SS1 CHEMISTRY 2ND TERM

SCHEME OF WORK

WEEK	TOPIC
1	Mole concept.
2.	Symbols and Formulae
3.	Writing and balancing of chemical equations.
4-5	Laws of Chemical combinations
6.	Chemical combinations
7.	Kinetic theory of matter
8-10	Gas laws
11-12	Revision & Examination

The Mole Concept

- **<u>LEARNING OUTCOMES</u>**: *At the end the lesson students should be able to: - Define the mole as an amount*;
- - Relate mole to formula mass of substance;
- - Write the formulae relating the mole to number of particles
- - Relate the moles to Avogadro number;
- - Solve related problems on mole concept and related topics.

The Mole Concept

DEFINITION:

- ***The mole is the amount of a substance that contains as many elementary entities (particles) as the number of atoms in exactly 12g of carbon-12***.
- The mole is the unit of measurement of the amount of particles in a given mass of substance. It is the standard used in chemistry.
- **Elementary entities of matter include**: *atoms, molecules, ions, electrons e.t.c*
- Through experimental work, it was found that **1 mole or 12 grams of Carbon-12 contains 6.02 x 10^{23} atoms**. This number is known as the Avogadro's number.

Avogadro's Number.

The number of atoms in a mole of a substance is known as Avogadro's number.

Avogadro's Number = 6.02 x 10^{23}

Note: *A mole is not a number; it is the S.I. unit of the amount of a chemical substance*.

Inter-relationship:

mass-mole-Avogadro's constant.

One mole contains 6.02 x 10^{23} atoms [molecules of particles]

One mole is equal to the Relative atomic mass of an element [Na = 23, Cl = 35.5, Pb = 207,

One mole of compound is equal to Relative molecular mass of an element

[e.g. NaOH = 40, H_2O = 18]

For gases,

One mole of a gas contains 22.4 dm^3 volume at s.t.p.

PERIODIC TABLE OF THE ELEMENTS

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1 H 1.0079																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.065	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.64	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57-71 La-Lu	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89-103 Ac-Lr	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Uun (281)	111 Uuu (272)	112 Uub (285)		114 Uuq (289)				

57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Assignment

- 1. State the S.I. unit of:
- (i). Mass (ii). Amount (iii). Molar mass
- 2. define Avogadro constant and give its numerical value.
- 3. How many particles are there in one mole of a substance?
- 4. How many atoms are in 65.0g of zinc? [Zn = 65.0]

Relative molecular mass (R.M.M)

Relative molecular mass of a compound is the sum of the masses of individual atoms in a molecule of a compound. It's unit is g/mol ($gmol^{-1}$)

Example

Calculate the relative molecular mass of the following compounds

[Pb = 207; Ca = 40; Cu = 64: N = 14: H = 1: O = 16: S = 32; Na = 23]

$CaCO_3$ =

$Pb(NO_3)_2$ =

Classwork:

Calculate the relative molecular mass of the following compounds

1. $CuSO_4$;
2. $CaSO_4$;
3. $NaHCO_3$
4. $Ca(HCO_3)_2$
5. Na_2CO_3

C=12; O = 16, H =1, Ca= 40, S = 32, Na= 23, Cu= 63.5

Calculations using the mole concept.

- 1. Proportion Method: Most chemical calculations are based on direct and inverse proportions.
- **Note**: **In solving mathematical problems in chemistry, follow the steps below**: -
- *List the required and given variables* and constants.
- - *Provide mathematical formula(e) required*.
- - *Do the correct substitution, and make the required variable subject of the formula*.
- - *Evaluate to 3 significant figures then, insert unit, if necessary*.

Worked Examples.

1. ***How many moles are in 15.0g of carbon atoms? [C = 12]***

Solution:

The required variable is amount of carbon n = ? mol.
Given variable is mass of carbon = 15.0g
Constant is molar mass of carbon = 12.0g mol^{-1}

<u>First Principle: (Method 1)</u>.

12g of carbon = 1 mole
1g of carbon = 1/12 mol.
15g of carbon = x

1/12 x 15
x = 1.25 mol.

Formula Method. The formula that connects mass-mole-molar mass is:

$$\textbf{Amount, n (in mol.)} = \frac{\textbf{Mass of substance, m (in grams)}}{\textbf{Molar mass, M (in g mol}^{-1}\textbf{)}}$$

Solution:

Given : mass, m = 15.0g

Molar mass = 12.0g mol^{-1} (constant).

Required = amount, n in mol.

Substituting in the formula:

Amount, n in mol. = 15/12 = 1.25 mol.

2. Calculate the mass of iron present in 0.025 mol. of iron fillings.
[Fe = 56.0]

Answer:

From first Principle: method 1

Given = amount in mol.

Molar mass of Fe. = 56.0 (constant)

Required = mass of iron = ?g

1 mol. of iron = 56.0g

0.025 of iron = $\frac{56 \times 0.025}{1}$g = 1.4g

Using Formula:

Amount (n) = $\frac{\text{mass in g}}{\text{Molar mass}}$

0.025 = $\frac{\text{mass in g}}{56.0}$

mass in g = 0.025 x 56.0 = 1.40g

CALCULATE THE PERCENTAGE OF EACH ELEMENT IN THE COMPOUNDS

1. $Ca(HCO_3)_2$
2. Na_2CO_3

C=12; O = 16, H =1, Ca= 40, S = 32, Na= 23, Cu= 63.5

Practical 1

Title: Filtration

Aim: To separate mixtures by filtration

Apparatus: funnel, test tube, filter paper

Procedure:

- Put solid sample into the test tube
- Add 10 cm3 of distilled water
- Prepare a filter paper/funnel filtration setup
- Pour mixture into filtration setup

Observation: mixture separates into residue (colourless/white) and Filtrate (colourless)

Conclusion: The mixture is composed of colourless residue and filtrate

Atom - Mass – Mole relationship: *Avogadro's constant*.

One mole of an element (substance) contains a fixed number of atoms; the Avogadro's constant L, i.e., 6.02×10^{23} atoms.

$$L = \frac{\text{Number of atom (N)}}{\text{Amount (n)}}$$

$$\text{Amount of element (mol.)} = \frac{\text{Number of atoms (N)}}{\text{Avogadro's constant}}$$

Example: What is the mass of 6.02×10^{24} atoms of magnesium?

[Mg = 24, Avogadro's constant = 6.02×10^{23}]

Answer: Amount in mol. $= \dfrac{6.02 \times 10^{24}}{6.02 \times 10^{23}}$

$= 10$ mol.

Mass = amount (n) x Molar mass = 10 x 24 = 240g.

Molecule – mass – Mole Relationship.

When the relative molecular mass of an element or compound is expressed in grams or any unit of mass, it represents the mass of one mole of the molecule, i.e., the molar mass.

For example, the formula of one mole of oxygen is O_2. That is, a diatomic gas. Its relative molecular mass is 32 (two atoms of oxygen).

This also contains 6.02×10^{23} molecules of oxygen.

Example: How many atoms are in 24.0g of oxygen gas?

[O = 16.0, Avogadro's constant = 6.02×10^{23} mol^{-1}]

<u>Answer</u>: *Recall* : Oxygen is diatomic (two atoms in a molecule)

Action: *First calculate number of molecules in 24 g of oxygen*.

32g of oxygen = 6.02×10^{23} molecules

24g of oxygen = $\frac{24 \times 6.02 \times 10^{23}}{32}$

= 4.52×10^{23} molecules, but 1 mol. of O_2 = 2 atoms

number of atoms of O_2 = $2 \times 4.52 \times 10^{23}$ = $\underline{9.04 \times 10^{23} \text{ atoms}}$

NB:

Avogadro's number (constant) is the number of atoms (particles or molecules) in one mole of a substance. It is given as 6.02×10^{23}

Assignment

1. What would be the mass of:

i. 1.51×10^{23} molecules of carbon (IV) oxide [CO_2]?

ii. 3.01×10^{22} molecules of water [H_2O]?

[H = 1, O = 16, C = 12, Avogadro’s constant = 6.02×10^{23} mol^{-1}]

2. Calculate the number of molecules in:

i. 6.4g of sulphur (IV) oxide. [SO_2]

ii. 1.5 moles of hydrogen chloride [HCl].

[H = 1, O = 16, S = 32, Cl = 35.5, L = 6.02×10^{23} mol^{-1}]

SYMBOLS, FORMULAE AND EQUATIONS

- **<u>LEARNING OUTCOMES</u>**: *At the end of the lesson, students should be able to*:
- - write names and symbols of common elements including the first twenty.
- - write formulae of elements, radicals and their valences.
- - write formulae of binary and other simple compounds.
- - formulate, write and balance chemical equations.

SYMBOLS OF ELEMENTS

- **Symbols of elements** are derived from initial letter of the names of the elements concerned e.g. O for oxygen, C for carbon, F for fluorine etc. *such letters are written in capital*.
- Some other symbols are derived from the initial letter and one other to represent the element e.g. Cl for chlorine, Ca for calcium, Mg for magnesium, Al for aluminum etc. *The first letter is written in capital and the other in small letter*.
- In the same way, symbols of some elements are derived from their Latin names e.g. Na for sodium (Natrium), Fe for iron (Ferrum), Cu for copper (Cuprum), Ag for silver (Argentum), Au for gold (Aurum), K for potassium (Kalium), etc.

