Stoichiometric Relationships Part three (answers)

IB CHEMISTRY SL/HL

25	16		6	2	25
Mn	S	J	C	He	Mn
Manganese	Sulfur		Carbon	Helium	Manganese
54.938045	32.065		12.0107	4.002602	54.938045

Syllabus objectives

Understandings

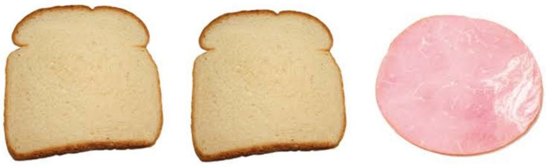
- Reactants can be either limiting or excess.
- The experimental yield can be different from the theoretical yield.
- Avogadro's law enables the mole ratio of reacting gases to be determined from volumes of the gases.
- The molar volume of an ideal gas is a constant at specified temperature and pressure.
- The molar concentration of a solution is determined by the amount of solute and the volume of solution.
- A standard solution is one of known concentration.

Applications and skills

- Solution of problems relating to reacting quantities, limiting and excess reactants, theoretical, experimental and percentage yields.
- Calculation of reacting volumes of gases using Avogadro's law.
- Solution of problems and analysis of graphs involving the relationship between temperature, pressure and volume for a fixed mass of an ideal gas.
- Solution of problems relating to the ideal gas equation.
- Explanation of the deviation of real gases from ideal behavior at low temperatures and high pressures.
- Obtaining and using experimental values to calculate the molar mass of a gas from the ideal gas equation.
- Solution of problems involving molar concentration, amount of solute and volume of solution.
- Use of the experimental method of titration to calculate the concentration of a solution by reference to a standard solution.

Limiting reactant and excess reactant

- The limiting reactant (reagent) is the reactant that limits the amount of product(s) that can be made.
- The excess reactant is the reactant that remains when the limiting reactant is consumed.



- How many sandwiches can be made with 12 pieces of bread and 7 slices of ham?
  6 sandwiches
- Which is the limiting reactant? The bread
- Which is the excess reactant? The ham

How to determine the limiting and excess reactant.

1. Determine the amount (in mol) of each reactant.
2. Divide the amount of each reactant by its coefficient in the balanced equation.
3. The lowest value is the limiting reactant and the highest is the excess reactant.

Exercises:

1. 50.0 g of N_2H_4 is reacted with 75.0 g of N_2O_4 according to the following equation.

$$2N_2H_{4(l)} + N_2O_{4(l)} \rightarrow 3N_{2(g)} + 4H_2O_{(g)}$$

a. Determine the limiting and excess reactants.

$n(N_2H_4) = \frac{50.0 \text{ g}}{32.06 \text{ g mol}^{-1}} = 1.56$ mol $\qquad n(N_2O_4) = \frac{75.0 \text{ g}}{92.02 \text{ g mol}^{-1}} = 0.815$ mol

$N_2H_4 = \frac{1.56}{2} = 0.780 \quad N_2O_4 = \frac{0.815}{1} = 0.815$

N_2H_4 is the limiting reactant and N_2O_4 is the excess reactant.

b. Determine the amount of excess reactant that remains at the end of the reaction.

Molar ratio of N_2H_4 to N_2O_4 is 2:1

1.56 mol $N_2H_4 \times \frac{1 \text{ mol } N_2O_4}{2 \text{ mol } N_2H_4} = 0.780$ mol N_2O_4

$n(N_2O_4)$ remaining = 0.815 – 0.780 = 0.0350 mol

2. 3.00 g of Zn is reacted with 50.0 cm^3 of 1.00 mol dm^{-3} HCl according to the following equation.

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

a. Determine the limiting and excess reactants.

$n(Zn) = \frac{3.00\text{ g}}{65.38\text{ g mol}^{-1}} = 0.0459$ mol $\quad n(HCl) = 1.00\text{ mol dm}^{-3} \times 0.0500\text{ dm}^3 = 0.0500$ mol

$Zn = \frac{0.0459}{1} = 0.0459 \quad HCl = \frac{0.0500}{2} = 0.0250$

HCl is the limiting reactant and Zn is the excess reactant.

b. Determine the amount of excess reactant that remains at the end of the reaction.

