ALKALI AND ALKALINE EARTH METALS

Explain the common features of alkali and alkaline earth metals

Features	Alkali metals	Alkaline earth metals
Physical state and Natural abundance	i) Alkali metals are highly reactive and are found in nature only as compounds. ii) Rubidium and Caesium are found associated in minute quantities with minerals of other alkali metals iii) Francium is radio active and its half life is 21 minutes	Beryllium is rare and radium is the rarest of all comprising only 10 % of igneous rocks. Magnesium and calcium are very common in the earth's crust, with calcium the fifth-most-abundant element, and magnesium the eighth. Magnesium and calcium are found in many rocks and minerals
Electronic configuration	ns^1	ns^2
Common oxidation state	+1 (M^+ ion)	+2 (M^{2+} ion)
Atomic and ionic radii	Alkali metals have the largest atomic and ionic radii in their respective periods. On moving down the group, there is an increase in the atomic and ionic radii.	Alkaline earth metals have the smaller atomic and ionic radii than alkali metals On moving down the group, there is an increase in the atomic and ionic radii.
Ionisation Enthalpy	i) Alkali metals have the lowest ionisation enthalpy compared to other elements present in the respective period. As we go down the group, the ionisation enthalpy decreases due to the increase in atomic size. ii) The second ionisation enthalpies of alkali metals are very high. The removal of an electron from the alkali metals gives monovalent cations having stable electronic configurations similar to the noble gas. Therefore, it becomes very difficult to remove the second electron from the stable configurations already attained.	Due to a fairly large size of the atoms, alkaline earth metals have low ionisation enthalpies when compared to 'p' block elements. Down the group the ionisation enthalpy decreases as atomic size increases. This is due to the addition of new shells as well as increase in the magnitude of the screening effect of inner shell electrons. Members of group 2 have higher ionization enthalpy values than group 1 because of their smaller size Although IE_1 values of alkaline earth metals are higher than that of alkali metals, the IE_2 values of alkaline earth metals are much smaller than those of alkali metals.

<table>
<tr>
<td>Hydration enthalpy</td>
<td>Lithium salts are more soluble than the salts of other metals of group 1. eg. $LiClO_4$ is up to 12 times more soluble than $NaClO_4$. $KClO_4$, $RbClO_4$ and $CsClO_4$ have solubilities only 10^{-3} times of that of $LiClO_4$. The high solubility of Li salts is due to strong solvation of small size of Li^+ ion.</td>
<td>Compounds of alkaline earth metals are more extensively hydrated than those of alkali metals, because the hydration enthalpies of alkaline earth metal ions are larger than those of alkali metal ions.

Like alkali metal ions, the hydration enthalpies of alkaline earth metal ions also decrease with increase in ionic size down the group.
Be > Mg > Ca > Sr > Ba</td>
</tr>
<tr>
<td>Electronegativity</td>
<td>Alkali metals have comparatively smaller value of electronegativity than the other elements in the respective period. When they react with other elements, they usually produce ionic compounds.</td>
<td>In alkaline earth metals the electronegativity values decrease as we go down the group as seen in the alkali metals.</td>
</tr>
<tr>
<td>Flame colour and the spectra</td>
<td>When the alkali metal salts moistened with concentrated hydrochloric acid are heated on a platinum wire in a flame, they show characteristic coloured flame as shown below.
<table>
<tr><th>Element</th><th>Colour</th></tr>
<tr><td>Lithium</td><td>Crimson red</td></tr>
<tr><td>Sodium</td><td>Yellow</td></tr>
<tr><td>Potassium</td><td>Lilac</td></tr>
<tr><td>Rubidium</td><td>Reddish violet</td></tr>
<tr><td>Caesium</td><td>Blue</td></tr>
</table>
</td>
<td>When the alkaline earth metal salts moistened with concentrated hydrochloric acid are heated on a platinum wire in a flame, they show characteristic coloured flame
<table>
<tr><th>Element</th><th>Colour</th></tr>
<tr><td>Calcium</td><td>Brick - Red</td></tr>
<tr><td>Strontium</td><td>Crimson</td></tr>
<tr><td>Barium</td><td>Apple Green</td></tr>
</table>
</td>
</tr>
<tr>
<td>Diagonal relationship</td>
<td>Similarity between the first member of group 1 (Li) and the diagonally placed second element of group 2 (Mg) is called diagonal relationship.
Eg.,
It is due to similar size ($r\ Li^+$ = 0.766 Å and Mg^{2+} = 0.72 Å) and comparable electronegativity values (Li = 1.0; Mg = 1.2).</td>
<td>As observed in alkali metals, beryllium (the first member of group 2) shows a diagonal relationship with aluminium. In this case, the size of these ions (rBe^{2+} = 0.45 Å and rAl^{3+} = 0.54 Å) is not as close. However, their charge per unit area is closer (Be^{2+} = 2.36 and Al^{3+} = 2.50). They also have same electronegativity values (Be = 1.5; Al = 1.5).</td>
</tr>
<tr>
<td>Distinctive behavior</td>
<td>The distinctive behaviour of Li^+ ion is due to its exceptionally small size, high polarising power, high hydration energy and non availability of d-orbitals.</td>
<td>The anomalous properties of beryllium is mainly due to its small size, high electronegativity, high ionisation energy and high polarising power compared to the other elements in the block.</td>
</tr>
</table>