FORMULAE OF ELEMENTS, RADICALS AND THEIR VALENCES.

A chemical formula represents a molecule of an element or a compound. It is represented by combination of symbols (and valences for compounds).

Element	Symbols	Valency
Sodium	Na	1
Aluminum	Al	3
Magnesium	Mg	2
Potassium	K	1
Chlorine	Cl	1
Hydrogen	H	1
Carbon	C	2 or 4
Copper	Cu	1 or 2
Sulphur	S	2, 4 or 6
Silver	Ag	1

Valency: *This is defined as the combining power of an element*.

Radical: *A radical is group of two or more different atoms reacting as one unit*. Radicals may be charged with oxidation number written as net charge on the radical.

Radical	Formulae	Valences
Ammonium	NH_4^+	1
Hydroxide	OH^-	1
Trioxonitrate(V)	NO_3^-	1
Dioxonitrate(III)	NO_2^-	1
Trioxocarbonate(IV)	CO_3^{2-}	2
Hydrogen trioxocarbonate(IV)	HCO_3^-	1
Tetraoxosulphate(VI)	SO_4^{2-}	2
Trioxosulphate(IV)	SO_3^{2-}	2
Oxide	O^{-2}	2

Writing Chemical Formulae From Valencies

Formulae of compounds can be deduced from the valencies of component elements. Generally, we write as shown:

Component element/radical	Valency	Formula.
A	x	
B	y	= A_yB_x

Rules for writing chemical formulae.

- write the symbols for the component elements and radicals.
- write the valencies above and to the right of the symbols.
- re-write the symbols with the valencies written below and to the right of symbols.

- Note: *Valencies or oxidation numbers could be used when writing formulae. However, when oxidation numbers are used, the charges are omitted in the last step*.

Examples:

		Valency	-1	-2	-3
		Element/ Radical	Cl	O	N
(i)	+1	Na	$NaCl$	Na_2O	Na_3N
(ii)	+2	Ca	$CaCl_2$	CaO	Ca_3N_2
(iii)	+3	Al	$AlCl_3$	Al_2O_3	Al_3N_3 or AlN

Further Examples

Valency		-1	-2	-3
	Radical/metal	OH	SO_4	PO_4
+1	NH_4	NH_4OH	$(NH_4)_2SO_4$	$(NH_4)_3PO_4$
+2	Cu	$Cu(OH)_2$	$CuSO_4$	$Cu_3(PO_4)_2$
+3	Fe	$Fe(OH)_3$	$Fe_2(SO_4)_3$	$FePO_4$

Note that when the number of a radical is more than one, it is put in parenthesis to avoid confusion with the valencies.

Percentage Mass of Elements in Compounds.

The mass of 1 mole of a compound is the sum of the masses of the moles of its component elements. For example:

Mass of 1 mole of $CaCO_3$ = mass of 1 mole of Ca, C and 3 atoms of O.

Where Ca = 40, C = 12, O = 16. so,

$CaCO_3 = 40 + 12 + 3(16) = 100$

% by mass of Ca = $\frac{40}{100} \times \frac{100}{1} = 40\%$

% by mass of C = $\frac{12}{100} \times \frac{100}{1} = 12\%$

% by mass of O = $\frac{48}{100} \times \frac{100}{1} = 48\%$

Class work

1. Calculate the percentage by mass of nitrogen in trioxonitrate(V) acid HNO_3.

(H=1, N=14, O=16)

Assignment

2. Calculate the percentage by mass of Oxygen in Aluminium tetraoxosulphate (VI) $Al_2(SO_4)_3$ [Al = 27, S= 32: O = 16]

Formulae.

There are two types of formulae in inorganic chemistry. These are:

- Empirical formula;
- Molecular formula.

Empirical formula (E.F): *The empirical formula of a compound shows the simplest ratio of the numbers of atoms of different elements in a molecule of the compound*.

Molecular formula (M.F): *The molecular formula of a compound shows the actual numbers of atoms of the different elements in one molecule of the compound*.

The two formulae are related by the simple formula:

(E.F) x n = M.F,

where n is a positive whole number. In other words, *molecular formula is a multiple of empirical formula*.

Example: ***A compound contains 7.2g of nitrogen and 1.0g of hydrogen. If the relative molecular mass of the compound is 32. determine its***

(i) Empirical formular (ii) molecular formula. (H=1, N=14)

Answer: Write out the symbols of the component elements:

	N	H
mass of elements		

(ii) (Empirical Formular) x n = Molecular Formular

i.e.; $(NH_2) \times n = 32$

$(14 + 2) \times n = 32$

$14n + 2n = 32$

$16n = 32$

$n = 2.$

$(NH_2) \times 2 = N_2H_4$

Molecular formula = N_2H_4

NB: if the vapour density is given, [2 x Vapour density = molecular mass]

Also: If the ratio of empirical atoms is 1:1.5 or 1: 2.5, do not approximate 0.5 to a whole number, rather multiply both ratios by 2 to give whole numbers.

e.g: 1 ; 1.5 = 2: 3, or 1: 2.5 = 2 : 5

Assignment

1. A compound contains 40.0% carbon, 6.7% hydrogen and 53.3% oxygen. Determine its (i) empirical formula and hence,

(ii) its molecular formula if its molar mass is 180. [C=12, O=16, H=1]

2. A hydrocarbon with a vapour density of 29 contains 82.76% carbon and 17.24% hydrogen.

Determine the:

I. empirical formula;

ii. molecular formula of the hydrocarbon.

[H = 1.00 C = 12.00]

CHEMICAL EQUATIONS

- **<u>LEARNING OUTCOMES</u>**: *At the end of the lesson, students should be able to*:-
- *- formulate chemical equations from chemical reactions;*
- *- balance chemical equations;*
- *- solve calculation problems from chemical equations*.

CHEMICAL EQUATIONS

A chemical equation is a precise quantitative statement which is used to summarize a chemical reaction.

A chemical equation shows:-

- the relative number of atoms and molecules taking part in the reaction;
- the state of the various substances taking part in the reaction i.e. state symbols:

  (s for solid, l for liquid, g for gas and aq for aqueous);
- reactants and products;
- heat change, ΔH.
- number of moles (masses) of reactants and products.

<u>Writing Chemical Equation</u>:-

- symbols or formulae of reactants (substances that combine, mix, join, etc.) are written on the left hand side while symbols or formulae of products (substances that are formed, appearing, etc.) are written on the right hand side.

- The reactants and products are linked by an arrow pointing from reactants to products.

REACTANTS $\longrightarrow$ PRODUCTS.

e.g. Hydrogen + Oxygen $\longrightarrow$ Water (word equation)

$H_{2(g)} + O_{2(g)} \longrightarrow H_2O_{(l)}$ (symbol equation)

Information that cannot be gotten from a chemical equation are:-

- speed of reaction;

- color of reactants and products.

BALANCING OF CHEMICAL EQUATION.

Balanced chemical equation gives complete description of what happens in a chemical reaction. Unbalanced equation violates the law of conservation of matter and cannot be used quantitatively and even qualitatively to determine what happened in a chemical reaction.

Steps in balancing equations:

i. Equations must be balanced through coefficients in front of the formulae and not by changing the subscripts and number within the formulae

ii. Common gases such as hydrogen, oxygen, chlorine, iodine and fluorine in the free state are diatomic

iii. Other elements such as K, Cu, and Fe in their free states are monoatomic and are represented by their symbols.

iv. The equation is balanced by making sure that the number of atoms of each element on the left hands side is equal to the number on the right hands.

Examples:

1. balance these chemical equation:

$$NH_{3(g)} + O_{2(g)} \longrightarrow NO_{(g)} + H_2O_{(l)}$$

2. $HCl + Zn \longrightarrow ZnCl_2 + H_2$

(3) $Fe_2O_3 + C \longrightarrow Fe + CO$

Assignment

Balance the following equations

1. $C + Al \longrightarrow Al_4C_3$
2. $I_2O_5 + CO \longrightarrow I_2 + CO_2$
3. $CuO + NH_3 \longrightarrow Cu + H_2O + N_2$
4. $Cu + HNO_3 \longrightarrow Cu(NO_3)_2 + H_2O + NO$

CALCULATIONS FROM CHEMICAL EQUATIONS

- A balanced chemical equation shows a quantitative relationship of the amounts of reactants and products in the reaction. Such equation is said to be *stoichiometric*. This is because the mole ratios of reactants and products can be obtained from it.
- **Stoichiometric equation** can be used to estimate quantities of reactants used up and products formed. The steps are:
- - *write balanced equations for the reaction*
- - *write the amounts, in moles, under each of the relevant reactants and products*
- - *convert the amounts (in moles) to mass in grams and kilograms.*
- - *find the quantity required by direct proportion.*
- - *express your final answer to three significant figures.*
- - *finally, insert the appropriate unit.*

Worked Examples:

1. How many molecules of hydrogen will combine with five molecules of oxygen? How many molecules of water is formed?

Solution:

Step 1: write balanced equation for the reaction;

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$$

2 mol. 1 mol. 2 mol.

