

PVT & FLOW (TPG 4145)

Note Title

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PETROLEUM NOMENCLATURE

QUANTITIES	SYMBOL	UNITS		
		SI	(Metric)	Field
Pressure	p	Pa kPa MPa	bar $= 10^5 \text{ Pa}$	psi ↑ lb _f ↓ pounds per square inch



Blaise Pascal

Two types of pressure:

Absolute (p^A)

Gauge (p^g)

$$p^A - p_{sc}^A = p^g$$

Standard conditions



$$p_{10m}^g = (2-1) 10^5 \text{ Pa}$$

$$\sim 1 \cdot 10^5 \text{ Pa}$$

$$\sim 1 \text{ bar}$$

$$(30-15) \text{ psi}$$

$$15 \text{ psi}$$

MPa abs

bara

MPa guage

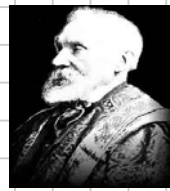
bang

psia
↓
psig

Pressure Units Conversion (App. A SPEPRM)

$$14.5037 \text{ psi} = \underline{1 \text{ bar}} = 10^5 \text{ Pa}$$

14.697 psi = 1 atm



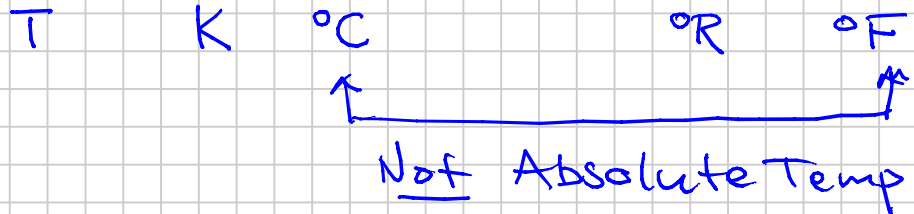
Lord Kelvin

$$\frac{1 \text{ atm}}{1 \text{ bar}} = \frac{14.697}{14.5037} = 1.01325$$

1 atm = 1.01325 · 10⁵ Pa

Use in PVT (EOS)

TEMPERATURE



Kata Celsius

Rankine (Bill)

$$T_K = T_{°C} + 273.15$$

$$T_{°R} = T_{°F} + 459.6$$

$$T_{°R} = 1.8 \cdot T_K$$

$$T_{°F} = 1.8 T_{°C} + 32$$

MANY PE correlations (equations) use °F

Water Boils	@ 100°C (1 atm)	212 °F
Freeze	@ 0°C (1 atm)	32 °F

VOLUME

V

m³ (cm³ or cc)

ft³ in³

prefix:

$$cm^3 = (10^{-2} m)^3 = 10^{-6} m^3$$

m	milli	10 ⁻³
c	centi	10 ⁻²
d	deci	10 ⁻¹
D	deca	10
k	kilo	10 ³
M	mega	10 ⁶
G	giga	10 ⁹

bbl (STB)
 barrels (liquid: oil)
 Gas @ 1 atm & 60°F
 1.0135 bara
 14.696 psia
 "scf"
 standard cubic

T tera 10^{12} feet

STB: stock = "sc"
tank
barrel

1 atm
60°F

Latin 1000

$Mscf = 10^3$ scf

$MMscf = 10^6$ scf
 $10^3 \cdot 10^3$

bcf = 10^9 scf

Tcf = 10^{12} scf

approaching a "giant" gas field
(large)

$Sm^3 =$ std m^3
↑
standard condition

Conversions: $35.31 \text{ ft}^3 = 1 \text{ m}^3$
 $6.28 \text{ bbl} = 1 \text{ m}^3$ } $5.615 \text{ ft}^3 = 1 \text{ bbl}$

Mass m kg (g) lbm oz

Conversion: $2.2046 \text{ lbm} = 1 \text{ kg}$
 $1000 \text{ g} = 1 \text{ kg}$
 $16 \text{ oz} = 1 \text{ lbm}$

Length m (cm) ft, in, mi
 $3.28 \text{ ft} = 1 \text{ m}$ $5280 \text{ ft} = 1 \text{ mi}$

Area A m^2 (cm^2) ft^2
 $\sim 10 \text{ ft}^2 = 1 \text{ m}^2$ Acre =
 43560 ft^2
→ Section: $1 \text{ mi} \times 1 \text{ mi}$
→ 640 Acres/section

