AS Chemistry

Calculations

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Relative Atomic Mass, Ar

Relative atomic mass is a weighted mean mass. It is the average mass of an element taking into account the contribution from all that element’s isotopes when compared to one-twelfth of the mass of carbon-12.

*Example 1*

*A sample of bromine contains 53.00% bromine-79 and 47.00% bromine-81. Determine the relative atomic mass of bromine.*

**Ar(Br) = 53.00/100 x 79 + 47.00/100 x 81**

 **= 41.87 + 38.07**

 **= 79.94**

1. Which two subatomic particles contribute to the mass of an atom?
2. What is the definition of relative atomic mass, Ar?
3. Calculate the relative atomic masses, Ar, of the following. Give your answers to 4 significant figures.
4. Boron: 19.77% 10B and 80.23% 11B
5. Silicon: 92.18% 28Si, 4.70% 29Si and 3.12% 30Si
6. Chromium: 4.31% 50Cr, 83.76% 52Cr, 9.55% 53Cr and 2.38% 54Cr

Relative Molecular Mass Mr

Relative molecular mass, Mr, is the weighted mean mass of a molecule compared with one-twelfth the mass of carbon-12

 This can be worked out by adding together the relative atomic masses of each atom making up a molecule.

*Example 2*

*Deduce the relative molecular mass of a molecule of chlorine.*

**Mr(Cl­2) = Ar x 2**

 **= 35.5 x 2**

 **= 71.0**

1. Using Ar ­values from the periodic table calculate the relative molecular mass of the following;
2. HCl
3. CO2
4. H2S
5. NH3
6. H2SO4

Relative Formula Mass

Substances such as NaCl do not exist as discrete molecules but instead as giant structures. Relative formula mass is the weighted mean mass of one formula unit compared with one-twelfth the mass of carbon-12.

 It can be calculated by adding up the Ar of all the atoms making up a formula unit.

*Example 3*

*Calculate the relative formula mass of CaBr2.*

**RfM(CaBr2) = 40.1 + (79.9 x 2)**

 **= 199.9**

1. Using Ar values from the periodic table calculate the relative formula masses of;
2. Fe2O3
3. Na2O
4. Pb(NO3)2
5. (NH4)2SO4
6. Ca3(PO4)2

Mass, Molar Mass& Moles

***n = m/M***

where; n = amount of substance, in mol

m = mass in g

M = molar mass in g mol-1

The Avogadro constant, NA, 6.02 x 1023 mol-1

The molar mass is the mass per mole of a substance. One mole always contains the same number of particles, 6.02 x 1023, which is equal to the number of atoms in a mole of the carbon-12 isotope. One mole of carbon-12 weighs exactly 12g.

1. What is the molar mass of;
2. Fe
3. Pb
4. O2
5. NaCl
6. H2O

To calculate the mass, in g, of a substance from an amount in moles you must know;

* The amount of substance, in moles.
* The molar mass of the substance in grams per mole.

*Example 4*

*Calculate the mass, in g, of 2.00mol of CH4.*

**Firstly we can work out the M of CH4 by deducing the Mr;**

 **Mr(CH4­) = 12 + (4 x 1)**

 **= 16**

**Therefore one mole of CH4 has a mass of 16g. (CH4 = 16 g mol-1)**

 **To work out the mass of two moles we use the formula m = M x n, so;**

 **m = 16 g mol-1 x 2 mol**

 **= 32g**

1. Calculate the mass, in g, of the following;
2. 1.00 mol of H
3. 1.00 mol of H2
4. 10.0 mol of Cl2
5. 3.00 mol of SiO2
6. 0.500 mol of Fe2O3
7. 0.250 mol Na2CO3

To calculate the amount of a substance, in moles, we must know;

* The mass of substance we have.
* The molar mass of the substance.

