PREVIOUS HSE QUESTIONS AND ANSWERS OF THE CHAPTER "SOME BASIC CONCEPTS"

1. State and illustrate law of multiple proportions. (3)

   Ans: It states that *if two elements combine to form more than one compound, the different masses of one of the elements that combine with a fixed mass of the other element, are in small whole number ratio.*

   Illustration: Hydrogen combines with oxygen to form two compounds – water and hydrogen peroxide.

   Hydrogen + Oxygen → Water
   2g 16g 18g
   Hydrogen + Oxygen → Hydrogen Peroxide
   2g 32g 34g

   Here, the masses of oxygen (i.e. 16 g and 32 g) which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e. 16:32 or 1: 2.

2. (i) Calculate the mass of $CO_2(g)$ in gram produced by the reaction between 3 mol of $CH_4(g)$ and 2 mol of $O_2(g)$ according to the equation : $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$ (2)
   (ii) Identify the limiting reagent in this reaction. (1) [December 2021]

   Ans: (i) $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$
   According to the equation, 2 mol O_2 requires only 1 mol of CH_4 for the complete combustion.
   So among 3 moles of CH_4, only 1 mol undergoes complete combustion, since there is only 2 mol of O_2.
   So no. of moles of CO_2 formed = 1 mol, according to the equation
   = 44 g CO_2
   (ii) Here CH_4 is the limiting reagent, since it is completely consumed.

3. (i) Write Avogadro number. (1)
   (ii) How many moles of water molecules are present in 180 g of water?
   (Molecular mass of water = 18 g). (2)

   Ans: (i) 6.022×10^{23}
   *(ii) No. of moles = Given mass in gram / Molar mass = 180/18 = **10 mol***

4. (i) Define Molarity. (1)
   (ii) State law of multiple proportions. (2) [September 2021]

   Ans: (i) Molarity is the no. of moles of solute present per litre of the solution.
   (ii) It states that *if two elements combine to form more than one compound, the different masses of one of the elements that combine with a fixed mass of the other element, are in small whole number ratio.*

5. (a) Who proposed the law of conservation of mass? (1)
   (b) Illustrate the above law by using a chemical reaction. (1)

   Ans: (a) Antoin Lavoisier
   (b) Consider the reaction $C(s) + O_2(g) \longrightarrow CO_2(g)$
   12g 32g 44g
   Total mass of reactants = 12+32 = 44g
   Total mass of products = 44 g
   i.e. Total mass of reactants = Total mass of products. This is law of conservation of mass.

6. Determine the empirical formula of an oxide of iron which has 69.9% iron (Fe) and 30.1% oxygen (O) by mass. [Hint: Atomic mass of Fe = 55.85]. (3) [December 2020]

   Ans:

Element	*Percentage*	*Atomic mass*	*Percentage/Atomic mass*	*Simple ratio*	*Simplest whole no. ratio*
Fe	*69.9*	*55.85*	*69.9/55.85 = 1.25*	*1.25/1.25 = 1*	*1 x 2 = 2*
O	*30.1*	*16*	*30.1/16 = 1.88*	*1.88/1.25 = 1.5*	*1.5 x 2 = 3*

   Empirical formula = Fe_2O_3

7. (a) Classify the following matter as homogeneous mixture, heterogeneous mixture, element and compounds. gold, air, muddy water, water (1)
   (b) Define limiting reagent of a reaction. (1)

   Ans: (a) Homogeneous mixture: air
   Heterogeneous mixture: muddy water
   Element: gold
   Compound: water
   (b) It is the reagent that limits a reaction or completely used up in a reaction.

8. (a) Hydrogen and oxygen combines to form H_2O and H_2O_2. Which law of chemical combination is illustrated here? (1)
   (b) The balanced chemical equation for combustion of CH_4 is $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$.
   Calculate the amount of water formed by the combustion of 32g of CH_4. (2) [March 2020]

   Ans: (a) Law of multiple proportion
   (b) $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$.
   16g 64g 44g 36g
   16g CH_4 *produces 36g water.*
   So the amount of water formed by the combustion of 32g CH4 = 36 x 32/16 = 72 g.