Molar ratio of Zn to HCl is 1:2

$0.0500\text{ mol HCl} \times \frac{1\text{ mol Zn}}{2\text{ mol HCl}} = 0.0250\text{ mol Zn}$

$n(Zn)$ remaining = 0.0459 – 0.0250 = 0.0209 mol

3. 30.0 g of ammonium nitrate (NH_4NO_3) and 50.0 g of sodium phosphate (Na_3PO_4) are reacted together. Determine the limiting and excess reactants.

$$3NH_4NO_3 + Na_3PO_4 \rightarrow (NH_4)_3PO_4 + 3NaNO_3$$

M (NH_4NO_3) = 80.04 g mol^{-1}

M (Na_3PO_4) = 163.94 g mol^{-1}

$n(NH_4NO_3) = \frac{30.0}{80.04} = 0.375$ mol $\quad 0.375 \div 3 = 0.125$

$n(Na_3PO_4) = \frac{50.0}{163.94} = 0.305$ mol $\quad 0.305 \div 1 = 0.305$

NH_4NO_3 is limiting reactant, Na_3PO_4 is excess reagent

Theoretical yield and percentage yield

- The theoretical yield is the maximum amount of product that can be formed in a chemical reaction (based on the stoichiometry of the reaction and amount of the limiting reactant).
- The actual yield is the actual amount of product that is formed in a chemical reaction.
- The percentage yield is the actual yield divided by the theoretical yield.

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$$

Exercises:

1. A 15.0 g sample of pure K_2O produces 7.62 g of K_2CO_3. Determine the percentage yield of K_2CO_3 in the reaction.

$$4K_2O_{(s)} + 2CO_{2(g)} \rightarrow 2K_2CO_{3(s)} + 3O_{2(g)}$$

$n(K_2O) = \frac{15.0}{94.20} = 0.159$ mol

$0.159 \div 2 = 0.0796$ mol K_2CO_3

$0.0796 \times 138.19 = 11.0$ g K_2CO_3

$\%\ \text{yield} = \frac{7.62}{11.0} = 69.3\%$

2. Aluminium reacts with excess oxygen according to the following equation. Determine the percentage yield if 20.0 g of Al reacts with oxygen to produce 32.7 g of Al_2O_3.

$$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$$

$n(Al) = \frac{20.0}{26.98} = 0.741$ mol

$0.741 \div 2 = 0.371$ mol Al_2O_3

$0.371 \times 101.96 = 37.8$ g Al_2O_3

$\%\ \text{yield} = \frac{32.7}{37.8} \times 100 = 86.5\%$

3. A 20.0 g sample of pure Fe_3O_4 produces 5.98 g of Fe. Determine the percentage yield of Fe in the reaction.

$$Fe_3O_{4(s)} + 4H_{2(g)} \rightarrow 3Fe_{(s)} + 4H_2O_{(l)}$$

$n(Fe_3O_4) = \frac{20.0}{231.54} = 0.0864$ mol

$0.0864 \times 3 = 0.259$ mol Fe

$0.259 \times 55.85 = 14.5$ g Fe

$\%\ \text{yield} = \frac{5.98}{14.5} = 41.2\%$

4. 100.0 g of iron(II) oxide is reacted with 100.0 g of carbon. 46.73 g of iron is produced. Calculate the % yield of Fe.

$$2Fe_2O_{3(s)} + 3C_{(s)} \rightarrow 4Fe_{(s)} + CO_{2(g)}$$

$n(Fe_2O_3) = \frac{100.0 \text{ g}}{159.70 \text{ g mol}^{-1}} = 0.6262$ mol

$n(C) = \frac{100.0 \text{ g}}{12.01 \text{ g mol}^{-1}} = 8.326$ mol

$Fe_2O_3 = \frac{0.6262}{2} = 0.3131$

$C = \frac{8.326}{3} = 2.775$

Fe_2O_3 is the limiting reactant and C is the excess reactant.

Molar ratio of Fe_2O_3 to Fe is 2:4

$0.6262 \text{ mol } Fe_2O_3 \times \frac{4 \text{ mol Fe}}{2 \text{ mol } Fe_2O_3} = 1.252$ mol Fe

$m(Fe) = 1.252 \text{ mol } \times 55.85 \text{ g mol}^{-1} = 69.92$ g

Theoretical yield of Fe = 69.92 g

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$$

$$\text{Percent yield} = \frac{46.73 \text{ g}}{69.92 \text{ g}} \times 100\ \%$$

$$\text{Percent yield} = 66.83\ \%$$

5. 15.0 g of $CaCO_3$ is reacted with 50.0 cm^3 of 2.00 mol dm^{-3} HCl. 1.85 g of CO_2 is produced. Calculate the % yield of CO_2.

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

$n(CaCO_3) = \frac{15.0 \text{ g}}{100.09 \text{ g mol}^{-1}} = 0.150$ mol

$n(HCl) = 2.00 \text{ mol dm}^{-3} \times 0.0500 \text{ dm}^3 = 0.100$ mol

$CaCO_3 = \frac{0.150}{1} = 0.150$

$HCl = \frac{0.100}{2} = 0.0500$

HCl is the limiting reactant and $CaCO_3$ is the excess reactant.

