Physical Behavior of Matter

Unit 1

Physical Behavior of Matter: Section A (separating matter and energy changes)

Chemistry: The study of matter and its changes

Matter: anything that has mass and takes up space

Substance: has a constant composition and constant properties throughout a given sample and from sample to sample

Element: a single type of matter, every atom of which has the same number of protons

- can not be broken down into simpler substances by chemical means
- have only one type of atom
- all atoms have the same atomic number (or number of protons)

Examples: sodium (^{11}Na), boron (^{5}B), carbon (^{6}C), scandium (^{21}Sc)

Compound: composed of two or more different elements chemically combined in a fixed ratio

- can be broken down into simpler substances by chemical means
- has a specific chemical identity
- has a specific formula and a unique name has a specific formula and a unique name
- binary compound

  Contains TWO kinds of elements

  Will have two capital letters in its formula

Examples (binary): salt (NaCl), water (H_2O), carbon dioxide (CO_2), ammonia (NH_3)

Examples (not binary): sulfuric acid (H_2SO_4) or glucose ($C_6H_{12}O_6$)

Mixture: two or more substances that are not combined chemically and the composition can vary from sample to sample and sometimes even within a sample

- can be separated into two or more substances by physical means
- are not chemically combined
- have no definite ratio of elements

Examples: salt and water ($NaCl_{(aq)}$), oil and vinegar, or air

How could we separate salt and sand?

1. Add water (salt dissolves, sand does not)
2. Carefully pour out the water
3. Let the water evaporate (leaving the salt)

Heterogeneous:

- not uniform throughout
- has phase boundaries

Examples: fog in air, oil and vinegar, minestrone soup

Homogeneous:

- uniform throughout
- has *no* phase boundaries
- usually called a solution
- can be solid in solid, solid in liquid, gas in liquid, or gas in gas

Examples: filtered air, alloys, salt solution ($NaCl_{(aq)}$)

Solutions and Table G

- unsaturated solution (below the line on Table G)
- saturated solution (on the line on Table G)
- supersaturated solution (above the line on Table G)

Example: in terms of saturation, describe a solution made by adding 19 g of NaCl to 50 grams of water at 30°C

Table G shows 38 g NaCl / 100 g H_2O at 30°C

$$\frac{19\ g_{NaCl}}{50\ g_{H_2O}} = \frac{38\ g_{NaCl}}{100\ g_{H_2O}} - \text{a saturated solution}$$

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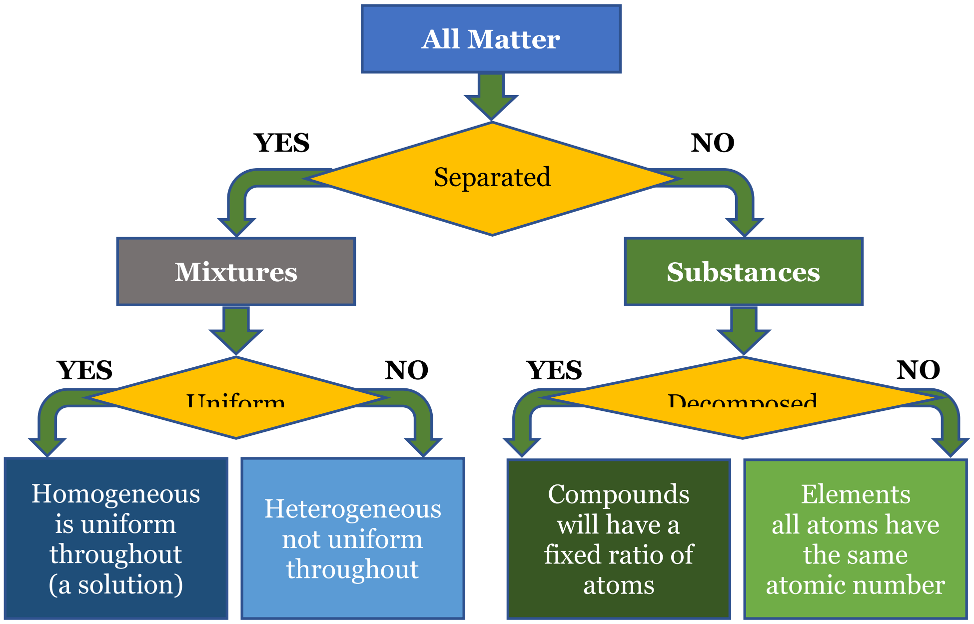
All matter can be classified as pure substances or mixtures of substances