9. Which of the following contains the maximum number of molecules?
   a) 1g N_2 b) 1g CO_2 c) 1g H_2 d) 1g NH_3 (1)

   Ans: c) 1g H_2 *(1)*

10. Calculate the mass of SO_3 (g) produced, if 500 g SO_2 (g) reacts with 200 g O_2 (g) according to the equation: $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(S)$. Identify the limiting reagent. (3) [July 2019]

    Ans: $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(S)$
    128 g 32g 160g
    128g SO_2 *requires 32g Oxygen for the complete reaction.*
    So 500g SO_2 *requires 32x500/128 = 125g* O_2
    Here there is 200g O_2*. So* SO_2 *is completely used up and hence it is the limiting reagent. (3)*

11. Round off 0.0525 to a number with two significant figures. (1)

    Ans: 0.052

12. A reaction mixture for the production of NH_3 gas contains 250 g of N_2 gas and 50 g of H_2 gas under suitable conditions. Identify the limiting reactant if any and calculate the mass of NH_3 gas produced. (3)
    [March 2019]

    Ans: Nitrogen reacts with Hydrogen to form ammonia according to the equation,
    $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
    28g 6g 34g
    i.e. 28g N_2 *requires 6g* H_2 *for the complete reaction.*
    So 250g N_2 *requires, 6x 250/28 = 53.57g* H_2*.*
    But here there is only 50g H_2*.*
    So we have to consider the reverse case.
    i.e. 6g H_2 *requires 28g* N_2*.*
    So 50g H_2 *requires 28 x 50/6 = 233.33g* N_2
    Here H_2 *is completely consumed. So it is the limiting reagent.*
    Amount of ammonia formed = 50+ 233.33 = ***283.33 g***

13. Which among the following measurements contains the highest number of significant figures?

a) 1.123×10^{-3} kg b) 1.2×10^{-3} kg c) 0.123×10^{3} kg d) 2×10^{5} kg (1)

Ans: a) 1.123×10^{-3} kg

14. State and illustrate the law of multiple proportions. (2)

Ans: It states that *if two elements combine to form more than one compound, the different masses of one of the elements that combine with a fixed mass of the other element, are in small whole number ratio.*

Illustration: Hydrogen combines with oxygen to form two compounds – water and hydrogen peroxide.

Hydrogen + Oxygen → Water
2g 16g 18g
Hydrogen + Oxygen → Hydrogen Peroxide
2g 32g 34g

Here, the masses of oxygen (i.e. 16 g and 32 g) which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e. 16:32 or 1: 2.

15. Calculate the amount of CO_2(g) produced by the reaction of 32g of CH_4 (g) and 32g of O_2 (g). (3)

[August 2018]

Ans: CH_4 (g) + $2O_2$ (g) $\longrightarrow$ CO_2 (g) + 2 H_2O (g)

16g 64g 44g 36g

64g O_2 requires 16g CH_4 for the complete reaction.

So, 32g O_2 requires 8g CH_4.

16g CH_4 combines with 64g O_2 to form 44g CO_2.

Therefore, 8g CH_4 combines with 32g Oxygen to form 22g CO_2.

16. The number of oxygen atoms present in 5 moles of glucose ($C_6H_{12}O_6$) is .... (1)

Ans: $30 \times 6.022 \times 10^{23}$ atoms

17. Find the molecular formula of the compound with molar mass 78 g mol^{-1} and empirical formula CH. (2)

Ans: Molar mass = 78 g/mol

Empirical Formula mass = 12+1 = 13

Molecular formula = Empirical formula x n

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = 78/13 = 6$$

So, Molecular formula = (CH) x 6 = C_6H_6

18. Calculate the mass of oxalic acid dihydrate ($H_2C_2O_4.2H_2O$) required to prepare 0.1M, 250 ml of its aqueous solution. (3) [March 2018]

Ans: Molarmass of Oxalic acid dihydrate = 126

Molarity = 0.1M

Volume of the solution = 250 mL

$$\text{Molarity} = \frac{\text{Mass of the solute x 1000}}{\text{Molar mass of the solute x volume of solution in mL}}$$

So Mass of the solute = (Molarity x Molar mass of the solute x volume of solution in mL)/1000