Molar ratio of HCl to CO_2 is 2:1

$0.100 \text{ mol HCl} \times \frac{1 \text{ mol } CO_2}{2 \text{ mol HCl}} = 0.0500$ mol CO_2

$m(CO_2) = 0.0500 \text{ mol } \times 44.01 \text{ g mol}^{-1} = 2.20$ g

Theoretical yield of CO_2 = 2.20 g

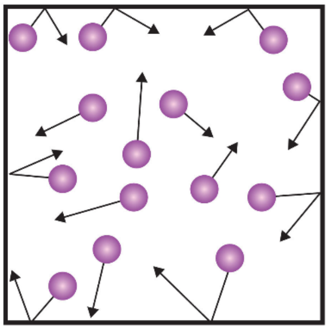
$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$$

$$\text{Percent yield} = \frac{1.85 \text{ g}}{2.20 \text{ g}} \times 100\ \%$$

$$\text{Percent yield} = 84.1\ \%$$

Real gases vs ideal gases

Pressure and temperature of gases



- The particles in a gas are in constant motion, colliding with the walls of the container.
- The forces exerted by the particles on the walls of the container gives rise to pressure.
- Increasing the temperature increases the average kinetic energy of the particles, causing them to move faster.

Ideal gas behavior

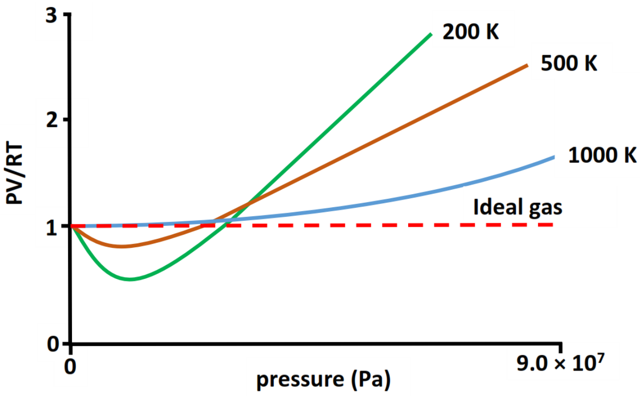
- An ideal gas is one which abides by the kinetic molecular theory (and the gas laws).
- According to the kinetic molecular theory of gases:

1. The particles of an ideal gas are in constant, random, straight line motion.
2. The collisions between ideal gas particles are elastic; total kinetic energy is conserved.
3. The volume occupied by ideal gas particles is negligible relative to the volume of the container.
4. There are no intermolecular forces acting between the particles in an ideal gas.

How does a real gas differ from an ideal gas?

- Under normal conditions, real gases behave very much like ideal gases.
- Real gases differ the most from ideal gases under two conditions; high pressures and low temperatures.
- For one mole of an ideal gas, the product of PV/RT is equal to one.
- Under conditions of high pressure and low temperature, the product of PV/RT is no longer equal to one.

$$\boldsymbol{n} = \frac{\mathbf{1.00 \times 10^5 \times 0.0227}}{\mathbf{8.31 \times 273}} = \mathbf{1.00\ mol}$$



Pressure

- At high pressures, the values of PV/RT are less than one, mainly because of the effects of intermolecular forces; intermolecular forces acting between gaseous particles cause the pressure inside the container to decrease.
- At very high pressures, the values of PV/RT are greater than one, mainly because of the effects of molecular volume; the volume of the gaseous particles becomes significant as the space between them decreases.

Temperature

- At high temperatures, the kinetic energy of the particles overcomes the intermolecular forces between the particles.
- At low temperatures, the particles have insufficient kinetic energy to overcome the intermolecular forces between the particles.

Molar volume of a gas, V_m

- The molar volume of a gas is the volume occupied by one mole of an ideal gas under standard conditions (STP).
- At STP (273 K and 1.00×10^5 Pa), one mole of an ideal gas occupies a volume of:

$$22.7 \text{ dm}^3 \textit{ or } 0.0227 \text{ m}^3$$

- The equations below can be used to calculate amount (in mol) of gas or the volume (in dm^3) of gas.

$$\boldsymbol{V}\ (\mathbf{dm^3}) = \boldsymbol{n}\ (\mathbf{mol}) \times \boldsymbol{V_m}\ (\mathbf{22.\ 7\ dm^3}) \qquad \boldsymbol{V} = \boldsymbol{n} \times \mathbf{22.\ 7}$$

$$\boldsymbol{n}\ (\mathbf{mol}) = \frac{\boldsymbol{V}\ (\mathbf{dm^3})}{\boldsymbol{V_m}\ (\mathbf{22.\ 7\ dm^3})} \qquad \boldsymbol{n} = \frac{\boldsymbol{V}}{\mathbf{22.\ 7}}$$

Examples:

1. Calculate the volume occupied by 16.00 g of O_2 at STP.

   1 mol of gas at STP = 22.7 dm^3

   Convert from mass of O_2 to mol of O_2

   $n = m \div M$

   n = 16.00 ÷ 32.00 = 0.5000 mol O_2

   0.5000 × 22.7 = 11.4 dm^3 O_2

2. Calculate the amount in mol (n) of 54.5 dm^3 of CH_4 at STP.

   1 mol of gas at STP = 22.7 dm^3

   54.5 ÷ 22.7 = 2.40 mol CH_4

Answers:

1.
   a. 0.754 × 22.7 = 17.1 dm^3
   b. 0.754 × 70.9 = 53.5 g Cl_2
   c. 0.754 × 6.02×10^{23} = 4.54×10^{23} Cl_2 molecules
   d. 4.54×10^{23} × 2 = 9.08×10^{23} Cl atoms

2.
   a. $3.01 \times 10^{23} \div 6.02 \times 10^{23}$ = 0.500 mol O_2
   b. 0.500 × 32.00 = 16.0 g O_2
   c. 0.500 × 22.7 = 11.4 dm^3
   d. 3.01×10^{23} × 2 = 6.02×10^{23} oxygen atoms

3.

a. $25.0 \div 28.02 = 0.892$ mol N_2

b. $0.892 \times 22.7 = 20.2$ dm^3 N_2

c. $0.892 \times 6.02 \times 10^{23} = 5.37 \times 10^{23}$ molecules N_2

4.

a. $5.72 \times 22.7 = 1.30 \times 10^2$ dm^3

b. $5.72 \times 6.02 \times 10^{23} = 3.44 \times 10^{24}$ molecules NH_3

c. $3.44 \times 10^{24} \times 3 = 1.03 \times 10^{25}$ atoms H

5. Balanced equation: $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

   $n(Mg) = 3.54 \div 24.31 = 0.146$ mol

   Ratio of Mg to H_2 is 1:1

   $n(H_2) = 0.146$ mol

   Volume of H_2 at STP $= 0.146 \times 22.7 = 3.31$ dm^3

6. Balanced equation: $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

   $n(CaCO_3) = (139 \div 100.09) = 1.39$ mol

   Ratio of $CaCO_3$ to CO_2 is 1:1

   $n(CO_2) = 1.39$ mol

   Volume of CO_2 at STP $= 1.39 \times 22.7 = 31.6$ dm^3

Avogadro's law

- Equal volumes of gases at the same temperature and pressure contain the same number of particles.

At STP (273 K and 100 kPa)

Amount (mol)	**1 mol H_2**	**1 mol N_2**	**1 mol O_2**
Volume (dm^3)	**22.7**	**22.7**	**22.7**
Mass (g)	**2.02**	**28.02**	**32.00**
Number of particles	**6.02×10^{23}**	**6.02×10^{23}**	**6.02×10^{23}**

Example: 40 cm^3 of CO reacts with 40 cm^3 of O_2. What volume of CO_2 is produced? What volume of the excess reactant remains?

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

Molar ratio of CO to O_2 is 2:1

40 cm^3 of CO reacts with 20 cm^3 of O_2 (molar ratio is 2:1)

Volume of O_2 is 40 cm^3, therefore O_2 is excess reactant.

Excess reactant remaining = 40 – 20 = 20 cm^3 of O_2

Exercises:

1. What volume of sulfur trioxide, in cm^3, can be prepared using 40 cm^3 sulfur dioxide and 20 cm^3 oxygen gas by the following reaction? Assume all volumes are measured at the same temperature and pressure.

$$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$$

Ratio of SO_2 to O_2 is 2:1

40 cm^3 of SO_2 would react with 20 cm^3 of O_2 to produce 40 cm^3 of SO_3

2. 5 dm^3 of carbon monoxide, $CO_{(g)}$, and 2 dm^3 of oxygen, $O_{2(g)}$, at the same temperature and pressure are mixed together. What is the maximum volume of carbon dioxide, $CO_{2(g)}$, in dm^3, that can be formed? What volume of the excess reactant remains?

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

Ratio of CO to O_2 is 2:1
5 dm^3 of CO would need 5 ÷ 2 = 2.5 dm^3 of O_2 to react completely.
Therefore, O_2 is the limiting reactant and CO is the excess reactant.
Ratio of O_2 to CO_2 is 1:2, therefore, 2 dm^3 of O_2 would produce 4 dm^3 of CO_2
2 dm^3 of O_2 reacts with 4 dm^3 of CO, 5-4 = 1 dm^3 of CO remains.

