

Physics of Gases and Fluids: A Clinical Perspective

2011 Edition

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OUTLINE

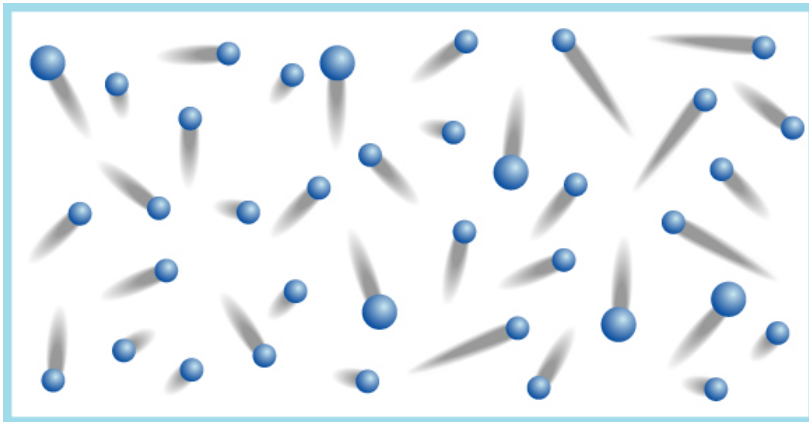
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I. GOALS AND OBJECTIVES

The principal goals and objective of this presentation are to understand the following concepts and their application to anesthesiology and clinical medicine:

- The kinetic theory of gases
- Pressure and its measurement
- Units of pressure measurement and their conversion
- Vapor pressure
- The ideal gas law
- Boyle's law
- Charles' Law
- Dalton's law
- Henry's law
- Fick's law of diffusion
- The van der Waals effect
- Laplace's law
- The Bernoulli effect
- Venturi devices
- Relationship between pressure, flow and resistance
- Poiseuille's law
- Laminar versus turbulent gas flow
- Helium / oxygen mixtures
- Flowmeter design

I. THE KINETIC THEORY OF GASES



On the microscopic scale, gases are modeled by the **kinetic theory of gases**. This model assumes that the molecules have very small sizes relative to the distance between them and that the molecules are in constant, random motion related to their **kinetic energy**, which is given by $e = \frac{1}{2} mv^2$, where m is the mass of the molecules and v is its

velocity. The molecules frequently collide with each other and with the walls of the container holding the gas molecules. Like all molecules, gas molecules have physical properties of mass, velocity, momentum, and energy. At the macroscopic level these properties are related to properties of density, pressure, and temperature. In the case of temperature, the temperature of a gas is related to the mean kinetic energy of the gas; the higher the temperature, the greater the molecular motion. Note that at Absolute Zero the molecules still move a bit (“zero-point energy”); to do otherwise would be in violation of the Heisenberg uncertainty principle.

II. PRESSURE

Pressure is force per unit area, and in the case of gases is the result of the kinetic properties of gas molecules. Discussions about physics in anesthesia may be confusing because of a variety of pressure units used in the clinical literature:

- atmospheres
- mm Hg
- cm H₂O
- pascals
- PSI (pounds per square inch)
- dynes/cm²

The data below provides conversion information:

$$1 \text{ N/m}^2 = 10 \text{ dynes/cm}^2 = 1 \text{ Pa} = 0.00751 \text{ mm Hg}$$

$$1 \text{ atmosphere} = 1000 \text{ millibars} = 760 \text{ mm Hg}$$

$$1 \text{ cm H}_2\text{O} = 0.736 \text{ mm Hg}$$

$$1 \text{ Pascal (Pa)} = 0.00751 \text{ mm Hg}$$

$$1 \text{ PSI} = 51.72 \text{ mm Hg}$$

$$1 \text{ dyne/cm}^2 = 0.1 \text{ Pa} = 0.000751 \text{ mm Hg}$$

American engineers often use PSI units. Scientists often use Pascal units (SI). Clinicians use mm Hg for blood pressures and cm H₂O for airway pressures and PEEP levels (and sometimes for CVP reports).

