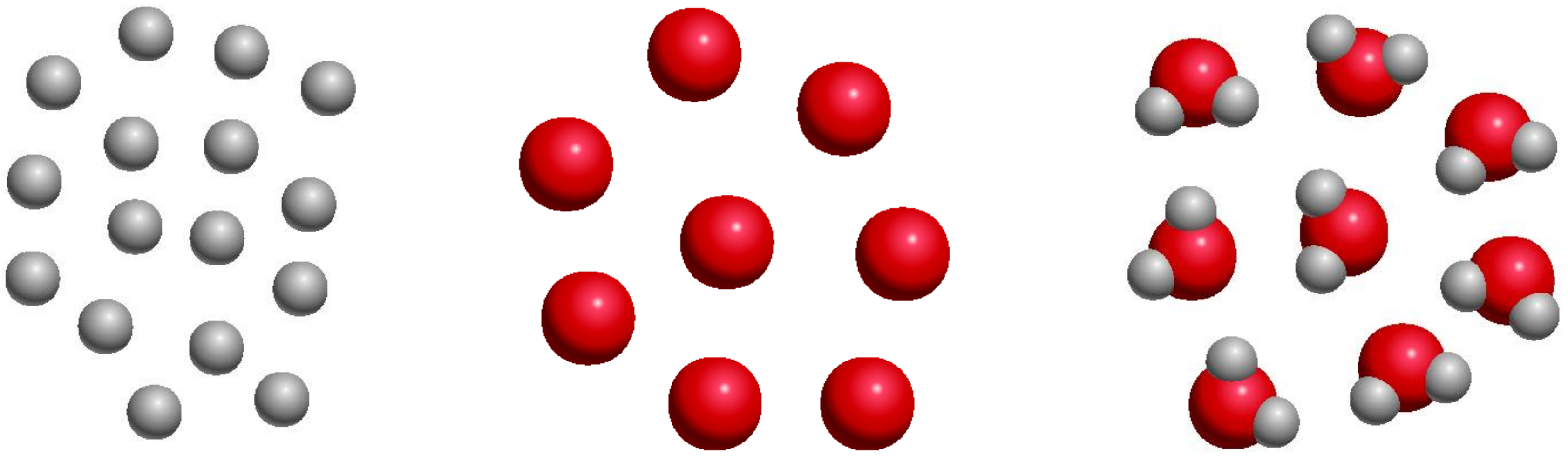
Carbon monoxide



Carbon dioxide



❖**Law of Conservation of Mass:** States that matter can be neither created nor destroyed.



Atoms of element X　　Atoms of element Y　　Compounds of elements X and Y

$$16\ X \quad + \quad 8\ Y \longrightarrow 8\ X_2Y$$

The Structure of the Atom

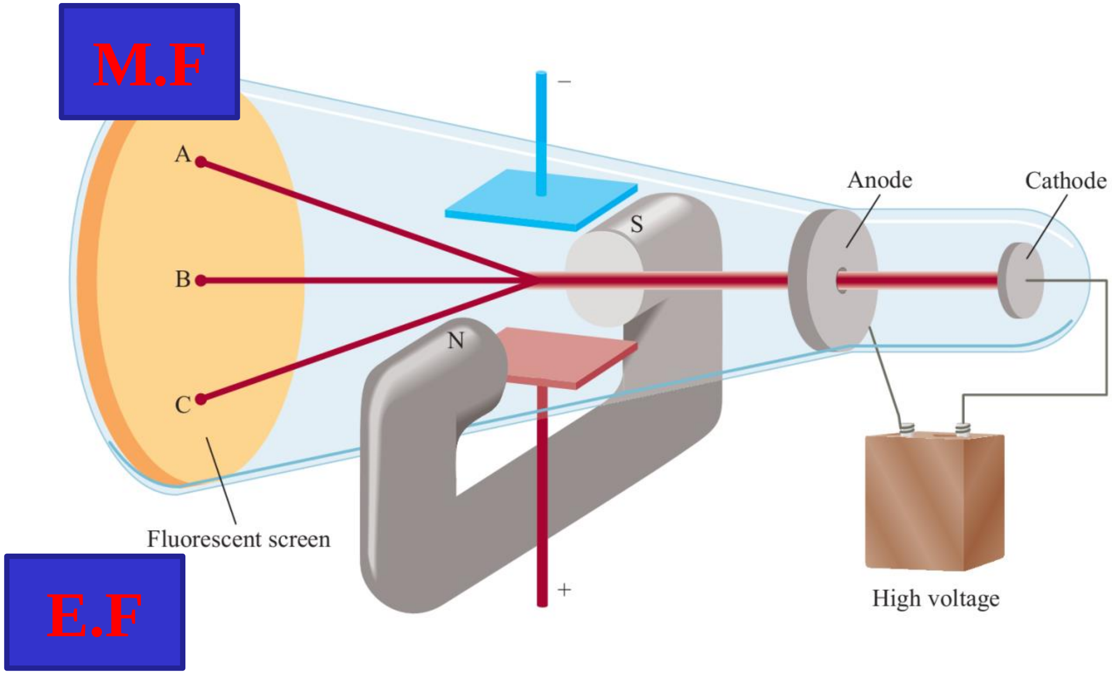
On the basis of Dalton's atomic theory:

✓an atom is the basic unit of an element that can enter into chemical combination

✓an atom is both extremely small and indivisible

➢However, a series of investigations that began in the 1850s and extended into the twentieth century clearly demonstrated that atoms actually possess internal structure; that is, they are made up of even smaller particles, which are called subatomic particles. This research led to the discovery of three such particles-electrons, protons, and neutrons.

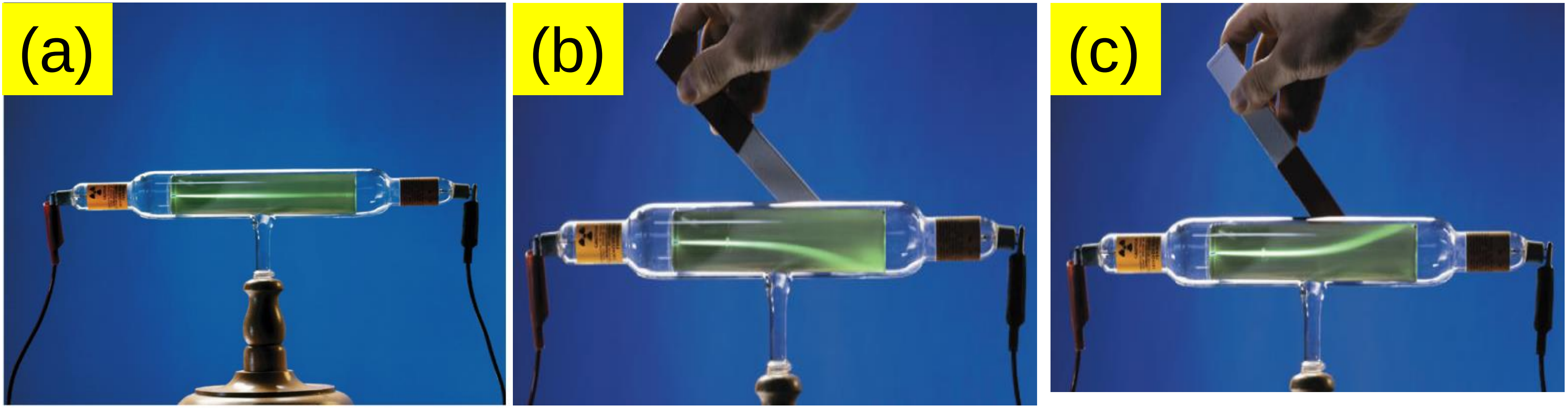
The Discovery of Electron



A cathode ray tube with an electric field perpendicular to the direction of the cathode rays and an external magnetic field. The symbols N and S denote the north and south poles of the magnet. The cathode rays will strike the end of the tube at A in the presence of a magnetic field, at C in the presence of an electric field, and at B when there are no external fields present or when the effects of the electric field and magnetic field cancel each other.

- Because the cathode ray is attracted by the plate bearing positive charges and repelled by the plate bearing negative charges, it must consist of negatively charged particles
- These negatively charged particles are electrons!

Effect of a bar magnet on the cathode ray



(a) A cathode ray produced in a discharge tube traveling from the cathode (left) to the anode (right). The ray itself is invisible, but the fluorescence of a zinc sulfide coating on the glass causes it to appear green. (b) The cathode ray is bent downward when the north pole of the bar magnet is brought toward it. (c) When the polarity of the magnet is reversed, the ray bends in the opposite direction

Cathode rays: properties

- ✓ Cathode rays originate from cathode
- ✓ Cathode rays travel in straight line
- ✓ Cathode rays are deflected by electric field
- ✓ Cathode rays are deflected by magnetic field
- ✓ These rays consist of material particles
- ✓ Cathode rays consist negatively charged particles called "electron"

❖J. J. Thomson, used a cathode ray tube and his knowledge of electromagnetic theory to determine the ratio of electric charge to the mass of an individual electron. The number he came up with is -1.76×10^{8} C/g, where C stands for coulomb, which is the unit of electric charge.

❖In a series of experiments carried out between 1908 and 1917, R. A. Millikan, an American physicist, found the charge of an electron to be -1.6022×10^{-19} C. From these data he calculated the mass of an electron

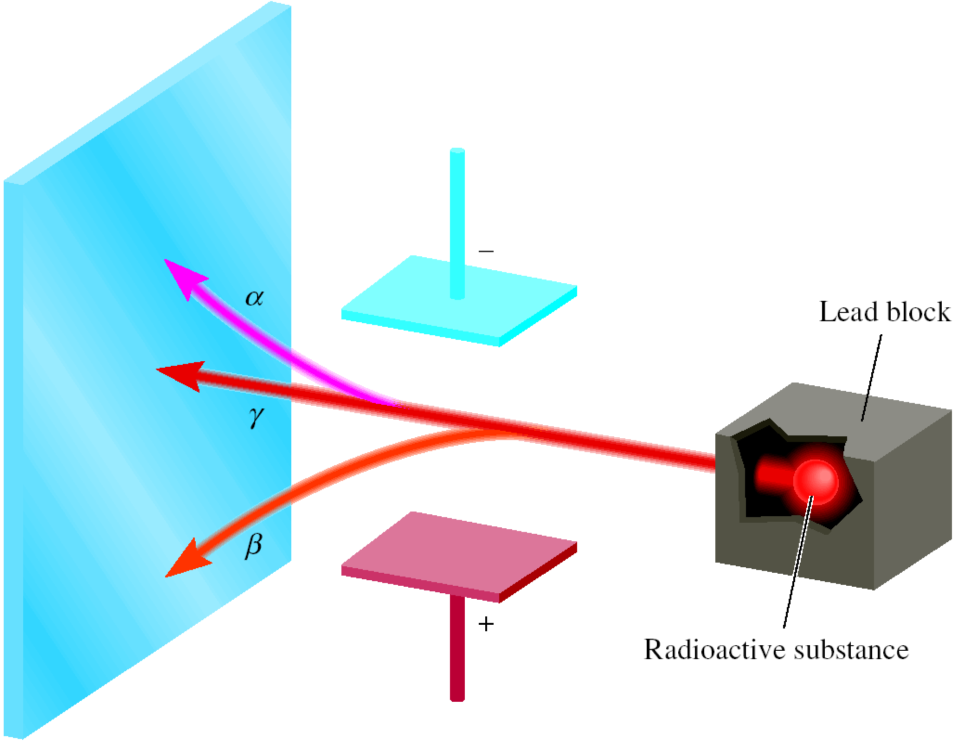
Radioactivity

❖In 1895, the German physicist Wilhelm Röntgen noticed that cathode rays caused glass and metals to emit very unusual rays. This highly energetic radiation penetrated matter, darkened covered photographic plates, and caused a variety of substances to fluoresce. Because these rays could not be deflected by a magnet, they could not contain charged particles as cathode rays do. Röntgen called them X rays.