= (0.1 x 126 x 250/1000 = 3.15 g

19. a) NO and NO_2 are two oxides of nitrogen.

i) Which law of chemical combination is illustrated by these compounds? (1)

ii) State the law. (1)

b) Calculate the mass of a magnesium atom in grams. (1)

c) What is molality? (1) [July 2017]

Ans: a) i) Law of multiple proportions

ii) It states that *if two elements combine to form more than one compound, the different masses of one of the elements that combine with a fixed mass of the other element, are in small whole number ratio.*

b) Atomic mass of magnesium = 24g

i.e. Mass of 6.022 x 10^{23} atoms of Magnesium = 24g

So, Mass of one magnesium atom = 24/(6.022x10^{23}) = 3.98x10^{-23} g

c) Molarity is the no. of moles of solute present per litre of the solution.

20. a) Determine the number of moles present in 0.55 mg of electrons.

i) 1 mole ii) 2 moles iii) 1.5 moles iv) 0.5 mole (1)

b) Give the empirical formula of the following.

$C_6H_{12}O_6$, C_6H_6, CH_3COOH, $C_6H_6Cl_6$ (2)

c) Two elements, carbon and hydrogen combine to form C_2H_6, C_2H_4 and C_2H_2. Identify the law illustrated here. (1) [March 2017]

Ans: (a) (i) 1 mol (Score 1)

[Explanation: Mass of one electron = 9.1 x 10^{-28} g = 9.1 x 10^{-25} mg.

Mass of 1 mol of electron = 9.1 x 10^{-25} x 6.022 x 10^{23} mg = 0.55 mg.

So, 0.55 mg of electron ≡ 1 mol of electron.]

(b) Empirical formulae are: CH_2O, CH, CH_2O, CHCl. (Score 2)

(c) Law of multiple proportions (Score 1)

21. Empirical formula represents the simplest whole number ratio of various atoms present in a compound.

a) Give the relation between empirical formula and molecular formula. (1)

b) An organic compound has the following percentage composition C = 12.36%, H = 2.13%, Br = 85%. Its vapour density is 94. Find its molecular formula. (2)

c) What is mole fraction? (1) [September 2016]

Ans: a) Molecular formula = Empirical formula x n (1)

b)

Element	Percentage	Atomic mass	Percentage/Atomic mass	Simple ratio	Simplest whole no. ratio
C	12.36	12	12.36/12 = 1.03	1.03/1.03 = 1	1
H	2.13	1	2.13/1 = 2.13	2.13/1.03 = 2	2
Br	85	80	85/80 = 1.06	1.06/1.03 = 1	1

Empirical Formula = CH_2Br

Empirical Formula Mass (EFM) = 12+2+80 = 94

Molar mass (MM) = 2 x vapour density = 2 x 94 = 188

n = MM/EFM = 188/94 = 2

Molecular formula = Empirical formula x n

= (CH_2Br) x 2 = $C_2H_4Br_2$ (2)

c) Mole fraction is the ratio of the number of moles of a particular component to the total number of moles of solution. (1)

22. a) When nitrogen and hydrogen combines to form ammonia, the ratio between the volumes of gaseous reactants and products is 1: 3: 2. Name the law of chemical combination illustrated here. (1)

b) A compound is made up of two elements A and B, has A = 70% and B = 30%. The relative number of moles of A and B in the compound are 1.25 and 1.88 respectively. If the molar mass of the compound is 160, find the molecular formula of the compound. (3) [March 2016]

Ans: a) Gay – Lussac's law of Gaseous volumes (1)

b) The relative number of moles means %/atomic mass.

Elements	*%*	*Relative no. of moles*	*Simple ratio*	*Simplest whole no. ratio*
A	*70*	*1.25*	*1.25/1.25 = 1*	*2*
B	*30*	*1.88*	*1.88/1.25 = 1.5*	*3*

Empirical formula is A_2B_3

(Here at. Mass is not given. But it can be find out from % composition and the no. of moles as follows)

Atomic mass of A = %/no. of moles = 70/1.25 = 56

Atomic mass of B = %/no. of moles = 30/1.88 = 15.96

So, emp. Formula mass = 56 x 2 + 15.96 x 3 = 159.88 = 160

n = Mol.mass/Emp. Formula Mass = 160/160 = 1

Molecular formula = (emp. Formula) x n = (A_2B_3)x 1 = A_2B_3 (3)

23. 12 g of ^{12}C contains Avogadro's number of carbon atoms.

a) Give the Avogadro's number. (1)

b) The mass of 2 moles of ammonia gas is ....