Matter can be separated by:

Physical change

- only rearranges particles of substances
- chemical identities (chemical formulas) remain the same

Examples: filtering, evaporating, melting, freezing, boiling

Chemical change

- forms new substances with new properties
- changes chemical identities (chemical formulas)

Example:

HCl	+	$NaOH$	→	$NaOH$	+	HOH
acid		base		salt		water
toxic		caustic		necessary for life		

Typical lab procedures for separating matter by physical means

Filtration: based on large differences in the size of particles to be separated
Small particles (water) pass through a filter paper but large particles (sand) do not

Distillation: based on differences in boiling points
Low boiling point particles (water) evaporate, high boiling point particles (salt) do not

Crystallization: based on solubility or concentration differences
Make a hot saturated solution and allow it to cool, some solute will crystallize

Sublimation: based on vapor pressure
Iodine will go from solid to gas if warmed, and from gas to solid if cooled

Chromatography: based on differences in molecular polarities
Like dissolves like – some ink is more attracted to paper while alcohol attracts others

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Five indicators of a chemical change

- Heat or light evolved
  - A natural gas flame
- Evolution of a gas
  - Baking soda and vinegar produce a salt and carbon dioxide gas
- Precipitation
  - Copper(II) nitrate and sodium hydroxide produce sodium nitrate and solid copper(II) hydroxide
- Color change
  - Black copper(II) oxide and hydrogen produce bright shiny copper and water vapor
- Odor changes
  - A rotting egg produces hydrogen sulfide

Phases of Matter

Solid

- definite shape and definite volume
- particles are close together in fixed positions, they vibrate but cannot move past one another (cannot flow)
- result from the strongest forces of attraction between particles
- crystalline in structure (have an orderly repeat pattern)

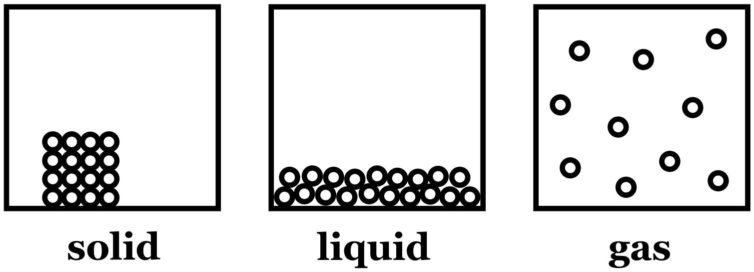
Liquid

- definite volume but no definite shape (takes the shape of the container)
- particles are close together but can move past one another while touching (can flow)
- result from fairly strong forces of attraction between particles (less strong than solids)
- have no crystalline in structure (have no regular arrangement of particles)

Gas

- no definite shape and no definite volume (completely fill a closed container)
- particles are far apart, they easily move past one another because they do not touch each other except in brief collisions (can flow)
- result from the weakest forces of attraction between particles
- totally disordered with particles moving in random directions

Particle diagrams



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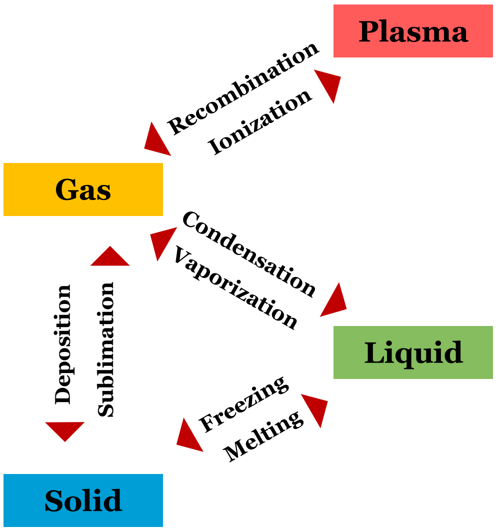
Phases of matter