❖Not long after Röntgen's discovery, Purely by accident, Antoine Becquerel, found that exposing thickly wrapped photographic plates to a certain uranium compound caused them to darken, even without the stimulation of cathode rays.

❖One of Becquerel's students, Marie Curie, suggested the name radioactivity to describe this spontaneous emission of particles and/or radiation. Consequently, any element that spontaneously emits radiation is said to be radioactive.

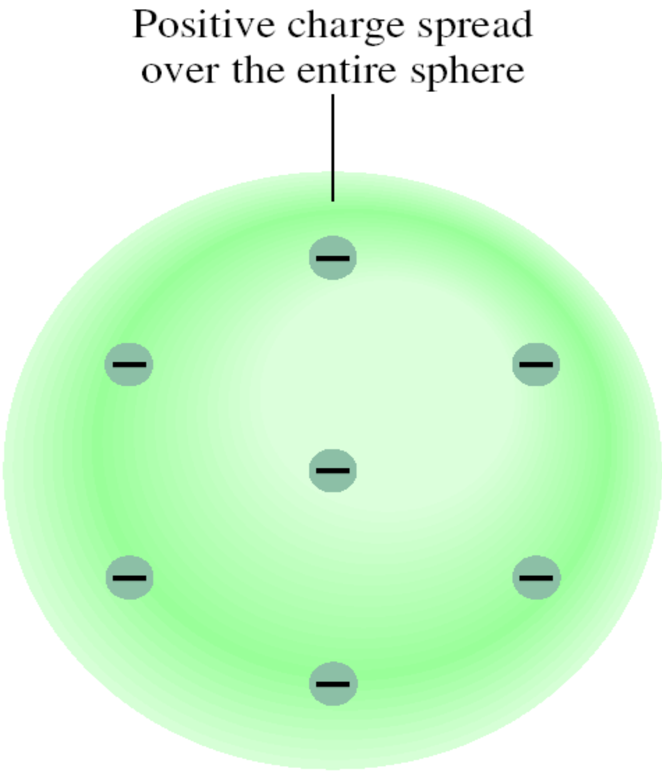
Three types of rays emitted by radioactive elements. β rays consist of negatively charged particles (electrons) and are therefore attracted by the positively charged plate. The opposite holds true for α rays—they are positively charged and are drawn to the negatively charged plate. Because γ rays have no charges, their path is unaffected by an external electric field.



Thomson's Atomic Model

❖By the early 1900s, two features of atoms had become clear:

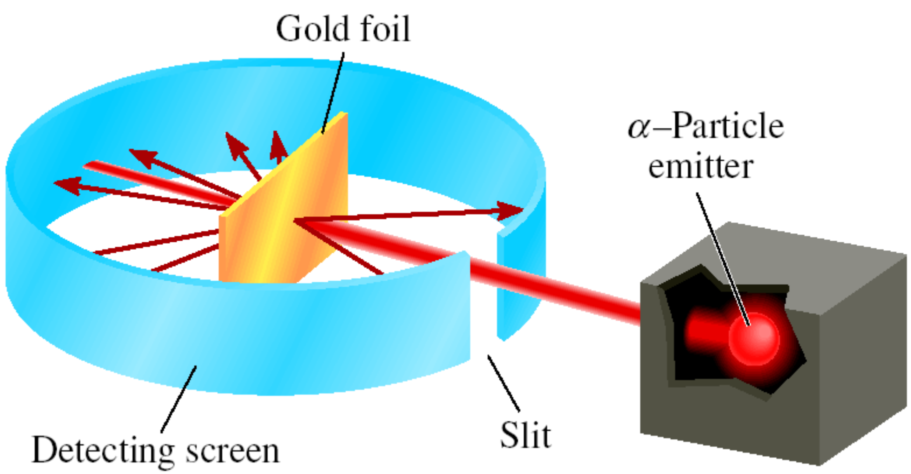
➢they contain electrons,
➢they are electrically neutral

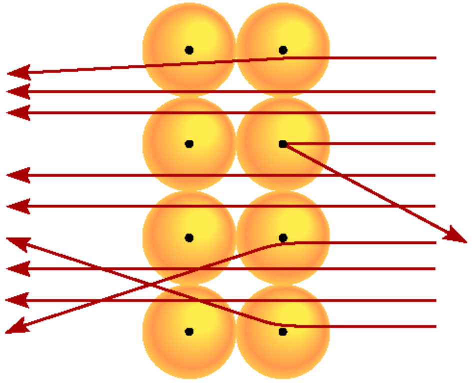


❖On the basis of this information, Thomson proposed that an atom could be thought of as a uniform, positive sphere of matter in which electrons are embedded

Rutherford's Experiment, Discovery of Proton

(1908 Nobel Prize in Chemistry)





α particle velocity ~ 1.4 x 10^7 m/s
(~5% speed of light)

Observations of Alpha Rays Scattering Experiment

1. Most of the α-particles passed through the metal foil with out any change in their path.

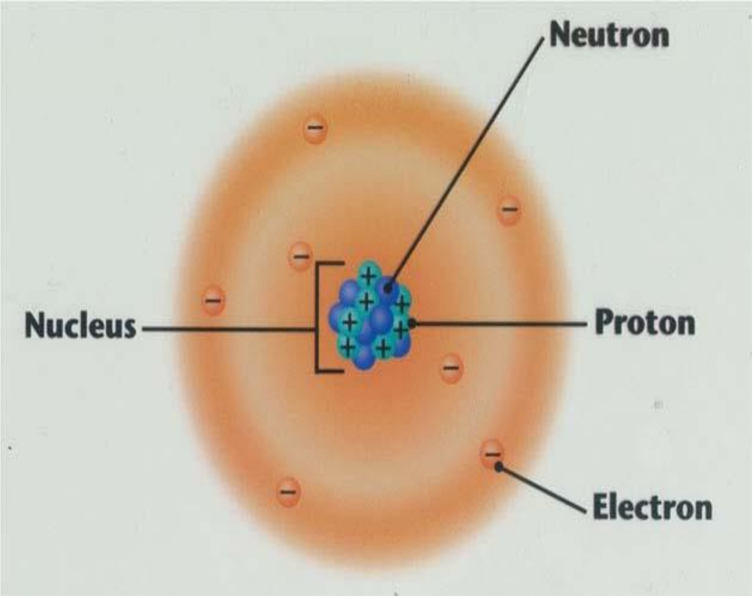
2. A few of the α-particles were deflected through small angles.

3. A very small number of the α-particles were deflected through such large angles that they almost retraced their original path. So,

- Positive charge is concentrated in the nucleus of an atom
- Proton is positively charged particle of an atom
- Most mass of an atom is concentrated in the nucleus
- Nucleus is a dense central core within the atom
- nucleus occupies only about $1/10^{13}$ of the volume of the atom

Rutherford's Model of Atom

- ***Total number of positive charges on the nucleus is equal to the number of electrons.***
- ***Almost the entire mass of the atom is concentrated in the nucleus.***
- ***The volume of the nucleus is very small compared to the volume of the atom.***
- ***Electrons are not stationary. They revolve round the nucleus at extremely high speed.***



"If the atom is the Adama stadium, then the nucleus is a coin on the middle of the stadium."

Chadwick's Experiment, Discovery of neutrons

- The ratio of the mass of a helium atom to that of a hydrogen atom should be 2:1 in reality, however, the ratio is 4:1

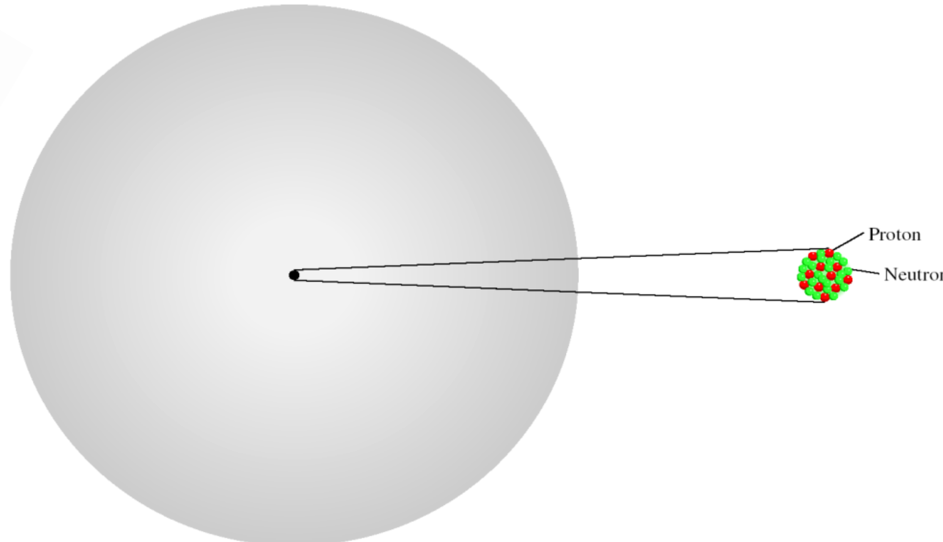
- Rutherford and others postulated that there must be another type of subatomic particle in the atomic nucleus

$$\alpha + {}^{9}Be \longrightarrow {}^{1}n + {}^{12}C + \text{energy } (\gamma\text{-ray})$$

- The γ -rays actually consisted of electrically neutral particles having a mass slightly greater than that of protons

- Chadwick named these particles neutrons (mass ~1.67 x 10^{-24} g)

The protons and neutrons of an atom are packed in an extremely small nucleus. Electrons are shown as "clouds" around the nucleus.



atomic radius ~ 100 pm = 1 x 10^{-10} m

nuclear radius ~ 5 x 10^{-3} pm = 5 x 10^{-15} m

❖Electron, proton, and the neutron are the three fundamental components of the atom that are important in chemistry