2. Chemical properties of alkali metals

Reaction with oxygen

All the alkali metals on exposure to air or oxygen burn vigorously, forming oxides on their surface. Lithium forms only monoxide, sodium forms the monoxide and peroxide and the other elements form monoxide, peroxide, and superoxides. These oxides are basic in nature.

$4\ Li + O_2 \rightarrow 2Li_2O$ (simple oxide)

$2\ Na + O_2 \rightarrow Na_2O_2$ (peroxide)

$M + O_2 \rightarrow MO_2$

(M= K, Rb, Cs; MO_2 - superoxide)

Reaction with hydrogen

All alkali metals react with hydrogen at about 673 K (lithium at 1073 K) to form the corresponding ionic hydrides. Reactivity of alkali metals with hydrogen decreases from Li to Cs.

$$2M + H_2 \rightarrow 2\ M^+H^- \quad (M = Li, Na, K, Rb, Cs)$$

The ionic character of the hydrides increases from Li to Cs and their stability decreases. The hydrides behave as strong reducing agents and their reducing nature increases down the group.

Reaction with liquid ammonia:

Alkali metals dissolve in liquid ammonia to give deep blue solutions that are conducting in nature. This happens because the alkali metal atom readily loses its valence electron in ammonia solution. Both the cation and the electron are ammoniated to give ammoniated cation and ammoniated electron.

$$M + (x + y)NH_3 \rightarrow [M(NH_3)_x]^+ + [e(NH_3)_y]^-$$

The blue colour of the solution is due to the ammoniated electron which absorbs energy in the visible region of light and thus imparts blue colour to the solution. The solutions are paramagnetic and on standing slowly liberate hydrogen resulting in the formation of an amide.

$$M^+ + e^- + NH_3 \rightarrow MNH_2 + \tfrac{1}{2}H_2$$

In concentrated solution, the blue colour changes to bronze colour and become diamagnetic.

Reaction with water:

Alkali metals react with water to give corresponding hydroxides with the liberation of hydrogen.

$$2\ Li + 2\ H_2O \rightarrow 2\ LiOH + H_2$$

They also react with alcohol, and alkynes which contain active hydrogens.

$$2\ Na + 2\ C_2H_5OH \rightarrow 2\ C_2H_5ONa + H_2$$

$$H\text{-}C \equiv C\text{-}H \xrightarrow{Na} H\text{-}C \equiv C\text{-}Na \xrightarrow{Na} Na\text{-}C \equiv C\text{-}Na$$

Reducing activity:

Alkali metals can lose their valence electron readily hence they act as good reducing agents.

$$M_{(s)} \rightarrow M^+_{(g)} + e^-$$

Reaction with carbon:

Lithium directly reacts with carbon to form the ionic compound, lithium carbide. Other metals do not react with carbon directly. However, when they are treated with compounds like acetylene they form acetelydes.

$$2\ Li + 2C \rightarrow Li_2C_2$$

Write the Uses of alkali metals.

i. Lithium metal is used to make useful alloys.