*Example 5*

*Calculate the amount, in moles, of CO2 in 96g of the substance.*

**Calculating the Mr of CO2 gives us;**

 **Mr(CO2) = 12 + (16 x 2)**

 **= 48**

**Therefore CO2 has a molar mass of 48 g mol-1. Using the formula n = m / M gives us;**

 **n = 96g / 48g mol-1**

 **= 2 mol**

1. Calculate the amount, in mol, of the following;
2. S in 32.10g
3. Pb in 413.60g
4. H2O in 9.00g

1. C6H6 in 195.00g
2. C2H5OH in 5.75g

Calculating Empirical Formulae

Empirical formulae are the simplest way of expressing a chemical formula. For example a salt crystal will contain many millions of atoms of Na and Cl arranged in a giant structure but we usually represent it using its empirical formula of NaCl.

We can deduce the empirical formula of a compound if we know the mass of each element present in it.

*Example 6*

*Analysis shows that 0.6075 g of Magnesium combines with 3.995g of bromine to form a compound [Ar: Mg, 24.3; Br 79.9].*

*Find the molar ratio of atoms;*

|  |  |  |
| --- | --- | --- |
|  | **Mg** | **Br** |
| **(mass / molar mass)** | **0.6075/24.3** | **3.995/79.9** |
|  | **0.025** | **0.050** |
| **Divide by the smallest****(0.25)** | **1**  | **: 2** |
| **Empirical formula is;** |  **MgBr2** |  |

*Example 7*

*Analysis of a compound showed the following percentage composition by mass:*

*Na: 74.19 % O: 25.81 % [Ar: Na, 23.0; O, 16.0]*

**We know that a hundred grams of the substance will contain 74.19g of Na and 25.81g of O.**

|  |  |  |
| --- | --- | --- |
|  | **Na** | **O** |
| **(mass / molar mass)** | **74.19/23.0** | **25.81/16** |
|  | **3.226** | **1.613** |
| **Divide by the smallest (1.613)** | **2** | **1** |
| **Empirical Formula Is;** | **Na2O** |

1. Determine the empirical formula of the following;
2. The compound formed when 6.54g of Zinc reacts with 1.60g of Oxygen.
3. The compound formed when 5.40g of aluminium reacts with 4.80g of oxygen.
4. The compound containing: Ag, 69.19%; S, 10.29%; and O, 20.52%.

Molecular Formula

Molecular formula are used for compounds that exist as simple molecules, they tell you the number of each type of atom that make up a molecule.

 They can be determined if we know the molecular formula and Mr of a substance.

*Example 8*

*A compound has the empirical formula CH2 and the Mr 56.0. What is the molecular formula?*

**Empirical formula mass of CH2: 12.0 + (1.0 x 2) = 14.0**

**Number of CH2 units in a molecule: 56.0/14 = 4**

**Molecular formula: (4 x CH2); C4H8**

1. Determine the molecular formula of the following;
2. The compound containing 0.49g of nitrogen combined with 1.12g of oxygen, Mr 92.0.
3. The compound consisting of carbon, hydrogen and oxygen atoms containing 1.80g of carbon combined with 0.30g of hydrogen and 1.20g of oxygen, Mr 88.0.
4. The compound of carbon, hydrogen and oxygen with the composition by mass: C, 40.0%; H, 6.7%; O, 53.3%; Mr 180.0.

 Moles and Gas Volumes

Avogadro’s Law states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules. This means that the same amount of any gas under the same conditions will occupy the same amount of space.

* One mole of gas occupies 24.0 dm3 (24 000 cm3) at room temperature and pressure (RTP).
* The molar volume of any gas at RTP is 24.0 dm3 mol-1

We can use this to calculate amounts using the equations:

1. n = (V / 24.0) if the volume is in dm3
2. n = (V / 24 000) is the volume is in cm3

Where n is the amount of gas in moles and V is the volume in either dm3 or cm3.

1. Calculate the amount of gas molecules, in moles, that are in the following volumes of gas at RTP;
2. 36.0 dm3
3. 1080.0dm3
4. 4.0 dm3

The equation can be rearranged to tell us how much volume certain amounts of gas will occupy at RTP.

V = n x (24 000 or 24)

1. What is the volume of the following at RTP?
2. 6.00 mol SO2(g)
3. 0.25 mol O2(g)
4. 20.7 g NO2(g)
5. What is the mass of the following at RTP?
6. 0.6dm3 N2
7. 1920 cm3 C3H8(g)
8. 84 cm3 N2O(g)
9. What is the volume of the following at RTP?
10. 1.282g SO2(g)
11. 1.485g HCN(g)
12. 1.260g C3H6(g)

Moles and Solution

Concentration is a measure of how much of a substance (the solute) is dissolved 1dm3 (1000 cm3) of a solvent.

