Structure 1.5 Answers

IB CHEMISTRY SL

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

Structure 1.5.1 and 1.5.2

Understandings:

- An ideal gas consists of moving particles with negligible volume and no intermolecular forces. All collisions between particles are considered elastic.
- Real gases deviate from the ideal gas model, particularly at low temperature and high pressure.

Learning outcomes:

- Recognize the key assumptions in the ideal gas model.
- Explain the limitations of the ideal gas model.

Ideal gas behavior

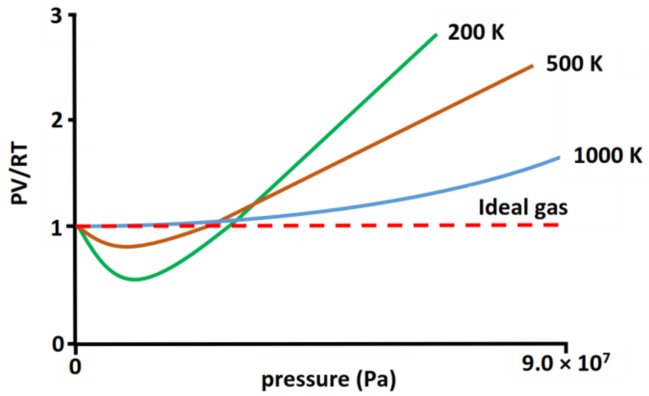
- An ideal gas is one which abides by the kinetic molecular theory (and the gas laws).
- According to the kinetic molecular (KMT) theory of gases:

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.

Structure 1.5.3

Understandings:

- The molar volume of an ideal gas is a constant at a specific temperature and pressure.

Learning outcomes:

- Investigate the relationship between temperature, pressure and volume for a fixed mass of an ideal gas and analyse graphs relating these variables.

Additional notes:

- The names of specific gas laws will not be assessed.
- The value for the molar volume of an ideal gas under standard temperature and pressure (STP) is given in the data booklet.

Linking questions:

- Reactivity 2.2 Graphs can be presented as sketches or as accurately plotted data points. What are the advantages and limitations of each representation?

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 dm^3 *or* 0.0227 m^3

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

$$V\ (\mathrm{dm}^3) = n\ (\mathrm{mol}) \times V_m\ (22.7\ \mathrm{dm}^3) \qquad V = n \times 22.7$$

$$n\ (\mathrm{mol}) = \frac{V\ (\mathrm{dm}^3)}{V_m\ (22.7\ \mathrm{dm}^3)} \qquad n = \frac{V}{22.7}$$

Exercises

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 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

3. A sample of gas at STP contains 0.754 mol of Cl_2. Calculate the following:

a. the volume occupied by the gas 0.754 × 22.7 = 17.1 dm^3
b. the mass of Cl_2 present 0.754 × 70.9 = 53.5 g Cl_2
c. the number of Cl_2 molecules in the sample of gas $0.754 \times 6.02 \times 10^{23} = 4.54 \times 10^{23}$ Cl_2 molecules
d. the number of Cl atoms present in the sample $4.54 \times 10^{23} \times 2 = 9.08 \times 10^{23}$ Cl atoms

4. A sample of O_2 gas at STP contains 3.01×10^{23} molecules. Calculate the following:

a. the amount of O_2 in mol $3.01 \times 10^{23} \div 6.02 \times 10^{23} = 0.500$ mol O_2
b. the mass of O_2 present $0.500 \times 32.00 = 16.0$ g O_2
c. the volume occupied by the gas $0.500 \times 22.7 = 11.4$ dm^3
d. the number of oxygen atoms present in the sample $3.01 \times 10^{23} \times 2 = 6.02 \times 10^{23}$ oxygen atoms