Time t s (s) D, hr, yr

Permeability (Area) m^2 (md) md, D

p V T

$$1 \mu\text{m}^2 = 1 \text{ D}$$
$$10^{-12} \text{ m}^2 = 1 \text{ D}$$

Ideal Gas Law

8.314

$$pV = nRT$$

abs. pressure \uparrow p

moles of the gas \uparrow n

Universal gas constant \uparrow R

abs. temp. \uparrow T

SI Units :

p	[Pa]
V	[m ³]
T	[K]
n	[kgmole]

} Pure SI

$R_{SI} = 8314.3$

$$M_{ci} = 16.04 \frac{\text{g}}{\text{gmole}} = 16.04 \frac{\text{kg}}{\text{kgmole}} = 16.04 \frac{\text{lbm}}{\text{lbmole}}$$

$$= 16.04 \frac{\text{oz}}{\text{ozmole}}$$

$$pV = nRT$$

Convert : $T [^{\circ}\text{R}]$

Find R_{Field} $p [\text{psia}]$

$V [\text{ft}^3]$

$n [\text{lb-mol}]$

$$1.8 \text{ K} = 1^{\circ}\text{R}$$

$$14.5037 \text{ psia} = 1 \text{ bara}$$

$$35.31 \text{ ft}^3 = 1 \text{ m}^3$$

$$2.204 \text{ lb} = 1 \text{ kg}$$

$$p \downarrow V = n R T$$

$$[Pa] [m^3] [kg] R_{SI} [K]$$

Input using:

$$p [psia] \quad V [ft^3] \quad n [lb] \quad T [^{\circ}R]$$

$$p [Pa] = \left\{ p [psia] \cdot \frac{\text{bara}}{14.50377 \text{ psia}} \cdot \frac{10^5 \text{ Pa}}{\text{bara}} \right\}$$

$$V [m^3] = \left\{ V [ft^3] \cdot \frac{m^3}{35.31 \text{ ft}^3} \right\}$$

$$n [kg] = \left\{ n [lb] \cdot \frac{kg}{2.204 \text{ lb}} \right\}$$

$$T [K] = \left\{ T [^{\circ}R] \cdot \frac{K}{1.8 \text{ } ^{\circ}R} \right\}$$

$$\left\{ p [psia] \cdot \frac{10^5}{14.5} \right\} \left\{ V [ft^3] \cdot \frac{1}{35.31} \right\} =$$

$$\left\{ n [lb] \cdot \frac{1}{2.2} \right\} \cdot \left\{ T [^{\circ}R] \cdot \frac{1}{1.8} \right\} \cdot R_{SI}$$

8314

$$p [psia] \cdot V [ft^3] = n [lb\text{-mole}] T [^{\circ}R] \cdot$$

$$\left\{ 8314 \cdot \frac{1}{2.2} \cdot \frac{1}{1.8} \cdot \frac{14.5}{10^5} \cdot 35.31 \right\}$$

R_{Field}

$$R_{Field} = 10.73146$$

Mole (unit)

From Wikipedia, the free encyclopedia

The **mole** is a unit of measurement used in chemistry to express amounts of a chemical substance, defined as the amount of any substance that contains as many elementary entities (e.g., atoms, molecules, ions, electrons) as there are atoms in 12 grams of pure carbon-12 (¹²C), the isotope of carbon with relative atomic mass of exactly 12 by definition. This corresponds to the

Avogadro constant, which has a value of $6.022\ 141\ 29(27) \times 10^{23}$ elementary entities of the substance. It is one of the base units in the International System of Units; it has the unit symbol **mol** and corresponds with the dimension symbol *n*.^[1]

The mole is widely used in chemistry instead of units of mass or volume as a convenient way to express amounts of reactants or of products of chemical reactions. For example, the chemical equation $2\ \text{H}_2 + \text{O}_2 \rightarrow 2\ \text{H}_2\text{O}$ implies that 2 mol of dihydrogen (H₂) and 1 mol of dioxygen (O₂) react to form 2 mol of water (H₂O). The mole may also be used to express the number of atoms, ions, or other elementary entities in a given sample of any substance. The concentration of a solution is commonly expressed by its molarity, defined as the number of moles of the dissolved substance per litre of solution.

