Chapter 1: Some Basic Concepts of Chemistry

Exercises — Exercises

Q1 Calculate the molar mass of the following:
(i) H_2O (ii) CO_2 (iii) CH_4

SOLUTION

Atomic masses used: H = 1.008 u, C = 12.011 u, O = 15.999 u

(i) Molar mass of H_2O
= 2 × 1.008 + 1 × 15.999
= 2.016 + 15.999
= 18.015 g mol^{-1}

(ii) Molar mass of CO_2
= 1 × 12.011 + 2 × 15.999
= 12.011 + 31.998
= 44.009 g mol^{-1}

(iii) Molar mass of CH_4
= 1 × 12.011 + 4 × 1.008
= 12.011 + 4.032
= 16.043 g mol^{-1}

Answer: (i) H_2O: 18.015 g mol^{-1} (ii) CO_2: 44.009 g mol^{-1} (iii) CH_4: 16.043 g mol^{-1}

Q2 Calculate the mass per cent of different elements present in sodium sulphate (Na_2SO_4).

SOLUTION

Atomic masses: Na = 22.990 u, S = 32.06 u, O = 15.999 u
Molar mass of Na_2SO_4 = 2(22.990) + 32.06 + 4(15.999)
= 45.980 + 32.06 + 63.996
= 142.036 g mol^{-1}

Mass % of Na = (2 × 22.990 / 142.036) × 100 = 32.37%
Mass % of S = (32.06 / 142.036) × 100 = 22.57%
Mass % of O = (4 × 15.999 / 142.036) × 100 = 45.06%

Verification: 32.37 + 22.57 + 45.06 = 100.00%

Answer: Na: 32.37%, S: 22.57%, O: 45.06%

Q3 Determine the empirical formula of an oxide of iron, which has 69.9% iron and 30.1% dioxygen by mass.

SOLUTION

Given: 69.9% Fe, 30.1% O by mass.
Step 1 — convert to moles (per 100 g sample):
Moles of Fe = 69.9 / 55.845 = 1.252 mol
Moles of O = 30.1 / 15.999 = 1.881 mol

Step 2 — divide by smallest (1.252):
Fe : O = 1.252/1.252 : 1.881/1.252 = 1.000 : 1.503

Step 3 — multiply by 2 to get whole numbers:
Fe : O = 2 : 3

Empirical formula = Fe_2O_3

Answer: Fe_2O_3

Q4 Calculate the amount of carbon dioxide that could be produced when
(i) 1 mole of carbon is burnt in air.
(ii) 1 mole of carbon is burnt in 16 g of dioxygen.
(iii) 2 moles of carbon are burnt in 16 g of dioxygen.

SOLUTION

Balanced equation: $C(s) + O_2(g) \rightarrow CO_2(g)$
1 mol C reacts with 1 mol O_2 to give 1 mol CO_2 (44.009 g).

(i) 1 mol C burnt in air (O_2 in excess):
O_2 is unlimited, so 1 mol C is the limiting reagent.
CO_2 produced = 1 mol = 44.009 g = 44 g

(ii) 1 mol C burnt in 16 g O_2:
Moles of O_2 = 16 / 32.00 = 0.5 mol
1 mol C needs 1 mol O_2, but only 0.5 mol O_2 is available.
O_2 is the limiting reagent.
CO_2 produced = 0.5 mol = 0.5 × 44.009 = 22.0 g

(iii) 2 mol C burnt in 16 g O_2:
Moles of O_2 = 0.5 mol (same as above).

2 mol C needs 2 mol O_2, but only 0.5 mol O_2 is available.
O_2 is still the limiting reagent.
CO_2 produced = 0.5 mol = 22.0 g

Answer: (i) 44 g CO_2 (ii) 22 g CO_2 (iii) 22 g CO_2

Q5 Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0245 g mol^{-1}.