For example with lead it is used to make 'white metal' bearings for motor engines, with aluminium to make aircraft parts, and with magnesium to make armour plates. It is used in thermonuclear reactions.

ii. Lithium is also used to make electrochemical cells.

iii. Lithium carbonate is used in medicines

iv. Sodium is used to make Na/Pb alloy needed to make $Pb(Et)_4$ and $Pb(Me)_4$ is used as anti-knock additives to petrol, but nowadays lead-free petrol in use.

v. Liquid sodium metal is used as a coolant in fast breeder nuclear reactors. Potassium has a vital role in biological systems.

vi. Potassium chloride is used as a fertilizer. Potassium hydroxide is used in the manufacture of soft soap. It is also used as an excellent absorbent of carbon dioxide.

vii. Caesium is used in devising photoelectric cells.

ALKALI AND ALKALINE EARTH METALS
ALKALI ELEMENTS

5 MARKS

1. Explain the common features of alkali elements.
2. Explain the chemical properties of alkali elements.
3. Write the uses of alkali elements.
4. Explain the preparation of sodium carbonate by solvey process.
5. Write the uses of sodium carbonate.
6. Write the commercial preparation of NaCl. Write its uses. (OR) How is NaCl obtain from brain solution? How is it purified? Write its uses.
7. How is NaOH obtained by Castner-Kellner process? Write its uses.
8. How is sodium carbonate prepared? Write its uses.
9. Explain the biological importance of sodium and potassium.
10. Explain the distinctive behavior of lithium with other elements of the group.
11. Write the similarities (OR) diagonal relationship between lithium and magnesium.

ALKALINE EARTH METALS

1. Explain the common features of alkaline earth metals.
2. Compare the properties of beryllium with other elements of the group (or) distinctive behavior of Beryllium.
3. Explain the similarities (OR) diagonal relationship between Be and Al.
4. Explain the nature of halides of alkaline earth metals.
5. How is quick lime prepared? Write the chemical properties of quick lime? Write its uses.
6. Explain the preparation, properties and uses of $Ca(OH)_2$.
7. Write the various properties of Gypsum.
8. Write the uses of Gypsum.
9. How is plaster of paris prepared? Write its uses.
10. Explain the Biological importance of Magnesium and calcium.

3 Marks

1. Why the solubilities of sulphates of alkaline earth metals decreases on going down the group?
2. Write and prove amphoteric Nature of $Be(OH)_2$.
3. How is BeF_2 & $BeCl_2$ prepared?
4. How is BeH2 prepared?
5. Why lithium carbonate unstable?

BASIC CONCEPTS OF CHEMISTRY AND CHEMICAL CALCULATIONS

QUESTIONS COLLECTIONS

2 & 3 MARKS

1. How can matter classified?
2. Define unified atomic mass.
3. Define relative atomic mass based on amu scale. (B.B)
4. Define relative molecular mass based on amu scale.
5. Define molecular mass.
6. Define the term mole. (B.B)
7. Define Avogadro Number.
8. Define molar mass.
9. Differentiate molar mass and molecular mass (B.B)
10. Define molar volume.
11. Define gram equivalent mass.
12. How can you calculate the equivalent mass of the acid, base, oxidizing agent or reducing agent?
13. What is empirical formula?
14. What is molecular formula?
15. What is the meaning of 'n'? How is it determined?
16. What is stoichiometry?
17. What is limiting reagent?
18. What are redox reactions? Give one example.
19. Define oxidation number?
20. Name the types of redox reactions.
21. Define redox reactions interms of oxidation number.
22. Distinguish between oxidation and reduction. (B.B)
23. What happen when a person consumes cyanide poison? Explain.
24. Hydrogen peroxide is an oxidising agent. It oxidises ferrous ion to ferric ion and reduced itself to water. Write a balanced equation.

5 marks

1. Explain the classification of matter.
2. Write the steps involving for determination of empirical formula.
3. Even a small amount of oxygen present in air leads to the rusting of iron. But Fe2+ ion present in haemoglobin which binds oxygen during transport of oxygen from lungs to tissues never gets oxidized. Why?
4. Write the rules for assigning oxidation number.
5. Explain the types of redox reactions.