Energy: The ability to do work

Energy changes must follow the Law of Conservation of Matter and Energy

Matter and energy can be converted from one form to another but are neither created nor destroyed

Mechanical energy is the sum of the:

Kinetic energy (energy of motion) – associated with temperature

- temperature is defined as the average kinetic energy of the sample
- When I say, "temperature," you say, "average kinetic energy."

Potential energy (stored energy) – includes chemical energy (energy stored in bonds)

- energy changes during phase changes (melting, freezing, boiling, or condensation) change the PE or entropy but do not change the temperature (average KE)
- entropy: the amount of disorder in a system, more disorder = higher entropy

Bonds are broken *only* when enough energy is added

Endothermic or energy absorbed (<u>en</u>dothermic – energy <u>en</u>ters)

Bond formation *always* releases energy

Exothermic or energy released (<u>ex</u>othermic – energy <u>ex</u>its)

Heat *always* flows from HOT to NOT hot

Thermometers

- measure changes in the average KE (or measure heat transfers)
- are fixed based on the freezing and boiling points of water

Freezing point: ice / water equilibrium temperature at 1 atm pressure (0°C)

Boiling point: water / steam equilibrium temperature at 1 atm pressure (100°C)

Temperature is measured in:

- Celsius degrees
- Kelvin units

(Never say degrees Kelvin)

A change of 1°C = a change of 1 K

To convert: K = °C + 273 (See Table T)

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Heat is a way to measure energy changes in a chemical reaction (measured in joules)

- joule (J) = unit of heat
- 4.18 J increases the temperature of water 1°C
- Refer to Tables B and D
- *total* KE of the particles in a system
- IS a form of energy

Temperature is a way to determine which way energy will flow (from high T to low T)

- *average* KE of the particles in a system
- is NOT a form of energy

If you still think heat and temperature are the same thing, think about which would hurt more, one drop of boiling water or one gallon of boiling water

- if both are boiling, they are at the same temperature, 100°C
- more water can transfer more energy (more heat)

Calculating Energy Changes

Reference Table T shows three equations for heat calculations

$q = mC\Delta T$

$q = H_f$

$q = H_v$

q = heat

m = mass

C = specific heat capacity

ΔT = change in temperature (°C or K)

H_f = heat of fusion

H_v = heat of vaporization

Knowing which equation to choose:

- *freezing* (or melting): use the H_f equation
- *vaporizing* (or condensing): use the H_v equation
- *temperature changes*: use the ΔT equation

Reference Table B shows the three Physical Constants for Water

Heat of Fusion	334 J/g
Heat of Vaporization	2260 J/g
Specific Heat Capacity of $H_2O_{(\ell)}$	4.18 J/g•K

Examples:

How much heat energy in joules is absorbed by 27.5 grams of liquid water as it is heated from 15.5°C to 73.6°C?

$q = ?$

$m = 27.5\text{ g}$

$\Delta T = 73.6°\text{C} - 15.5°\text{C} = 58.1°\text{C}$

$q = mC\Delta T$

$= (27.5\text{ g})(4.18\text{ J/g}\cdot\text{K})(58.1\text{ K})$

$= 6678.595\text{ J or } 6.68\text{ kJ}$

How much heat energy in joules is released when 83.2 grams of liquid water freezes at 0°C?

$q = ?$

$m = 83.2\text{ g}$

$q = mH_f$

$= (83.2\text{ g})(334\text{ J/g})$

$= 27\,788.8\text{ J or } 27.8\text{ kJ}$

How much heat energy in joules is absorbed by 27.5 grams of liquid water as it is vaporized to steam at the boiling point of water?