SOLUTION

Given: Volume = 500 mL = 0.500 L, Molarity = 0.375 M
Molar mass of CH_3COONa = 82.0245 g mol^{-1}

Moles of CH_3COONa required = Molarity × Volume (L)
= 0.375 mol L^{-1} × 0.500 L = 0.1875 mol

Mass required = moles × molar mass
= 0.1875 mol × 82.0245 g mol^{-1}
= 15.38 g

Answer: 15.38 g of sodium acetate (CH_3COONa)

Q6 Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g mL^{-1} and the mass per cent of nitric acid in it being 69%.

SOLUTION

Given: density = 1.41 g mL^{-1}, mass % HNO_3 = 69%
Molar mass of HNO_3 = 1.008 + 14.007 + 3(15.999) = 63.012 g mol^{-1}

Mass of 1 L (1000 mL) of solution = 1000 × 1.41 = 1410 g
Mass of HNO_3 in 1 L = 69/100 × 1410 = 972.9 g

Moles of HNO_3 in 1 L = 972.9 / 63.012 = 15.44 mol

Molarity = 15.44 mol L^{-1}

Answer: 15.44 mol L^{-1} (approximately 15.4 M)

Q7 How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)?

SOLUTION

Molar mass of $CuSO_4$:
= 63.546 + 32.06 + 4(15.999)
= 63.546 + 32.06 + 63.996 = 159.602 g mol^{-1}

Moles of $CuSO_4$ in 100 g = 100 / 159.602 = 0.6266 mol
Each mole $CuSO_4$ contains 1 mol Cu.
Mass of Cu = 0.6266 × 63.546 = 39.81 g

Alternatively: mass % of Cu = (63.546/159.602) × 100 = 39.81%
Mass of Cu from 100 g = 39.81 g

Answer: 39.81 g of copper

Q8 Determine the molecular formula of an oxide of iron, in which the mass per cent of iron and oxygen are 69.9 and 30.1, respectively.

SOLUTION

Step 1 — find empirical formula (same composition as $Q_1.3$):
Moles of Fe = 69.9/55.845 = 1.252 mol
Moles of O = 30.1/15.999 = 1.881 mol
Ratio Fe : O = 1.252 : 1.881 => 1 : 1.503 => 2 : 3
Empirical formula = Fe_2O_3

Step 2 — empirical formula mass of Fe_2O_3:
= 2(55.845) + 3(15.999) = 111.690 + 47.997 = 159.687 g mol^{-1}

Step 3 — since molar mass is not separately given, the molecular formula equals the empirical formula (n = 1).
Molecular formula = Fe_2O_3

Answer: Fe_2O_3

Q9 Calculate the atomic mass (average) of chlorine using the following data:

	% Natural Abundance	Molar Mass
35Cl	75.77	34.9689
37Cl	24.23	36.9659

SOLUTION

Average atomic mass = sum of (fractional abundance × atomic mass) for each isotope

Average atomic mass of Cl
= (75.77/100) × 34.9689 + (24.23/100) × 36.9659
= 0.7577 × 34.9689 + 0.2423 × 36.9659
= 26.496 + 8.957
= 35.453 u

Answer: 35.453 u

Q10 In three moles of ethane (C_2H_6), calculate the following:
(i) Number of moles of carbon atoms.
(ii) Number of moles of hydrogen atoms.
(iii) Number of molecules of ethane.

SOLUTION

Molecular formula of ethane = C_2H_6
Each molecule contains 2 C atoms and 6 H atoms.

(i) Moles of C atoms = 3 mol C_2H_6 × 2 mol C / 1 mol C_2H_6 = 6 mol C atoms

(ii) Moles of H atoms = 3 mol C_2H_6 × 6 mol H / 1 mol C_2H_6 = 18 mol H atoms

(iii) Number of C_2H_6 molecules = 3 mol × 6.022×10^{23} mol^{-1}
= $18.066 \times 10^{23} = 1.8066 \times 10^{24}$ molecules

Answer: (i) 6 mol carbon atoms (ii) 18 mol hydrogen atoms (iii) 1.807×10^{24} molecules

Q11 What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2 L?