Step 2: From equation: 2mol of H_2 reacts with 1mol of O_2.

x mol. of H_2 would combines with 5 mol. of O_2

x = 5x2 = <u>10 molecules</u>.

2. What mass of calcium oxide will be obtained by heating 20.0g of calcium trioxocarbonate (IV) to constant mass? [Ca=40, C=12; O=16]

Answer: [This is a mass-mass relationship]

$CaCO_{3(s)} \longrightarrow CaO_{(s)} + CO_{2(g)}$ (balanced)

1 mol. 1 mol.

(40+12+3x16) (40+16)

100g 56g (given: 20.0g $CaCO_3$; required mass of CaO)

Method 1: *First principle*.

100g of $CaCO_3$ will give 56g of CaO

20g $CaCO_3$ will give x

$$X = \frac{56 \times 20}{100} = 11.2g$$

Assignment

1. *What amount of copper will be deposited when 28g of iron filings react with excess $CuSO_4$ solution according to the following equation?*

$$Fe_{(s)} + CuSO_{4(aq)} \longrightarrow FeSO_{4(aq)} + Cu_{(s)}\ [Fe=56]$$

2. In the reaction:

$$Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2(g)}.$$

(i). Interpret the reaction in terms of moles.

(ii). How many moles of the metal will react with three moles of hydrochloric acid.

LAWS OF CHEMICAL COMBINATIONS.

- **Learning Outcomes**: *At the end of the lesson, students should be able to:*
- *Define the chemical laws;*
- *Illustrate the chemical laws with examples;*
- *Write balanced equations where necessary;*
- *Describe experiments to illustrate the laws.*
- *Solve simple problems on the topic and related topics*.

Law of Conservation of mass (indestructibility of matter).

The law states that '*matter cannot be created nor destroyed during a chemical reaction but can be transformed from one form to another*'.

In a chemical reaction i.e. the total number of atoms of reactants is always equal to the total number of atoms of products, although recombination of atoms would have occurred.

Verification of the law: The law may be verified as explained below:

- in a conical flask, put about $10cm^3$ of silver trioxonitrate (v) solution.

in a test-tube, put about $10cm^3$ of sodium chloride solution.

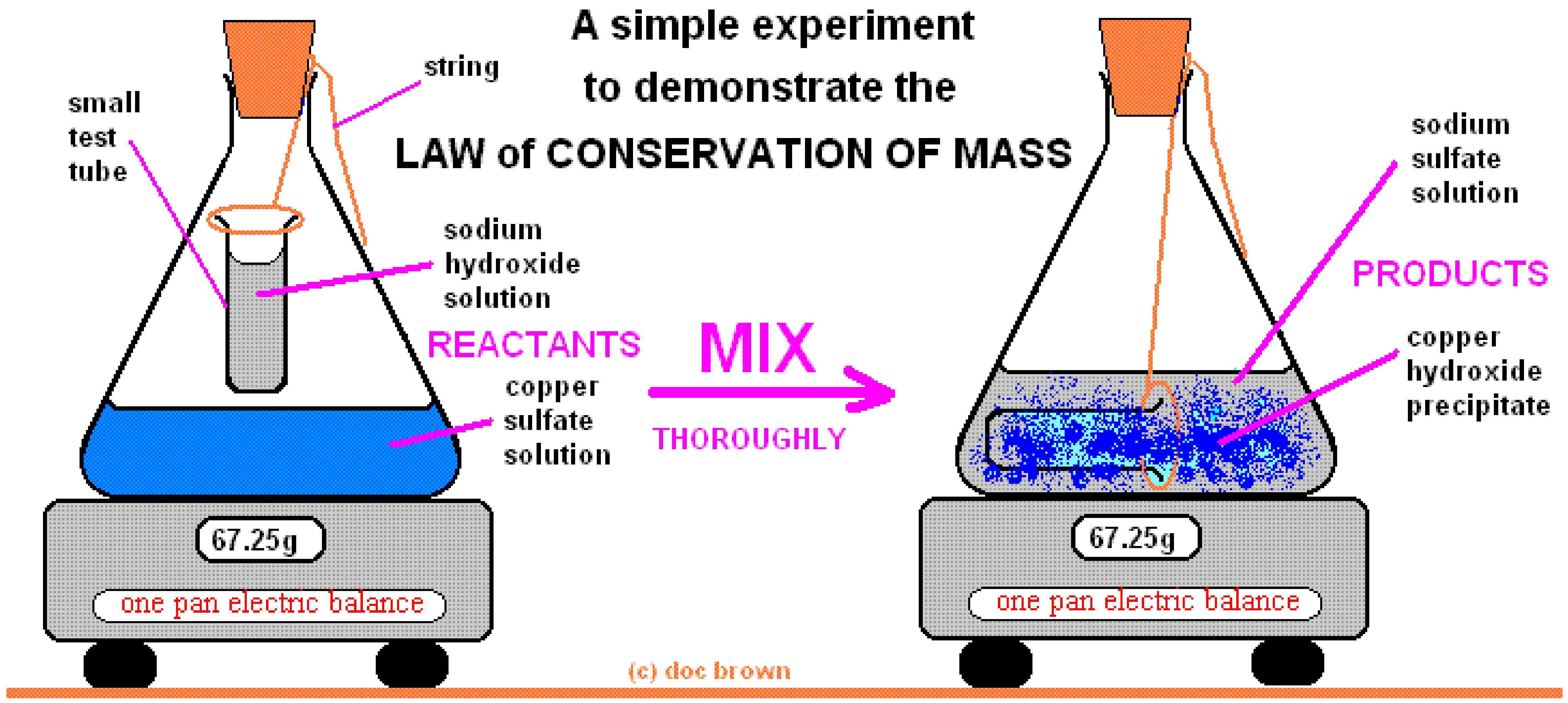
- with a thread, carefully suspend the test tube into the flask.

- stopper the flask with a cork and place it on a balance. Take the mass.

- carefully shake the flask to allow the two solutions mix. A white precipitate is an indication that reaction has occurred.

- weigh the flask again.

Conclusion: The mass of the flask before and after reaction remains unchanged so, the law is verified.



A simple experiment
to demonstrate the
LAW of CONSERVATION OF MASS
small
test
tube
string
sodium
hydroxide
solution
REACTANTS
copper
sulfate
solution
MIX
THOROUGHLY
sodium
sulfate
solution
PRODUCTS
copper
hydroxide
precipitate
67.25g
one pan electric balance
67.25g
one pan electric balance
(c) doc brown

<u>Law of Definite Proportions or Constant Composition</u>.

It states that: *All pure samples of the same chemical compound contain the same elements combined in the same proportion by mass*.

<u>Experimental Illustration</u>: Pure samples of black copper(ii) oxide is prepared in *three* different ways:

a. **<u>From copper(II) trioxonitrate (V)</u>**:

When a blue crystals of copper (II) trioxonitrate (V) are heated strongly to a constant mass, a black sample of copper (II) Oxide is obtained (sample A).

$$2Cu(NO_3)_{2(s)} \longrightarrow 2CuO_{(s)} + 4NO_{2(g)} + O_{2(g)}$$

b **<u>From copper (II) trioxocarbonate (IV)</u>**: When powdered copper (II) trioxocarbonate(IV), dirty green is heated, it decomposes to give a black solid as residue. (sample B).

$$CuCO_{3(s)} \longrightarrow CuO_{(s)} + CO_{2(g)}$$

c. <u>From copper(II) hydroxide</u>: When a sample of blue copper (II) is heated strongly, a black residue is formed. (sample C).

$$Cu(OH)_{2(s)} \longrightarrow CuO_{(s)} + H_2O_{(l)}$$

- **<u>Analysis of Samples A, B and C</u>**. Each sample is weighed in a porcelain boat separately, and the three reduced in a steam of hydrogen, to produce a reddish-brown residue of copper. Allow to cool, and re-weigh.
- **<u>Precautions</u>**:
- - The hard glass tube in which the samples are heated must be tilted.
- (see diagram). *This is to prevent water from sucking back which could crack the glass*.
- - continuous flow of hydrogen gas should be allowed until it is cold. *This is to prevent re-oxidation by air*.

ANALYSIS

Experimental Data and Calculations:

	Sample A	Sample B	Sample C
Mass of CuO (a)	0.95g	1.15g	2.11g
Mass of Cu (b)	0.76g	0.92g	1.69g
% of Cu	$\frac{0.76 \times 100}{0.95}$	$\frac{0.92 \times 100}{1.15}$	$\frac{1.69 \times 100}{2.11}$
	= 80.0%	= 80.0%	= 80.1%

Conclusion: The % by mass of Cu is approximately 80 in all the samples. Thus, the law is obeyed.