If we know the concentration we can work out the amount of substance in any volume using the equation;

n = c x V (in dm3) or n = c x V/1000 (in cm3)

Where n is amount of substance in moles, c is concentration and V is volume in either dm3 or cm3.

1. Find the amount, in moles, of solute dissolved in the following solutions;
2. 4dm3 of a 2 mol dm-3 solution.
3. 25.0cm3 of a 0.150 mol dm-3 solution.
4. 24.35cm3 of a 0.125 mol dm-3 solution.
5. Find the concentration, in mol dm-3, for the following solutions (*you will have to rearrange the equation to find concentration*);
6. 6 moles dissolved in a 2dm3 solution.
7. 0.500 moles dissolved in 250cm3 of solution.
8. 8.75 x 10-3 moles dissolved in 250cm3 of solution.

The mass concentration is the mass of a solute, in g, dissolved in 1dm3 of solution. They are measured in units of g dm-3.

1. Find the mass concentration, in g dm-3, for the following solutions;
2. 0.042 moles of HNO3 dissolved in 250cm3 of solution.
3. 0.500 moles of HCl dissolved in 4dm3 of solution.
4. 3.56 x 10-3 moles of H2SO4 dissolved in 25cm3 of solution.

Exam Questions

1. (i) Complete the table below to show the number of sub-atomic particles in an atom of iodine-127 and iodine-131. [2]

|  |  |  |  |
| --- | --- | --- | --- |
|  | **Protons** | **Neutrons** | **Electrons** |
| **Iodine-127** |  |  | **53** |
| **Iodine-131** |  |  | **53** |

1. In the human body iodide ions, I-, are necessary for the thyroid gland to function correctly. Some countries add potassium iodide, KI, to table salt as a source of iodide ions.

The guideline daily amount, GDA, of iodide ions is 70.0µg

(1 µg = 1x10-6g).

Calculate the mass of KI, in µg, that would be needed to supply the GDA of iodide ions. [2]

1. A student knew that calcium hydroxide could be made by adding calcium to water.

The student added 0.00131mol of calcium to a beaker containing about 100cm3 of water.

A reaction took place as shown in the equation below.

All the calcium hydroxide formed was soluble.

Ca(s) + 2H2O(l) 🡪 Ca(OH)2(aq) + H2(g)

1. Calculate the mass of calcium that the student added [1]
2. Calculate the volume of hydrogen gas, in dm3, produced in this reaction at room temperature and pressure [1]
3. The student transferred the contents of the beaker to a 250cm3 volumetric flask and water was added to make the solution up to 250cm3.

Calculate the concentration, in mol dm-3, of hydroxide ions in the 250cm3 solution [2]

1. The student repeated the experiment using the same mass of pure barium.

The student found that a smaller volume of hydrogen gas was produced, measured at RTP.

Explain why. [3]

1. (a) A compound containing magnesium, silicon and oxygen is present in rocks in Italy. A sample of this compound weighing 5.27g was found to have the following composition by mass:

Mg, 1.82g; Si, 1.05g; O, 2.40g

1. Calculate the empirical formula of the compound. [2]

(b) Pharmacists sell tablets containing magnesium hydroxide, Mg(OH)2 to combat indigestion.

A student carried out an investigation to find the percentage mass of Mg(OH)2 in an indigestion tablet. The student reacted the tablet with dilute hydrochloric acid.

Mg(OH)2(s) + 2HCl(aq) 🡪MgCl2(aq) + 2H2O(l)

The student found that 32.00cm3 of 0.500mol dm-3 HCl was needed to react with the Mg(OH)2 in a 500mg tablet. [1g = 1000mg]

1. Calculate the amount, in mol, of HCl used. [1]
2. Determine the amount, in mol, of Mg(OH)2 present in the tablet. [1]
3. Determine the percentage by mass of Mg(OH)2 present in the tablet. [3]

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