Physics of gases

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Abstract

This article will explore the physical principles of gases relevant to anaesthetic practice, including the kinetic theory of gases, phase transition, the application of the gas laws and factors affecting flow. It will also explain the Bernoulli principle and the Coanda effect.

Keywords Adiabatic change; Bernoulli principle; critical pressure; critical temperature; gas flow; gas laws; phase transition

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Learning Objectives

After reading this article, you should be able to:

- explore the molecular theory of states of matter and transition between solids, liquids and gases
- define latent heat, critical temperature and critical pressure
- explain the gas laws and illustrate them in graphical form
- understand the characteristics of laminar and turbulent flow patterns and the factors affecting flow rate
- explain the Bernoulli principle, Venturi effect and Coanda effect

States of matter

All substances are made up of atoms or molecules. The commonly described states, or phases, of matter are solids, liquids and gases. A phase is determined by the bonds between the molecules, and a transition from one phase to another requires a transfer of energy. A solid is composed of a lattice of molecules, held together in a fixed shape by strong forces or bonds. As heat energy is applied to the molecules, they vibrate and move apart, exerting a weaker force on neighbouring molecules. The lattice structure breaks down and the substance changes into a liquid. The heat energy required to break these bonds, changing the substance from a solid to a liquid, is known as the latent heat of fusion (Figure 1). There are still weak attractive forces, known as van der Waals forces, between the molecules. A liquid occupies a finite volume, but does not have a defined shape. If further heat energy is added to the liquid, the molecules gain kinetic energy which overcomes the van der Waals forces, allowing molecules to escape from the liquid to become a gas. In a gas, the molecules move freely and have no finite shape or volume, but will fill the space in which they are contained. The heat energy required to change a liquid into a gas is referred to as the latent heat of vaporization (Figure 1). It is important to note that when a change in phase occurs, the temperature of the substance does not increase, as the absorbed heat energy is transferred to break the bonds.

Gases and vapours

A change in phase can also occur due to a change in pressure. Below a specific temperature, a substance in a gaseous form can be compressed into a liquid. As the temperature of a substance rises, the molecules have more kinetic energy and so an increasing amount of pressure is required to liquefy it. The critical temperature is the temperature above which a substance can no longer be liquefied by the application of pressure alone. Above its critical temperature, a substance is known as a gas; at or below this temperature, it is known as a vapour, which exists in equilibrium with the liquid phase. The critical pressure is the pressure required to liquefy a vapour at its critical temperature.

Saturated vapour pressure

When a liquid in a closed container is heated, molecules will escape from the surface of the liquid (evaporate) to form a vapour. Molecules in the vapour phase will also condense, returning to the liquid phase. When the rate of evaporation is equal to the rate of condensation, the vapour is 'saturated', and the pressure exerted by the molecules in the vapour phase on the wall of the container is the saturated vapour pressure (SVP). The SVP depends on what the liquid is, and it increases in a non-linear relationship with increasing temperature. When the SVP reaches atmospheric pressure, the liquid completely evaporates, and this is referred to as the 'boiling point'.

The 'ideal' gas

The perfect or 'ideal' gas is described by the following concepts:

- It contains a large number of particles, but these are so widely spread that the volumes of the particles themselves are negligible.
- The particles move at random speeds, in random directions, with no attractive or repulsive forces acting between them.
- Collisions between the particles or with the wall of a container are perfectly elastic, so that no kinetic energy is transferred, and the particles travel in straight lines between collisions.

A perfect gas does not actually exist; hydrogen is the closest to this ideal, as it has the lowest molecular weight. A perfect gas will also obey the three gas laws. The anaesthetic vapours obey these laws closely.