(i) 2 g (ii) 1.2 x 10^{22}g (iii) 17 g (iv) 34g (1)

c) Calculate the volume of ammonia gas produced at STP when 140 g of nitrogen gas reacts with 30 g of hydrogen gas. (Atomic mass: N = 14u, H = 1 u) (2) [October 2015]

Ans: a) 6022 x 10^{23} (1)

b) (iv) 34g (1)

c) No. of moles of N_2 = Mass in gram/ molar mass = 140/28 = 5 mol

No. of moles of H_2 = Mass in gram/ molar mass = 30/2 = 15 mol

N_2 combines with H_2 to form NH_3 according to the equation: $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

From the equation, it is clear that 1 mol nitrogen reacts with 3 mol H_2 to form 20mol NH_3.

So, 5 mol N_2 reacts with 15 mol H_2 to form 10 mol NH_3.

Volume of 1 mol of ammonia at STP = 22.4 L

So volume of 10 mol of ammonia at STP = 224 L (2)

24. 'A given compound always contains exactly the same proportion of elements by weight.'

a) (i) Name the above law. (1)

(ii) Write the name of the Scientist who proposed this law. (1)

b) Calculate the number of molecules in each of the following:

i) 1 g N_2 ii) 1 g CO_2 (Given that N_A is 6.022 x 10^{23}, molecular mass of N_2 = 28 and CO_2 = 44) (2)

[March 2015]

Ans: a) (i) Law of definite (constant) proportions (1)

(ii) Joseph Proust (1)

b) No of molecules = $\frac{\text{Given mass in gram x NA}}{\text{Molar mass}}$

(i) (1 x 6.022 x 10^{23})/28 = 0.0357 x 10^{23} molecules (1)

(ii) (1 x 6.022x10^{23})/44 = 0.0227 x 10^{23} molecules (1)

25. a) How many moles of dioxygen are present in 64g of dioxygen? (Molar mass of dioxygen is 32). (1)

b) The following data were obtained when dinitrogen (N_2) and dioxygen (O_2) react together to form different compounds.

Mass of N_2	Mass of O_2
14 g	16 g
14 g	32 g
28 g	32 g
28 g	80 g

Name the law of chemical combination obeyed by the above experimental data. (1)

c) Define empirical formula. How is it related to the molecular formula of a compound? (2)

[March 2014]

Ans: a) No. of moles = Mass in gram / Molar mass = 64/32 = 2 moles (1)

b) Law of multiple proportion (1)

c) Empirical formula is the simplest formula which gives only the ratio of different elements present in the compound. It is related to molecular formula as Molecular formula = Empirical formula x n (2)

26. Hydrogen combines with oxygen to form two different compounds, namely water (H_2O) and hydrogen peroxide (H_2O_2).

a) Which law is obeyed by this combination? (1)

b) State the law. (2)

c) How many significant figures are present in the following?

i) 0.0025 ii) 285 (1) [August 2014]

Ans: a) Law of Multiple proportions (1)

b) It states that if two elements combine to form more than one compound, the different masses of one of the elements that combine with a fixed mass of the other element, are in small whole number ratio.

c) i) 2 ii) 3 (1)

27. a) Atoms have very small mass and so usually the mass of atoms are given relative to a standard called atomic mass unit. What is atomic mass unit (amu)? (1)

b) In a reaction $A + B_2 \rightarrow AB_2$, identify the limiting reagent in the reaction mixture containing 5mol A and 2.5 mol B. (1)

c) Calculate the mass of NaOH required to make 500 ml of 0.5M aqueous solution. (Molar mass of NaOH = 40) (2) [October 2013]

Ans: a) 1/12th the mass of a C^{12} atom is called atomic mass unit. (1)

b) B (1)

c) $$\text{Molarity} = \frac{\text{Mass of the solute x 1000}}{\text{Molar mass of the solute x volume of solution in mL}}$$