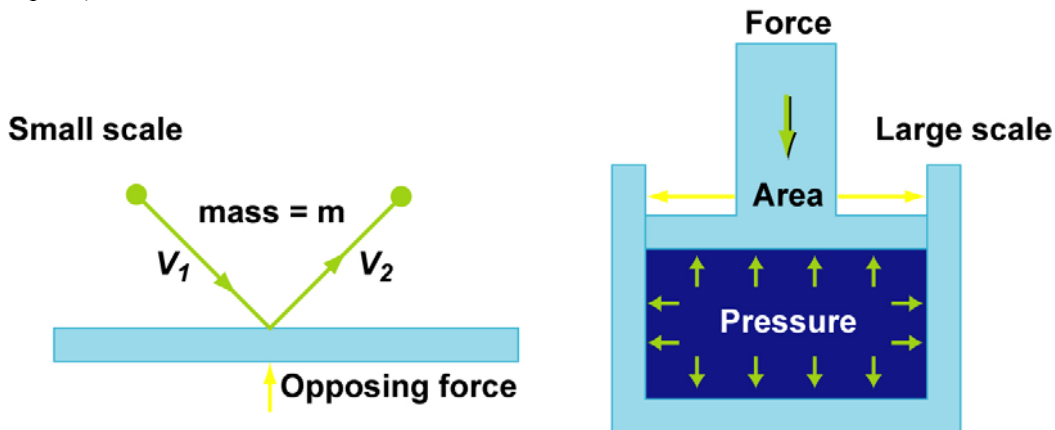


Figure 1: The concept of pressure illustrated on microscopic (left) and macroscopic (right) scales.

III. TEMPERATURE

There are three commonly used temperature scales. Americans and Europeans use the Fahrenheit and the Celsius scales respectively, with temperatures measured in degrees Fahrenheit ($^{\circ}\text{F}$) or degrees Celsius ($^{\circ}\text{C}$). On the Fahrenheit scale, water freezes at a temperature of 32°F and boils at 212°F (at one atmosphere pressure). Absolute zero on this scale is not at 0°F , but at -459°F . Using the Celsius scale the freezing point of water is 0°C and the boiling point is 100°C , while, Absolute Zero corresponds to a temperature of $-273.15^{\circ}\text{Celsius}$.

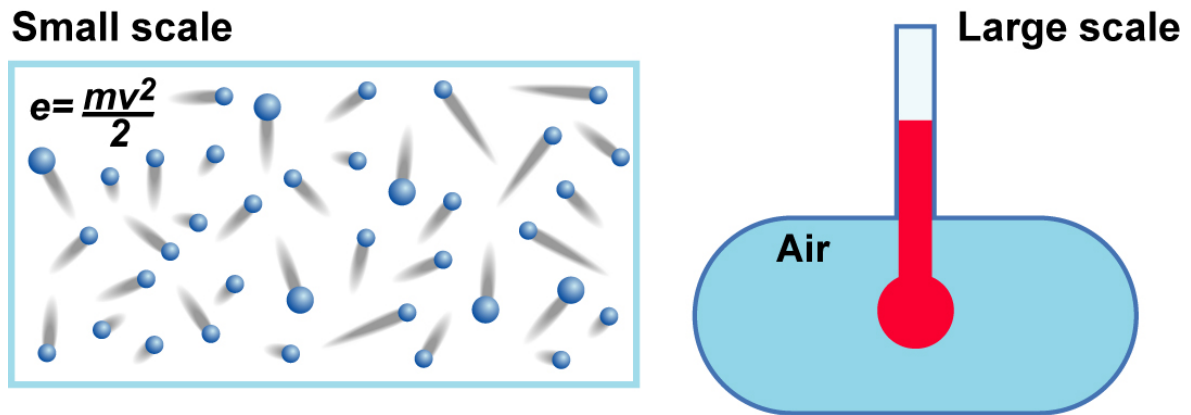


Figure 2: The concept of temperature illustrated on microscopic and macroscopic scales. Since the element mercury is liquid in the temperature range of -38.9°C to 356.7°C it is very useful in thermometer design. While a liquid, mercury expands linearly as it gets warmer.

<p>$^{\circ}\text{C} = 5/9 (^{\circ}\text{F} - 32)$</p> <p>$^{\circ}\text{F} = (9/5) ^{\circ}\text{C} + 32$</p> <p>$^{\circ}\text{C} = ^{\circ}\text{K} - 273.15$</p> <p>$^{\circ}\text{K} = ^{\circ}\text{C} + 273.15$</p> <p>Zero degrees Kelvin is Absolute Zero = -273.15°C</p> <p>Examples: $80^{\circ}\text{C} = 176^{\circ}\text{F}$</p> <p>$14^{\circ}\text{F} = -10^{\circ}\text{C}$</p> <p>Need more information? Visit http://www.temperatureworld.com/</p>
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Figure 3: Temperature conversions can easily be performed using the formulas shown above.

