The Mole Concept and Stoichiometry Discussion Questions Solutions

Calculation of Relative Molecular Mass

1 Bromine consists of two isotopes, ^{79}Br and ^{81}Br, in the relative abundance ratio of 1:1. Bromine exists as diatomic molecule, Br_2, at room temperature and pressure.
Calculate the possible relative molecular masses, M_r, of Br_2 molecules formed by these two isotopes and their relative abundance ratio.

[1:2:1]

Molecular Formula	$^{79}Br^{79}Br$	$^{79}Br^{81}Br$ or $^{81}Br^{79}Br$	$^{81}Br^{81}Br$
M_r	158	160	162
Relative Ratio	$\left(\frac{1}{2}\right)\left(\frac{1}{2}\right)=\frac{1}{4}$	$2\left(\frac{1}{2}\right)\left(\frac{1}{2}\right)=\frac{1}{2}$	$\left(\frac{1}{2}\right)\left(\frac{1}{2}\right)=\frac{1}{4}$
Abundance Ratio	1	2	1

Calculation involving Parts Per Million

2 Gardeners sometimes fumigate their greenhouses to rid them of pests and moulds by burning a sulfur 'candle'. A gaseous concentration of sulfur dioxide of 50 ppm (parts per million) by volume is effective.

Calculate how many grams of sulfur a gardener needs to burn in order to produce a concentration of 50 ppm of SO_2 in a greenhouse that measures 2 m × 3 m × 4 m. Assume room conditions.

[1.61 g]

Equation for reaction: $S(s) + O_2(g) \rightarrow SO_2(g)$

Volume of greenhouse = (2)(3)(4) = 24 m^3 = 24000 dm^3 [Recall: 1 m^3 = 10^3 dm^3]

Given that concentration of SO_2 = 50 ppm

$\Rightarrow$ In 10^6 dm^3 of air, there is 50 dm^3 of SO_2,

$\Rightarrow$ In 24000 dm^3 of air, there is $\frac{50}{10^6}$ x 24000 = 1.20 dm^3 of SO_2

$$\boxed{\frac{x}{24000} \times 10^6 = 50 \qquad x = 1.20 \text{ dm}^3}$$

Amt of SO_2 = 1.20 ÷ 24.0 = 0.0500 mol [Recall room conditions = r.t.p = 24.0 dm^3 mol^{-1}]

Mass of S = 0.0500 x 32.1 = 1.605 = 1.61 g (to 3 sf)

Calculation using Mole Concept

3 Phosgene, $COCl_2$, was once used as a war gas. It is poisonous because when inhaled, it reacts with water in the lungs to produce carbon dioxide gas and hydrochloric acid which causes severe lung damage, leading to death ultimately.

(a) Calculate the percentage by mass of chlorine in $COCl_2$. [71.7%]

$$\text{\% by mass of } Cl \text{ in } COCl_2 = \frac{2(35.5)}{12.0+16.0+2(35.5)} \times 100\% = \underline{\mathbf{71.7\,\%}}$$

(b) Write a balanced equation for the reaction between $COCl_2$ and H_2O.

$COCl_2 + H_2O \rightarrow CO_2 + 2HCl$

(c) Calculate the amount of HC*l* that will be produced by the complete reaction of 0.430 mol of $COCl_2$. [0.860]

Since $COCl_2$: HC*l* = 1 : 2,

amount of HC*l* produced $= \mathbf{2} \times 0.430 = \underline{\mathbf{0.860\ mol}}$

(d) Identify the limiting reagent when 0.200 mol of $COCl_2$ is mixed with 6.20 g of H_2O. Hence, calculate the amount of HC*l* produced at the end of the reaction. [0.400]

$$\text{amount of } H_2O \text{ used} = \frac{6.20}{2(1.0)+16.0} = \underline{\mathbf{0.344\ mol}}$$

amount of $COCl_2$ used $= 0.200$ mol

0.200 mol of $COCl_2$ will require 0.200 mol of H_2O,

Since 0.344 mol of H_2O is present (excess), $COCl_2$ is the limiting reagent.

Since $COCl_2$: HC*l* is 1 : 2,

amount of HC*l* produced $= \mathbf{2} \times 0.200 = \underline{\mathbf{0.400\ mol}}$

Calculation involving % purity & % yield

4 The mineral phosphorite, $Ca_3(PO_4)_2$, exists as phosphate rock in its impure form. Elemental phosphorus can be prepared from phosphate rock by reduction using carbon in the presence of sand, SiO_2. The reduction of phosphorite also produces solid **A** and carbon monoxide.