Law of Multiple Proportions:

- The law states that *"if two elements, A and B combine to form more than one chemical compound, then the various masses of one element, A, which combine separately with a fixed mass of the other element, B, are in simple multiple ratio"*.
- Some paired elements that form more than one compound include:
- - Copper and Oxygen: They combine to form black CuO and red Cu_2O oxides.
- -Iron and Chlorine: Which combine with chlorine to form green $FeCl_2$ and yellow $FeCl_3$.
- - Lead and Sulphur: They combine to form black PbS and PbS_2.
- Any of these can be used to verify the law.
- Experiment: Weigh samples of black CuO and red Cu_2O oxides respectively in already weighed clean porcelain boats. Reduce each as explained in previous experiment on law of definite proportions.

Experimental Data and Calculation:

	Copper(I) oxide, Cu_2O	Copper(II) oxide, CuO
mass of oxide	3.04g	1.91g
mass of Cu	2.55g	1.38g
mass of O	0.49g	0.53g

Mass of copper that separately combine with fixed mass of O (1g) in each compound is:

Cu_2O : 0.49g of O combine with 2.55g of Cu.

so, 1.0g of O will combine with $\frac{1.0 \times 2.55}{0.49}$ = 5.2g of Cu

CuO: 0.53g of O combine with 1.38g of Cu;

so, 1.0g of O will combine with $\frac{1.0 \times 1.38}{0.53}$ = 2.6g of Cu.

Summary:	CuO_2	CuO
	5.2g	2.6g
Divide by smallest:	$\frac{5.2g}{2.6}$	$\frac{2.6g}{2.6}$
	2	1

Therefore, the masses of copper which combine separately with a fixed mass of oxygen in copper(I) oxide and copper(II) oxide are in simple multiple ratio 2:1, thus, verifying the law.

Assignment

- State the law of multiple proportions. A metal X forms two different chlorides. If 12.7g of chloride A and 16.3g of chloride B contain 7.1g and 10.7g of chlorine respectively, show that the figures agree with the law of multiple proportions.

- A metal M forms two oxides containing 11.1% and 20.0% of oxygen. Show that this result is consistent with the law of multiple proportion

CHEMICAL COMBINATIONS.

- **<u>Learning Outcomes</u>**: ***At the end of lesson, students should be able to***:
- *- state types of chemical combinations;*
- *- explain each type of combination with examples;*
- *- use diagrams and equations to illustrate each type of combination;*
- *- write the properties of electrovalent and covalent compounds*.

CHEMICAL COMBINATIONS

Why do elements combine?

- The noble or rare gases have stable electronic configuration as such, they are very stable and unreactive. It is the tendency of other elements to attain the duplet or octet configuration of the rare gases that is responsible for combinations or reactions.
- **<u>Types of chemical Combinations</u>**:

Types of chemical combinations are:

- (a) Electrovalent(ionic) combination, and

- (b) Covalent combination, which include:

(i) Ordinary covalent combination

(ii). Dative or coordinate covalent combination.

Other types of combination are

(c) Metallic bonding

(d) Hydrogen bonding

(e) Van der Waal's forces

Intermolecular forces

Electrovalent(ionic) Combination:

This is the type of chemical combination in which there is a complete transfer of electron(s) from an atom, usually(metal) to another atom, usually(non-metal) to form ions.

The atom that loses electron(s) becomes positively charged while the atom that gains becomes negatively charged.

It is the electrostatic attraction between the oppositely charged ions that constitute the ionic or electrovalent bond.

Examples:

1. Formation of sodium chloride, (Na^+Cl^-).

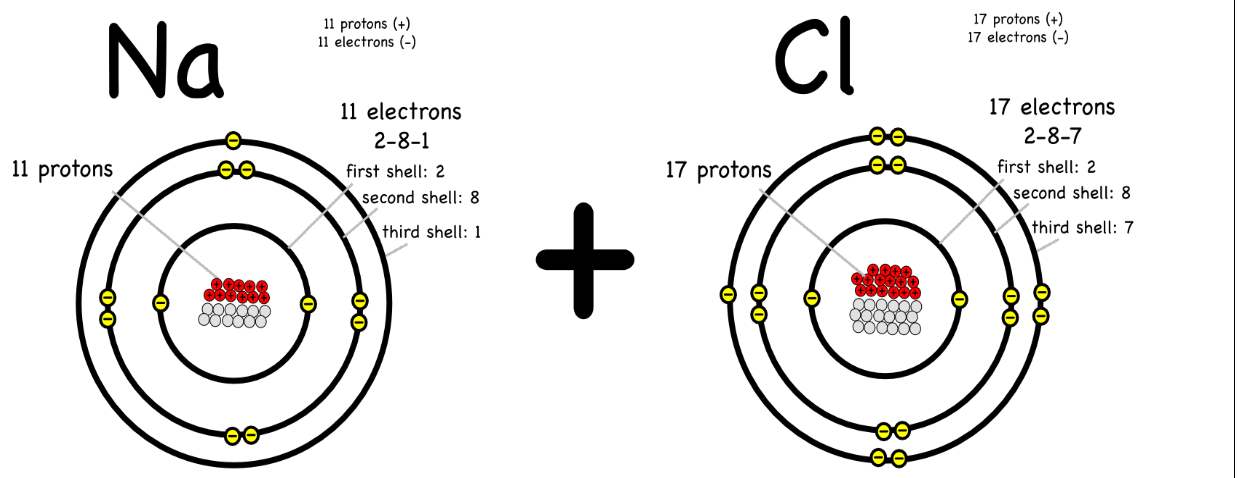
	Before combination		After combination	
	sodium atom	chlorine atom	sodium ion	chloride ion
Proton	11	17	11	17
Electron	2,8,1	2,8,7	2,8	2,88

Equations of reaction 1. $Na \longrightarrow Na^+ + e^-$

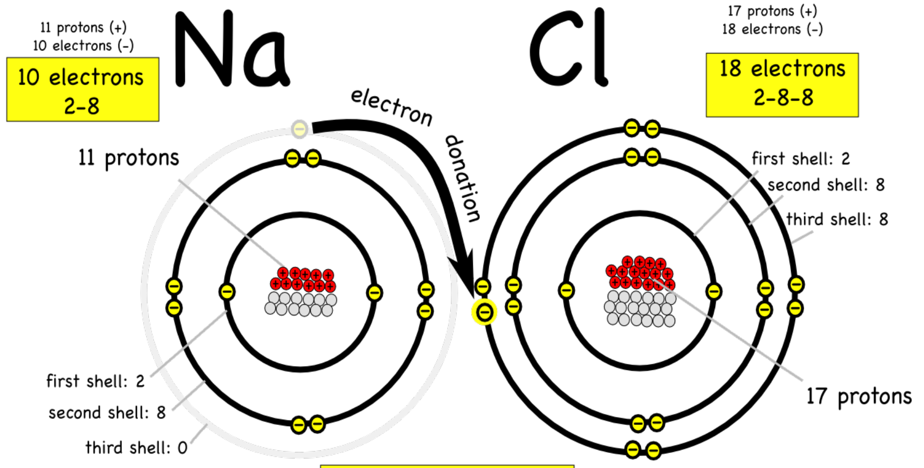
2. $Cl + e^- \longrightarrow Cl^-$

add equation 1 and 2: $Na + Cl + e^- \longrightarrow Na^+ + Cl^- + e-$ which gives:

$$Na + Cl \longrightarrow Na^+ + Cl^-$$



Sodium (Na) donates its outer-shell electron to chlorine (Cl)