Problems (2 mark & 3 Mark)

1. Calculate the molar mass of the following.
(i) Ethanol(C_2H_5OH)
(ii) Potassium permanganate ($KMnO_4$)
(iii) Potassium dichromate ($K_2Cr_2O_7$)
(iv) Sucrose ($C_{12}H_{22}O_{11}$)
2. Calculate the mass of single atom or molecule for the following.
a) C b) glucose c) potassium di chromate
d) potassium per manganate
3. Calculate the amount of acid neutralized by an antacid that contains 250mg if aluminium hydroxide and magnesium hydroxide respectively when 0.1M of gastric acid present in our stomach.
4. Calculate the number of moles present in 9 g of ethane.
5. Calculate the number of molecules of oxygen gas that occupies a volume of 224 ml at 273 K and 3 atm pressure.
6. Calculate the equivalent mass of i) H_2SO_4 ii) KOH iii) $KMnO_4$
7. 0.456 g of a metal gives 0.606 g of its chloride. Calculate the equivalent mass of the metal.
8. Calculate the equivalent mass of potassium dichromate. The reduction half-reaction in acid medium is,
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2\, Cr^{3+} + 7H_2O$
9. How many moles of hydrogen is required to produce 10 moles of ammonia ?
10. Calculate the amount of water produced by the combustion of 32 g of methane.
11. How much volume of carbon dioxide is produced when 50 g of calcium carbonate is heated completely under standard conditions?
12. How much volume of chlorine is required to form 11.2 L of HCl at 273 K and 1 atm pressure ?
13. The balanced equation for a reaction is given below

$$2x+3y \rightarrow 4l + m$$

When 8 moles of x react with 15 moles of y, then

i) Which is the limiting reagent?
ii) Calculate the amount of products formed.
iii) Calculate the amount of excess reactant left at the end of the reaction.

14. Calculate the oxidation number for the following
i) CO_2 ii) H_2SO_4 iii) C_2H_2 iv) $Cr_2O_7^{2-}$ v) CH_2F_2
vi) SO_2

15. Balance the following equation using oxidation number method
$As_2S_3 + HNO_3 + H_2O \rightarrow H_3AsO_4 + H_2SO_4 + NO$

16. Calculate the molar mass of the following compounds.
i) urea $[CO(NH_2)_2]$ ii) acetone $[CH_3COCH_3]$
iii) boric acid $[H_3BO_3]$ iv) sulphuric acid $[H_2SO_4]$

17) The density of carbon dioxide is equal to 1.965 kgm^{-3} at 273 K and 1 atm pressure. Calculate the molar mass of CO_2.

18) Which contains the greatest number of moles of oxygen atoms
i) 1 mol of ethanol ii) 1 mol of formic acid
iii) 1 mol of H_2O

19. Calculate the average atomic mass of naturally occurring magnesium using the following data

Isotope	Isotopic atomic mass	Abundance (%)
Mg^{24}	23.99	78.99
Mg^{26}	24.99	10.00
Mg^{25}	25.98	11.01

20) Mass of one atom of an element is 6.645 x 10^{-23} g. How many moles of element are there in 0.320 kg.

21) What is the empirical formula of the following ?
i) Fructose ($C_6H_{12}O_6$) found in honey
ii) Caffeine ($C_8H_{10}N_4O_2$) a substance found in tea and coffee.

22. How many moles of ethane is required to produce 44 g of CO_2 (g) after combustion.

5 marks

23. In a process, 646 g of ammonia is allowed to react with 1.144 kg of CO_2 to form urea.

i) If the entire quantity of all the reactants is not consumed in the reaction
ii)which is the limiting reagent ?

24) Calculate the quantity of urea formed and unreacted quantity of the excess reagent.

25) Experimental analysis of a compound containing the elements x,y,z on analysis gave the following data. x = 32 %, y = 24 %, z = 44 %. The relative number of atoms of x, y and z are 2, 1 and 0.5, respectively. (Molecular mass of the compound is 400 g) Find out.
i) The atomic masses of the element x,y,z.
ii) Empirical formula of the compound and
iii) Molecular formula of the compound.