The gas laws

The gas laws describe the relationship between volume, pressure and temperature of gases (Figure 2). They can be explained by

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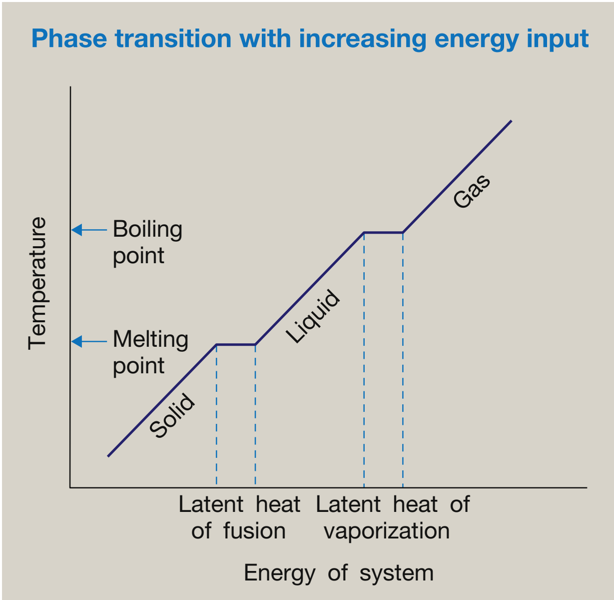

Figure 1

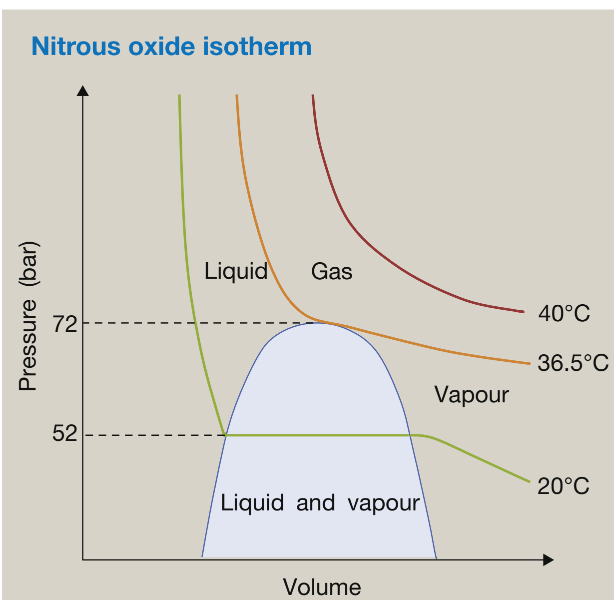

Figure 2

imagining a gas with a fixed number of particles within a closed container. The particles move freely within the container; the kinetic energy of these particles and therefore the frequency of collisions is determined by the temperature. The force of the molecules colliding over the area of the container wall is defined as pressure.

Boyle's law

$$PV = K$$

If the same number of particles, with the same energy of motion, are contained within a smaller volume, they will collide with the walls of the container more frequently, and thus exert a greater pressure. At a constant temperature, the volume of a given mass of gas is inversely proportional to its absolute pressure (Figure 3a).

Charles' law

$$V/T = K$$

As the molecules of the gas are heated, they will have more kinetic energy and thus move further apart and collide more frequently. This explains why a gas 'expands' on heating. If the absolute temperature of the container is doubled, the volume of the container must also be doubled to keep the number of collisions, or pressure, constant. At a constant pressure, the volume of a fixed mass of gas is directly proportional to its absolute temperature (Figure 3b).

Gay Lussac's law (The Third Gas Law)

$$P/T = K$$

By the same notion, if the volume of the container is constant, the pressure exerted by the particles will increase as the temperature is increased. If the volume of a fixed mass of gas is constant, absolute temperature and pressure are directly proportional (Figure 3c).

Avogadro's hypothesis

Avogadro's hypothesis states that equal volumes of any gas at the same temperature and pressure contain the same number of molecules. One mole (the SI unit of measurement for amount of substance) of any gas, which by definition contains 6.023×1023 molecules, will occupy 22.4 litres at standard temperature and pressure (273.15K and 101.325kPa).

Universal gas law

If the three gas laws are combined:

$$PV/T = \text{constant}$$

For one mole of gas, this constant is known as the universal gas constant, R (equal to 8.31 J/K/mol). From this, a generally applied formula can be applied, where n is the number of moles of gas.

$$PV = nRT$$

Dalton's law

Dalton's law states that in a mixture of gases, each gas exerts the same pressure as it would if it were alone in the container. The total pressure exerted by the mixture of gases is the sum of the partial pressures of the component gases.

$$P_{tot} = P_1 + P_2 + \ldots + P_n$$

Adiabatic change

According to the gas laws, heat energy must be added to or removed from the gas to change its volume or pressure. If a gas is

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