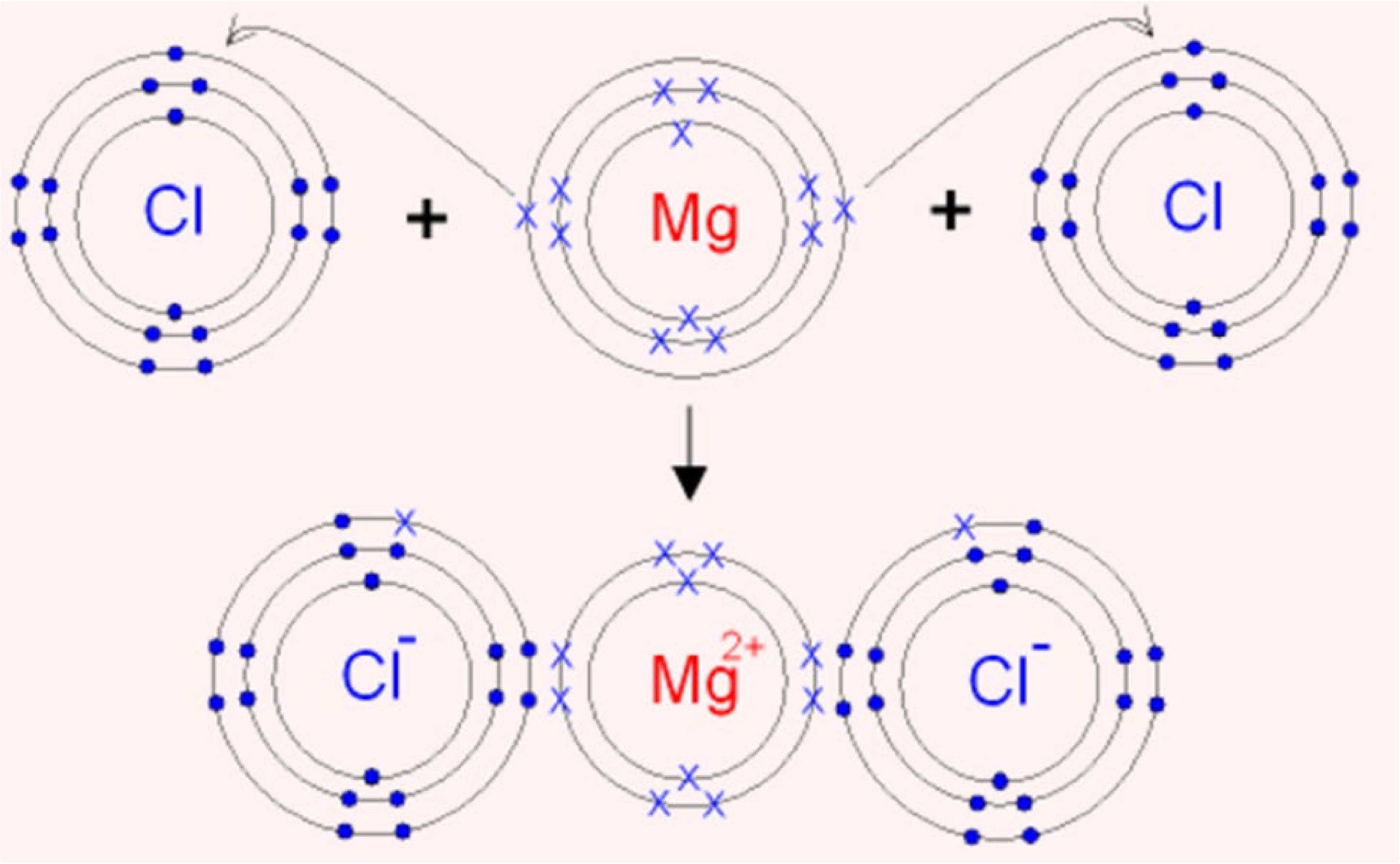


Chemical Formula: **NaCl**

- **<u>Formation of Magnesium Chloride</u> $MgCl_2$:**

	Before combination			After combination		
	Mg atom	Cl	Cl atoms	Mg	Cl	Cl ions
Protons	12	17	17	12	17	17
Elec. Config.	2,8,2	2,8.7	2,8,7	2,8	2,8,8	2,8,8

- Equation for reaction:
- $Mg + Cl_2 \longrightarrow Mg^{2+} + 2Cl^-$



Assignment

- *Illustrate the formation of the following:*
- *(i). CaO*
- *(ii). $AlCl_3$.   Show shell diagrams*

- *The atomic numbers of each element*
  - *Ca= 20:*
  - *Al = 13*
  - *O= 8*
  - *Cl=17*

Properties of Electrovalent (ionc) compounds:

Electrovalent or ionic compounds are compounds formed as a result of electrovalent combination. The following are properties of electrovalent compounds:

1. - They are mostly solids at room temperature.
2. - They are made up of aggregate ions.
3. - They are soluble in water and other polar solvents
4. - they are electrolytes(conduct electricity in molten or aqueous solution.
5. - They have high melting and boiling points.

Solubility of ionic compounds:

- When ionic compound such as NaCl is dissolved in water, energy is required to break the ionic bond(lattice energy). Then the ions are surrounded by water molecules liberating heat energy (hydration energy). The solubility of a substance depends on the net value of the two forms of energy.
- An ionic compound will dissolve if: Hydration energy > Lattice energy
- An ionic compound will not dissolve if Hydration Energy < Lattice energy

Covalent Combination: This involves sharing of electrons by the combining atoms: there are two types of covalent combination:

- Ordinary covalent combination and
- Dative or co-ordinate combination.

1. **Ordinary covalent combination**:

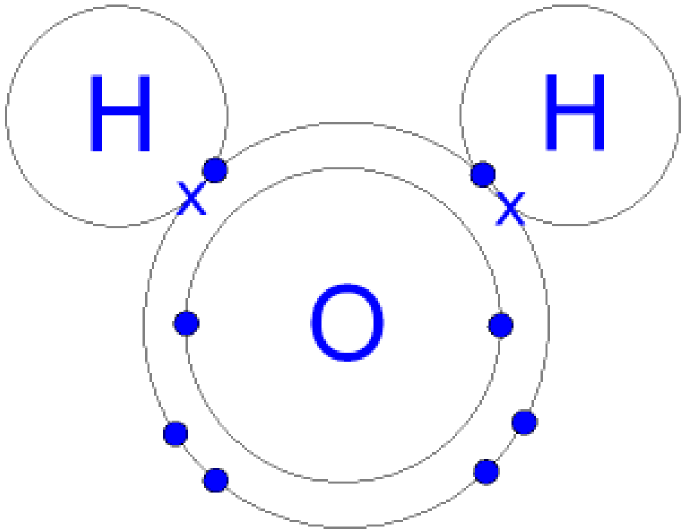
Covalent combination involves the sharing of electrons between atom of the same or different elements such that, each combining atom contributes a number of electron(s) (equal to its valency) to a shared pair in order to attain octet structure.

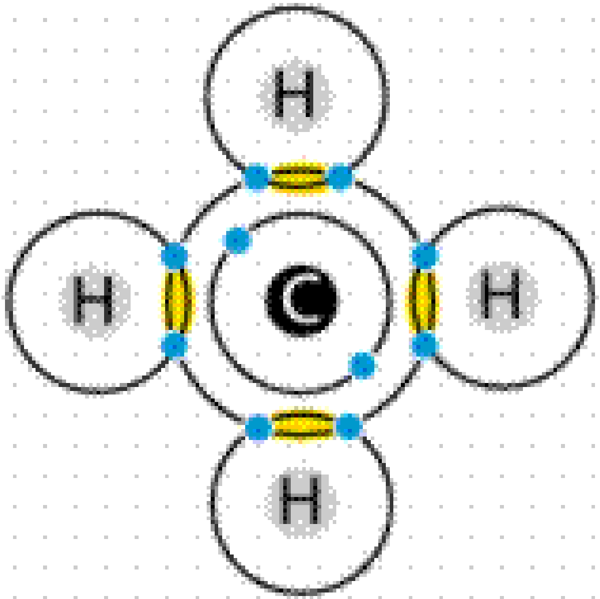
The bond formed is called covalent bond while the compound is covalent compound.

Example of covalent compounds are Cl_2, CO_2, H_2O, HCl, CH_4 e.t.c

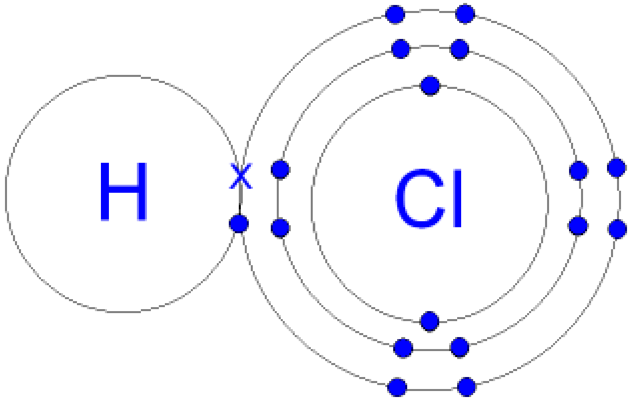
Diagram of covalent molecules

Water H_2O





Methane CH_4



Hydrogen chloride HCl

Assignment

Illustrate the formation of the following covalent compounds

(i). CO_2

(ii). Cl_2.   Show shell diagrams

Properties of covalent compounds

- They exist as molecules which have a definite shape
- They exist as gases or volatile liquids at room temperature but as solids with low melting points e.g camphor and Iodine
- They have low melting and boiling points due to weak intermolecular forces
- They are soluble in non-polar solvents but insoluble in polar solvents (water)
- They do not conduct electricity and hence, they are non-electrolytes due to absence of ions

Co-ordinate covalent (Dative) bonding

- This is a type of covalent bonding in which the electron shared is contributed by one of the participating atoms.
- It occurs between an atom that has a lone-pair of electrons and an outer shell of another that has an empty valence orbital and therefore requires one pair or electron to attain octet or duplet state.
- Example of Coordinate covalent compound is: ammonium ion [NH_4^+], hydronium ion [H_3O^+], hydrated copper (II) ion, $[Cu(H_2O)_4]^+$ Phosphorus oxochloride $POCl_3$, tetraamine copper (II) ion $[Cu(NH_3)_4]^{2+}$