(a) **A** has the following composition by mass:

Ca: 34.2% Si: 24.5% O: 41.3%

Calculate the empirical formula for solid **A**. [$CaSiO_3$]

elements	Ca	Si	O
% by mass	34.2	24.5	41.3
amount / mol	$\frac{34.2}{40.1} = 0.853$	$\frac{24.5}{28.1} = 0.872$	$\frac{41.3}{16.0} = 2.58$
Simplest whole number mole ratio	$\frac{0.853}{0.853} = 1$	$\frac{0.872}{0.853} \approx 1$	$\frac{2.58}{0.853} \approx 3$

∴ The empirical formula of solid **A** is $CaSiO_3$.

(b) Write a balanced equation for the reaction.

$Ca_3(PO_4)_2 + 5C + 3SiO_2 \rightarrow 2P + 3CaSiO_3 + 5CO$

(c) A 30.0 g sample of phosphate rock was subjected to the above reaction and produced 5.3 g of phosphorus. Calculate the percentage purity of phosphorite in the rock sample. [88.4%]

$$\text{amount of phosphorus} = \frac{5.3}{31.0} = 0.171 \text{ mol}$$

Since $Ca_3(PO_4)_2$: P is 1 : 2

amount of $Ca_3(PO_4)_2$ = ½ × 0.171 = 0.0855 mol

mass of $Ca_3(PO_4)_2$ = 0.0855 × [3(40.1) + 2(31.0) + 8(16.0)] = 26.53 g

% purity of $Ca_3(PO_4)_2$ in the rock sample $= \frac{26.53}{30.0} \times 100\% = 88.4\%$

5 In the Solvay process, ammonia is recovered by the reaction:

$$2NH_4Cl\ (aq) + CaO\ (s) \rightarrow CaCl_2\ (aq) + H_2O\ (l) + 2NH_3\ (g)$$

(a) What is the maximum volume of ammonia that can be recovered, at s.t.p., from 20.0 g of NH_4Cl and 4.50 g of CaO?
State the limiting reagent, if any, and assume that any impurity present is unreactive.

[3.64 dm^3]

Amount of NH_4Cl = $\frac{20.0}{53.5}$ = 0.3738 mol

Amount of CaO = $\frac{4.50}{56.1}$ = 0.08021 mol

0.08021 mol of CaO will require 2(0.08021) = 0.1604 mol of NH_4Cl

Since 0.3738 mol of NH_4Cl is present (excess), hence CaO is the limiting reagent.

Since CaO:NH_3 is 1 : 2,
Amount of NH_3 = 2 × 0.08021 = 0.1604 mol
Vol of NH_3 = 0.1604 × 22.7 = 3.64 dm^3

(b) What is the percentage yield if only 3.03 dm^3 of NH_3 is obtained experimentally at s.t.p.?

[83.2%]

% yield = $\frac{3.03}{3.64}$ × 100 = 83.2 %

Calculation involving % composition by mass

6 In small quantities, nicotine in tobacco is addictive. In large quantities, it is a deadly poison. Determine the molecular formula of nicotine, $C_xH_yN_z$, if 4.38 mg of nicotine burns to form 11.9 mg of carbon dioxide and 3.41 mg of water. [M_r of nicotine = 162.0] [$C_{10}H_{14}N_2$]

Mass of C in nicotine = 11.9 × $\frac{12.0}{44.0}$ = 3.245 mg

Mass of H in nicotine = 3.41 × $\frac{2.0}{18.0}$ = 0.3789 mg

Mass of N in nicotine = 4.38 – 3.245 – 0.3789 = 0.7561 mg

Element	C	H	N
Mass / mg	3.245	0.3789	0.7561
Amount / mol	0.0002704	0.0003789	5.401 x 10^{-5}
Simplest whole number mole ratio	5	7	1

∴ Empirical formula of nicotine is C_5H_7N.

Let the molecular formula be $(C_5H_7N)n$
M_r of $(C_5H_7N)n$ = 162.0
{(5(12.0) + 7(1.0) + 14.0)}n = 162.0
n = 2
Molecular formula of nicotine is $C_{10}H_{14}N_2$.