3. 100 cm^3 of ethene, C_2H_4, is burned in 400 cm^3 of oxygen, producing carbon dioxide and some liquid water. Some oxygen remains unreacted (excess).

$$C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(g)}$$

Calculate the volume of carbon dioxide produced and the volume of oxygen remaining.
Molar ratio of reactants and products is 1:3:2:2
100 cm^3 of C_2H_4 reacts with 300 cm^3 of O_2 to produce 200 cm^3 of CO_2 and 200 cm^3 of H_2O
Volume of O_2 remaining: 400 – 300 = 100 cm^3

Ideal gas equation

$$PV = nRT$$

***P* is pressure in Pa**

***V* is volume in m^3**

***n* is amount in mol**

***R* is the gas constant (8.31 J K^{-1} mol^{-1})**

***T* is temperature in kelvin (K)**

- The ideal gas equation can be rearranged to calculate amount (in mol), volume (in m^3), pressure (in Pa), temperature (in K) or molar mass (g mol^{-1}).

$$n = \frac{PV}{RT} \qquad V = \frac{nRT}{P} \qquad P = \frac{nRT}{V}$$

$$T = \frac{PV}{nR} \qquad M = \frac{mRT}{PV}$$

Unit conversions

- Temperature in kelvin (K): ^{o}C + 273

  25 ^{o}C = 298 K

- Pressure in Pa: 1.00 × 10^5 Pa = 100 kPa
- 1 cm^3 = 1 × 10^{-3} dm^3 = 1 × 10^{-6} m^3
- 1 m^3 = 1 × 10^3 dm^3 = 1 × 10^6 cm^3
- 1 atm = 101325 Pa

Convert the following quantities:

a. 100 cm^3 to m^3

b. 5 dm^3 to m^3

c. 12 m^3 to cm^3

d. 0 ^{o}C to K

e. 300 K to ^{o}C

f. 34 ^{o}C to K

Exercises:

1. Calculate the volume in (m^3) occupied by one mole of a gas at 25.0 °C and 100.0 kPa.

$$V = \frac{nRT}{P} \quad V = \frac{1 \times 8.31 \times 298}{1.00 \times 10^5} = 0.0248\ m^3$$

2. Calculate the pressure of a gas (in Pa) given that 0.200 moles of the gas occupy 10.0 dm^3 at 20.0 °C.

$$P = \frac{nRT}{V} \quad P = \frac{0.200 \times 8.31 \times 293}{0.0100} = 4.87 \times 10^4\ Pa$$

3. Calculate the amount (in mol) of carbon dioxide which occupies 20.0 dm^3 at 27.0 °C and 100.0 kPa.

$$n = \frac{PV}{RT} \quad n = \frac{1.00 \times 10^5 \times 0.0200}{8.31 \times 300} = 0.802\ mol\ CO_2$$

4. Calculate the molar mass of a gas if a 500.0 cm^3 sample at 20.0 °C and 1.00 atm (101325 Pa) has a mass of 0.666 g.

$$M = \frac{mRT}{PV} \quad M = \frac{0.666 \times 8.31 \times 293}{101325 \times 5.00 \times 10^{-4}} = 32.0\ g\ mol^{-1}$$

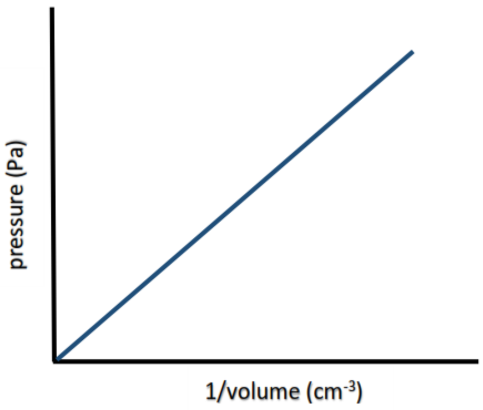
The gas laws

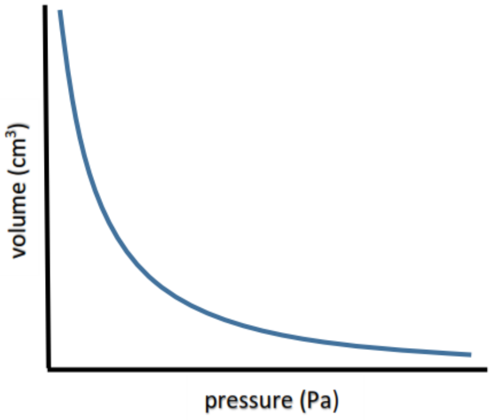
Boyle’s law – the relationship between volume and pressure

- The volume occupied by a gas is inversely proportional to its pressure (at constant temperature).
- If the pressure of a fixed mass of gas is doubled (at constant temperature) then the volume of the gas will halve.