SOLUTION

Molar mass of $C_{12}H_{22}O_{11}$:
= 12(12.011) + 22(1.008) + 11(15.999)
= 144.132 + 22.176 + 175.989
= 342.297 g mol^{-1}

Moles of sugar = 20 g / 342.297 g mol^{-1} = 0.05843 mol

Molarity = moles / volume = 0.05843 mol / 2 L = 0.02921 mol L^{-1}

Answer: 0.02921 mol L^{-1} (approximately 0.0293 M)

Q12 If the density of methanol is 0.793 kg L^{-1}, what is its volume needed for making 2.5 L of its 0.25 M solution?

SOLUTION

Given: density of methanol = 0.793 kg L^{-1} = 793 g L^{-1}
Molar mass of CH_3OH = 12.011 + 4(1.008) + 15.999 = 32.042 g mol^{-1}

Moles of methanol required = Molarity × Volume
= 0.25 mol L^{-1} × 2.5 L = 0.625 mol

Mass of methanol = 0.625 mol × 32.042 g mol^{-1} = 20.026 g

Volume of methanol = mass / density = 20.026 g / 793 g L^{-1}
= 0.02525 L = 25.25 mL

Answer: 25.25 mL of methanol

Q13 Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal is as shown below: 1 Pa = 1 N m^{-2}. If mass of air at sea level is 1034 g cm^{-2}, calculate the pressure in pascal.

SOLUTION

Given: mass per unit area = 1034 g cm^{-2}

Convert to SI units:
1034 g cm^{-2} = 1034 × 10^{-3} kg / $(10^{-2}$ m$)^2$
= 1034 × 10^{-3} kg / 10^{-4} m^2
= 1034 × 10 kg m^{-2} = 10340 kg m^{-2}

Pressure P = (mass per area) × g
= 10340 kg m^{-2} × 9.8 m s^{-2}
= 101332 Pa
≈ 1.013 × 10^5 Pa

Answer: 1.013 × 10^5 Pa

Q14 What is the SI unit of mass? How is it defined?

SOLUTION

The SI unit of mass is the kilogram (symbol: kg).

Definition: The kilogram is defined by taking the fixed numerical value of the Planck constant h to be 6.62607015 × 10^{-34} when expressed in the unit J s (= kg m^2 s^{-1}), where the metre and the second are defined in terms of the speed of light c and the caesium frequency V_Cs.

Prior to 2019, it was defined as the mass of a platinum-iridium (Pt-Ir) cylinder kept at the International Bureau of Weights and Measures in Sevres, France. The redefinition in 2019 linked the kilogram to a fundamental physical constant (Planck constant), making it independent of a physical artefact.

Answer: The SI unit of mass is the kilogram (kg), defined by fixing the numerical value of Planck's constant h = 6.62607015 × 10^{-34} J s.

Q15 Match the following prefixes with their multiples:

Prefixes	Multiples
(i) micro	10^6
(ii) deca	10^9
(iii) mega	10^{-6}
(iv) giga	10^{-15}
(v) femto	10

SOLUTION

Standard SI prefix values:
micro (μ) = 10^{-6}
deca (da) = 10^1 = 10
mega (M) = 10^6
giga (G) = 10^9

femto (f) = 10^{-15}

Matching (prefix -> correct multiple from the given list):
(i) micro -> 10^{-6}
(ii) deca -> 10
(iii) mega -> 10^{6}
(iv) giga -> 10^{9}
(v) femto -> 10^{-15}

Answer: (i) micro–10^{-6} (ii) deca–10 (iii) mega–10^{6} (iv) giga–10^{9} (v) femto–10^{-15}

Q16 What do you mean by significant figures?

SOLUTION

Significant figures are the meaningful digits in a measured or calculated quantity that are known with certainty, plus one estimated (uncertain) digit.

Rules for counting significant figures:
1. All non-zero digits are significant (e.g., 285 has 3 sig figs).
2. Zeros between non-zero digits are significant (e.g., 2005 has 4 sig figs).
3. Zeros preceding the first non-zero digit are not significant (e.g., 0.0052 has 2 sig figs).
4. Zeros at the end of a number are significant only if a decimal point is present (e.g., 0.200 has 3; 100 has 1; 100. has 3).
5. Exact numbers (counted quantities) have infinite significant figures.