So, Mass of the solute = (Molarity x Molar mass of the solute x volume of solution in mL)/1000

= (0.5 x 40 x 500)/1000 = 10 g (2)

28. The mole concept helps in handling a large number of atoms and molecules in stoichiometric calculations.

a) Define 1 mol. (1)

b) What is the number of hydrogen atoms in 1 mole of methane (CH_4)? (1)

c) Calculate the amount of carbon dioxide formed by the complete combustion of 80g of methane as per the reaction:

$$CH_4\,(g) + 2O_2\,(g) \longrightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

(Atomic mass of C = 12.01u, H = 1.008u, O = 16u) (2) [March 2013]

Ans: a) 1 mol is the amount of substance that contains as many particles as there are atoms in exactly 12g C^{12} isotope.   (1)

b) 1 mol CH_4 contains 4 mol hydrogen = 4 x 6.022 x 10^{23} H atoms.   (1)

c) $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$

According to the equation, 16 g CH_4 gives 44 g CO_2

So 80 g CH_4 gives 44 x 80/16 = 220 g CO_2 (2)

29. a) Mole is a very large number to indicate the number of atoms, molecules etc. Write another name for one mole.   (1)

b) i) How the molecular formula is different from that of empirical formula?   (1)

ii) An organic compound on analysis gave the following composition. Carbon = 40%, Hydrogen = 6.66% and oxygen = 53.34%. Calculate its molecular formula if its molecular mass is 90. (2) [September 2012]

Ans: a) Avogadro's Number or Avogadro's constant (1)

b) (i) Empirical formula is the simplest formula which gives only the ratio of different elements present in the compound. But molecular formula is the actual formula that gives the exact number of different elements present in the compound.   (1)

(ii)

Element	*Percentage*	*Atomic mass*	*Percentage/Atomic mass*	*Simple ratio*	*Simplest whole no. ratio*
C	*40*	*12*	*40/12 = 3.33*	*3.33/3.33 = 1*	*1*
H	*6.66*	*1*	*6.66/1 = 6.66*	*6.66/3.33 = 2*	*2*
O	*53.34*	*16*	*53.34/16 = 3.33*	*3.33/3.33 = 1*	*1*

Empirical Formula = CH_2O

Empirical Formula Mass (EFM) = 12+2+16 = 30

Molar mass (MM) = 90

n = MM/EFM = 90/30 = 3

Molecular formula = Empirical formula x n

= (CH_2O) x 3 = $C_3H_6O_3$ (2)

30. The combination of elements to form compounds is governed by the laws of chemical combination.

a. Hydrogen combines with oxygen to form compounds, namely water and hydrogen peroxide. State and illustrate the related law of chemical combination.   (2)
b. What is mean by limiting reagent in a chemical reaction?   (1)
c. 28 g of nitrogen is mixed with 12 g of hydrogen to form ammonia as per the reaction, $N_2 + 3H_2 \longrightarrow 2NH_3$. Which is the limiting reagent in this reaction?   (1)   [March 2012]

Ans: a) Law of multiple proportions (Answer of Qn. No. 6)   (2)

b) A reagent that is completely consumed in a reaction.   (1)

c) No. of mol of N_2 = 28/28 = 1 mol

No. of mol of H_2 = 12/2 = 6 mol.

Here N_2 is completely used up. So it is the limiting reagent.   (1)

31. The laws of chemical combination govern the formation of compounds from elements.

a) State the law of conservation of mass. Who put forward this law?   (1½)

b) The following data are obtained when dinitrogen and dioxygen react together to form different compounds.

Sl. No.	Mass of dinitrogen (in g)	Mass of dioxygen (in g)
1	14	16
2	14	32
3	28	48
4	28	80

Which law of chemical combination is illustrated by the above experimental data? Explain? (2½)

[October 2011]

Ans: a) It states that matter can neither be created nor destroyed. Or, in a chemical reaction, the total mass of reactants = the total mass of products. This law was put forward by Antoine Lavoisier.

b) Law of Multiple proportion (See the answer of Qn. No. 6)

32. The laws of chemical combination are the basis of the atomic theory.

a) Name the law of chemical combination illustrated by the pair of compounds, CO and CO_2. (1)

b) State and explain the law of conservation of mass. (1½)

c) Calculate the molarity of a solution containing 8 g of NaOH in 500 mL of water. (1½) [March 2011]

Ans: a) Law of multiple proportions

b) It states that matter can neither be created nor destroyed. Or, in a chemical reaction, the total mass of reactants = the total mass of products.