Mass and Charge of Subatomic Particles

		Charge	
Particle	Mass (g)	Coulomb	Charge Unit
Electron*	9.10938×10^{-28}	-1.6022×10^{-19}	-1
Proton	1.67262×10^{-24}	$+1.6022 \times 10^{-19}$	$+1$
Neutron	1.67493×10^{-24}	0	0

mass p ≈ mass n ≈ 1840 x mass e^-

Subatomic Particle	Charge	Location
Proton	Positive (+)	Nucleus or "Core"
Neutron	No Charge (0)	Nucleus or "Core"
Electron	Negative (-)	Electron Cloud

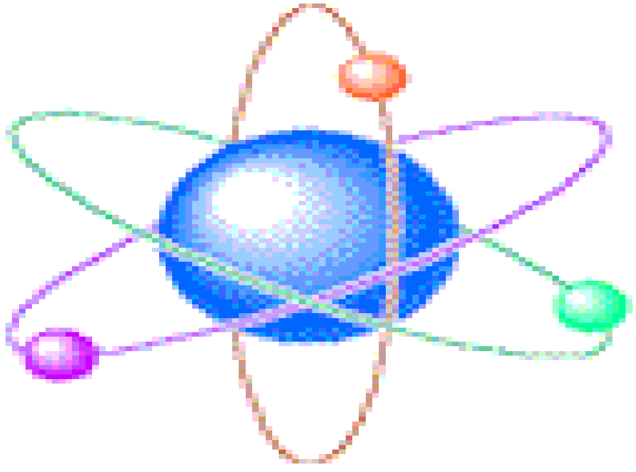
Rutherford-Bohr Model of Atom

- ***Electrons revolve around the nucleus in definite orbits. These are called Stationary states.***

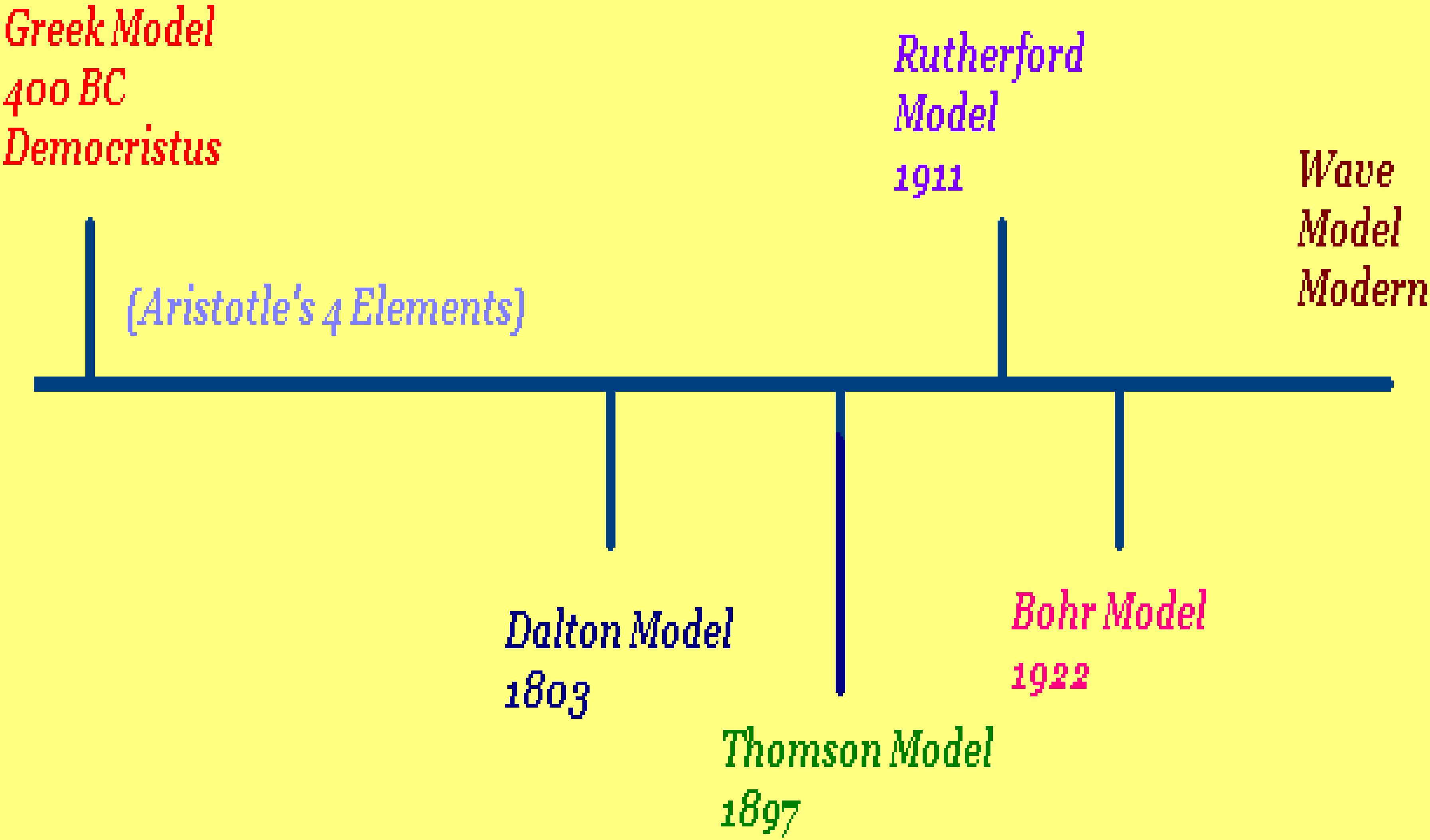
- ***Each stationary state is associated with a definite quantity of energy. Hence these stationary states are also called Energy levels.***

- ***As long as electrons are moving in these stationary states , they do not lose or gain energy.***

- ***Energy is lost or gained by an electron whenever it jumps from one energy level to another .***



Timeline of Atomic Theory



	Indivisible	Electron	Nucleus	Orbit	Electron Cloud	
Greek	X					
Dalton	X					
Thomson		X				
Rutherford		X	X			
Bohr		X	X	X		
Wave		X	X		X	

Atomic number, Mass number and Isotopes

Atomic number (Z) = number of protons in nucleus

Mass number (A) = number of protons + number of neutrons

= atomic number (Z) + number of neutrons

Isotopes are atoms of the same element (X) with different numbers of neutrons (mass number) in their nuclei

Mass Number → A

Atomic Number → Z

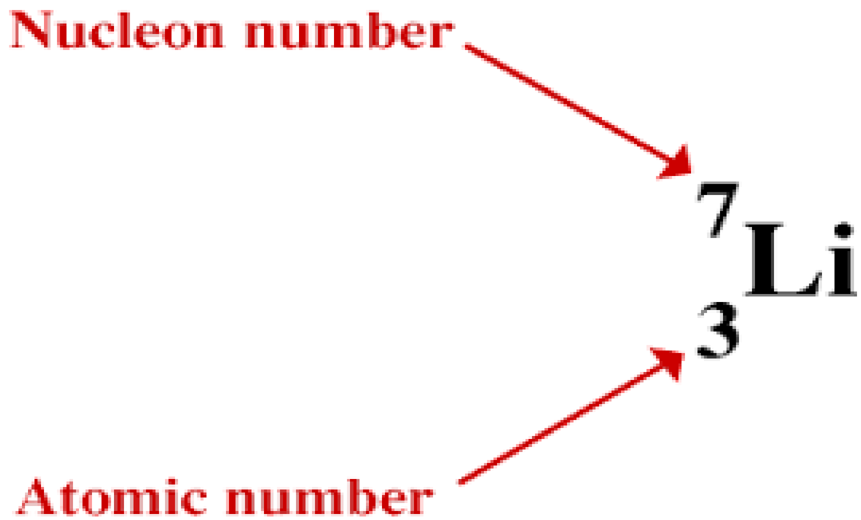
$${}^{A}_{Z}X$$

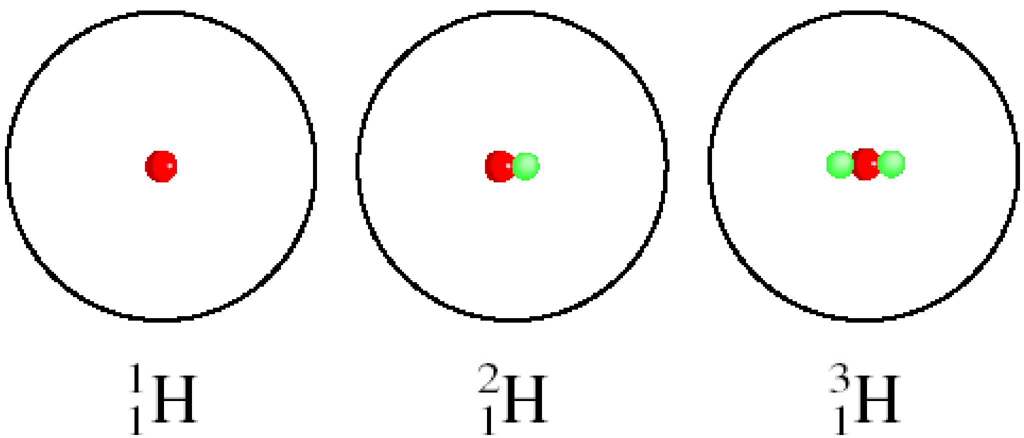
← Element Symbol (Short hand representation of an element)

$${}^{1}_{1}H \qquad {}^{2}_{1}H\ (D) \qquad {}^{3}_{1}H\ (T)$$

$${}^{235}_{92}U \qquad {}^{238}_{92}U$$

The Isotopes of Hydrogen





protium (A = 1) deuterium (A = 2) tritium (A = 3)

How many protons, neutrons, and electrons are in $^{14}_{6}C$?

6 protons, 8 (14 - 6) neutrons, 6 electrons

How many protons, neutrons, and electrons are in $^{12}_{6}C$?