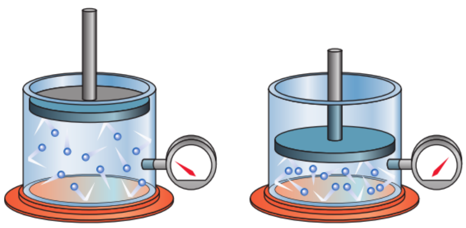
$$PV = k \quad P \propto \frac{1}{V}$$

$$P_1V_1 = P_2V_2$$





Exercise: Explain what happens to the pressure when the volume of the gas in the container is halved.

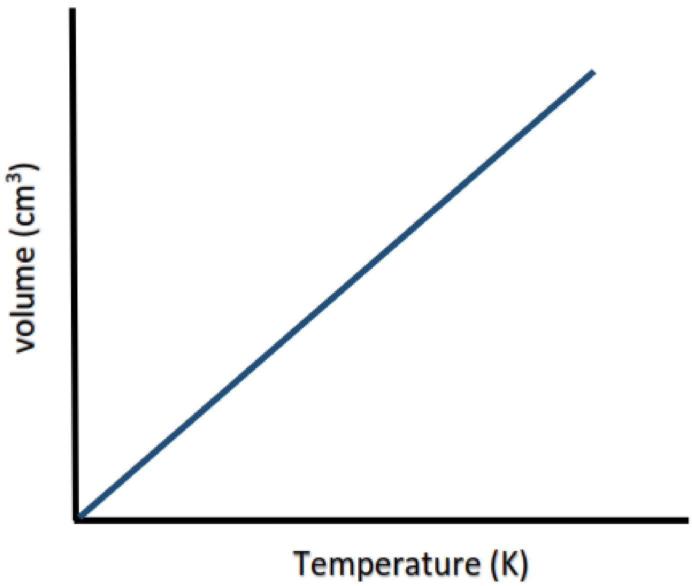


According to Boyle’s law, volume and pressure are inversely proportional at constant temperature. If the volume of the gas is halved, the pressure is doubled.

Charles's law – the relationship between volume and temperature

- The volume occupied by a gas is directly proportional to its absolute temperature (at constant pressure).
- If the temperature of a fixed mass of a gas is doubled, the volume also doubles (at constant pressure).

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad V \propto T \qquad \frac{V}{T} = k$$



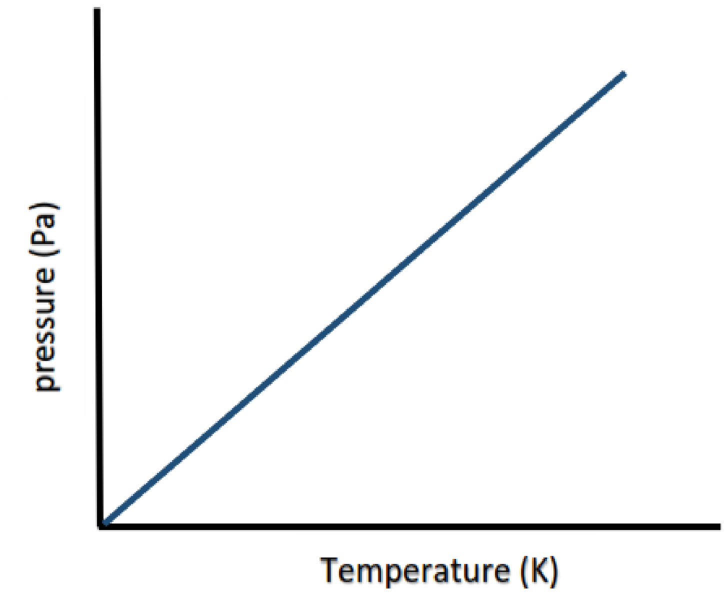
Exercise:
Imagine a balloon filled with a gas. Explain what happens when the balloon is placed into a freezer (at constant pressure).

According to Charles's law, volume and temperature are directly proportional at constant pressure. If the temperature is decreased, the volume also decreases.

Gay Lussac's law – the relationship between temperature and pressure

- The pressure of a gas is directly proportional to its absolute temperature (at constant volume).
- If the temperature of a fixed mass of gas is doubled, the pressure of the gas is also doubled.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \qquad P \propto T \qquad \frac{P}{T} = k$$



Exercise: Explain why the pressure inside a car tyre increases on a hot day.

According to Gay Lussac's law, pressure and temperature are directly proportional at constant volume. If the temperature increases, the pressure also increases.

The combined gas law

$$P \propto T \qquad V \propto T \qquad P \propto \frac{1}{V}$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Example: the molar volume of a gas is 22.7 dm^3 at STP. Calculate the volume occupied by a gas at 25^{o}C.

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

$$V_2 = \frac{100000 \times 22.7 \times 298}{273 \times 100000} = 24.8 \text{ dm}^3$$

- Note that this value is greater than the molar volume of a gas at 273 K (0 ^{o}C).