Calculation involving Volumes of gas

7 Buckminsterfullerene, C_{60}, is a molecule with 60 carbon atoms arranged in pentagons and hexagons that are similar to those in a football. C_{60} reacts with hydrogen to form hydrofullerenes with the molecular formula $C_{60}H_n$.
When 60 g of C_{60} is reacted with hydrogen gas at 273 K and 1 bar, the volume of hydrogen gas is decreased by 34 dm^3. Find the value of n in the formula of hydrofullerene, $C_{60}H_n$. [36]

Amount of C_{60} = $\frac{60}{60 \times 12.0}$ = 0.08333 mol

Amount of H_2 = $\frac{34}{22.7}$ = 1.498 mol

Mole ratio of C_{60} : H_2 = 0.08333 : 1.498
1 : 18

Mole ratio of C_{60} to H = 1 : 36

Thus, molecular formula of the hydrofullerene formed is $C_{60}H_{36}$. n= 36

Calculation using Volumes of gas in Combustion reactions

8 20 cm^3 of a gaseous hydrocarbon **A**, C_xH_y, was exploded with an excess of oxygen. Upon cooling to room temperature, there was a contraction in volume of 50 cm^3. When the products were treated with excess potassium hydroxide, there was a further contraction of 60 cm^3.

Deduce the molecular formula of **A**. All volumes were measured at r.t.p. [C_3H_6]

Before reaction	After reaction & cooled to r.t.p		After reaction with KOH(aq)	
$V_{C_xH_y} + V_{O_2}$ added **20 cm^3 + excess O_2**	$\longrightarrow$ **decrease in vol. of 50 cm^3**	$V_{CO_2} + V_{O_2}$ **unreacted**	$\longrightarrow$ ***further* decrease of 60 cm^3**	V_{O_2} **unreacted**

Vol of CO_2 = 60 cm^3 (2nd Contraction)
Initial vol – Final vol = 50 cm^3 (1st Contraction)
(Vol C_xH_y + Total Vol O_2) – (Vol CO_2 + Vol unreacted O_2) = 50 cm^3
(20 + Total Vol O_2) – (60 + Vol unreacted O_2) = 50 cm^3
Total Vol O_2 –Vol unreacted O_2 = 90 cm^3
Vol O_2 reacted = 90 cm^3

	$C_xH_y(g)$	+ $(x + \frac{y}{4})\,O_2(g)$	→ $x\,CO_2(g)$	+ $\frac{y}{2}\,H_2O(l)$
Reacting vol / cm^3:	20	90	60	–
Mole ratio	$\frac{20}{20} = 1$	$\frac{90}{20} = 4.5$	$\frac{60}{20} = 3$	

x = 3; y = 6

∴The molecular formula of **A** is C_3H_6

9 A 20 cm^3 mixture containing methane (CH_4) and ethane (C_2H_6) was burnt completely in excess oxygen. On passing the residual gas through aqueous sodium hydroxide, there was a reduction in volume by 25 cm^3. All volumes were measured at room temperature and pressure.

Calculate the percentage by mass of methane in the original gaseous mixture. [61.5%]

$CH_4\ (g) + 2O_2\ (g) \rightarrow CO_2\ (g) + 2H_2O\ (l)$

$C_2H_6\ (g) + 7/2\ O_2\ (g) \rightarrow 2CO_2\ (g) + 3H_2O\ (l)$

Let volume of methane be x cm^3, then vol of ethane = $(20 - x)$ cm^3

Total volume of CO_2 produced = $x + 2(20 - x) = 40 - x = 25$

$\Rightarrow x = 15\ cm^3$

Mass of CH_4 in the mixture = $\dfrac{15}{24000} \times 16.0 = 0.0100$ g

Mass of C_2H_6 in the mixture = $\dfrac{5}{24000} \times 30.0 = 6.25 \times 10^{-3}$ g

Percentage by mass of methane in the original gaseous mixture

$= \dfrac{0.01}{0.01 + 6.25 \times 10^{-3}} \times 100\% = 61.5\ \%$

Calculation using Mole Concept in Solutions

10 **(a)** What is the mass of NaOH(s) required to prepare 250 cm^3 of 0.120 mol dm^{-3} NaOH(aq)? [1.20 g]

Amount of NaOH required = $\dfrac{0.120}{1000}$ x 250 = 0.0300 mol □ □ □

Mass of NaOH needed = 0.0300 x (23.0 + 16.0 + 1.0) = 1.20 g □ □□□□

(b) What volume of water must be added to 900 cm^3 of 0.120 mol dm^{-3} NaC*l* solution to dilute it to 0.100 mol dm^{-3}? [180 cm^3]

Amount of solute in original solution = Amount of solute in diluted solution

c_1V_1 (before dilution) = c_2V_2 (after dilution)

$(0.120)(900) = (0.100)(V_2)$

$V_2 = 1080\ cm^3$

Final vol of NaC*l*(aq) after dilution = 1080 cm^3

Hence, volume of water to be added = 1080 – 900 = 180 cm^3

□ □