6 protons, 6 (12 - 6) neutrons, 6 electrons

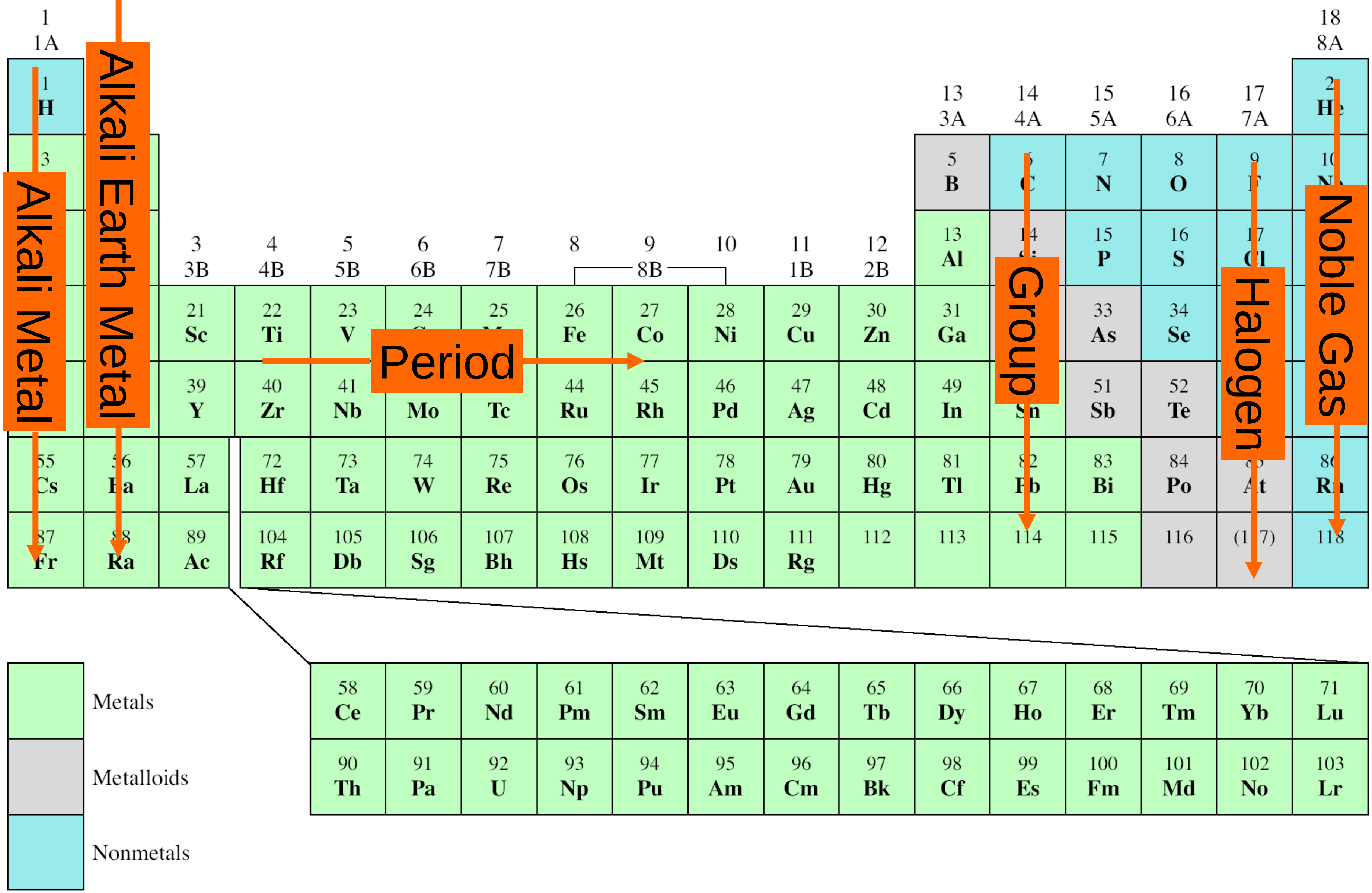
How many protons and electrons are in $^{27}_{13}Al^{3+}$?

13 protons, 10 (13 – 3) electrons

How many protons and electrons are in $^{78}_{34}Se^{2-}$?

34 protons, 36 (34 + 2) electrons

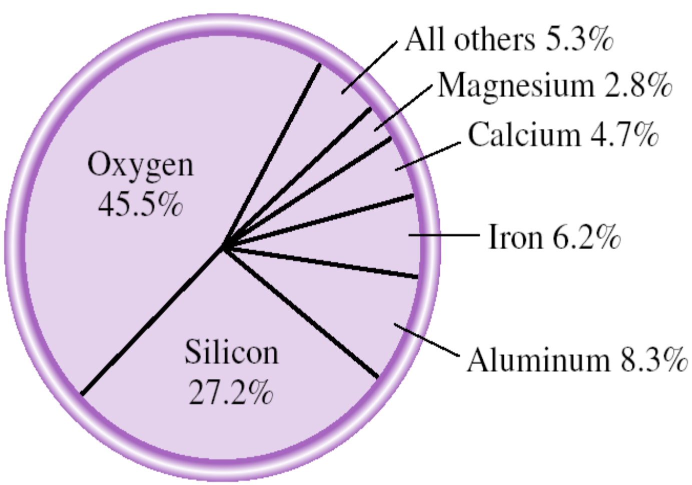
The Modern Periodic Table

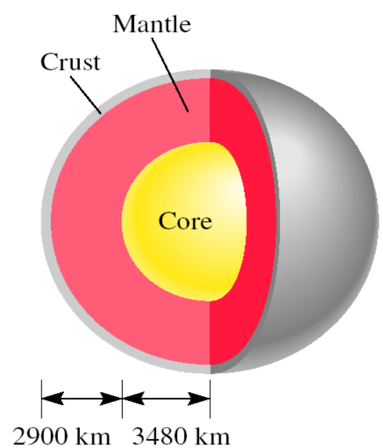


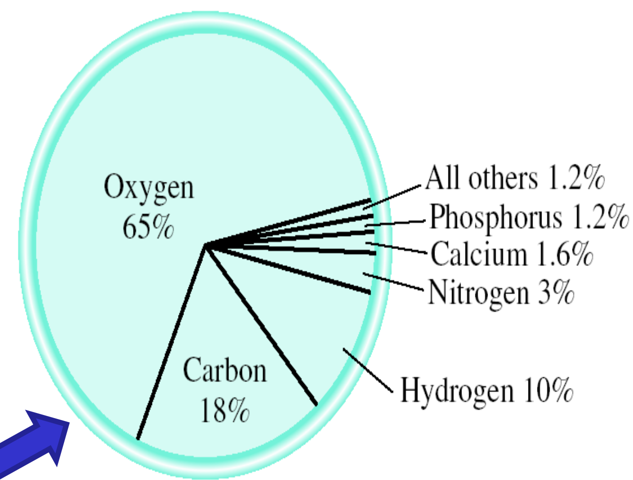
Molecules and Ions

Of all the elements, only the six noble gases in Group 8A of the periodic table (He,Ne, Ar, Kr, Xe, and Rn) exist in nature as single atoms. For this reason, they are called monatomic (meaning a single atom) gases. Most matter is composed of molecules or ions formed by atoms.

Natural abundance of elements in Earth's crust

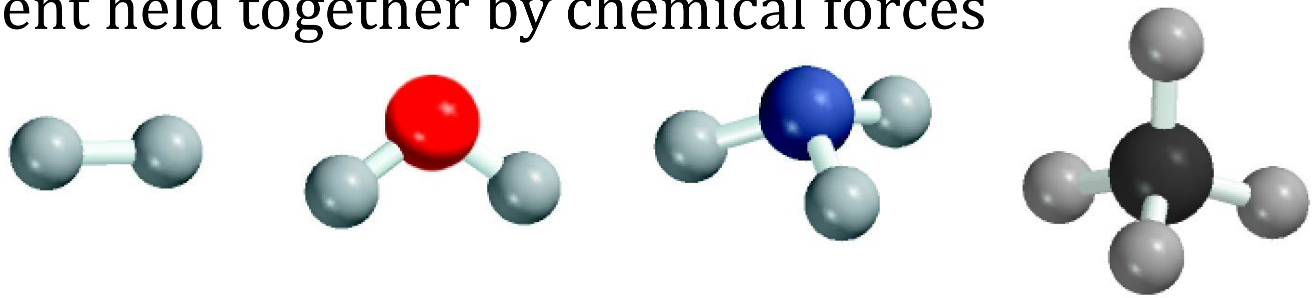






Natural abundance of elements in human body

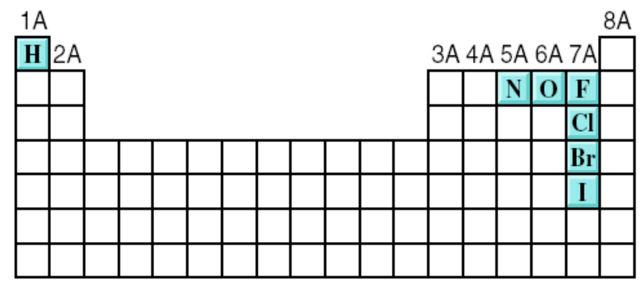
A **molecule** is an aggregate of two or more atoms in a definite arrangement held together by chemical forces



H_2 H_2O NH_3 CH_4

A **diatomic molecule** contains only two atoms

H_2, N_2, O_2, Br_2, HCl, CO, NO



diatomic elements

A **polyatomic molecule** contains more than two atoms

O_3, H_2O, NH_3, CH_4

An ***ion*** is an atom, or group of atoms, that has a net positive or negative charge.

cation – ion with a positive charge

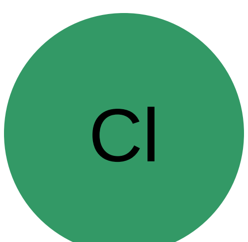
If a neutral atom **loses** one or more electrons it becomes a cation.

11 protons
11 electrons

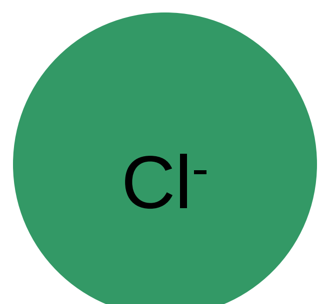
11 protons
10 electrons

anion – ion with a negative charge

If a neutral atom **gains** one or more electrons it becomes an anion.