Additional practice examples

1) What is the final volume if the pressure of 10 dm^3 of gas is doubled at constant temperature?
   5 dm^3

2) The absolute temperature of a gas at 100.0 kPa is doubled at constant volume. What is the new pressure of the gas?
   200.0 kPa

3) The absolute temperature of 150 dm^3 of gas is doubled at constant pressure. What is the new volume of the gas?
   300 dm^3

4) What happens to the volume of a fixed mass of gas when its pressure and its absolute temperature are both doubled?
   Pressure is doubled, volume is halved;
   Temperature is doubled, volume is doubled;
   Volume is halved and then doubled, therefore the volume stays the same.

5) The volume of an ideal gas at 27.0 °C is increased from 3.00 dm^3 to 6.00 dm^3. At what temperature, in °C, will the gas have the original pressure?

   Convert to kelvin (K): 27.0 + 273 = 300 K
   Volume is doubled, pressure is halved;
   Double absolute temperature (to 600 K), pressure is doubled (back to original pressure);
   Convert back to ^oC: 600 – 273 = 327^oC

Solutions

- Solutions are homogeneous mixtures.
- A solution is composed of a solute (usually a solid) dissolved in a solvent (usually water).
- Solutions can be dilute (less solute, more solvent) or concentrated (more solute, less solvent).

Standard solutions

- A standard solution is a solution that has a concentration that is known accurately.
- A primary standard solution is prepared using a substance of high purity which is dissolved in a known volume of solvent.
- A secondary standard solution refers to a solution that has its concentration determined by titration with a primary standard solution.
- When making up a standard solution it is important that the correct mass of substance is accurately measured. In addition, it is important that all the solute is transferred to the volumetric flask used to make up the solution.

Calculating the concentration of a solution

- The concentration of a solution can be expressed in mol dm^{-3} or g dm^{-3}
- The equation for calculating concentration in mol dm^{-3} is shown below.
- In this equation, **volume must be in dm^3** (to convert from cm^3 to dm^3, divide by 1000).

$$c\ (\text{mol dm}^{-3}) = \frac{\text{amount of solute (mol)}}{\text{volume of solution (dm}^3\text{)}}$$

$$c = \frac{n}{V}$$

c = concentration in mol dm^{-3}
n = amount in mol
V = volume in dm^3

$$n = cV \qquad V = \frac{n}{c}$$

- The equations for calculating concentration in g dm^{-3} and ppm are shown below.

$$c\,(\text{g dm}^{-3}) = \frac{\text{mass of solute (g)}}{\text{volume of solution (dm}^3\text{)}}$$

$$c\,(\text{ppm}) = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 10^6$$

Example:

50.0 g of NaCl are dissolved in 100 cm^3 of water which is then made up to 500.0 cm^3 in a volumetric flask. Calculate the concentration of the solution in mol dm^{-3} and g dm^{-3}.

In g dm^{-3}

c = mass of solute ÷ volume of solution

$c = 50.0 \div (500.0 \div 1000)$

c = 100.0 g dm^{-3}

In mol dm^{-3}

Convert from mass to amount in mol

$n = m \div M$

n = 50.0 ÷ 58.44 = 0.856 mol NaCl

$c = n \div V$

$c = 0.856 \div (500 \div 1000)$

c = 1.71 mol dm^{-3}

Exercises:

1. Calculate the concentration (in mol dm^{-3} and g dm^{-3}) of these solutions:

   a. 10.6 g of sodium carbonate (Na_2CO_3) in 1.00 dm^3 of solution.

   $n = m \div M$

   n = 10.6 ÷ 105.99 = 0.100 mol Na_2CO_3

   $c = n \div V$

   $c = 0.100 \div (1.00)$

   c = 0.100 mol dm^{-3}

   $c = 10.6 \div (1.00)$

   c = 10.6 g dm^{-3}

   b. 117 g of sodium chloride (NaCl) in 5.00 dm^3 of solution.

   $n = m \div M$

   n = 117 ÷ 58.44 = 2.00 mol NaCl

   c = n ÷ V

   $c = 2.00 \div 5.00$

   c = 0.400 mol dm^{-3}

   $c = 117 \div 5.00$

   c = 23.4 g dm^{-3}

c. 0.830 g of potassium iodide (KI) in 25.0 cm^3 of solution.

$n = m \div M$
n = 0.830 ÷ 166.00 = 5.00×10^{-3} mol KI
$c = n \div V$
$c = 5.00 \times 10^{-3} \div (25.0 \div 1000)$
$c = 0.200$ mol dm^{-3}

$c = 0.830 \div (25.0 \div 1000)$
$c = 33.2$ g dm^{-3}

2. Calculate the amount (in mol) of solute in each of the following solutions:

a. 0.250 dm^3 of 0.400 mol dm^{-3} ammonium chloride solution.