Significant figures indicate the precision of a measurement: more significant figures imply a more precise measurement.

Answer: Significant figures are all meaningful digits (certain + one uncertain) in a measurement. They indicate the precision of the value.

Q17 A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).
(i) Express this in per cent by mass.
(ii) Determine the molality of chloroform in the water sample.

SOLUTION

Given: contamination = 15 ppm (by mass)
15 ppm means 15 g $CHCl_3$ per 10^6 g of solution.

(i) Mass per cent:
15 ppm = $15 / 10^6 \times 100 = 1.5 \times 10^{-3}$ % by mass

(ii) Molality:
Molar mass of $CHCl_3$ = 12.011 + 1.008 + 3(35.453)
= 12.011 + 1.008 + 106.359 = 119.378 g mol^{-1}

In 10^6 g solution: 15 g $CHCl_3$, rest water = $(10^6 - 15)$ g ≈ 10^6 g = 1000 kg
Moles of $CHCl_3$ = 15 / 119.378 = 0.12565 mol

Molality = 0.12565 mol / 1000 kg = 1.26×10^{-4} mol kg^{-1}

Answer: (i) 1.5×10^{-3} % by mass (ii) 1.26×10^{-4} mol kg^{-1}

Q18 Express the following in the scientific notation:
(i) 0.0048
(ii) 234,000
(iii) 8008
(iv) 500.0
(v) 6.0012

SOLUTION

Scientific notation: N × 10^n, where 1 ≤ N < 10

(i) 0.0048 = 4.8×10^{-3}
(ii) 234,000 = 2.34×10^5
(iii) 8008 = 8.008×10^3
(iv) 500.0 = 5.000×10^2
(v) 6.0012 = 6.0012×10^0 (already in proper form)

Answer: (i) 4.8×10^{-3} (ii) 2.34×10^5 (iii) 8.008×10^3 (iv) 5.000×10^2 (v) 6.0012×10^0

Q19 How many significant figures are present in the following?
(i) 0.0025
(ii) 208
(iii) 5005
(iv) 126,000
(v) 500.0
(vi) 2.0034

SOLUTION

(i) 0.0025: leading zeros are not significant; only 2 and 5 are. -> 2 significant figures
(ii) 208: all digits including the zero between non-zero digits are significant. -> 3 significant figures
(iii) 5005: zeros between non-zero digits are significant. -> 4 significant figures
(iv) 126,000: trailing zeros without a decimal point are ambiguous; conventionally -> 3 significant figures
(v) 500.0: the decimal point makes all trailing zeros significant. -> 4 significant figures
(vi) 2.0034: all digits are significant (zero between non-zero digits). -> 5 significant figures

Answer: (i) 2 (ii) 3 (iii) 4 (iv) 3 (v) 4 (vi) 5

Q20 Round up the following upto three significant figures:
(i) 34.216
(ii) 10.4107
(iii) 0.04597
(iv) 2808

SOLUTION

To round to 3 significant figures, retain 3 sig figs and round the last retained digit based on the digit after it.

(i) 34.216: 3 sig figs -> 34.2 (drop 1, which is < 5, so 2 stays)
(ii) 10.4107: 3 sig figs -> 10.4 (drop 107... first dropped digit is 1 < 5)
(iii) 0.04597: leading zeros not significant; 3 sig figs from 4,5,9,7 -> 0.0460 (the 7 rounds 9 up; 9+1=10 carries; 5+1=6 -> 0.0460)
(iv) 2808: 3 sig figs -> 2810 (drop 8, which is > 5, so 0 rounds up to 1, giving 2810

= 2.81 × 10^3)

Answer: (i) 34.2 (ii) 10.4 (iii) 0.0460 (iv) 2810 (= 2.81 × 10^3)

Q21 The following data are obtained when dinitrogen and dioxygen react together to form different compounds:

	Mass of dinitrogen	Mass of dioxygen
(i)	14 g	16 g
(ii)	14 g	32 g
(iii)	28 g	32 g
(iv)	28 g	80 g