17 protons
17 electrons



17 protons
18 electrons

A **monatomic ion** contains only one atom

Na^+, Cl^-, Ca^{2+}, O^{2-}, Al^{3+}, N^{3-}

A **polyatomic ion** contains more than one atom

OH^-, CN^-, NH_4^+, NO_3^-

Common Ions Shown on the Periodic Table

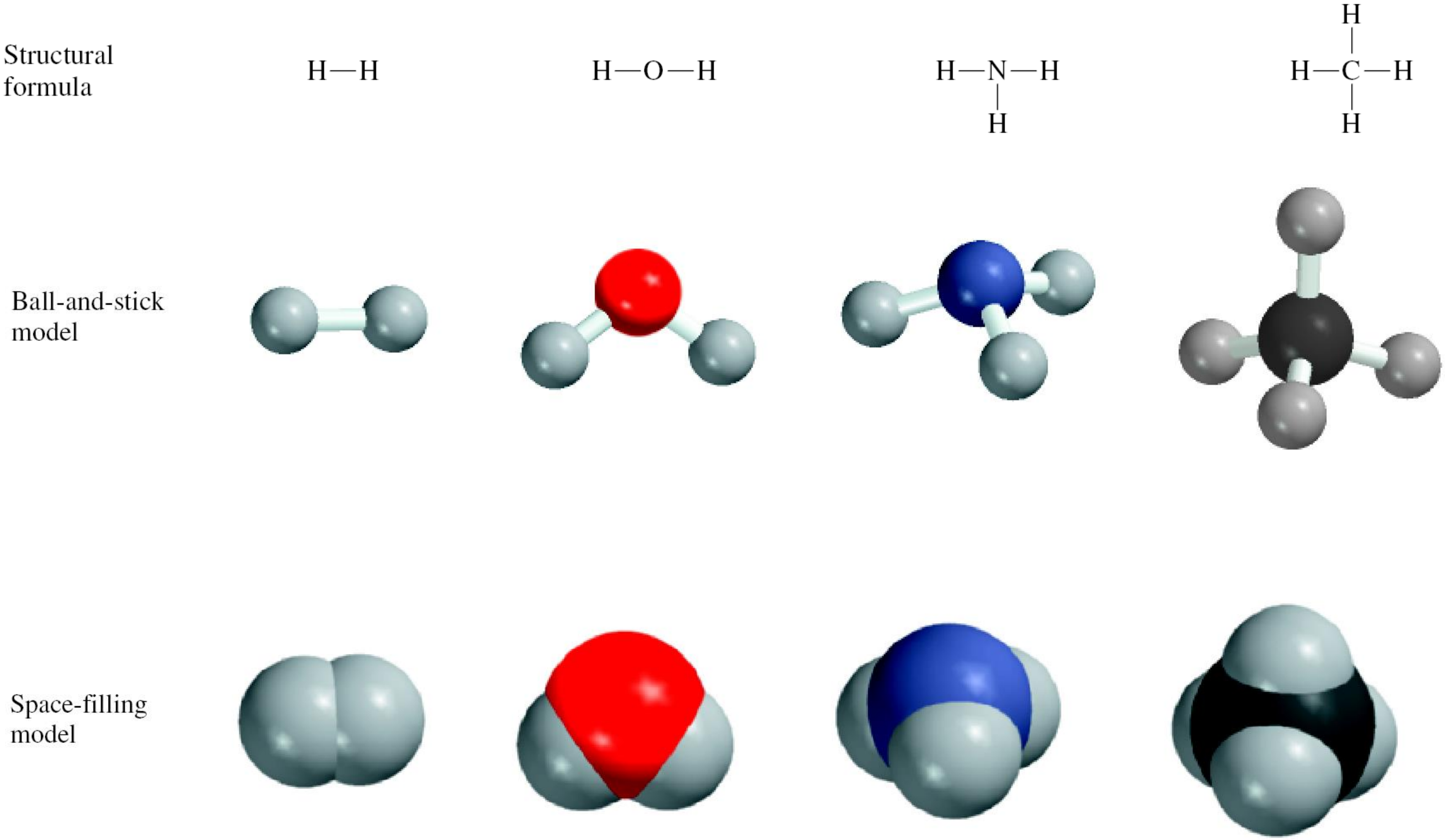
1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
Li^+													C^{4-}	N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}											Al^{3+}		P^{3-}	S^{2-}	Cl^-	
K^+	Ca^{2+}				Cr^{2+} Cr^{3+}	Mn^{2+} Mn^{3+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+} Ni^{3+}	Cu^+ Cu^{2+}	Zn^{2+}				Se^{2-}	Br^-	
Rb^+	Sr^{2+}									Ag^+	Cd^{2+}		Sn^{2+} Sn^{4+}		Te^{2-}	I^-	
Cs^+	Ba^{2+}									Au^+ Au^{3+}	Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}				

Chemical Formulas

- **Chemical formula**: short hand representation of molecules/ chemical compounds.
- Used to express the composition of molecules and ionic compounds in terms of chemical symbols.
- Composition we mean not only the elements present but also the ratios in which the atoms are combined.
- A **chemical equation** is the symbolic representation of a chemical reaction in the form of symbols and formula.

Formulas and Models

	Hydrogen	Water	Ammonia	Methane
Molecular formula	H_2	H_2O	NH_3	CH_4



Structural formula
H—H
H—O—H
Ball-and-stick model
Space-filling model

Molecular formula shows the exact number of atoms of each element in the smallest unit of a substance

Empirical formula shows the simplest whole-number ratio of the atoms in a substance

molecular	**empirical**
H_2O	H_2O
$C_6H_{12}O_6$	CH_2O
O_3	O
N_2H_4	NH_2

A structural formula uses lines to represent covalent bonds, and shows how the atoms in a molecule are joined together: H—O—O—H H—O—H O=C=O.

Calculating Empirical Formula

An oxide of aluminum is formed by the reaction of 4.151 g of aluminum with 3.692 g of oxygen. Calculate the empirical formula.

1. Determine the number of grams of each element in the compound.

4.151 g Al and 3.692 g O

2. Convert masses to moles.

$$\frac{4.151 \text{ g Al} \mid 1 \text{ mol Al}}{\mid 26.98 \text{ g Al}} = 0.1539 \text{ mol Al}$$

$$\frac{3.692 \text{ g O} \mid 1 \text{ mol O}}{\mid 16.00 \text{ g O}} = 0.2308 \text{ mol O}$$

3. Find ratio by dividing each element by smallest amount of moles.

$$\frac{0.1539 \text{ moles Al}}{0.1539} = 1.000 \text{ mol Al}$$

$$\frac{0.2308 \text{ moles O}}{0.1539} = 1.500 \text{ mol O}$$

4. Multiply by common factor to get whole number. (cannot have fractions of atoms in compounds)

O = 1.500 x 2 = 3

Al = 1.000 x 2 = 2

4. Use the whole numbers obtained as subscript for the corresponding element in simplest whole number therefore, $\mathbf{Al_2O_3}$

When a 2.00 g sample of iron metal is heated in air, it reacts with oxygen to achieve a final mass of 2.573 g. Determine the empirical formula.

Fe = 2.000 g   O = 2.573 g - 2.000 g = 0.5730 g

$$\frac{2.000\text{ g Fe}}{} \left| \frac{1\text{ mol Fe}}{55.85\text{ g Fe}} \right. = 0.03581\text{ mol Fe}$$

$$\frac{0.573\text{ g O} \quad 1\text{ mol O}}{16.00\text{ g}} = 0.03581\text{ mol Fe}$$, 1 : 1   **FeO**

Q. Vitamin C (ascorbic acid) contains 40.92% C, 4.58% H, and 54.50% O by mass. What is the empirical formula of ascorbic acid?

A sample of lead arsenate, an insecticide used against the potato beetle, contains 1.3813 g lead, 0.00672g of hydrogen, 0.4995 g of arsenic, and 0.4267 g of oxygen. Calculate the empirical formula for lead arsenate.

$$\frac{1.3813\text{ g Pb} \mid 1\text{ mol Pb}}{\mid 207.2\text{ g Pb}} = 0.006667\text{ mol Pb};$$

$$\frac{0.00672\text{ gH} \mid 1\text{ mol H}}{\mid 1.008\text{ g H}} = 0.00667\text{ mol H}$$

$$\frac{0.4995\text{ g As} \mid 1\text{ mol As}}{\mid 74.92\text{ g As}} = 0.006667\text{ mol As}$$

$$\frac{0.4267\text{g O} \mid 1\text{ mol O}}{\mid 16.00\text{ g O}} = 0.02667\text{ mol O}$$

$PbHAsO_4$

A white powder is analyzed and found to have an empirical formula of P_2O_5. The compound has a molar mass of 283.88g. What is the compound's molecular formula?

Step 1: Molar Mass

$P = 2 \times 30.97\ g = 61.94g$
$O = 5 \times 16.00g = \underline{80.00\ g}$
$141.94\ g$

Step 2: Divide MM by Empirical Formula Mass

$$\frac{238.88\ g}{141.94g} = 2$$

Step 3: Multiply

$MF = (EF)n$

$MF = (P_2O_5)_2$

$MF = P_4O_{10}$

Suppose that in one experiment the combustion of 11.5 g of ethanol produced 22.0 g of CO_2 and 13.5 g of H_2O. We can calculate the mass of carbon and hydrogen in the original 11.5 g sample of ethanol as follows:

$$\text{mass of C} = 22.0\ \cancel{\text{g CO}_2} \times \frac{1\ \cancel{\text{mol CO}_2}}{44.01\ \cancel{\text{g CO}_2}} \times \frac{1\ \cancel{\text{mol C}}}{1\ \cancel{\text{mol CO}_2}} \times \frac{12.01\text{ g C}}{1\ \cancel{\text{mol C}}}$$
$$= 6.00\text{ g C}$$

$$\text{mass of H} = 13.5\ \cancel{\text{g H}_2\text{O}} \times \frac{1\ \cancel{\text{mol H}_2\text{O}}}{18.02\ \cancel{\text{g H}_2\text{O}}} \times \frac{2\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol H}_2\text{O}}} \times \frac{1.008\text{ g H}}{1\ \cancel{\text{mol H}}}$$
$$= 1.51\text{ g H}$$

Thus, 11.5 g of ethanol contains 6.00 g of carbon and 1.51 g of hydrogen. The remainder must be oxygen, whose mass is

$$\begin{aligned}\text{mass of O} &= \text{mass of sample} - (\text{mass of C} + \text{mass of H})\\ &= 11.5\text{ g} - (6.00\text{ g} + 1.51\text{ g})\\ &= 4.0\text{ g}\end{aligned}$$

The number of moles of each element present in 11.5 g of ethanol is

$$\text{moles of C} = 6.00\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 0.500\ \text{mol C}$$

$$\text{moles of H} = 1.51\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}} = 1.50\ \text{mol H}$$

$$\text{moles of O} = 4.0\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.25\ \text{mol O}$$

The formula of ethanol is therefore $C_{0.50}H_{1.5}O_{0.25}$ (we round off the number of moles to two significant figures). Because the number of atoms must be an integer, we divide the subscripts by 0.25, the smallest subscript, and obtain for the empirical formula C_2H_6O.

A sample of a compound contains 1.52 g of nitrogen (N) and 3.47 g of oxygen (O). The molar mass of this compound is between 90 g and 95 g. Determine the molecular formula and the accurate molar mass of the compound.

Strategy To determine the molecular formula, we first need to determine the empirical formula. How do we convert between grams and moles? Comparing the empirical molar mass to the experimentally determined molar mass will reveal the relationship between the empirical formula and molecular formula.

Solution We are given grams of N and O. Use molar mass as a conversion factor to convert grams to moles of each element. Let n represent the number of moles of each element. We write

$$n_N = 1.52\ \cancel{\text{g N}} \times \frac{1\ \text{mol N}}{14.01\ \cancel{\text{g N}}} = 0.108\ \text{mol N}$$

$$n_O = 3.47\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.217\ \text{mol O}$$

Thus, we arrive at the formula $N_{0.108}O_{0.217}$, which gives the identity and the ratios of atoms present. However, chemical formulas are written with whole numbers. Try to convert to whole numbers by dividing the subscripts by the smaller subscript (0.108). After rounding off, we obtain NO_2 as the empirical formula.

The molecular formula might be the same as the empirical formula or some integral multiple of it (for example, two, three, four, or more times the empirical formula). Comparing the ratio of the molar mass to the molar mass of the empirical formula will show the integral relationship between the empirical and molecular formulas. The molar mass of the empirical formula NO_2 is

$$\text{empirical molar mass} = 14.01\text{ g} + 2(16.00\text{ g}) = 46.01\text{ g}$$

Next, we determine the ratio between the molar mass and the empirical molar mass

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{90\text{ g}}{46.01\text{ g}} \approx 2$$

The molar mass is twice the empirical molar mass. This means that there are two NO_2 units in each molecule of the compound, and the molecular formula is $(NO_2)_2$ or N_2O_4.