26. A Compound on analysis gave the following percentage composition C=54.55%, H=9.09%, O=36.36%. Determine the empirical formula of the compound.

27. Calculate the empirical and molecular formula of a compound containing 76.6% carbon, 6.38 % hydrogen and rest oxygen its vapour density is 47.

28. A Compound on analysis gave Na = 14.31% S = 9.97% H= 6.22% and O= 69.5% calculate the molecular formula of the compound if all the hydrogen in the compound is present in combination with oxygen as water of crystallization. (molecular mass of the compound is 322).

29) Balance the following equations by oxidation number method
i) $K_2Cr_2O_7 + KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + I_2 + H_2O$
ii) $KMnO_4 + Na_2SO_3 \rightarrow MnO_2 + Na_2SO_4 + KOH$
iii) $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$
iv) $KMnO_4 + H_2C_2O_4 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + CO_2 + H_2O$

30) Balance the following equations by ion electron method.
i) $KMnO_4 + SnCl_2 + HCl \rightarrow MnCl_2 + SnCl_4 + H_2O + KCl$
ii) $C_2O_4^{2-} + Cr_2O_7^{2-} \rightarrow Cr^{3+} + CO_2$ (in acid medium)
iii) $Na_2S_2O_3 + I_2 \rightarrow Na_2S_4O_6 + NaI$ (in acid medium)
iv) $Zn + NO_3^{-} \rightarrow Zn^{2+} + NO$

GASEOUS STATE

Book back questions and answers

26. State Boyle's law.

At a given temperature the volume occupied by a fixed mass of a gas is inversely proportional to its pressure.

Mathematically, the Boyle's law can be written as

V α 1/ P

(T and n are fixed, T-temperature, n- number of moles)

V = k x 1/P

k - proportionality constant

When we rearrange above equation

PV = k at constant temperature and mass

27. A balloon filled with air at room temperature and cooled to a much lower temperature can be used as a model for Charle's law.

According to Charle's law the volume of given mass of gas is directly proportional to temperature at constant pressure.

V α T (at constant P,n)

If we take a balloon filled with air at room temperature and then cooled to a much lower temperature, the size of balloon is reduced.

This is because the gas molecule inside the balloon move slower due to decrease in temperature and hence the volume decreases.

This will prove the Chrle's law and it can be used as a model for Charle's law.

28. Name two items that can serve as a model for Gay Lusaac' law and explain.

1. Firing a bullet. When gunpowder burns, it creates a significant amount of superheated gas. The high pressure of the hot gas behind the bullet forces it out of the barrel of the gun.

2. Heating a closed aerosol cane. The increased pressure may cause the container to explode. You don't toss an "empty" can of hairspray into a fire.

29. Give the mathematical expression that relates gas volume and moles. Describe in words what the mathematical expression means.

- The Mathematical expression between the volume of gas and number of moles is
  V α n

$V_1/n_1 = V_2/n_2$ = constant

- where V_1 & n_1 are the volume and number of moles of a gas and V_2 & n_2 are a different set of values of volume and number of moles of the same gas at same temperature and pressure.
- The above relation is expressed as " Equal volumes of all gases under the same condition of temperature and pressure contain equal number of molecules"
-

30. What are ideal gases? In what way real gases differ from ideal gases.

- An ideal gas a gas which obeys the gas laws and ideal gas equation PV=nRT.
- An ideal gas is defined as one in which all collisions between atoms or molecules are perfectly elastic forces.

- The difference between ideal gas and real gas is

Sl.No	Ideal gas	Real gas
1	It obeys gas laws under all conditions of temperature and pressure	It obeys gas laws only under low pressure and high temperature
2.	No gas is ideal	All gases are real
3.	Volume occupied by the molecules is negligible as compared to the total volume occupied by the gas.	Volume occupied by the molecules is not negligible as compared to the total volume occupied by the gas.
4.	The forces of attraction among the molecules of the gas are negligible	The forces of attraction among the molecules cannot be neglected at high pressure and low temperature
5.	It obeys ideal gas equation PV = nRT	It obeys Van der Waals equation $\left(P+\frac{an^2}{V^2}\right)$ (V-nb) = nRT

31. Can a Van der Waals gas with a=0 be liquefied? explain.

- The van der Waals constant 'a' is a measure of the attractive forces among the molecules of the gas
- Greater the value of 'a' larger the intermolecular force of attraction and the gas can be liquefied.
- Here a=0 means there is intermolecular attraction and the gas **cannot be liquefied.**

32. Suppose there is a tiny sticky area on the wall of a container of gas. Molecules hitting this area stick there permanently. Is the pressure greater or less than on the ordinary area of walls?