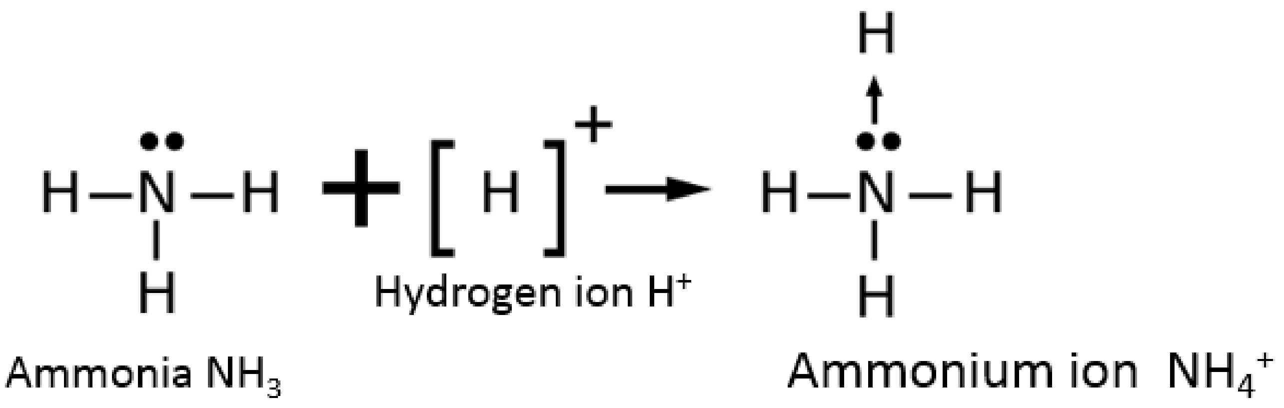
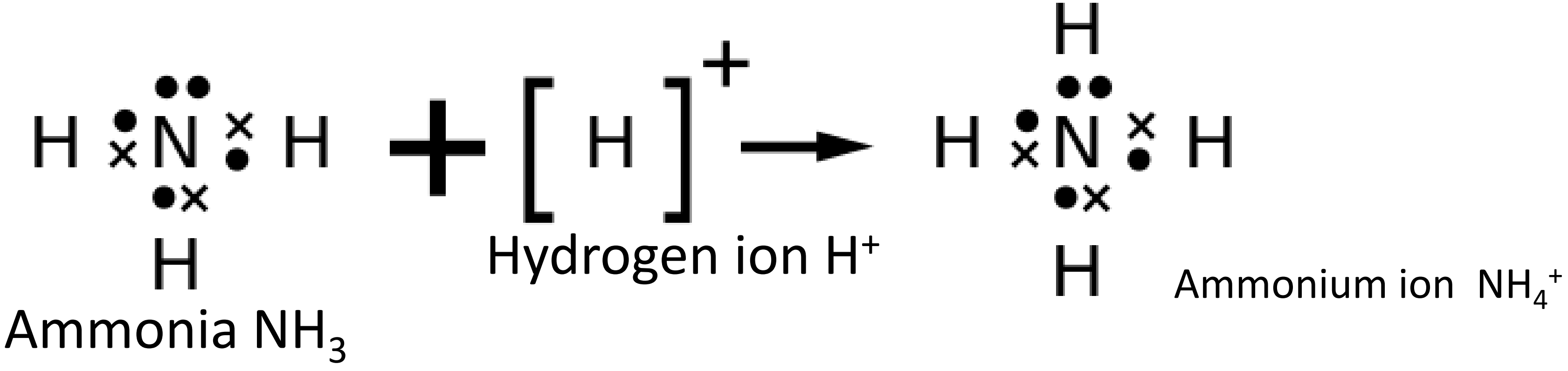
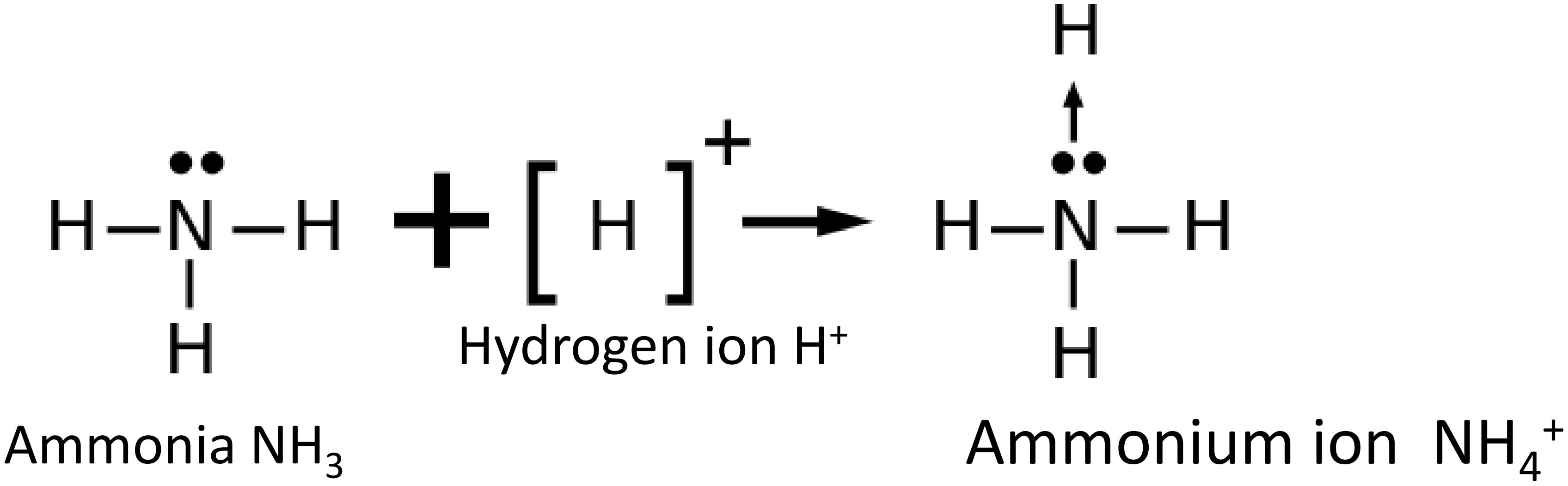


Diagram of ammonium ion (co-ordinate covalent compound



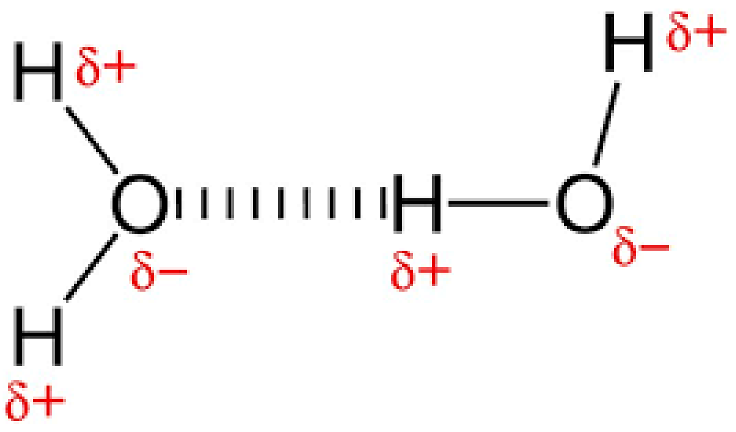


Hydrogen bonding

- Hydrogen bonding is a dipole-dipole intermolecular attraction which occurs when hydrogen is covalently to highly electronegative elements of small atomic size.
- Hydrogen bond is present in the following compounds: HF, H_2O
- **Importance of hydrogen bond**:

I. It accounts for the solubility of some compounds containing O, N and F in water

II. It provides the attractive force that keeps water molecules together



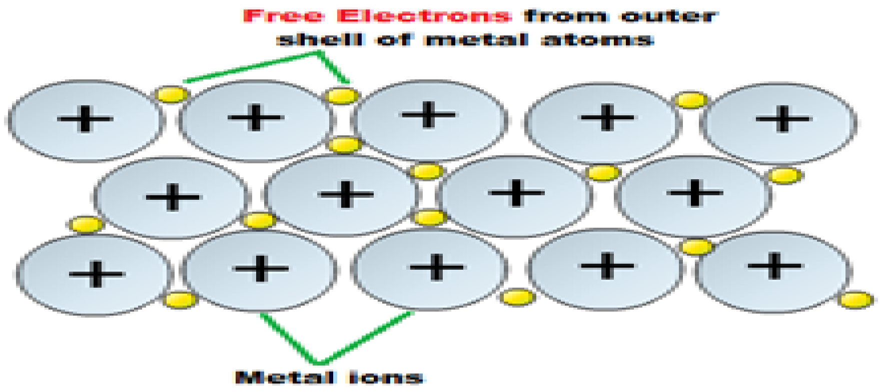
Vander Waal's forces

- This is the weakest type of bonding when compared to ionic and covalent.
- It is the attractive force which makes it possible for non-polar molecules like Nitrogen gas, carbon (IV) oxide, oxygen to form Solids and liquids.
- Vander Waal's forces increase with increase in the number of electrons. They are stronger in Iodine than bromine and less in chlorine.

Metallic bonding

Metallic bonding is an attraction between the positive nuclei of a metal and the sea of mobile electrons of another metallic nuclei.

It is stronger in metals like Iron but weaker in metals such as sodium



Assignment

- 1. Give three main difference between covalent and electrovalent compounds
- 2. An element X with electronic configuration 2, 8, 1 ionizes to 2,8 when it combines with another element Y of configuration 2, 8, 7.
- (a) state the type of bonding formed between the two elements
- (b) How many atom(s) of Y are needed for the combination
- (c) Write the formular of the compound formed.