IV. GAS PHYSICS

A. IDEAL GASES

Air is a fluid. Understanding the fundamentals of basic fluid mechanics is essential in grasping the concepts of airway flow. Since air is also a gas, it is important to understand the laws that govern its gaseous behavior. Gases are usually described in terms of pressure, volume, and temperature. Pressure is most often quantified clinically in terms of mm Hg (or torr), volume in ml, and temperature in degrees Celsius. Perhaps the most important law of gas flow in airways is the Ideal (or Perfect) Gas Law, which can be written as:

$$PV = nRT \quad (1)$$

where

P = pressure of gas (Pascals or mm Hg)

V = volume of gas (m^3 or cm^3 or ml)

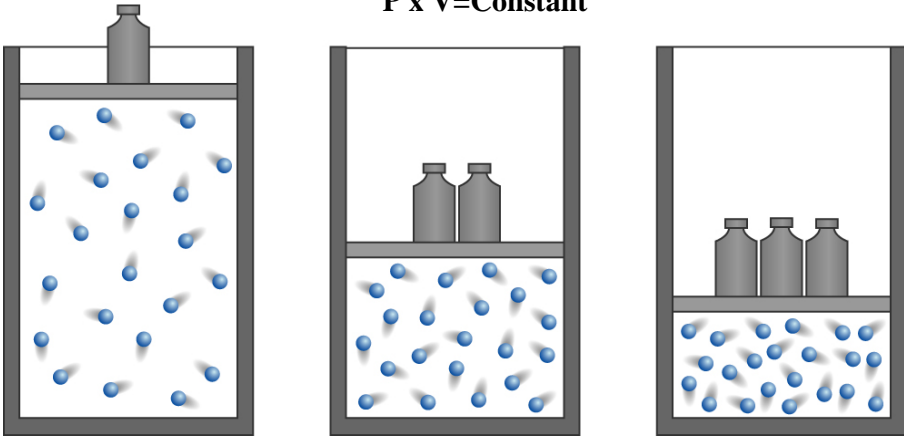
n = number of moles of the gas in volume V

R = gas constant ($8.3143 \text{ J g-mol}^{-1} \text{ K}^{-1}$, assuming P in Pascals, V in m^3)

T = absolute temperature (in Kelvins or K, $273.16 \text{ K} = 0^\circ\text{C}$)

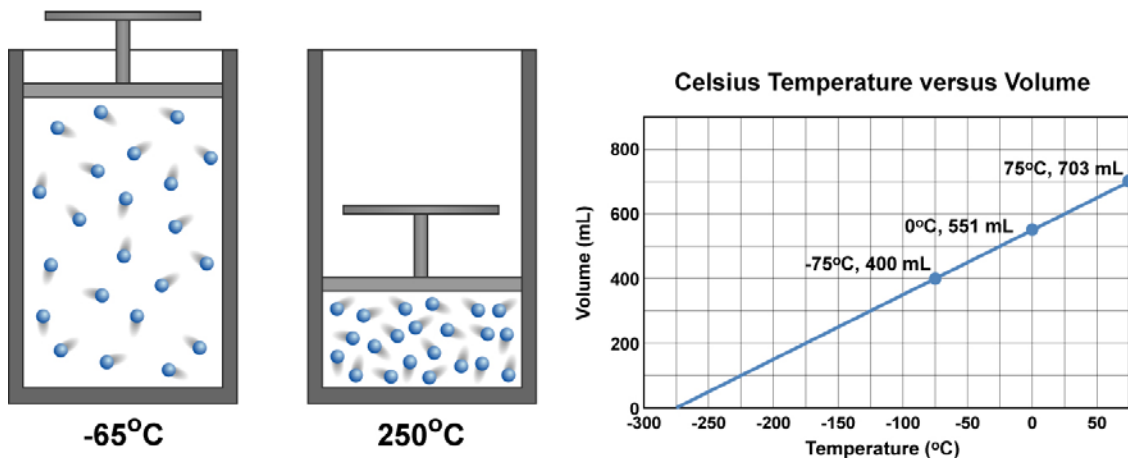
A mole of gas contains 6.023×10^{23} molecules, and this quantity is represented as **Avogadro's number**. One mole of an ideal gas takes up 22.4138 liters at standard temperature and pressure (STP) (STP: 273.16 K at 1 atmosphere [760 mm Hg]). Avogadro also stated that equal volumes of all ideal gases at the same temperature and pressure contain the same number of molecules.