(a) Which law of chemical combination is obeyed by the above experimental data? Give its statement.
(b) Fill in the blanks in the following conversions:
(i) 1 km = ... mm = ... pm
(ii) 1 mg = ... kg = ... ng
(iii) 1 mL = ... L = ... dm^3

SOLUTION

(a) Fix the mass of dinitrogen at 14 g and find corresponding mass of dioxygen:
(i) 14 g N_2, 16 g O_2 -> O_2 = 16 g
(ii) 14 g N_2, 32 g O_2 -> O_2 = 32 g
(iii) 28 g N_2, 32 g O_2 -> 14 g N_2 corresponds to 16 g O_2
(iv) 28 g N_2, 80 g O_2 -> 14 g N_2 corresponds to 40 g O_2

Ratio of O_2 masses (for fixed 14 g N_2) = 16 : 32 : 16 : 40 = 2 : 4 : 2 : 5
These are simple whole-number ratios.

Law obeyed: Law of Multiple Proportions (Dalton, 1803).
Statement: When two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in the ratio of small whole numbers.

(b) Unit conversions:
(i) 1 km = 10^3 m × (10^3 mm/m) = 10^6 mm
1 km = 10^3 m × (10^{12} pm/m) = 10^{15} pm
So: 1 km = 10^6 mm = 10^{15} pm

(ii) 1 mg = 10^{-3} g = $10^{-3} \times 10^{-3}$ kg = 10^{-6} kg
1 mg = 10^{-3} g = $10^{-3} \times 10^{9}$ ng = 10^{6} ng
So: 1 mg = 10^{-6} kg = 10^{6} ng

(iii) 1 mL = 10^{-3} L
1 L = 1 dm^3, so 1 mL = 10^{-3} dm^3
So: 1 mL = 10^{-3} L = 10^{-3} dm^3

Answer: (a) Law of Multiple Proportions; ratio of O_2 = 2:4:2:5. (b)(i) 10^{6} mm, 10^{15} pm (ii) 10^{-6} kg, 10^{6} ng (iii) 10^{-3} L, 10^{-3} dm^3

Q22 If the speed of light is 3.0 × 10^{8} m s^{-1}, calculate the distance covered by light in 2.00 ns.

SOLUTION

Given: speed of light c = 3.0 × 10^{8} m s^{-1}
Time t = 2.00 ns = 2.00 × 10^{-9} s

Distance = speed × time
= 3.0 × 10^{8} m s^{-1} × 2.00 × 10^{-9} s
= 6.0 × 10^{-1} m
= 0.600 m

Answer: 0.600 m

Q23 In a reaction A + B_2 -> AB_2, identify the limiting reagent, if any, in the following reaction mixtures.
(i) 300 atoms of A + 200 molecules of B_2
(ii) 2 mol A + 3 mol B_2
(iii) 100 atoms of A + 100 molecules of B_2
(iv) 5 mol A + 2.5 mol B_2
(v) 2.5 mol A + 5 mol B_2

SOLUTION

Balanced equation: A + B_2 -> AB_2
Molar ratio: 1 mol A : 1 mol B_2 (1:1 stoichiometry)

(i) 300 atoms A + 200 molecules B_2:
200 B_2 needs exactly 200 A. Only 200 of 300 A are consumed.
Limiting reagent: B_2 (200 molecules used; 100 atoms A remain).

(ii) 2 mol A + 3 mol B_2:
2 mol A needs 2 mol B_2. Only 2 mol B_2 used; 1 mol B_2 remains.
Limiting reagent: A

(iii) 100 atoms A + 100 molecules B_2:
Equal amounts; both are completely consumed. No limiting reagent (exact stoichiometry).

(iv) 5 mol A + 2.5 mol B_2:
5 mol A needs 5 mol B_2, but only 2.5 mol B_2 available.
Limiting reagent: B_2

(v) 2.5 mol A + 5 mol B_2:
2.5 mol A needs 2.5 mol B_2. Only 2.5 mol B_2 used; 2.5 mol B_2 remains.
Limiting reagent: A

Answer: (i) B_2 (ii) A (iii) No limiting reagent (both fully consumed) (iv) B_2 (v) A

Q24 Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:
$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
(i) Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen reacts with 1.00×10^3 g of dihydrogen.
(ii) Will any of the two reactants remain unreacted?
(iii) If yes, which one and what would be its mass?