Consider the reaction $2H_2 + O_2 \rightarrow 2H_2O$

Here 4 g of H_2 *combines with 32 g of* O_2 *to form 36 g of water.*

Total mass of reactants = 4 + 32 = 36g

Total mass of products = 36 g

c) $\text{Molarity} = \dfrac{\text{Mass of the solute x 1000}}{\text{Molar mass of the solute x volume of solution in mL}}$

= (8 x 1000)/40x500 = 0.4 M

33. One mole is the amount of substance that contains as many particles as 12 g of C^{12} isotope of carbon.

a) What do you mean by molar mass of a compound? (1)

b) Calculate the number of moles in 1 L of water (Density of water 1 g/mL). Also calculate the number of water molecules in 1 L water. (3) [September 2010]

Ans: a) It is the mass of 1 mol of any substance. (1)

b) Since density of water = 1 g/mL, 1 L H_2O *= 1 kg* H_2O *= 1000g* H_2O *(mass = density x volume)*

No. of moles = Mass in gram / Molar mass = 1000/18 = 55.55 mol

No of molecules = no. of moles x $N_A = 55.55 \times 6.022 \times 10^{23}$ *molecules. (3)*

34. If the mass percent of various elements of a compound is known, its empirical formula can be calculated.

a) What is mass percent? (1)

b) A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% chlorine. Its molecular mass is 98.96. What are the empirical and molecular formulae? (3) [March 2010]

Ans: a) $\text{Mass \%} = \dfrac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$

OR, $\text{Mass percent} = \dfrac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}} \times 100$ *(1)*

b)

Element	*Percentage*	*Atomic mass*	*Percentage/Atomic mass*	*Simple ratio*	*Simplest whole no. ratio*
C	*24.27*	*12*	*24.27/12 = 2.02*	*2.02/2.02 = 1*	*1*
H	*4.07*	*1*	*4.07/1 = 4.07*	*4.07/2.02 = 2*	*2*
Cl	*71.65*	*35.5*	*71.65/35.5 = 2.02*	*2.02/2.02 = 1*	*1*

Empirical Formula = CH_2Cl

Empirical Formula Mass (EFM) = 12+2+35.5 = 49.5

Molar mass (MM) = 98.96

n = MM/EFM = 98.96/49.5 = 2

Molecular formula = Empirical formula x n

= $(CH_2Cl) \times 2 = C_2H_4Cl_2$ *(3)*

35. Calculate the number of moles of oxygen required to produce 240 g of MgO by burning Mg metal. (Atomic mass Mg = 24, O = 16) (4) [March 2009]

Ans: Here mass of MgO = 240 g

Molecular mass of MgO = 24 + 16 = 40 g/mol

No. of moles of MgO = Mass of MgO/Molar mass of MgO = 240/40 = 6 mol

The chemical equation for the burning of Mg metal can be represented as:

$2\ Mg + O_2 \longrightarrow 2\ MgO$

i.e. 2 mol MgO require 1 mol O_2

So, 6 mol MgO require ½ x 6 = 3 mol O_2. *(4)*

36. One gram atom of an element contains 6.02×10^{23} atoms.

a) Find the number of oxygen atoms in 4 g of O_2. (1)

b) Which is heavier, one oxygen atom or 10 hydrogen atoms? (1) [February 2008]

Ans: a) No. of moles of $O_2 = 4/32 = 0.125$ *mol*

1 mol O_2 *contains* $2 \times 6.022 \times 10^{23}$ *Oxygen atoms.*

So, 0.125 mol O_2 *contains* $0.125 \times 2 \times 6.022 \times 10^{23}$ *Oxygen atoms* *(1)*

b) Mass of one oxygen atom = 16u

Mass of 10 hydrogen atoms = 10u

So one oxygen atom is heavier. *(1)*

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