**Reacting masses and volumes**

Learning objectives

* Solve problems involving masses of substances
* Calculate the theoretical and percentage yield in a reaction
* Understand the terms limiting reactant and reactant in excess and solve problems involving these.

**Using moles**

There are three main steps in a moles calculation.

* Work out the number of moles of anything you can.
* Use the chemical (stoichiometric) equation to work out the number of moles of the quantity you require.
* Convert moles to the required quantity – volume, mass etc.

**Solve the following:**

1. **Consider the reaction of sodium with oxygen:**

 4Na(s) + O2(g) → 2Na2O(s)

1. How much sodium reacts exactly with 3.20 g of oxygen?
2. What mass of Na2O is produced?
3. **Consider the following equation:**

2NH3 + 3CuO → N2 + 3H2O + 3Cu

If 2.56 g of ammonia (NH3) is reacted with excess CuO, calculate the mass of copper produced.

Formula for solving mole equations involving masses

$$\frac{m1}{n1m1}=\frac{m2}{n2m2}$$

1. **The following equation represents the combustion of butane:**

2C4H10(g) + 13O2(g) → 8CO2(g) + 10H2O(l)

If 10.00 g of butane is used, calculate the mass of oxygen required for an exact reaction.

1. **Calculate the mass of carbon dioxide produced from the complete combustion of 1.00 g of methane.**
2. **Iodine chloride, ICl, can be made by the following reaction:**

 2I2 + KIO3 + 6HCl → 5ICl + KCl + 3H2O

Calculate the mass of iodine, I2, needed to prepare 28.60 g of ICl by this reaction.

**Calculating the yield of a chemical reaction-** The theoretical yield is determined by the limiting reactant.

In any commercial process it is very important to know the yield (the amount of desired product) of a chemical reaction. The reactant that determines the quantity of product is known as the **limiting reactant.** Other reactants will therefore not be fully used, and are said to be in **excess. The yield of a chemical reaction is usually quoted as a percentage.**

**Example: Consider the preparation of 1,2-dibromoethane (C2H4Br2)**

 **C2H4(g) + Br2(l) → C2H4Br2(l)**

10.00 g of ethene (C2H4) will react exactly with 56.95 g of bromine. The theoretical yield for this reaction is 66.95 g – this is the maximum possible yield that can be obtained. The actual yield of C2H4Br2 may be 50.00 g.

 % yield =$\frac{50.00}{66.95}$ × 100 = 74.68%

1. **C2H5OH(l) + CH3COOH(l) → CH3COOC2H5(l) + H2O(l)**

 (ethanol) (ethanoic acid) (ethyl ethanoate) (water)

If the yield of ethyl ethanoate obtained when 20.00 g of ethanol is reacted with excess ethanoic acid is 30.27 g, calculate the percentage yield.

1. **Consider the reaction between magnesium and nitrogen:**

3Mg(s) + N2(g) → Mg3N2(s)

10.00 g of magnesium is reacted with 5.00 g of nitrogen. Which is the limiting reactant?

**Exam tip**

To do a moles question you need to know the mass of just one of the reactants. If you are given the masses of more than one reactant, you must consider that one of these reactants will be the limiting reactant and you must use this one for all calculations.

**Calculations involving volumes of gases**

**Learning objectives**

* Understand Avogadro’s law and use it to calculate reacting volumes of gases
* Use the molar volume of a gas in calculations at standard temperature and pressure
* Understand the relationships between pressure, volume and temperature for an ideal gas
* Solve problems using the equation $\frac{P1V1}{T1}=\frac{P2V2}{T2}$
* Solve problems using the ideal gas equation PV=Nrt

An ‘ideal gas’ is a concept invented by scientists to approximate (model) the behaviour of real gases. Two assumptions we make when defining the ideal gas are that the molecules themselves have no volume (they are point masses) and that no forces exist between them (except when they collide). This means that the volume occupied by a gas at a certain temperature and pressure depends only on the number of molecules present and not on the nature of the gas.

Volume of gas ∝ number of moles of the gas

Gases deviate most from ideal behaviour at high pressure and low temperature.

**Using volumes of gases**

***Avogadro’s law: equal volumes of ideal gases measured at the same temperature and pressure contain the same number of molecules.***

In other words 100 cm3 of H2 contains the same number of molecules at 25 °C and 100 kPa as 100 cm3 of NH3, if we assume that they both behave as ideal gases.

This means that volumes can be used directly (instead of moles) in equations involving gases.

TOK

*The ideal gas concept is an approximation which is used to model the behaviour of real gases. Why do we learn about ideal gases when they do not exist? What implications does the ideal gas concept have on the limits of knowledge gained from this course?*

1. Consider the following reaction for the synthesis of methanol:

 CO(g) + 2H2(g) → CH3OH(g)

1. What volume of H2 reacts exactly with 2.50 dm3 of CO?
2. What volume of CH3OH is produced?

Converting volumes of gases to number of moles

The volume occupied by one mole of a gas under certain conditions is called the molar volume.

 STP= standard temperature and pressure = 273 K, 100 kPa (1 bar).

 100 kPa = 1.00 × 105 Pa

Molar volume of an ideal gas at STP = 22.7 dm3 mol−1 or 2.27 × 10−2 m3 mol−1

 number of moles = $\frac{volume}{ molar volume}$

A change of 1 °C is the same as a change of 1 K, and 0 °C is equivalent to 273 K

Because 1 dm3 (1 litre) is equivalent to 1000 cm3 to convert cm3 to dm3 we divide by 1000.