Boyle's Law
 $P \times V = \text{Constant}$



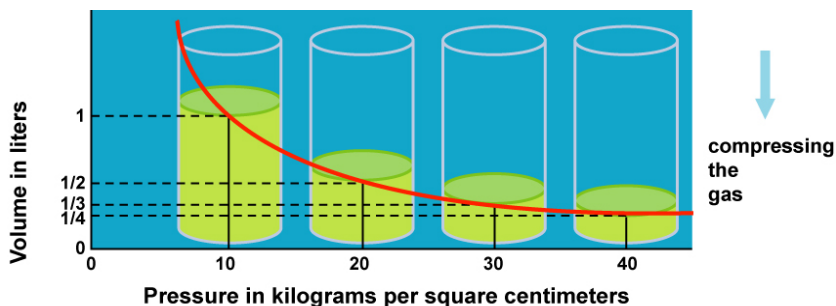
The Ideal Gas Law incorporates the laws of Boyle and Charles. **Boyle's law** states that, at a constant temperature, the product of pressure and volume is equal to a constant, that is, $P \times V = \text{Constant}$ (at constant T). Hence, P is proportional to $1/V$. Gases do not obey Boyle's law at temperatures approaching their point of liquefaction (the state where the gas becomes a liquid). Boyle's law concerns perfect gases and is not obeyed by real gases over a wide range of pressures (see the next section for non-ideal gases). However, at infinitely low pressures, all gases obey Boyle's law. Boyle's law does not apply to anesthetic gases and many other gases because of the van der Waals attraction between molecules (i.e., they are non-ideal gases).

Charles' Law

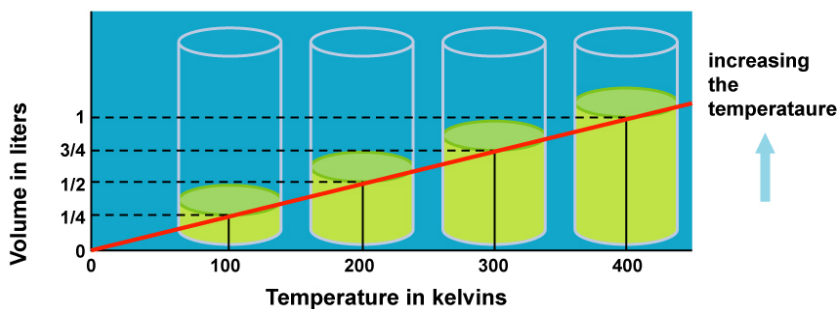


Charles' law states that, at a constant pressure, volume is proportional to temperature, that is, V is proportional to T (at constant P). **Gay-Lussac's law** states that, at a constant volume, pressure is proportional to temperature, that is, $P \propto T$ (at constant V). When a gas obeys both Charles' law and Boyle's law, it is said to be an ideal gas and obeys the Ideal Gas Law.

Boyle's Law



Charles' Law



In clinical situations, gases are typically mixtures of several "pure" gases. Quantifiable properties of mixtures may be determined using **Dalton's law of partial pressures**. Dalton's law states that the pressure exerted by a mixture of gases is the sum of the individual pressures exerted by each "pure" gas:

$$P_{\text{total}} = P_A + P_B + P_C + \dots + P_N \quad (2)$$

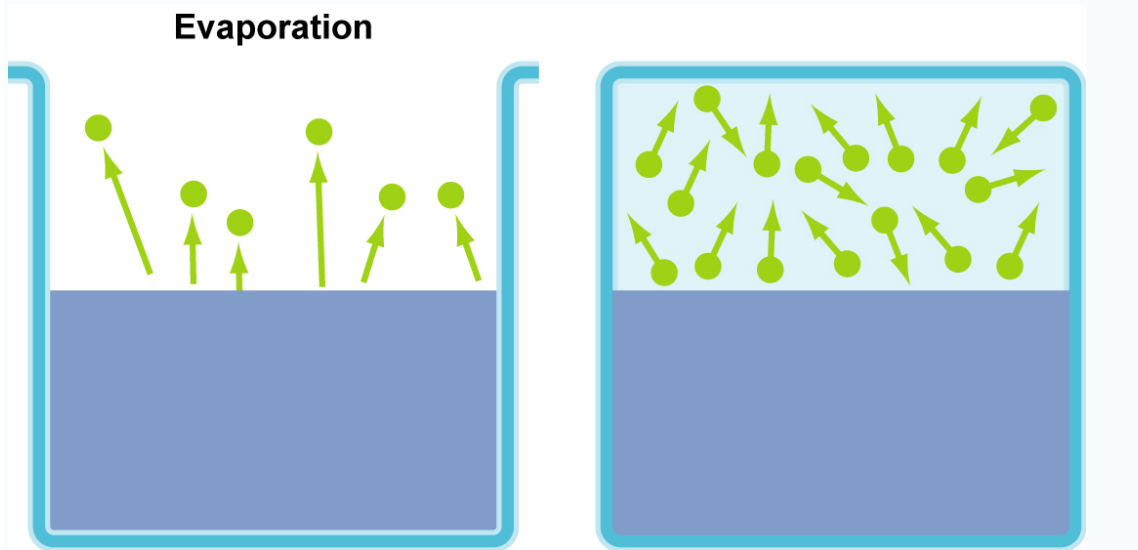
where P_A , P_B , and P_C are the partial pressures of "pure" ideal gases.