The actual molar mass of the compound is two times the empirical molar mass, that is, 2(46.01 g) or 92.02 g, which is between 90 g and 95 g.

Q1. A compound with an empirical formula of C_2OH_4 and a molar mass of 88 grams per mole. What is the molecular formula of this compound?

Q2. Nitrogen and oxygen form an extensive series of oxides with the general formula N_xO_y. One of them is a blue solid that comes apart, reversibly, in the gas phase. It contains 36.84% N. What is the empirical formula of this oxide?

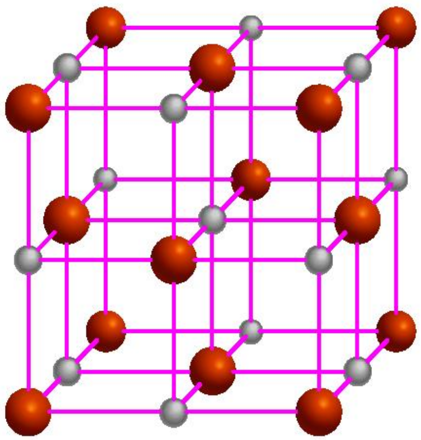
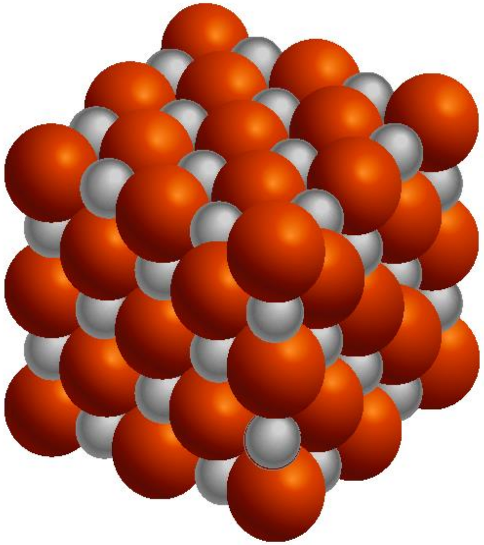
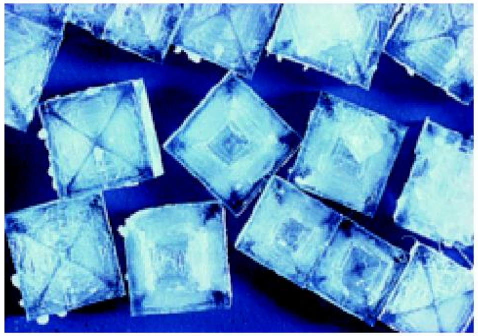
Q3. An unknown compound was found to have a percent composition as follows: 47.0 % potassium, 14.5 % carbon, and 38.5 % oxygen. What is its empirical formula? If the true molar mass of the compound is 166.22 g/mol, what is its molecular formula?

Q4. What are the empirical and molecular formulas of caffeine that contains by mass composition of 49.5% C, 5.15% H, 28.9% N and 16.5 % O and the molecular mass is about 195 g/mol?

Ionic compounds consist of a combination of cations and anions

- The formula is usually the same as the empirical formula because ionic compounds do not consist of discrete molecular units

- The sum of the charges on the cation(s) and anion(s) in each formula unit must equal zero

The ionic compound NaCl

Formula of Ionic Compounds

$2 \times +3 = +6$ $\quad$ $3 \times -2 = -6$

Al_2O_3

Al^{3+} $\quad$ O^{2-}

$1 \times +2 = +2$ $\quad$ $2 \times -1 = -2$

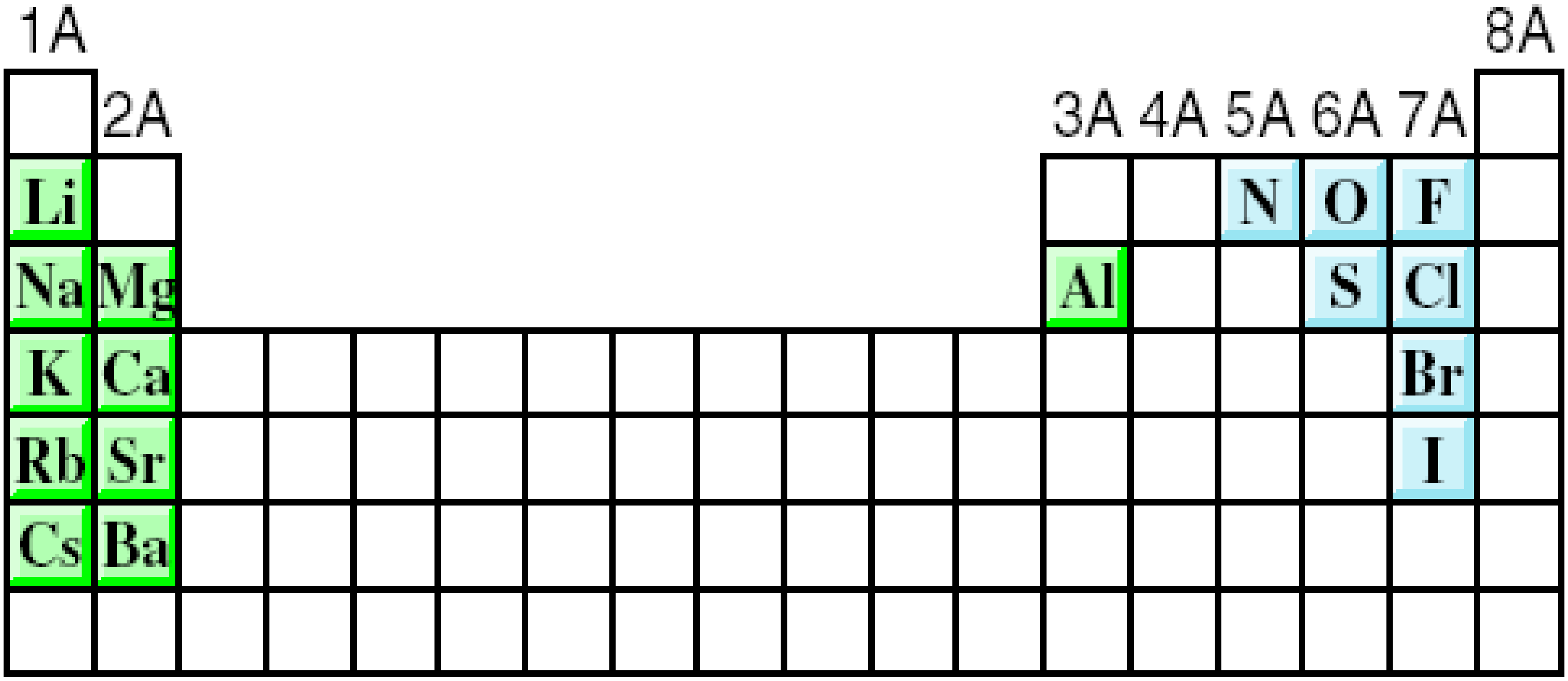
$CaBr_2$

Ca^{2+} $\quad$ Br^{-}

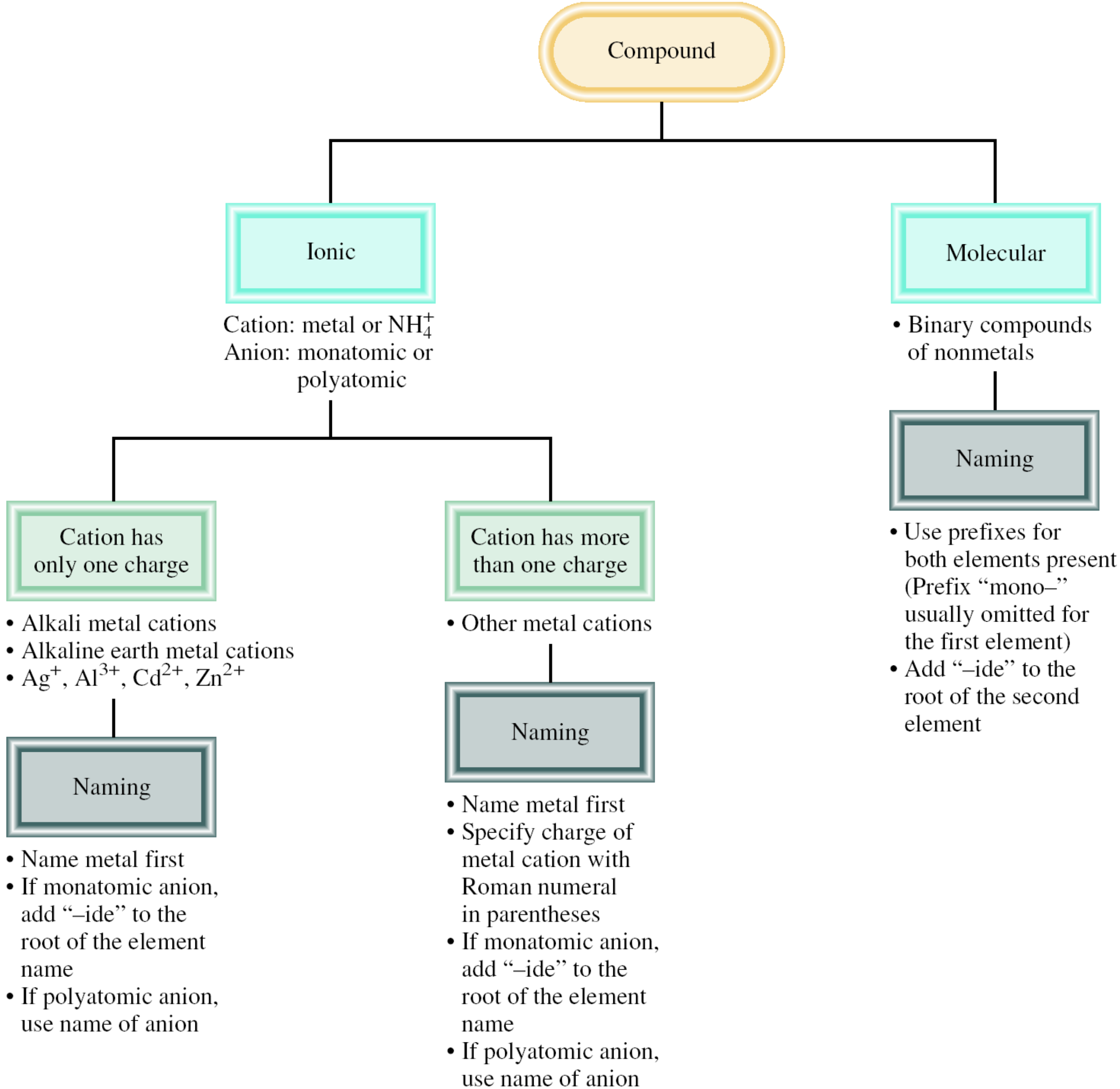
$1 \times +2 = +2$ $\quad$ $1 \times -2 = -2$

Na_2CO_3

Na^{+} $\quad$ CO_3^{2-}



The most reactive **metals** (green) and the most reactive **nonmetals** (blue) combine to form ionic compounds.