$n = cV$

$n = 0.400 \times 0.250$

$n = 0.100$ mol

b. 200.0 cm^3 of 0.800 mol dm^{-3} sodium carbonate solution.

$n = cV$

$n = 0.800 \times (200.0 \div 1000)$

$n = 0.160$ mol

c. 300.0 cm^3 of 4.00 mol dm^{-3} sodium hydroxide solution.

$n = cV$

$n = 4.00 \times (300.0 \div 1000)$

$n = 1.20$ mol

3. Calculate the mass of solute in the following solutions:

a. 2.00 dm^3 of 0.200 mol dm^{-3} potassium hydroxide (KOH) solution.

$n = cV$

$n = 0.200 \times 2.00$

$n = 0.400$ mol

$m = nM$

$m = 0.400 \times 56.11$

$m = 22.6$ g KOH

b. $200.0\ cm^3$ of $0.100\ mol\ dm^{-3}$ sodium carbonate (Na_2CO_3) solution.

$n = cV$

$n = 0.100 \times (200.0 \div 1000)$

$n = 0.0200$ mol

$m = nM$

$m = 0.0200 \times 105.99$

$m = 2.12$ g Na_2CO_3

c. $25.0\ cm^3$ of $0.0500\ mol\ dm^{-3}$ copper(II) sulphate ($CuSO_4 \bullet 5H_2O$) solution.

$n = cV$

$n = 0.0500 \times (25.0 \div 1000)$

$n = 1.25 \times 10^{-3}$ mol

$m = nM$

$m = 1.25 \times 10^{-3} \times 249.72$

$m = 0.312$ g $CuSO_4 \bullet 5H_2O$

Parts per million (ppm) exercises

1. 25.0 g of a chemical is dissolved in 75.0 g of water. Calculate the concentration of the solution in ppm.
   ppm = $(25.0 \div 100.0) \times 10^6 = 250000$ ppm

2. 17.0 g of sucrose is dissolved in 183 g of water. Calculate the concentration of the solution in ppm.
   ppm = $(17.0 \div 200) \times 10^6 = 85000$ ppm

3. 35.0 g of ethanol is dissolved in 115 g of water. Calculate the concentration of the solution in ppm.
   ppm = $(35.0 \div 140) \times 10^6 = 233333$ ppm

4. The solubility of NaCl is 284 g per 100.0 g of water. Calculate the concentration of the solution in ppm.
   ppm = $(284 \div 384) \times 10^6 = 7400000$ ppm

5. The solubility of AgCl is 0.008 g per 100.0 g of water. Calculate the concentration of the solution in ppm.
   ppm = $(0.008 \div 100.008) \times 10^6 = 80$ ppm

More practice examples

1. What volume of 2.00 mol dm^{-3} HCl reacts completely with 5.00 g of magnesium? What volume of hydrogen gas will be produced at STP?

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

n (Mg) = 5.00 ÷ 24.31 = 0.206 mol
Ratio of Mg to HCl is 1:2
0.206 mol of Mg reacts with 2 × 0.206 = 0.412 mol of HCl
$V = n \div c$
$V = 0.412 \div 2.00$
V = 0.206 dm^3 (206 cm^3) HCl

Ratio of Mg to H_2 is 1:1
0.206 mol of Mg produces 0.206 mol of H_2
1 mol of gas at STP = 22.7 dm^3
0.206 × 22.7 = 4.68 dm^3 H_2

2. What volume of 1.00 mol dm^{-3} HCl reacts completely with 10.00 g of calcium carbonate ($CaCO_3$)? What volume of carbon dioxide gas will be produced at STP?

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

n ($CaCO_3$) = 10.00 ÷ 100.09 = 0.100 mol
Ratio of $CaCO_3$ to HCl is 1:2
0.100 mol of $CaCO_3$ reacts with 2 × 0.100 = 0.200 mol of HCl
$V = n \div c$
$V = 0.200 \div 1.00$
V = 0.200 dm^3 (200 cm^3) HCl

Ratio of $CaCO_3$ to CO_2 is 1:1
0.100 mol of $CaCO_3$ produces 0.100 mol of CO_2
1 mol of gas at STP = 22.7 dm^3
0.100 × 22.7 = 2.27 dm^3 CO_2

3. Sodium hydroxide (NaOH) reacts with hydrochloric acid (HCl) according to the following equation:

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

What volume of 0.500 mol dm^{-3} HCl reacts with 25.0 cm^3 of 2.00 mol dm^{-3} NaOH solution?

n (NaOH) = 2.00 × (25.0 ÷ 1000) = 0.0500 mol
Ratio of NaOH to HCl is 1:1
0.0500 mol of NaOH reacts with 0.0500 mol of HCl
$V = n \div c$
$V = 0.0500 \div 0.500$
$V = 0.100\ dm^3$ (100 cm^3) HCl