SOLUTION

Equation: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
Molar masses: N_2 = 28.014 g mol^{-1}, H_2 = 2.016 g mol^{-1}, NH_3 = 17.031 g mol^{-1}

Step 1 — moles of reactants:
Moles of N_2 = 2000 g / 28.014 g mol^{-1} = 71.39 mol
Moles of H_2 = 1000 g / 2.016 g mol^{-1} = 496.0 mol

Step 2 — identify limiting reagent:
For 71.39 mol N_2, H_2 needed = 71.39 × 3 = 214.2 mol
Available H_2 = 496.0 mol > 214.2 mol => H_2 is in excess.
N_2 is the limiting reagent.

(i) Moles of NH_3 = 71.39 mol N_2 × (2 mol NH_3 / 1 mol N_2) = 142.8 mol
Mass of NH_3 = 142.8 × 17.031 = 2431 g ≈ 2.431 kg

(ii) H_2 remains unreacted (N_2 is consumed completely).

(iii) H_2 consumed = 71.39 mol × 3 = 214.2 mol
H_2 remaining = 496.0 - 214.2 = 281.8 mol
Mass of excess H_2 = 281.8 × 2.016 = 568.2 g

Answer: (i) 2431 g (≈ 2.43 kg) NH_3 (ii) Yes, H_2 remains unreacted (iii) 568 g of H_2 remains

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Q25 How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?

SOLUTION

0.50 mol Na_2CO_3 refers to a definite amount (number of moles) of substance.
Molar mass of Na_2CO_3 = 2(22.990) + 12.011 + 3(15.999) = 105.988 g mol^{-1}
Mass of 0.50 mol Na_2CO_3 = 0.50 × 105.988 = 52.99 g
This fixed mass can be present in any volume of solution.

0.50 M Na_2CO_3 refers to a concentration: 0.50 mol of Na_2CO_3 dissolved per 1 litre of solution.
This specifies how many moles are present in a given volume of solution, not a fixed total amount.

Key difference: 0.50 mol is an amount (moles) independent of volume, while 0.50 M is a concentration (moles per litre) and the number of moles depends on the volume taken.

Answer: 0.50 mol Na_2CO_3 = 52.99 g (fixed amount, no volume specified). 0.50 M Na_2CO_3 = 0.50 mol dissolved in 1 L solution (a concentration; amount depends on volume used).

Q26 If 10 volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

SOLUTION

Balanced equation: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

By Gay-Lussac's Law, gases react in simple ratios by volume at the same T and P.

Volume ratio: $H_2 : O_2 : H_2O = 2 : 1 : 2$

Given 10 volumes H_2 and 5 volumes O_2:

Check stoichiometry: 10 vol H_2 needs 10/2 = 5 vol O_2. Exactly 5 vol O_2 available. Both reactants are completely consumed (no limiting reagent).

Volume of H_2O produced = 10 volumes (same ratio as H_2)

Answer: 10 volumes of water vapour

Q27 Convert the following into basic units:
(i) 28.7 pm
(ii) 15.15 pm
(iii) 25365 mg

SOLUTION

The basic SI units are metre (m) for length and kilogram (kg) for mass.

(i) 28.7 pm: 1 pm = 10^{-12} m
28.7 pm = 28.7 × 10^{-12} m = 2.87 × 10^{-11} m

(ii) 15.15 pm:
15.15 pm = 15.15 × 10^{-12} m = 1.515 × 10^{-11} m

(iii) 25365 mg: 1 mg = 10^{-3} g = 10^{-6} kg
25365 mg = 25365 × 10^{-6} kg = 2.5365 × 10^{-2} kg

Answer: (i) 2.87 × 10^{-11} m (ii) 1.515 × 10^{-11} m (iii) 2.5365 × 10^{-2} kg

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Q28 Which one of the following will have the largest number of atoms?
(i) 1 g Au (s)
(ii) 1 g Na (s)
(iii) 1 g Li (s)
(iv) 1 g of Cl_2(g)