KINETIC THEORY OF MATTER AND GAS LAWS

Terms in Kinetic theory of matter.

Evaporation: it is the process of vaporization of liquids at all temperatures. Rate of evaporation is lower in electrovalent liquids than covalent liquids. This is due to the force of attraction which is greater between positive and negative ions (unlike neutral covalent molecules)

Critical temperature: This is the temperature abo ve which a gas cannot be liquefied by pressure alone.

Vapour Density: This is the number of times a given volume of a gas is as heavy as an equal volume of hydrogen at constant temperature and pressure

Brownian motion: this is the continuous random movement of small particles suspended in a fluid which arise from collision with the fluid molecules.

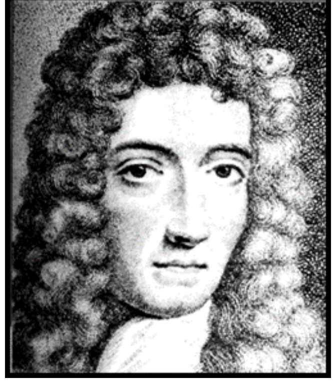
POSTULATES OF KINETIC THEORY OF GASES(IDEAL GASES

1. Gas molecules are constant random motion and move in straight lines
2. Molecules of gases collide with each other and the walls of the container
3. The volume occupied by gas molecules is small and negligible compared to the volume of the container
4. The force of attraction between gas molecules is negligible
5. Collision between gas molecules is perfectly elastic i.e. there is no loss of kinetic energy
6. The average kinetic energy of gas molecules is directly proportional to the absolute temperature

Phenomenon that support the kinetic theory of gases are: Diffusion, Osmosis, Brownian motion

Reasons why real gases deviate from ideal gas postulates

1. Molecules of real gases have significant attractive forces between them.
2. The volume of the molecules of real gases is not negligible.
3. Collision between the molecules are not perfectly elastic



Boyle's law states that at constant temperature, the volume of a fixed mass of gas is inversely proportional to its pressure

Mathematically, Boyle's law can be stated as

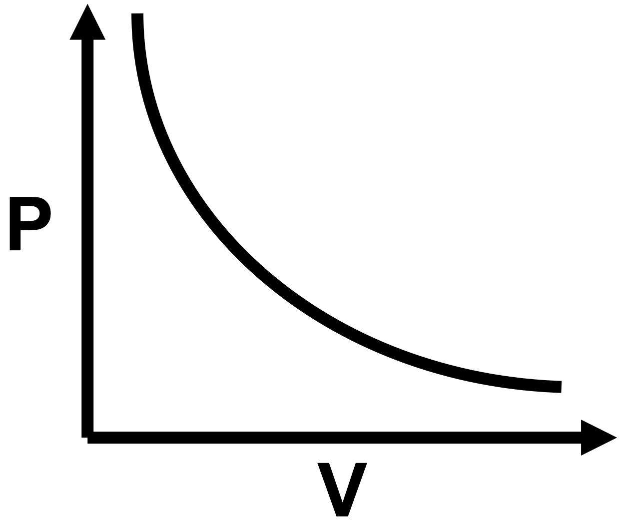
$$\mathbf{P \propto 1/V,}$$

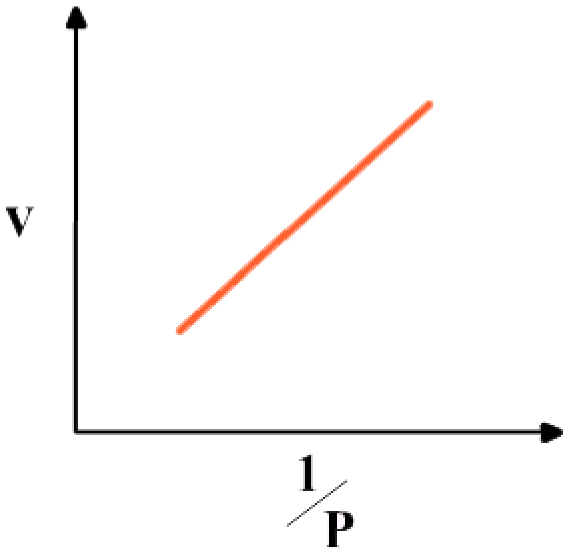
$$\mathbf{PV = k}$$

$$\mathbf{P_1V_1 = P_2V_2}$$

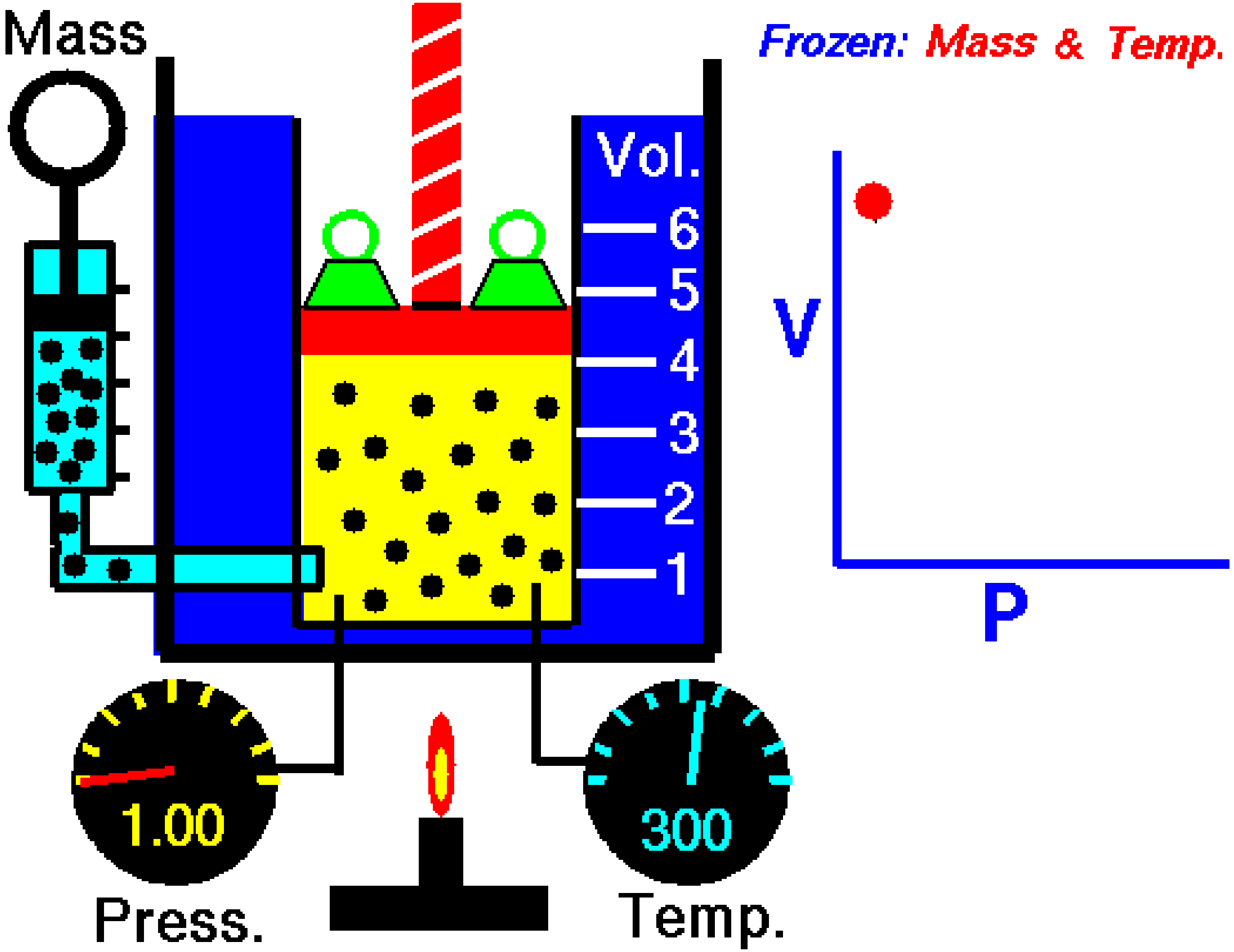
where *P* is the pressure of the gas, *V* is the volume of the gas, and *k* is a constant.