The number of molecules in a mole (known as Avogadro's constant) is defined such that the mass of one mole of a substance, expressed in grams, is exactly equal to the substance's mean molecular mass. For example, the mean molecular mass of natural water is about 18.015, so one mole of water is about 18.015 grams. Making use of this equation considerably simplifies many chemical and physical computations.

The term **gram-molecule** was formerly used for essentially the same concept.^[1] The term **gram-atom** (abbreviated **gat.**) has been used for a related but distinct concept, namely a quantity of a substance that contains Avogadro's number of *atoms*, whether isolated or combined in molecules. Thus, for example, 1 mole of MgB₂ is 1 gram-molecule of MgB₂ but 3 gram-atoms of MgB₂.^{[2][3]}

In honour of the unit, some chemists celebrate October 23 (a reference to the 10²³ part of the Avogadro constant) as "Mole Day". Some also do the same for February 6 and June 2.

<i>Mole</i>	
Unit system	SI base unit
Unit of	Amount of substance
Symbol	mol

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Definition and related concepts

As of 2011, the mole is defined by BIPM to be the amount of substance of a system which contains the same number of elementary entities (e.g. atoms, molecules, ions, electrons) as atoms in 0.012 kilograms of carbon-12 (¹²C), the isotope of carbon with relative atomic mass 12.^[1] Thus, by definition, one mole of pure ¹²C has a mass of *exactly* 12 g. It also follows from the definition that *X* moles of any substance will contain the same number of molecules as *X* moles of any other substance.

The mass per mole of a substance is called its molar mass. Since the standard unit for expressing the mass of molecules or atoms (atomic mass unit or the dalton) is defined as 1/12 of the mass of a ¹²C atom, it follows that the molar mass of a substance, measured in grams per mole, is exactly equal to its mean molecular or atomic mass, measured in unified atomic mass units or daltons; which is to say, to the substance's mean molecular or relative atomic mass.

The number of elementary entities in a sample of a substance is technically called its (chemical) amount. Therefore, the mole is a convenient unit for that physical quantity. One can determine the chemical amount of a known substance, in moles, by dividing the sample's mass by the substance's molar mass.^[4] Other methods include the use of the molar volume or the measurement of electric charge.^[4]

The mass of one mole of a substance depends not only on its molecular formula, but also on the proportion of the isotopes of each element present in it. For example, one mole of calcium-40 is 39.96259098 ± 0.00000022 grams, whereas one mole of calcium-42 is 41.95861801 ± 0.00000027 grams, and one mole of calcium with the normal isotopic mix is 40.078 ± 0.004 grams.

Since the definition of the gram is not (as of 2011) mathematically tied to that of the atomic mass unit, the number *N_A* of molecules in a mole (Avogadro's number) must be determined experimentally. The value adopted by CODATA in 2010 is *N_A* = 6.02214129 × 10²³ ± 0.00000027 × 10²³.^[5] In 2011 the measurement was refined to 6.02214078 × 10²³ ± 0.00000018 × 10²³.^[6]

We can find number of moles in a given mass (in grams) by

$$n = \frac{\text{given mass}}{\text{molecular mass}}$$

History

The history of the mole is intertwined with that of molecular mass, atomic mass unit, Avogadro's number and related concepts.

The first table of relative atomic mass (atomic weight) was published by John Dalton (1766–1844) in 1805, based on a system in which the relative atomic mass of hydrogen was defined as 1. These relative atomic masses were based on the stoichiometric proportions of chemical reactions and compounds, a fact that greatly aided their acceptance: It was not necessary for a chemist to subscribe to atomic theory (an unproven hypothesis at the time) to make practical use of the tables. This would lead to some confusion between atomic masses (promoted by proponents of atomic theory) and equivalent weights (promoted by its opponents and which sometimes differed from relative atomic masses by an integer factor), which would last throughout much of the nineteenth century.

Jöns Jacob Berzelius (1779–1848) was instrumental in the determination of relative atomic masses to ever-increasing accuracy. He was also the first chemist to use oxygen as the standard to which other masses were referred. Oxygen is a useful standard, as, unlike hydrogen, it forms compounds with most other elements, especially metals. However, he chose to fix the atomic mass of oxygen as 100, an innovation that did not catch on.