Vapor pressure describes the partial pressure contributed by various liquids. When a liquid's vapor pressure equals atmospheric pressure, a liquid boils. The concept of vapor pressure is important in the design of anesthetic vaporizers. The vapor pressure of a liquid increases with its temperature.

Vapor Pressure (mmHg) at 20°C

Halothane	243
Isoflurane	239
Enflurane	175
Methoxyflurane	23

Evaporation is the process whereby atoms or molecules in a liquid state gain sufficient kinetic energy to overcome the surface tension and become a gas. This process is important in understanding how anesthetic vaporizers work. The figure below shows how evaporation of gases to the atmosphere can occur in an open system (left panel), but does not occur with a closed container (right panel).



B. NONIDEAL GASES: THE VAN DER WAALS EFFECT

Ideal gases have no forces of interaction. Real gases, however, have intramolecular attraction, which requires that the pressure-volume gas law be written as:

$$\left[P + \frac{an^2}{V^2} \right] (V - nb) = nRT$$

where

P = pressure of gas (Pascals or mm Hg)

V = volume of gas (m³ or cm³ or ml)

n = number of moles of the gas in volume V

R = gas constant (8.3143 J g·mol⁻¹ K⁻¹ assuming P in Pascals, V in m³)

T = absolute temperature (K)

a,b = physical constants for a given gas

The terms *a* and *b* for a given gas may be found in physical chemistry textbooks and other sources. This formulation, provided by van der Waals, accounts for intramolecular forces fairly well.

C. DIFFUSION OF GASES

Clinically, diffusing of gases through a membrane is most applicable to gas flow across lung and placental membranes. The most commonly used relation to govern diffusion is **Fick's first law of diffusion**, which states that the rate of diffusion of a gas across a barrier is proportional to the concentration gradient for the gas. Fick's law may be expressed mathematically as:

$$\text{Flux} = -D \frac{\Delta C}{\Delta X} \quad (4)$$

where the Flux is the number of molecules/cm²/s crossing the membrane, ΔC is the concentration gradient (molecules/cm³), ΔX is the diffusion distance (cm), and D is the diffusion coefficient (cm²/s) whose value is generally inversely proportional to the gas's molecular weight as well as intrinsic properties of the membrane.

Since gases partially dissolve when they come into contact with liquid, **Henry's law** becomes important in some instances. It states that the mass of a gas dissolved in a given amount of liquid is proportional to the pressure of the gas at constant temperature. That is, gas concentration (in solvent) = constant \times pressure (at constant T).

D. PRESSURE, FLOW, AND RESISTANCE

The laws of fluid mechanics dictate an intricate relationship among pressure, flow, and resistance. Pressure is defined as a force per unit area and, as mentioned previously, is usually measured clinically as mm Hg or cm H₂O. However, it is most commonly measured scientifically in Pascals (newtons force per square meter).

Flow (or the rate of flow) is equal to the change in pressure (pressure drop or pressure difference) divided by the resistance experienced by the fluid. For example, if the flow is 100 ml/s at a pressure difference of 100 mm Hg, the resistance is 100 mm Hg/100 ml/s = 1 mm Hg/ml/s. In *laminar* flow systems only, the resistance is constant, independent of the flow rate.