$q = ?$

$m = 27.5\text{ g}$

$q = mH_v$

$= (27.5\text{ g})(2260\text{ J/g})$

$= 62\,150\text{ J or } 62.2\text{ kJ}$

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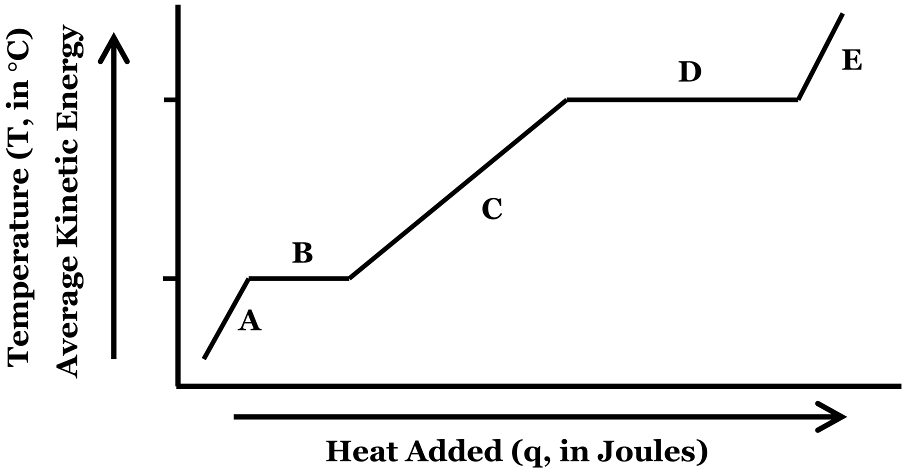
Heating and Cooling Curves

Heating Curve

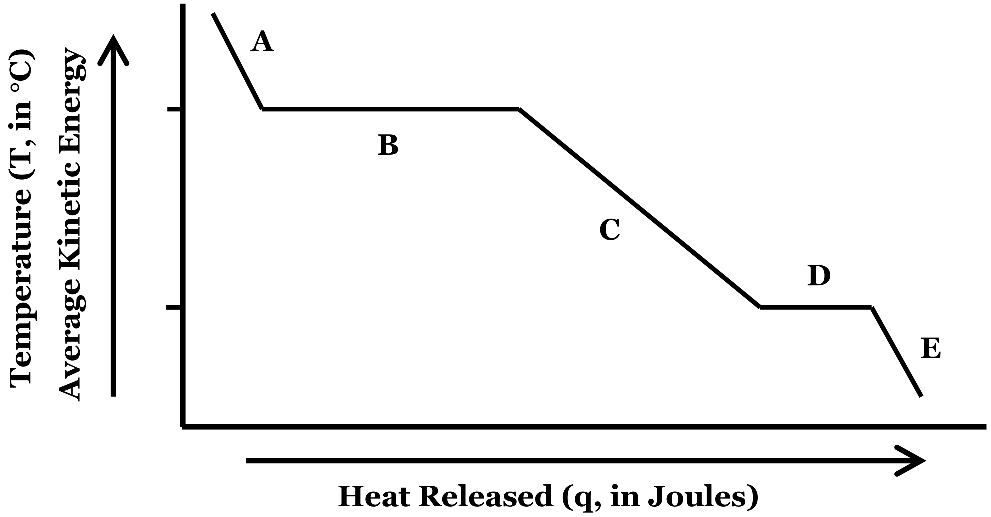


Remember: if heat is being used to change the temperature then it cannot melt or vaporize a substance at the same time

Heat is stupid: it can only do one thing at a time

Region A: all solid, temperature is rising as heat is added
Region B: solid and liquid, solid melting but temperature is constant (the MP)
Region C: all liquid, temperature is rising as heat is added
Region D: liquid and gas, liquid vaporizing but temperature is constant (the BP)
Region E: all gas, temperature is rising as heat is added