SOLUTION

Number of atoms = (mass / molar mass) × NA [for monoatomic; ×2 for diatomic Cl_2]
NA = 6.022 × 10^{23} mol^{-1}

Molar masses: Au = 196.97 g mol^{-1}, Na = 22.99 g mol^{-1}, Li = 6.941 g mol^{-1}, Cl_2 = 70.91 g mol^{-1}

(i) 1 g Au: (1/196.97) × 6.022×10^{23} = 3.057 × 10^{21} atoms
(ii) 1 g Na: (1/22.99) × 6.022×10^{23} = 2.619 × 10^{22} atoms

(iii) 1 g Li: $(1/6.941) \times 6.022 \times 10^{23} = 8.676 \times 10^{22}$ atoms
(iv) 1 g Cl_2: $(1/70.91) \times 6.022 \times 10^{23} \times 2 = 1.699 \times 10^{22}$ atoms

Largest: 1 g Li with 8.676×10^{22} atoms (Li has the smallest molar mass among the options).

Answer: (iii) 1 g Li has the largest number of atoms (8.676×10^{22} atoms)

Q29 Calculate the molarity of a solution of ethanol in water, in which the mole fraction of ethanol is 0.040 (assume the density of water to be one).

SOLUTION

Given: mole fraction of ethanol (xE) = 0.040, so mole fraction of water (xW) = 0.960
Density of water = 1 g mL^{-1} => 1 L water has mass 1000 g
Molar mass of water = 18.015 g mol^{-1}

Moles of water in 1 L = 1000 / 18.015 = 55.51 mol

From mole fraction: xE = nE / (nE + nW)
0.040 = nE / (nE + 55.51)
0.040 nE + 0.040 × 55.51 = nE
0.040 × 55.51 = nE (1 - 0.040) = 0.960 nE
nE = (0.040 × 55.51) / 0.960 = 2.313 mol

Assuming solution volume ≈ 1 L (dilute aqueous solution, density ≈ 1 g/mL):
Molarity = 2.313 mol / 1 L = 2.31 mol L^{-1}

Answer: 2.31 mol L^{-1} (approximately 2.31 M)

Q30 What will be the mass of one 12C atom in g?

SOLUTION

By definition, 1 mole of 12C atoms = 12 g (exact).
1 mole contains NA = $6.02214076 \times 10^{23}$ atoms.

Mass of one 12C atom = 12 g / $(6.02214076 \times 10^{23})$
= 1.9926×10^{-23} g

Answer: 1.9926×10^{-23} g

Q31 How many significant figures should be present in the answer of the following calculations?
(i) (0.02856 × 298.15 × 0.112) / 0.5785
(ii) 5 × 5.364
(iii) 0.0125 + 0.7864 + 0.0215

SOLUTION

(i) (0.02856 × 298.15 × 0.112) / 0.5785
Significant figures in each number: 0.02856 = 4, 298.15 = 5, 0.112 = 3, 0.5785 = 4
For multiplication/division, the result has the same number of sig figs as the factor with the fewest sig figs.
Fewest = 3 (from 0.112). Answer should have 3 significant figures.

(ii) 5 × 5.364
5 is an exact integer (infinite sig figs); 5.364 has 4 sig figs.
Answer should have 4 significant figures.

(iii) 0.0125 + 0.7864 + 0.0215
For addition/subtraction, result is rounded to the least number of decimal places.
All three numbers have 4 decimal places, so the answer has 4 decimal places (4 significant figures after the decimal).
0.0125 + 0.7864 + 0.0215 = 0.8204 (4 decimal places)

Answer: (i) 3 significant figures (ii) 4 significant figures (iii) 4 decimal places (4 sig figs)

Q32 Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes:

Isotope	Isotopic molar mass	Abundance
36Ar	35.96755 g mol^{-1}	0.337%
38Ar	37.96272 g mol^{-1}	0.063%
40Ar	39.9624 g mol^{-1}	99.600%