An important relation which quantifies the relationship of pressure, flow, and resistance in laminar flow systems, is given by the Hagen-Poiseuille (or simply, Poiseuille) equation. Poiseuille's law states that the fluid flow rate through a horizontal straight tube of uniform bore is proportional to the pressure gradient and the fourth power of the radius and is related inversely to the viscosity of the gas and the length of the tube. This law, *which is valid for laminar flow only*, may thus be stated as:

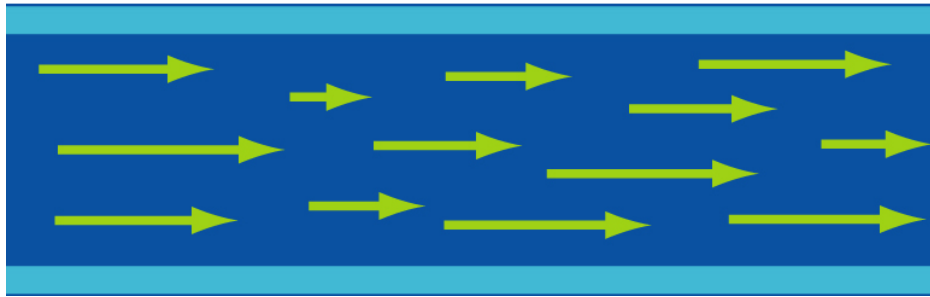
$$\Delta P = \frac{8\mu L}{\pi r^4} \times \text{Flow} \quad (5)$$

where μ is the fluid viscosity (poise [g/cm · s]) and L is the length of the tube (cm). Note that this relationship is somewhat similar to Ohm's law for electricity:

$$E \text{ (voltage)} = R \text{ (resistance to flow)} \times I \text{ (electron flow or current)} \quad (6)$$

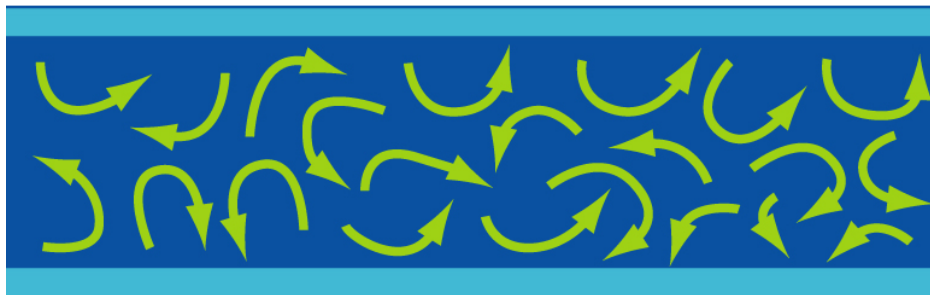
When the flow rate exceeds a *critical velocity* (the flow velocity below which flow is laminar), the flow loses its laminar parabolic velocity profile, becomes disorderly, and is termed *turbulent* (). When turbulent flow exists, the relationship between pressure drop and flow is no longer governed by the Hagen-Poiseuille equation. Instead, the pressure gradient required (or the resistance encountered) during turbulent flow varies as the square of the flow rate.

Laminar Flow

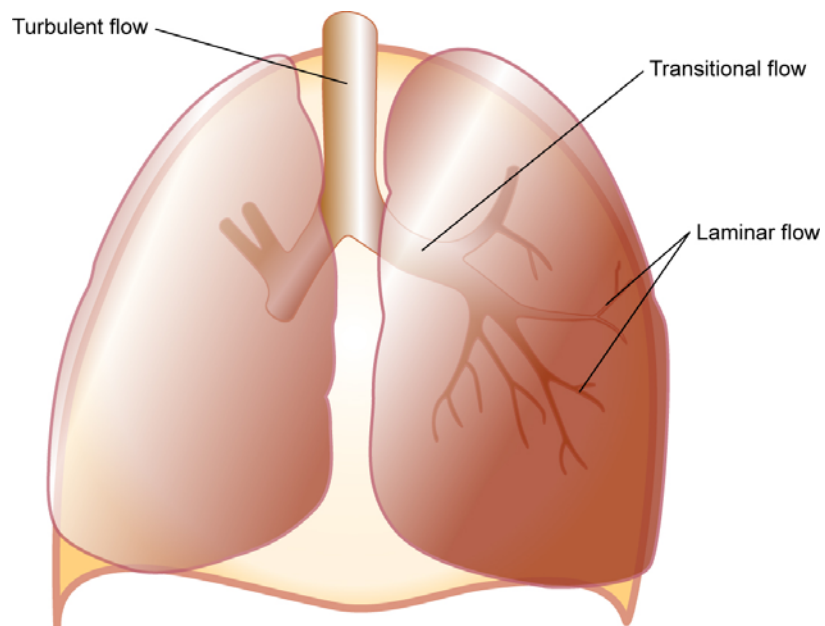


Laminar Flow. When flow is low velocity and through long narrow tubes, it tends to be more orderly and streamlined and to flow in a straight line. This type of flow is called laminar flow. Laminar flow is directly proportional to the driving pressure. During quiet breathing, laminar flow exists from the medium-sized bronchi down to the level of the bronchioles. During exercise, when the air flow is more rapid, laminar flow may be confined to the smallest airways.