Charles Frédéric Gerhardt (1816–56), Henri Victor Regnault (1810–78) and Stanislao Cannizzaro (1826–1910) expanded on Berzelius' works, resolving many of the problems of unknown stoichiometry of compounds, and the use of atomic masses attracted a large consensus by the time of the Karlsruhe Congress (1860). The convention had reverted to defining the atomic mass of hydrogen as 1, although at the level of precision of measurements at that time — relative uncertainties of around 1% — this was numerically equivalent to the later standard of oxygen = 16. However the chemical convenience of having oxygen as the primary atomic mass standard became ever more evident with advances in analytical chemistry and the need for ever more accurate atomic mass determinations.

Developments in mass spectrometry led to the adoption of oxygen-16 as the standard substance, in lieu of natural oxygen. The current definition of the mole, based on carbon-12, was approved during the 1960s.^{[1][7]} The four different definitions were equivalent to within 1%.

Scale basis	Scale basis relative to ¹² C = 12	Relative deviation from the ¹² C = 12 scale
Atomic mass of hydrogen = 1	1.00794(7)	−0.788%
Atomic mass of oxygen = 16	15.9994(3)	+0.00375%
Relative atomic mass of ¹⁶ O = 16	15.9949146221(15)	+0.0318%

The name *mole* is an 1897 translation of the German unit *Mol*, coined by the chemist Wilhelm Ostwald in 1894 from the German word *Molekül* (molecule).^{[8][9][10]} However, the related concept of equivalent mass had been in use at least a century earlier.^[11]

The mole was made the seventh SI base unit in 1971 by the 14th CGPM.^[12]

The mole as a unit

Since its adoption into the International System of Units in 1971, there have been a number of criticisms of the concept of the mole as a unit like the metre or the second:

- the number of molecules, etc. in a given lump of material is a fixed dimensionless quantity that can

- be expressed simply as a number, so does not require its own base unit;^[7]
- the SI thermodynamic mole is irrelevant to analytical chemistry and could cause avoidable costs to advanced economies;^[13]
- the mole is not a true metric (i.e. measuring) unit, rather it is a *parametric* unit and amount of substance is a *parametric* base quantity;^[14]
- the SI defines numbers of entities as quantities of dimension one, and thus ignores the ontological distinction between *entities* and *units of continuous quantities*.^[15]

In chemistry, it has been known since Proust's law of definite proportions (1794) that knowledge of the mass of each of the components in a chemical system is not sufficient to define the system. Amount of substance can be described as mass divided by Proust's "definite proportions", and contains information that is missing from the measurement of mass alone. As demonstrated by Dalton's law of partial pressures (1803), a measurement of mass is not even necessary to measure the amount of substance (although in practice it is usual). There are many physical relationships between amount of substance and other physical quantities, the most notable one being the ideal gas law (where the relationship was first demonstrated in 1857). The term "mole" was first used in a textbook describing these colligative properties.

Other units called "mole"

Chemical engineers use the concept extensively, but the unit is rather small for industrial use.^[16] For convenience in avoiding conversions in the Imperial (or American Customary Units), some engineers adopted the **pound-mole** (notation **lb-mol** or **lbmol**), which is defined as the number of entities in 12 lb of ¹²C. One lb-mol is equal to 453.59237 mol.^[17]

In the metric system, chemical engineers once used the **kilogram-mole** (notation **kg-mol**), which is defined as the number of entities in 12 kg of ¹²C, and often referred to the mole as the **gram-mole** (notation **g-mol**), when dealing with laboratory data.^[17]

Late 20th century chemical engineering practice came to use the **kilomole** (kmol), which is numerically identical to the kilogram-mole, but whose name and symbol adopt the SI convention for standard multiples of metric units - thus kmol means 1000 moles. This is analogous to the use of kg instead of g. The use of kmol is not only for "magnitude convenience" but also makes the equations used for modelling chemical engineering systems coherent. For example, the conversion of a flowrate of kg/s to kmol/s only requires the molecular mass not the factor 1000 unless the basic SI unit of mol/s were to be used. Indeed, the appearance of any conversion factors in a model can cause confusion and is to be avoided; possibly a definition of coherence is the absence of conversion factors in sets of equations developed for modelling.

Concentrations expressed as kmol/m³ are numerically the same as those in mol/dm³ i.e. the molarity conventionally used by chemists for bench measurements; this equality can be convenient for scale-up.

Proposed future definition

In 2011, the 24th meeting of the General Conference on Weights and Measures (CGPM) agreed a plan for a possible revision of the SI base unit definitions on an as yet undetermined date. This plan, set forward in the meeting's first resolution, included a proposal to redefine the mole in a way that will fix "the Avogadro