Compound
Ionic
Cation: metal or NH_4^+
Anion: monatomic or polyatomic
Cation has only one charge
• Alkali metal cations
• Alkaline earth metal cations
• Ag^+, Al^{3+}, Cd^{2+}, Zn^{2+}
Naming
• Name metal first
• If monatomic anion, add "–ide" to the root of the element name
• If polyatomic anion, use name of anion
Cation has more than one charge
• Other metal cations
Naming
• Name metal first
• Specify charge of metal cation with Roman numeral in parentheses
• If monatomic anion, add "–ide" to the root of the element name
• If polyatomic anion, use name of anion
Molecular
• Binary compounds of nonmetals
Naming
• Use prefixes for both elements present (Prefix "mono–" usually omitted for the first element)
• Add "–ide" to the root of the second element

Percent Composition

- **Percent Composition –**percentage by mass of each element in a compound.

$$\textbf{Percent} = \frac{\textbf{Part}}{\textbf{Whole}} \times 100\%$$

$$\textbf{\% composition of a cpd} = \frac{\text{Mass of element in1mol}}{\text{Mass of 1 mol}} \times 100\%$$

Example: What is the percent composition of Potassium $KMnO_4$?

%K = (39.1/158) X 100% = 24.75%

%Mn =(54.9/158) x 100% = 34.75%

%O = (64/158) X 100% = 40.50%

Phosphoric acid (H_3PO_4) is a colorless, syrupy liquid used in detergents, fertilizers, toothpastes, and in carbonated beverages for a "tangy" flavor. Calculate the percent composition by mass of H, P, and O in this compound.

Strategy Recall the procedure for calculating a percentage. Assume that we have 1 mole of H_3PO_4. The percent by mass of each element (H, P, and O) is given by the combined molar mass of the atoms of the element in 1 mole of H_3PO_4 divided by the molar mass of H_3PO_4, then multiplied by 100 percent.

Solution The molar mass of H_3PO_4 is 97.99 g. The percent by mass of each of the elements in H_3PO_4 is calculated as follows:

$$\%H = \frac{3(1.008\text{ g})\text{ H}}{97.99\text{ g H}_3\text{PO}_4} \times 100\% = \boxed{3.086\%}$$

$$\%P = \frac{30.97\text{ g P}}{97.99\text{ g H}_3\text{PO}_4} \times 100\% = \boxed{31.61\%}$$

$$\%O = \frac{4(16.00\text{ g})\text{ O}}{97.99\text{ g H}_3\text{PO}_4} \times 100\% = \boxed{65.31\%}$$

Check Do the percentages add to 100 percent? The sum of the percentages is $(3.086\% + 31.61\% + 65.31\%) = 100.01\%$. The small discrepancy from 100 percent is due to the way we rounded off.

Q1. What is the mass percentage of Cl in the chlorofluorocarbon CCl_2F_2 (Freon-12)?

$$\text{Mass \% of Cl} = \frac{2 \times \text{atomic mass of Cl x}}{\text{molarmass of } CCl_2F_2} \times 100$$

$$= \frac{2 \times 35.453 \text{ g/molx}}{120.91 \text{ g/mol } CCl_2F_2} \times 100 = 58.64\%$$

Q2. Hydrocarbons are organic compounds composed entirely of hydrogen and carbon. A 0.1647gram sample of a pure hydrocarbon was burned in a C-H combustion train to produce 0.4931 gram of CO2 and 0.2691 gram of H2O. Determine the masses of C and H in the sample and the percentages of these elements in this hydrocarbon.

Step1: We use the observed mass of CO_2, 0.4931 grams, to determine the mass of carbon in the original sample. There is one mole of carbon atoms, 12.01 grams, in each mole of CO_2, 44.01 grams; we use this information to construct the unit factor. $\frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2}$

Step 2: Likewise, we can use the observed mass of H_2O, 0.2691 grams, to calculate the amount of hydrogen in the original sample. We use the fact that there are two moles of hydrogen atoms, 2.016 grams, in each mole of H_2O, 18.02 grams, to construct the unit factor

$$\frac{2.016 \text{ g H}}{18.02 \text{ g } H_2O}$$

Step 3: Then we calculate the percentages by mass of each element in turn, using the relationship

$$\% \text{ element} = \frac{\text{g element}}{\text{g sample}} \times 100\%$$

Solution

Step 1: $\underline{\ ?\ }$ g C = 0.4931 g $CO_2 \times \dfrac{12.01\ \text{g C}}{44.01\ \text{g } CO_2} = \boxed{0.1346\ \text{g C}}$

Step 2: $\underline{\ ?\ }$ g H = 0.2691 g $H_2O \times \dfrac{2.016\ \text{g H}}{18.02\ \text{g } H_2O} = \boxed{0.03010\ \text{g H}}$

Step 3:

$$\%\ C = \frac{0.1346\ \text{g C}}{0.1647\ \text{g sample}} \times 100\% = \boxed{81.72\%\ \text{C}}$$

$$\%H = \frac{0.03010\ \text{g H}}{0.1647\ \text{g sample}} \times 100\% = \boxed{18.28\%\ \text{H}}$$

Total = 100.00%

A 24.81-g sample of a gaseous compound containing only carbon, oxygen, and chlorine is determined to contain 3.01 g C, 4.00 g O, and 17.81 g Cl. What is this compound's percent composition?

Chemical Reactions and Chemical Equations

Chemical reaction-a process in which a substance (or substances) is changed into one or more new substances.

A chemical equation- uses chemical symbols to show what happens during a chemical reaction.

starting materials in a chemical reaction

reactants $\longrightarrow$ products

substance formed as a result of a chemical reaction

Two hydrogen molecules	+	One oxygen molecule	$\longrightarrow$	Two water molecules
$2H_2$	+	O_2	$\longrightarrow$	$2H_2O$

$$2H_2 \quad + O_2 \longrightarrow 2H_2O$$

Two molecules + one molecule $\longrightarrow$ two molecules

2 moles + 1 mole $\longrightarrow$ 2 moles

$$\underbrace{2(2.02\text{ g}) = 4.04\text{ g} + 32.00\text{ g}}_{36.04\text{ g reactants}} \longrightarrow \underbrace{2(18.02\text{ g}) = 36.04\text{ g}}_{36.04\text{ g product}}$$

Reaction Yield

- ✓ Yield, also referred to as reaction yield, is the amount of product obtained in a chemical reaction.
- ✓ The absolute yield can be given as the weight in grams or in moles
- ✓ **Limiting Reactant-**The reactant used up first in a reaction
- ✓Determines maximum amount of product to be formed.
- ✓ **Excess reactants-** are the reactants present in quantities greater than necessary to react with the quantity of the limiting reagent

The food we eat is degraded, or broken down, in our bodies to provide energy for growth and function. A general overall equation for this very complex process represents the degradation of glucose ($C_6H_{12}O_6$) to carbon dioxide (CO_2) and water (H_2O):

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O$$

If 856 g of $C_6H_{12}O_6$ is consumed by a person over a certain period, what is the mass of CO_2 produced?

Strategy Looking at the balanced equation, how do we compare the amount of $C_6H_{12}O_6$ and CO_2? We can compare them based on the *mole ratio* from the balanced equation. Starting with grams of $C_6H_{12}O_6$, how do we convert to moles of $C_6H_{12}O_6$? Once moles of CO_2 are determined using the mole ratio from the balanced equation how do we convert to grams of CO_2?

Step 1: The balanced equation is given in the problem.

Step 2: To convert grams of $C_6H_{12}O_6$ to moles of $C_6H_{12}O_6$, we write

$$856 \cancel{\text{g } C_6H_{12}O_6} \times \frac{1 \text{ mol } C_6H_{12}O_6}{180.2 \cancel{\text{g } C_6H_{12}O_6}} = 4.750 \text{ mol } C_6H_{12}O_6$$

Step 3: From the mole ratio, we see that 1 mol $C_6H_{12}O_6 \simeq$ 6 mol CO_2. Therefore, the number of moles of CO_2 formed is

$$4.750 \cancel{\text{mol } C_6H_{12}O_6} \times \frac{6 \text{ mol } CO_2}{1 \cancel{\text{mol } C_6H_{12}O_6}} = 28.50 \text{ mol } CO_2$$

Step 4: Finally, the number of grams of CO_2 formed is given by

$$28.50 \cancel{\text{mol } CO_2} \times \frac{44.01 \text{ g } CO_2}{1 \cancel{\text{mol } CO_2}} = 1.25 \times 10^3 \text{ g } CO_2$$

After some practice, we can combine the conversion steps

$$\text{grams of } C_6H_{12}O_6 \longrightarrow \text{moles of } C_6H_{12}O_6 \longrightarrow \text{moles of } CO_2 \longrightarrow \text{grams of } CO_2$$

into one equation:

$$\text{mass of } CO_2 = 856\ \cancel{g\ C_6H_{12}O_6} \times \frac{1\ \text{mol}\ \cancel{C_6H_{12}O_6}}{180.2\ \cancel{g\ C_6H_{12}O_6}} \times \frac{6\ \cancel{\text{mol}\ CO_2}}{1\ \cancel{\text{mol}\ C_6H_{12}O_6}} \times \frac{44.01\ g\ CO_2}{1\ \cancel{\text{mol}\ CO_2}}$$

$$= \boxed{1.25 \times 10^3\ g\ CO_2}$$

Consider the formation of nitrogen dioxide (NO_2) from nitric oxide (NO) and oxygen:

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

Suppose initially we have 8 moles of NO and 7 moles of O_2 . One way to determine which of the two reactants is the limiting reagent is to calculate the number of moles of NO_2 obtained based on the initial quantities of NO and O_2. From the preceding definition, we see that only the limiting reagent will yield the smaller amount of the product. Starting with 8 moles of NO, we find the number of moles of NO_2 produced is

$$8\ \cancel{\text{mol NO}} \times \frac{2\ \text{mol NO}_2}{2\ \cancel{\text{mol NO}}} = 8\ \text{mol NO}_2$$

and starting with 7 moles of O_2, the number of moles of NO_2 formed is

$$7\ \cancel{\text{mol O}_2} \times \frac{2\ \text{mol NO}_2}{1\ \cancel{\text{mol O}_2}} = 14\ \text{mol NO}_2$$

Because NO results in a smaller amount of NO_2, it must be the limiting reagent. Therefore, O_2 is the excess reagent.

Urea $[(NH_2)_2CO]$ is prepared by reacting ammonia with carbon dioxide:

$$2NH_3(g) + CO_2(g) \longrightarrow (NH_2)_2CO(aq) + H_2O(l)$$

In one process, 637.2 g of NH_3 are treated with 1142 g of CO_2. (a) Which of the two reactants is the limiting reagent? (b) Calculate the mass of $(NH_2)_2CO$ formed. (c) How much excess reagent (in grams) is left at the end of the reaction?