Turbulent Flow



Turbulent Flow. When air flows at higher velocities, especially through an airway with irregular walls, flow is generally disorganized, even chaotic, and tends to form eddies. This is called turbulent flow, and is found mainly in the largest airways, like the trachea.



Locations of flow patterns

E. FLOW THROUGH AN ORIFICE

Flow through an orifice (defined as involving flow through a tube whose length is smaller than its radius) is always somewhat turbulent. Clinically, airway obstructing conditions such as epiglottitis or swallowed obstructions are often best viewed as breathing through an orifice. Under such conditions, the approximate flow across the orifice varies inversely with the square root of the gas **density**. This is in contrast to laminar flow conditions, where gas flow varies inversely with gas **viscosity**. The viscosity values for helium and oxygen are similar, but their densities are very different (Table 1). Table 2 provides useful data to allow comparison of gas flow rates through an orifice.

1. Helium-oxygen mixtures

The low density of helium allows it to play a significant clinical role in the management of some forms of airway obstruction. The usual available mixtures of helium and oxygen are 20% O₂: 80% He and 30% O₂: 70% He and are usually administered by a nonrebreathing face mask in patients who face an increased work-of-breathing effort because of the presence of airway pathology (e.g., edema) but in whom endotracheal intubation is preferably withheld at this time.

Table 1. Viscosity and density differences of anesthetic gases

	Viscosity @ 300 K	Density @ 20° C
Air	18.6 $\mu\text{Pa} \times \text{s}$	1.293 g/L
Nitrogen	17.9 $\mu\text{Pa} \times \text{s}$	1.250 g/L
Nitrous oxide	15.0 $\mu\text{Pa} \times \text{s}$	1.965 g/L
Helium	20.0 $\mu\text{Pa} \times \text{s}$	0.178 g/L
Oxygen	20.8 $\mu\text{Pa} \times \text{s}$	1.429 g/L

Table 2. Gas flow rates through an orifice

	%	Density (g/L)	Relative flow
Air	100	1.293	1.0
Oxygen	100	1.429	0.96
Helium (He)	100	0.179	2.68
He-oxygen	20/80	1.178	1.048
He-oxygen	60/40	0.678	1.381
He-oxygen	80/20	0.429	1.73

From Rudow M, Hill AB, Thompson NW et al: *Can Anaesth Soc J* 33:498, 1986.

2. Clinical vignette

Rudow et al. provide the following story (*Rudow M, Hill AB, Thompson NW et al: Helium-oxygen mixtures in airway obstruction due to thyroid carcinoma, Canadian Anaesthetists' Society Journal 33:498, 1986.*) A 78-year-old woman with both breast cancer and ophthalmic melanoma developed airway obstruction from a thyroid carcinoma that extended into her mediastinum and compressed her trachea. She had a 2-month history of worsening dyspnea, especially when positioned supine. On examination, inspiratory and expiratory stridor was present. Noted on the chest x-ray were a large superior mediastinal mass and pulmonary metastases. A solid mass was identified on a thyroid ultrasound scan. Computed tomography revealed a large mass at the thoracic inlet and extending caudally. Clinically, the patient was exhausted and in respiratory distress.

Almost instant relief was obtained by giving the patient a mixture of 78% He: 22% O₂, with improvements in measured tidal volume and oxygenation. Later, a thyroidectomy was carried out to relieve the obstruction. Here, anesthesia was conducted by applying topical anesthesia to the airway with awake laryngoscopy and intubation performed in the sitting position. Once the airway was secured using an armored tube, the patient was given a general anesthetic with an intravenous induction. Following the surgery, extubation occurred without complication.

F. BERNOULLI EFFECT

The fall in pressure at points of flow constriction (where the flow velocity is higher) is known as the Bernoulli effect. This phenomenon is used in apparatus employing the Venturi principle, for example, gas nebulizers, Venturi flowmeters, and some oxygen face masks. The lower pressure due to the Bernoulli effect sucks in (entrains) air to mix with oxygen. It can also be used as a source of laboratory suction.