- Molecule hitting the tiny sticky area on the wall of the container of gas moves faster as they get closer to adhesive surface, but this effect is not permanent.
- The pressure on the sticky wass is greater than on the ordinary area of walls.

33. Explain the following observations

a) Aerated water bottles are kept under water during summer

- In aerated water bottles the CO_2 is passed through the aqueous solution under pressure.
- The solubility of gas is decreases with increase of temperature. In summer season the temperature is raise the solubility is decreases.
- Due to this will increase very high pressure above the surface of the liquid inside the bottle and bottle will not able to withstand the pressure and bottle may explode.
- To avoid this Aerated water bottles are kept under water during summer.
- As a result, the temperature decrease and solubility of CO_2 is increases in aqueous solution resulting the pressure inside the bottle decreases.

b) Liquid ammonia bottle is cooled before opening the seal

- Liquid ammonia bottle contains the gas under very high pressure. If the bottle is opened, the sudden decrease in pressure will increase the volume of gas.
- As result, the gas is come out the bottle with greater force will cause breakage of bottle and accident.
- The pressure of the gas over liquid ammonia is decreased by dipping H_2O.
- Once it attains the temperature of water, it can be opened so that no gas will come out of the mouth of the tube with force.

c) The tyre of an automobile is inflated to slightly lesser pressure in summer than in winter.

- In summer due to high temperature the air expands and hence to avoid tyre burst, the pressure is kept less.
- As a result, when the vehicle runs and the tyre air gets heated and expands, resulting increased pressure is still kept below the bursting limit.

d) The size of a weather balloon becomes larger and larger as it ascends up into larger altitude.

- According Boyle's law the volume of gas is inversely proportional to the pressure at a given temperature.
- As the weather balloon ascends, the atmospheric pressure is less, pressure of the gas tends to decrease and so volume as well as the size of the balloon increases.

34. Give suitable explanation for the following facts about gases.

a) Gases don't settle at the bottom of a container

- Gases are less denser than solids and liquids. They have negligible intermolecular force of attraction between free particles, so are free to move.
- Hence gases don't settle at bottom of the container.

b) Gases diffuse through all the space available to them

- Gases have the tendency to mix with one another spontaneously and form a homogeneous mixture.
- This is due to the fact that gas particles are mainly in random at very high velocities and there in so much of inter molecular empty space in the volume of any gas.
- This permit them to mix spontaneous and this phenomenon is knwn as diffusion.

35. Suggest why there is no hydrogen (H_2) in our atmosphere. Why does the moon have no atmosphere?

- Under ordinary conditions on earth, hydrogen exist as diatomic (H_2). Because of its light weight, which enable into escape from earth gravity more easily than heavier gases. So, no hydrogen is there in atmosphere.
- Moon has no atmosphere because the value of acceleration due to gravity (g) on the surface of moon is small. The molecules of atmospheric gases on the surface of the moon have thermal velocities greater than escape velocity.

36. Explain whether a gas approaches ideal behavior or deviates from ideal behaviour if

a) it is compressed to a smaller volume at constant temperature.

When the gas is compressed, there is a decrease in volume and molecules are close to each other and hence inter molecular attraction becomes more and hence it deviates from ideal behavior.

b) the temperature is raised at while keeping the volume constant

- When volume constant, the pressure is directly proportional to the temperature.
- The temperature is raised the pressure is always increases which leads the intermolecular force of attraction between the gas molecules.
- So it deviates from the ideal behavior

c) more gas is introduced into the same volume and at the same temperature.

- If more gas is introduced in the same volume and temperature is constant, more will be the pressure.
- The increase in pressure is always increases which leads the intermolecular force of attraction between the gas molecules.
- So it deviates from the ideal behavior

37. Which of the following gases would you expect to deviate from ideal behaviour under conditions of low temperature F_2, Cl_2, or Br_2? Explain.