1. Calculate the number of moles in 250 cm3 of O2 at STP.

250cm^3=0.24dm^3

0.25/22.7=0.01moles

1. Calculate the volume of 0.135 mol CO2 at STP.

0.135\*22.7=3.06345dm^3

1. Calculate the volume of carbon dioxide (collected at STP) produced when 10.01 g of calcium carbonate decomposes according to the equation:

 CaCO3(s) → CaO(s) + CO2(g)

10.01/(40+12+16\*3)=0.1001moles

0.1001\*22.7=2.27dm^3

**Gas Laws**

The gas laws describe pressure, volume, and temperature relationships for all gases.

1. **Relationship between volume and pressure- Boyle’s law**

*At a constant temperature, the volume of a fixed mass of an ideal gas is inversely proportional to its pressure.*

By use of graphs show the relationship between*volume of a fixed mass of an ideal gas and pressure at constant temperature.*

1. **The relationship between volume and temperature (Charles’ law)**

If the temperature is in kelvin, the following relationship exists between the volume and the temperature:

*The volume of a fixed mass of an ideal gas at constant pressure is directly proportional to its kelvin temperature. V ∝ T*

**Graph:** The relationship between the volume and temperature (in kelvin) of a fixed mass of an ideal gas at constant pressure.

1. **The relationship between pressure and temperature**

Draw a graph to show the relationship between the pressure and temperature (kelvin) of a fixed mass of an ideal gas at constant volume.

For a fixed mass of an ideal gas at constant volume, the pressure is directly proportional to its absolute temperature: P ∝ T.

**The overall gas law equation**

An ideal gas is one that obeys all of the above laws exactly.

The three relationships above can be combined to produce the following equation:

 $\frac{P1V1}{T1}=\frac{P2V2}{T2}$

The temperature must be kelvin.

**Problem solving**

1. If the volume of an ideal gas collected at 0 °C and 100 kPa, i.e. at STP, is 50.0 cm3, what would be the volume at 60 °C and 108 kPa? Ans (56.5 cm3)

Volume = 0.0564 dm3=56.5 cm3

1. What temperature (in °C) is required to cause an ideal gas to occupy 1.34 dm3 at a pressure of 200 kPa if it occupies 756 cm3 at STP? Ans (695°C)

Temperature = 695°C use ideal gas equation

1. What happens to the volume of a fixed mass of gas when its pressure and its absolute temperature are both doubled? Ans (V1=V2)

The ideal gas equation

If the relationships between P, V and T are combined with Avogadro’s law, the ideal gas equation is obtained:

 $\frac{PV}{T}$ = constant

The value of the constant is directly proportional to the fixed mass of gas, or the number of moles, n.

 So $\frac{PV}{T}$ $α n$

This can be made into an equation by introducing a constant, R, known as the universal gas constant.

 $\frac{PV}{T}$ = nR, usually written as PV = nRT, R = 8.31 N m K–1 mol–1 or 8.31 J K–1 mol–1

This equation is known as the ideal gas equation.

1. A helium party balloon has a volume of 18.0 dm3. At 25 °C the internal pressure is 108 kPa. Calculate the mass of helium in the balloon. (3.14g)

108\*18/(25+273)=n\*8.31

108\*18/((25+273)\*8.31)=n

4\*108\*18/((25+273)\*8.31)=3.14

1. A sample of gas has a volume of 445 cm3 and a mass of 1.500 g at a pressure of 95 kPa and a temperature of 28 °C. Calculate its molar mass. (88.8 g mol-1)

95\*0.445/((28+273)\*8.31)=n

1.5/(95\*0.445/((28+273)\*8.31))=88.8

1. A gas has a density of 1.65 g dm–3 at 27 °C and 92.0 kPa. Determine its molar mass.

 density data ⇒ 1.65 g occupies 1.00 dm3 ( 44.7gmol-1)

92\*1/((27+273)\*8.31)=n=0.0369

1.65/0.0369=44.7

1. An ideal gas occupies 590 cm3 at 120 °C and 202 kPa. What amount of gas (in moles) is present? (0. 0365 mol)

(202)(0.590)/((120+273)\*8.31)=n=0.0365mol

**Real gases show a deviation from ideal behavior**

An ideal gas is defined as one that obeys the ideal gas law PV = nRT under all conditions.

This means that for one mole of gas, the relationship PV/RT should be equal to 1.

Graph of $\frac{PV}{RT}$ against P for one mole of an ideal gas

All gases known as real gases deviate to some extent from ideal gas behavior.

* Real gas behaves most like an ideal gas at low pressure and shows the greatest deviation at high pressure
* Real gas behaves most like an ideal gas at high temperature and shows the greatest deviation at low temperature.

Questioning the validity of the two assumptions for an ideal gas

1. The volume of the gas particles is negligible

At relatively low pressure, such as 1 × 105 Pa (STP), the volume occupied by the particles of a typical gas is only about 0.05% of the total volume of the molecules (negligible). With increased pressure of 5 × 105 Pa, the volume of the particles is about 20% of the total volume (not negligible). So, the volume of a real gas at high pressure is larger than that predicted from the ideal gas law and PV/nRT > 1.

1. There are no attractive forces between the particles.

At moderately low pressure, the particles are so widely spaced that interactive forces are highly unlikely, so this assumption is valid. But at high pressure as the particles approach more closely, attractive forces strengthen between them. These have the effect of reducing the pressure of the gas, so PV/nRT < 1.

Low temperatures increase this deviation because the lower kinetic energy of the particles increases the strength of inter-particle forces.

**Real gases deviate most from ideal behaviour at high pressure and low temperature.**