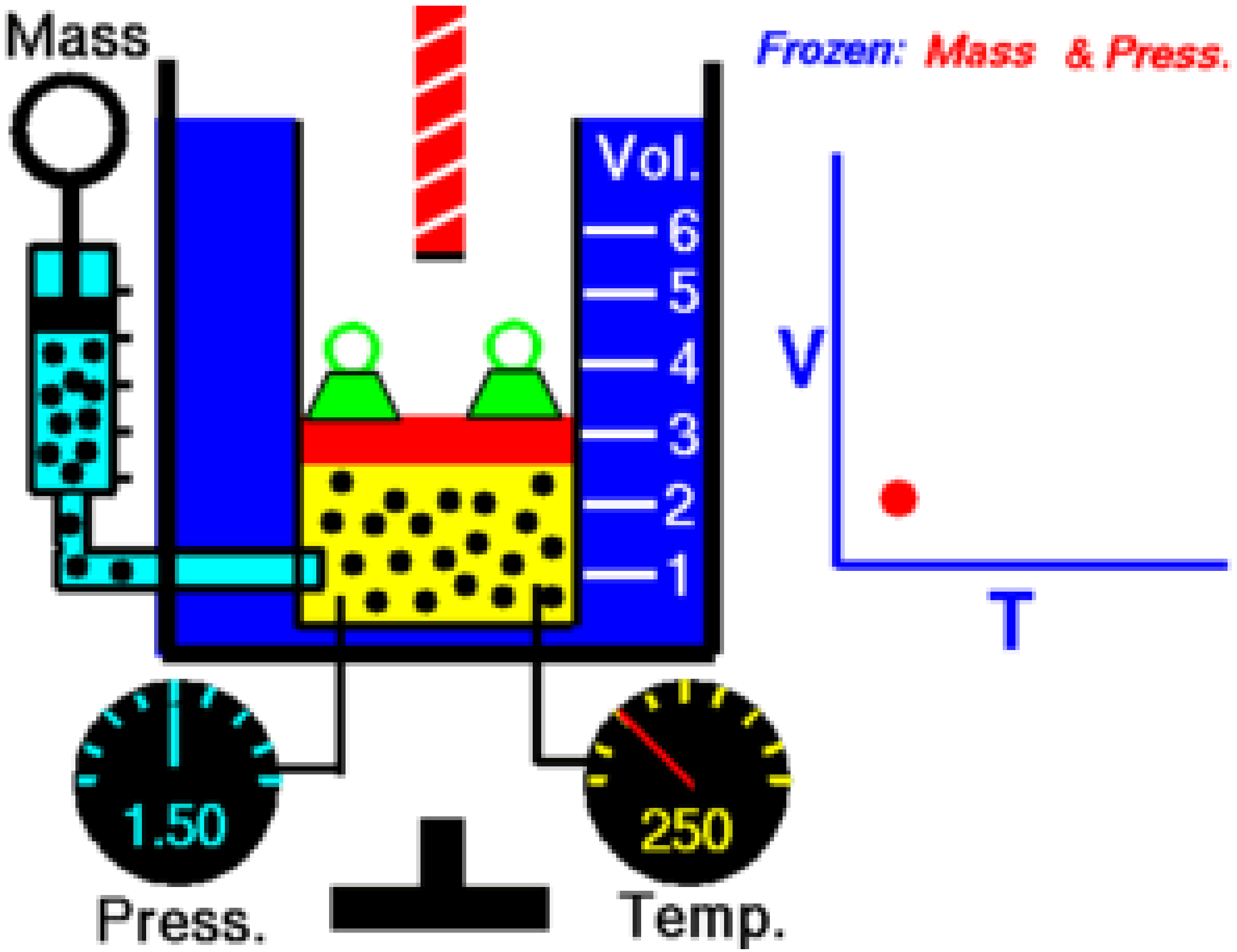
- Graphically,





- **NB: As pressure increases, volume decreases**


Mass
Vol.
6
5
4
3
2
1
1.00
Press.
300
Temp.
Frozen: Mass & Temp.
V
P

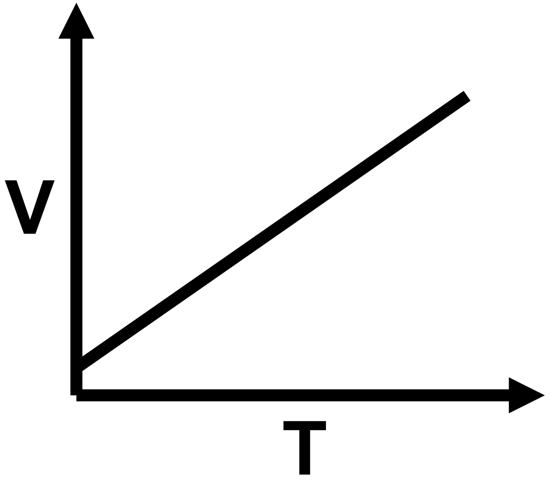


A decrease in temperature will lead to a decrease in volume.

Charle's Law

- It states that the volume(V) of a fixed mass of a gas is directly proportional to absolute temperature(T in kelvin) at constant pressure,

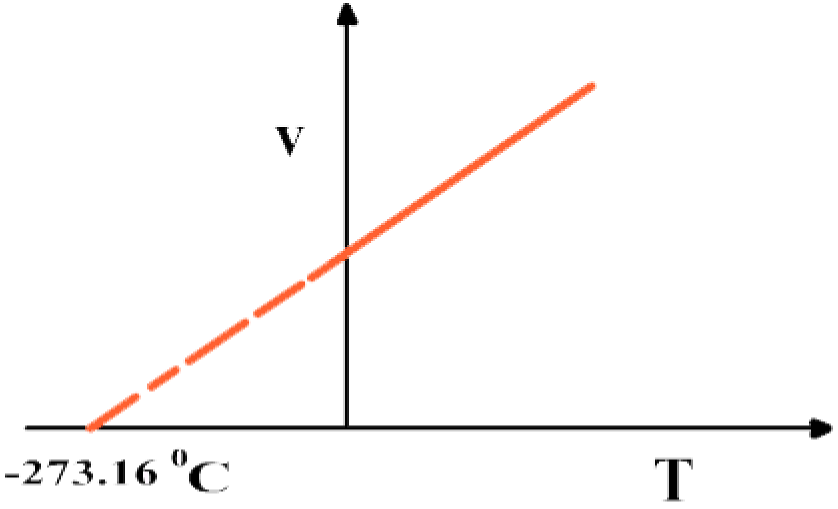
Graphically,



- Mathematically, V **α T:** $\frac{V}{T}$ = k

- $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

V is the volume of the gas,
T is the temperature of the gas (measured in kelvins),
k is a constant



T (K) = 273 + T⁰C

The advantage of Kelvin scale is that it eliminates zero or negative temperatures.

Ideal Gas equation

- PV = nRT
- When n=1, and R is a constant.
- The equation translates as
- $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
- n=number of moles

UNIVERSAL GAS CONSTANT
R=0.0821 L·atm/mol·K or
R=8.315 dm^3·kPa/mol·K

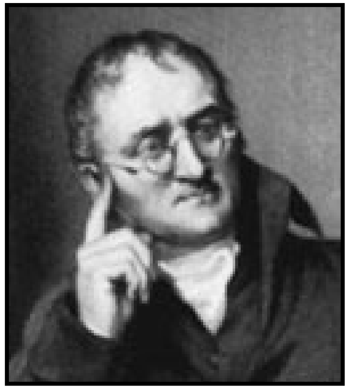
E. Gas Law Problems

- A gas occupies 473 cm^3 at 36°C. Find its volume at 94°C.

CHARLES' LAW

GIVEN: **T↑ V↑**	WORK:
$V_1 = 473\ cm^3$ $T_1 = 36°C = 309K$ $V_2 = ?$ $T_2 = 94°C = 367K$	$\cancel{P}_1 V_1 T_2 = \cancel{P}_2 V_2 T_1$ $(473\ cm^3)(367\ K) = V_2(309\ K)$ **$V_2 = 562\ cm^3$**

Dalton’s Law



- Dalton’s Law of partial pressure states that the total pressure of a mixture of gases equals the sum of the partial pressures of the individual gases.

$$P_{total} = P_1 + P_2 + \ldots$$

Avogadro’s law states that Equal volumes of gases contain equal numbers of molecules at constant temperature & pressure.

B. Ideal Gas Law

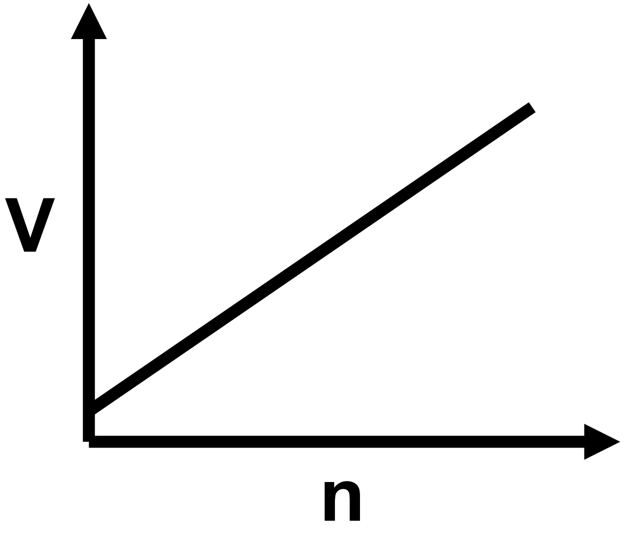
$PV=nRT$

UNIVERSAL GAS CONSTANT

$R=0.0821\ L\cdot atm/mol\cdot K$

$R=8.315\ dm^3\cdot kPa/mol\cdot K$

A. Avogadro's Law



$$\frac{V}{n} = k$$