SOLUTION

Average molar mass = sum of (fractional abundance × isotopic molar mass)
= (0.337/100)(35.96755) + (0.063/100)(37.96272) + (99.600/100)(39.9624)
= 0.00337 × 35.96755 + 0.00063 × 37.96272 + 0.99600 × 39.9624

= 0.12121 + 0.023917 + 39.8025
= 39.948 g mol^{-1}

Answer: 39.948 g mol^{-1}

Q33 Calculate the number of atoms in each of the following:
(i) 52 moles of Ar
(ii) 52 u of He
(iii) 52 g of He

SOLUTION

$N_A = 6.022 \times 10^{23}$ mol^{-1}
Atomic mass of He = 4.003 u, Atomic mass of Ar = 39.948 u

(i) 52 mol Ar:
Number of atoms = $52 \times 6.022 \times 10^{23} = 3.131 \times 10^{25}$ atoms

(ii) 52 u of He:
Mass of 1 He atom = 4.003 u
Number of atoms = 52 u / 4.003 u = 12.99 ≈ 13 atoms

(iii) 52 g of He:
Moles of He = 52 g / 4.003 g mol^{-1} = 12.99 mol
Number of atoms = $12.99 \times 6.022 \times 10^{23} = 7.823 \times 10^{24}$ atoms

Answer: (i) 3.131×10^{25} atoms (ii) 13 atoms (iii) 7.823×10^{24} atoms

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Q34 A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

SOLUTION

Molar masses: CO_2 = 44.009, H_2O = 18.015, C = 12.011, H = 1.008 g mol^{-1}

Step 1 — find mass of C and H from combustion products:
Mass of C = 3.38 × (12.011/44.009) = 0.9226 g
Mass of H = 0.690 × (2×1.008/18.015) = 0.07717 g

Step 2 — moles of each element:

Moles of C = 0.9226 / 12.011 = 0.07681 mol
Moles of H = 0.07717 / 1.008 = 0.07656 mol

Step 3 — ratio C : H = 0.07681 : 0.07656 ≈ 1.00 : 0.997 ≈ 1 : 1
Empirical formula = CH

(ii) Molar mass from density at STP (22.4 L mol^{-1} at old STP):
Molar mass = (11.6 g / 10.0 L) × 22.4 L mol^{-1} = 25.98 g mol^{-1} ≈ 26 g mol^{-1}

(iii) Empirical formula mass of CH = 12.011 + 1.008 = 13.019 g mol^{-1}
n = molar mass / empirical formula mass = 26 / 13 = 2
Molecular formula = C_2H_2 (ethyne / acetylene)

Answer: (i) CH (ii) 26 g mol^{-1} (iii) C_2H_2 (ethyne)

Q35 Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction:
$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$
What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl?

SOLUTION

Given: Volume of HCl = 25 mL = 0.025 L, Molarity of HCl = 0.75 M
Molar mass of $CaCO_3$ = 40.078 + 12.011 + 3(15.999) = 100.086 g mol^{-1}

Moles of HCl = 0.75 mol L^{-1} × 0.025 L = 0.01875 mol

From stoichiometry: 1 mol $CaCO_3$ reacts with 2 mol HCl
Moles of $CaCO_3$ = 0.01875 / 2 = 0.009375 mol

Mass of $CaCO_3$ = 0.009375 × 100.086 = 0.9383 g ≈ 0.938 g

Answer: 0.938 g of $CaCO_3$

Q36 Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction:
$4HCl(aq) + MnO_2(s) \rightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$
How many grams of HCl react with 5.0 g of manganese dioxide?

SOLUTION

Molar mass of MnO_2 = 54.938 + 2(15.999) = 86.936 g mol^{-1}
Molar mass of HCl = 1.008 + 35.453 = 36.461 g mol^{-1}

Moles of MnO_2 = 5.0 g / 86.936 g mol^{-1} = 0.05751 mol

From stoichiometry: 1 mol MnO_2 reacts with 4 mol HCl

Moles of HCl = 0.05751 × 4 = 0.2300 mol

Mass of HCl = 0.2300 × 36.461 = 8.39 g

Answer: 8.39 g of HCl