(a) Strategy The reactant that produces fewer moles of product is the limiting reagent because it limits the amount of product that can be formed. How do we convert from the amount of reactant to amount of product? Perform this calculation for each reactant, then compare the moles of product, $(NH_2)_2CO$, formed by the given amounts of NH_3 and CO_2 to determine which reactant is the limiting reagent.

Solution We carry out two separate calculations. First, starting with 637.2 g of NH_3, we calculate the number of moles of $(NH_2)_2CO$ that could be produced if all the NH_3 reacted according to the following conversions:

$$\text{grams of } NH_3 \longrightarrow \text{moles of } NH_3 \longrightarrow \text{moles of } (NH_2)_2CO$$

Combining these conversions in one step, we write

$$\begin{aligned}\text{moles of } (NH_2)_2CO &= 637.2\ \cancel{\text{g } NH_3} \times \frac{1\ \cancel{\text{mol } NH_3}}{17.03\ \cancel{\text{g } NH_3}} \times \frac{1\text{ mol } (NH_2)_2CO}{2\ \cancel{\text{mol } NH_3}} \\ &= 18.71\text{ mol } (NH_2)_2CO\end{aligned}$$

Second, for 1142 g of CO_2, the conversions are

$$\text{grams of } CO_2 \longrightarrow \text{moles of } CO_2 \longrightarrow \text{moles of } (NH_2)_2CO$$

The number of moles of $(NH_2)_2CO$ that could be produced if all the CO_2 reacted is

$$\begin{aligned}\text{moles of } (NH_2)_2CO &= 1142\ \cancel{\text{g } CO_2} \times \frac{1\ \cancel{\text{mol } CO_2}}{44.01\ \cancel{\text{g } CO_2}} \times \frac{1\text{ mol } (NH_2)_2CO}{1\ \cancel{\text{mol } CO_2}} \\ &= 25.95\text{ mol } (NH_2)_2CO\end{aligned}$$

It follows, therefore, that NH_3 must be the limiting reagent because it produces a smaller amount of $(NH_2)_2CO$.

(b) Strategy We determined the moles of $(NH_2)_2CO$ produced in part (a), using NH_3 as the limiting reagent. How do we convert from moles to grams?

Solution The molar mass of $(NH_2)_2CO$ is 60.06 g. We use this as a conversion factor to convert from moles of $(NH_2)_2CO$ to grams of $(NH_2)_2CO$:

$$\text{mass of } (NH_2)_2CO = 18.71 \cancel{\text{mol } (NH_2)_2CO} \times \frac{60.06 \text{ g } (NH_2)_2CO}{1 \cancel{\text{mol } (NH_2)_2CO}}$$
$$= \boxed{1124 \text{ g } (NH_2)_2CO}$$

(c) Strategy Working backward, we can determine the amount of CO_2 that reacted to produce 18.71 moles of $(NH_2)_2CO$. The amount of CO_2 left over is the difference between the initial amount and the amount reacted.

Solution Starting with 18.71 moles of $(NH_2)_2CO$, we can determine the mass of CO_2 that reacted using the mole ratio from the balanced equation and the molar mass of CO_2. The conversion steps are

$$\text{moles of } (NH_2)_2CO \longrightarrow \text{moles of } CO_2 \longrightarrow \text{grams of } CO_2$$

so that

$$\text{mass of } CO_2 \text{ reacted} = 18.71 \cancel{\text{mol } (NH_2)_2CO} \times \frac{1 \cancel{\text{mol } CO_2}}{1 \cancel{\text{mol } (NH_2)_2CO}} \times \frac{44.01 \text{ g } CO_2}{1 \cancel{\text{mol } CO_2}}$$
$$= 823.4 \text{ g } CO_2$$

The amount of CO_2 remaining (in excess) is the difference between the initial amount (1142 g) and the amount reacted (823.4 g):

$$\text{mass of } CO_2 \text{ remaining} = 1142 \text{ g} - 823.4 \text{ g} = \boxed{319 \text{ g}}$$

Theoretical yield- the amount of product that would result if all the limiting reactant reacted (obtained from balanced chemical reaction)

Actual yield- the amount of product actually obtained from a reaction (experimentally obtained yield!)

✓almost always less than the theoretical yield

Percent yield- describes the proportion of the actual yield to the theoretical yield,

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoritical yield}} \times 100\%$$

Why actually yield is less than theoretical yield?

Q1 - What is the % yield of H_2O if 138 g H_2O is produced from 16 g H_2 and excess O_2?

Step 1: write the balanced chemical equation

$$2H_2 + O_2 \rightarrow 2H_2O$$

Step 2: determine actual and theoretical yield. Actual is given, theoretical is calculated:

$$\#\ \text{g } H_2O = 16\ \cancel{\text{g } H_2} \times \frac{1\ \cancel{\text{mol } H_2}}{2.02\ \cancel{\text{g } H_2}} \times \frac{2\ \cancel{\text{mol } H_2O}}{2\ \cancel{\text{mol } H_2}} \times \frac{18.02\ \text{g } H_2O}{1\ \cancel{\text{mol } H_2O}} = 1\underline{4}3\ \text{g}$$

Step 2: Calculate % yield

$$\%\ \text{yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\% = \frac{138\ \text{g } H_2O}{1\underline{4}3\ \text{g } H_2O} \times 100\% = 9\underline{6}.7\%$$

- The reaction from your lab is conducted with a 2.40 g sample of $NaHCO_3$. In the lab, the reaction produces 1.57 g of NaCl. What is the percent yield?

$$2.40\ \cancel{g\ NaHCO_3} \times \frac{1\ \cancel{mol\ NaHCO_3}}{84.0\ \cancel{g/mol}} \times \frac{1\ \cancel{mol\ NaCl}}{1\ \cancel{mol\ NaHCO_3}} \times \frac{58.5\ g\ NaCl}{1\ \cancel{mol\ NaCl}} = 1.67\ g\ NaCl$$

$$\%\ Yield = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

$$= \frac{1.57\ g}{1.67\ g} \times 100 = 94.0\%$$

Phosphorous reacts with bromine to form phosphorous tribromide. If 35.0 grams of bromine are reacted and 27.9 grams of phosphorous tribromide are formed, what is the percent yield?

$$2\,P + 3\,Br_2 \rightarrow 2\,PBr_3$$

$$\frac{35.0\ g\ Br_2}{1}\ x\ \frac{1\ mole\ Br_2}{159.808\ g\ Br_2}\ x\ \frac{2\ moles\ PBr_3}{3\ moles\ Br_2}\ x\ \frac{270.686\ g\ PBr_3}{1\ mole\ PBr_3} = 39.5\ g\ g\ PBr_3$$

$$\%\ yield = \frac{27.9\ g\ g\ PBr_3}{39.5\ g\ g\ PBr_3}\ x\ 100 = 70.63\%$$

Silver Nitrate reacts with Magnesium Chloride to produce Silver Chloride and Magnesium Nitrate. If 305 grams of silver nitrate are reacted in an excess of magnesium chloride producing 23.7 grams of magnesium nitrate, what is the percent yield?

$$2\,AgNO_3 + MgCl_2 \rightarrow 2\,AgCl + Mg(NO_3)_2$$

$$\frac{305\ g\ AgNO_3}{1}\ x\ \frac{1\ mole\ AgNO_3}{169.872\ g\ AgNO_3}\ x\ \frac{2\ moles\ Mg(NO_3)_2}{2\ moles\ AgNO_3}\ x\ \frac{148.313\ g\ Mg(NO_3)_2}{1\ mole\ Mg(NO_3)_2} = 266\ g\ Mg(NO_3)_2$$

$$\%\ yield = \frac{23.7\ g\ Mg(NO_3)_2}{266\ g\ Mg(NO_3)_2}\ x\ 100 = 8.91\%$$

Zinc is reacted with Hydrochloric acid to form zinc chloride and hydrogen gas. The reaction is carried out in a glass container that has a mass of 14.7 grams. After placing the zinc in the glass container, the mass is 29.5 grams. The hydrochloric acid is poured into the container and the zinc chloride is formed. The excess hydrochloric acid is removed leaving the glass container and the zinc chloride, which together have a mass of 37.5 grams. What is the percent yield?

$$Zn + 2\ HCl_{(aq)} \rightarrow ZnCl_2 + H_{2(g)}$$

$$Amt\ ZnCl_2\ actually\ produced = 37.5\ grams - 14.7\ grams = 22.8\ grams\ ZnCl_2$$

$$Amt\ of\ Zinc\ reacted = 29.5\ grams - 14.7\ grams = 14.8\ grams\ Zn$$

$$Amount\ of\ ZnCl_2\ that\ could\ have\ been\ produced:$$

$$\frac{14.8\ g\ Zn}{1} x \frac{1\ mole\ Zn}{65.39\ g\ Zn} x \frac{1\ mole\ ZnCl_2}{1\ mole\ Zn} x \frac{136.296\ g\ ZnCl_2}{1\ mole\ ZnCl_2} = 30.8\ g\ ZnCl_2$$

$$\%\ yield = \frac{22.8\ grams\ ZnCl_2}{30.8\ g\ ZnCl_2} x\ 100 = 74.03\%$$

Q1 70.90 g/mol 26.98 g/mol 133.33 g/mol

$$2\ Cl_{2(g)} + 3\ Al_{(s)} \rightarrow 2\ AlCl_{3(s)}$$

1a) Calculate the theoretical yield of aluminum chloride (in grams) that can be produced from 10.00 grams of aluminum metal.

1b) A student performed this experiment and obtained 25.23 grams of aluminum chloride. Determine the percent yield of aluminum chloride.

Q2. Assume the reaction given below:

50.94 g/mol 32.00 g/mol 149.88 g/mol

$$4\ V_{(s)} + 3\ O_{2(g)} \xrightarrow{\Delta} 2\ V_2O_{3(s)}$$

Calculate the theoretical yield of vanadium(III) oxide, assuming you begin with 200.00 grams vanadium metal.

When this experiment is performed by a researcher, an experimental yield of 183.2 grams is produced. Calculate the percent yield for this experiment.

Tin (Sn) exists in Earth's crust as SnO_2. Calculate the percent composition by mass of Sn and O in SnO_2.

Cinnamic alcohol is used mainly in perfumery, particularly in soaps and cosmetics. Its molecular formula is $C_9H_{10}O$. (a) Calculate the percent composition by mass of C, H, and O in cinnamic alcohol. (b) How many molecules of cinnamic alcohol are contained in a sample of mass 0.469 g?

When heated, lithium reacts with nitrogen to form lithium nitride:

$$6Li(s) + N_2(g) \longrightarrow 2Li_3N(s)$$

What is the theoretical yield of Li_3N in grams when 12.3 g of Li are heated with 33.6 g of N_2? If the actual yield of Li_3N is 5.89 g, what is the percent yield of the reaction?