- These molecules are held together by a weak van der Waals forces.
- The forces of attraction between the molecules with increase in the size of the molecule.
- Br_2 deviate from ideal behavior, since the Br_2 has the biggest size provides maximum attraction between bromine molecules.

38. Distinguish between diffusion and effusion.

S.No	Diffusion	Effusion
1.	The spreading of the molecules of a gas throughout the available space or second substance is called diffusion	Effusion is the escape of gas molecules through a very small hole.
2.	Diffusion refers to the ability of the gases to mix with each other	Effusion is a ability of a gas to travel through a small pin-hole.

39. Aerosol cans carry clear warning of heating of the can. Why?

- Aerosols are colloids in which air (gas) is dispensed in liquid. On heating the can, the pressure of the gas increases and it can burst out.
- Hence they carry clear warning that they should not be heated or kept in near fire.

40. When the driver of an automobile applies brake, the passengers are pushed toward the front of the car but a helium balloon is pushed toward back of the car. Upon forward acceleration the passengers are pushed toward the front of the car. Why?

- When breaks are applied, the automobile comes to rest but the passenger due to inertia of motion tend to continue to move in forward direction hence the passengers are pushed toward the front of the car.
- The movement of helium balloon in the opposite direction is due to the difference in the density of surrounding air. When the car stops suddenly, the air moves forward due to inertia of motion.

41. Would it be easier to drink water with a straw on the top of Mount Everest?

- Drinking through a straw is slightly more difficult on the top a mountain.
- This is because the atmospheric pressure and the temperature is low at top mountain.
- When you drink through a straw you are decreasing the pressure in your mouth, so the atmospheric pressure pushes the liquid up the straw.
- On mountain there is less pressure so there is less pressure to push the water into the straw.

42. Write the Van der Waals equation for a real gas. Explain the correction term for pressure and volume

Refer book page number 171 and 172.

43. Derive the values of van der Waals equation constants in terms of critical constants.

Refer Book page No. 174 & 175

44. Why do astronauts have to wear protective suits when they are on the surface of moon?

- Astronauts must wear space suits since the surface of moon, ther is no air to breath and no air pressure.
- Space is extremely cold and filled with dangerous radiation.
- Space suits are specially designed to protect astronauts from the cold, radiation and low pressure in space. It also provide air to breathe.

45. When ammonia combines with HCl, NH_4Cl is formed as white dense fumes. Why do more fumes appear near HCl?

$$NH_3 + HCl \rightarrow NH_4Cl$$

- The rate of diffusion is inversely proportional to the molecular weight of the gas.
- Lower the molecular weight faster is the diffusion.
- The molar mass of HCl is 36.5 $g.mol^{-1}$ while the molar mass of NH_3 is 17 $g.mol^{-1}$.
- Hence NH_3 diffuses faster than HCl. Hence white fumes appear near HCl.

HYDROGEN

1. What are isotopes? Write examples.
2. Explain ortho and para hydrogen.
3. Differentiate ortho and para hydrogen.
4. How will you convert ortho hydrogen into para hydrogen.
5. Write a note on industrial preparation of hydrogen.
6. What is water gas shift reaction?
7. Write a note on preparation of heavy water.
8. How is Tritium prepared
9. Write the exchange reaction of deuterium.
10. Compare the structure of H_2O and H_2O_2.
11. Differentiate H_2O, D_2O and T_2O.
12. Prove amphoteric nature of water.
13. Discuss the importance of heavy water in nuclear reactor.
14. Temporary hard water becomes soft on boiling. Explain.
15. Write a note on industrial preparation of H_2O_2.
16. Hard water produces les foam with detergent. Why?
17. Which causes the permanent hardness of water? How is it removed?
18.
19. Write the uses of heavy water.
20. Why H_2O_2 is not stored in glass bottle?
21. Write the oxidizing properties of H_2O_2.
22. What are hydrides? Explain the different types of hydrides with example.
23. Define hydrogen bonding.
24. Write the condition for hydrogen bonding.
25. What are inter and intra molecular hydrogen bonding? Give examples.