G. LAPLACE'S LAW

One final consideration that is important in anesthesia is Laplace's law for a sphere. It states that, for a sphere with one air-liquid interface (e.g., an alveolus), the equation relating the transmural pressure difference, surface tension, and sphere radius is:

$$P = \frac{2T}{R} \quad (7)$$

where

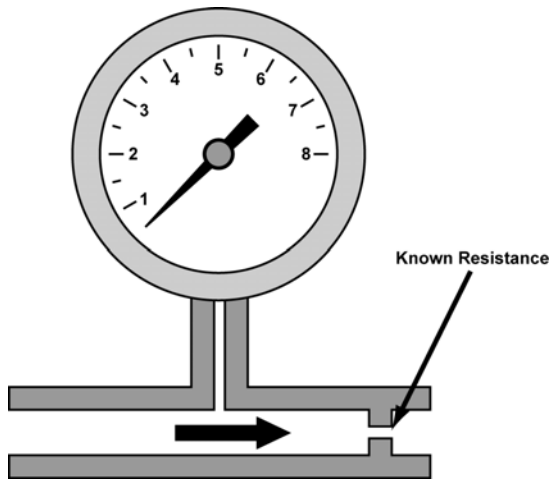
P = transmural pressure difference (dynes/cm²) [1 dyne/cm² = 0.1 Pa = 0.000 751 torr]

T = surface tension (dynes/cm)

R = sphere radius (cm)

The key point in Laplace's law is that the smaller the sphere radius, the higher the transmural pressure. However, real (in vivo) alveoli do *not* obey Laplace's law because of the action of pulmonary surfactant, which acts to decrease the surface tension disproportionately than what is predicted using physical principles. When pulmonary surfactant is missing from the lungs, the lungs take on the behavior described by Laplace's law. This is why surfactant is so important clinically.

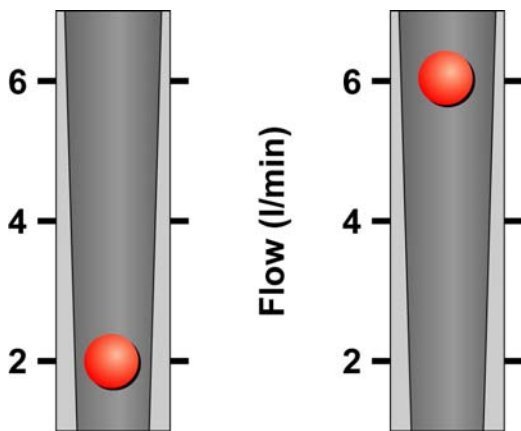
VI. FLOWMETERS



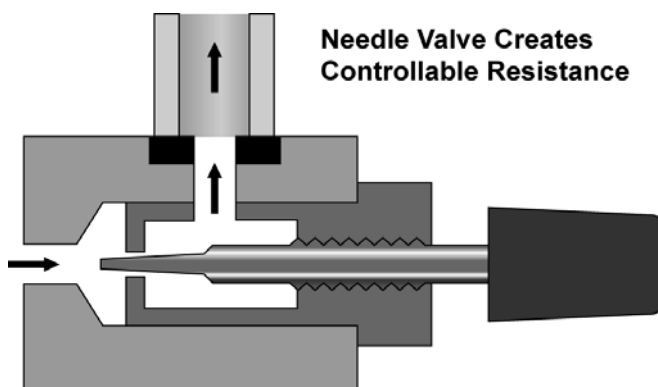
Many flowmeters are just disguised pressure gauges employing a flow restrictor or orifice with a known resistance.

Flow is determined from pressure using the following relationship:

$$\text{Flow} = \text{Pressure} \times \text{Resistance}$$



Flowmeters in many anesthesia machines and oxygen delivery systems employ a sapphire ball or similar indicator in a Thorpe tube, which is wider at the top than at the bottom. The higher the flow, the higher the ball rises. Flow is read at the center of the ball. The flow is controlled using a needle valve, which creates a variable resistance that may be adjusted by the user.



NEED MORE INFORMATION? TRY THESE RESOURCES:

D. John Doyle and K. O'Grady. Physics and biomechanics of the airway. In: J. L. Benumof, Editor: Airway Management: Principles and Practice, C.V. Mosby, 1995.

G.D. Parbrook, P.D. Davis, and E.O. Parbrook. Basic Physics and Measurement in Anesthesia. Butterworth-Heinemann; 3rd edition; 1991

Jerry A. Dorsch and Susan E. Dorsch. Understanding Anesthesia Equipment. Lippincott Williams & Wilkins; 4th edition; 1999 ISBN: 0683304879