5. A sample of N_2 gas at STP has a mass of 25.0 g. Calculate the following:

a. the amount of N_2 in mol

$25.0 \div 28.02 = 0.892$ mol N_2

b. the volume occupied by the gas

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

c. the number of nitrogen molecules present in the sample

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

6. A sample of gas at STP contains 5.72 mol of NH_3. Calculate the following:

a. the volume occupied by the gas

$5.72 \times 22.7 = 1.30 \times 10^{2}$ dm^3

b. the number of NH_3 molecules present in the sample

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

c. the number of hydrogen atoms present in the sample

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

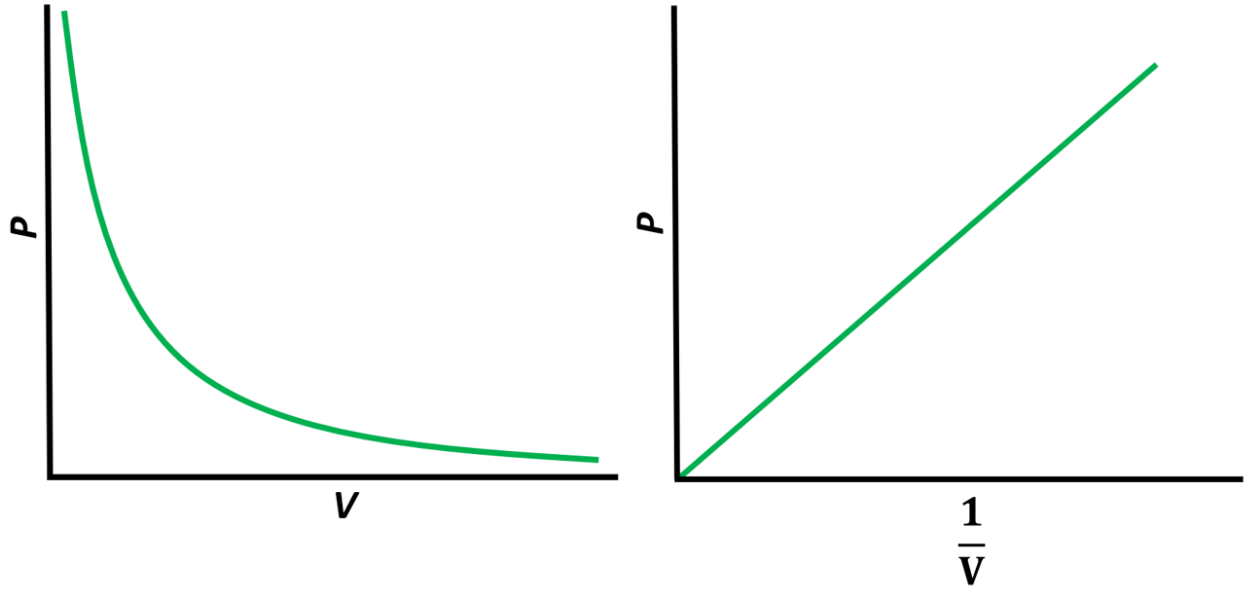
The gas laws

Boyle's law – the relationship between volume and pressure at constant temperature

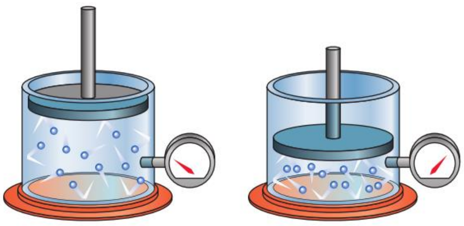
- The volume occupied by a gas is inversely proportional to its pressure (at constant temperature).
- If the pressure of a fixed mass of gas is doubled (at constant temperature) then the volume of the gas will halve.

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

$$P_1V_1 = P_2V_2$$



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

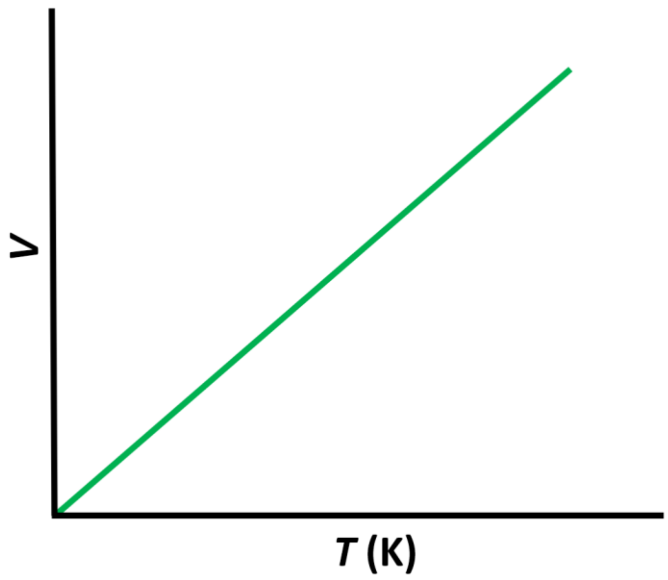


According to Boyle's law, volume and pressure are inversely proportional at constant temperature. If the volume of the gas is halved, the pressure is doubled.

Charles's law – the relationship between volume and temperature at constant pressure

- 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).

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



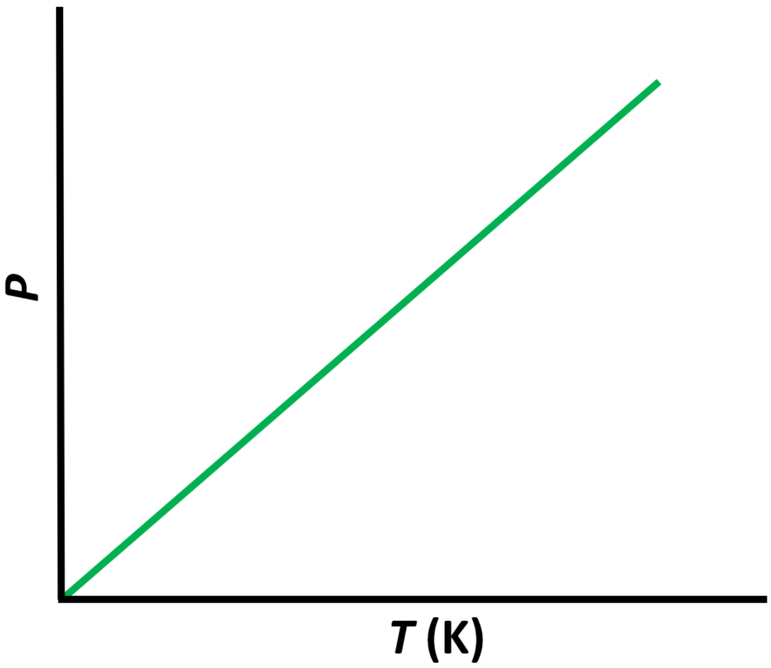
Exercise: Imagine a balloon filled with gas. Explain what happens when the balloon is placed into a freezer (at constant pressure).

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 at constant volume

- 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.

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



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.

Exercises:

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

The combined gas law

- The Combined gas law combines Boyle's law, Charles's law and Gay-Lussac's law.

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

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.

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

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

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

Exercise: 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

a. 1×10^{-4} m^3 **b.** 5×10^{-3} m^3 **c.** 1.2×10^6 cm^3

d. 273 K **e.** 573 ^{o}C **f.** 307 K

Exercises:

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

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

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 ^oC.

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

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

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

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

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