Cooling Curve



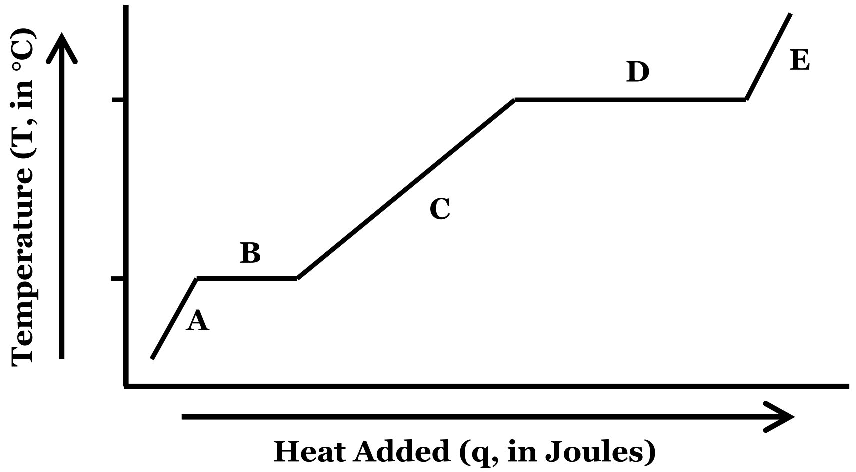
Heat is stupid: it can only do one thing at a time

Region A: all gas, temperature is falling as heat is removed
Region B: gas and liquid, gas condensing but temperature is constant (the BP)
Region C: all liquid, temperature is falling as heat is removed
Region D: liquid and solid, liquid freezing but temperature is constant (the FP)
Region C: all solid, temperature is falling as heat is removed

Sample Regents questions:

Starting as a solid, a sample of a molecular substance is heated until the entire sample of the substance is a gas. The graph below represents the relationship between the temperature of the sample and the elapsed time.

Heating Curve



Compare the average kinetic energy of the molecules of the sample during interval *B* to the average kinetic energy of the molecules of the sample during interval *D*.

(1) The average kinetic energy at interval *B* is greater than at interval *D*.
(2) The average kinetic energy at interval *B* is the same as during interval *D*.
[✓ 3] The average kinetic energy at interval *B* is less than at interval *D*.
(4) There is not enough information given to determine the average kinetic energy.

The boiling point of the substance occurs at

(1) interval *A*
(2) interval *B*
(3) interval *C*
[✓ 4] interval *D*

Entropy changes occur at intervals

(1) *A* and *B*
(2) *A* and *C*
[✓ 3] *B* and *D*
(4) *D* and *E*

The kinetic energy of the sample changes at intervals

(1) *A*, *C*, and *E*
(2) *B* and *D*
[✓ 3] *A*, *B*, and *D*
(4) *C*, *D* and *E*

The substance exists as a solid *only* at interval

[✓ 1] *A*
(2) *A* and *B*
(3) *B*
(4) *B* and *D*

The substance exists as *both* a liquid and a gas at interval

(1) *B*
(2) *C*
[✓ 3] *D*
(4) *E*

Physical Behavior of Matter: Section B (solids, liquids, gases, KMT, and gas laws)

Solids:

- definite shape and definite volume
- crystalline (usually, but amorphous solids do occur)
- particles vibrate but do not move past each other (do not flow)
- have a unique melting point (solid and liquid phases in equilibrium) Table S
- may sublime (change from solid directly to gas, e.g., dry ice or CO_2)

Liquids:

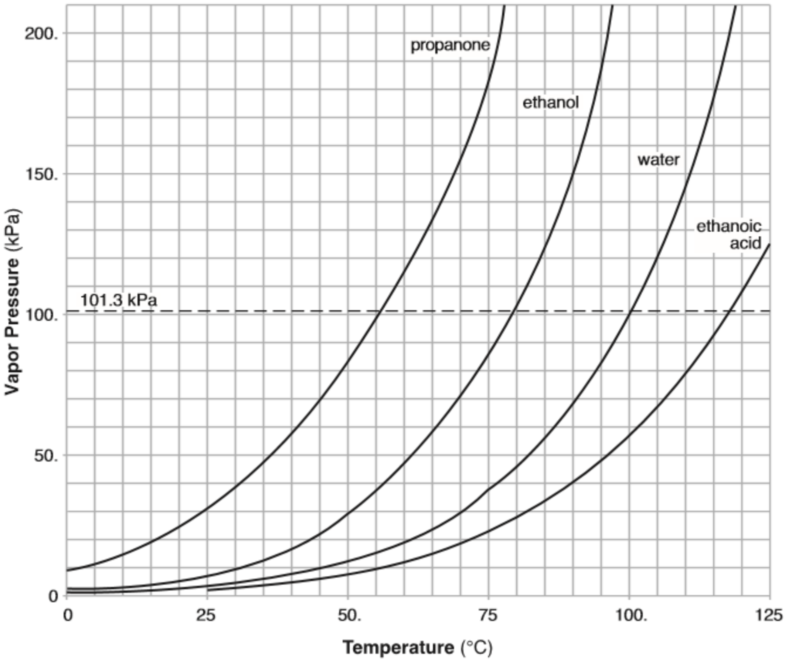
- definite volume but take the shape of their container
- particles vibrate and can move past each other (can flow)
- have unique freezing point (solid and liquid in equilibrium, see MP above)
- have a unique boiling point (liquid and gas phases in equilibrium at 1 atm) Table S
- evaporate or change from liquid to gas below the BP

Vapor pressure: the equilibrium pressure exerted by the molecules of a gas that evaporate from a liquid at a given temperature

- characteristic (or unique) for a given substance
- varies exponentially with the temperature
- varies inversely with intermolecular forces (IMFs)
  stronger IMFs require more energy to break bonds and become a gas

See Table H

Table H
Vapor Pressure of Four Liquids



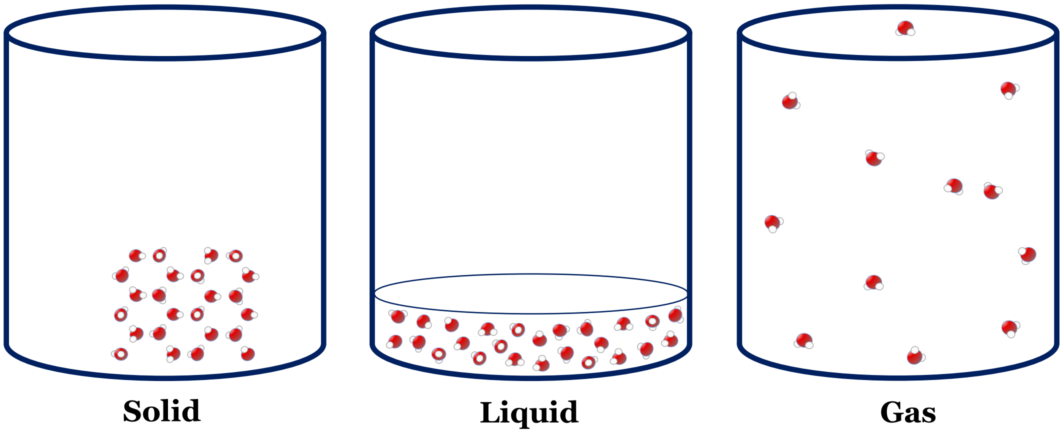
Examples:

What is the vapor pressure of ethanol at 75°C?
Answer: 85 kPa

What is the normal boiling point of ethanol?
Answer: 79°C

Gases:

- no definite volume and take the shape of their container
- particles are far apart and move in random straight lines, they flow easily
- have unique condensation point (liquid and gas in equilibrium, see BP above)
- some gases go directly from gas to solid state (deposition, e.g., CO_2 and iodine)

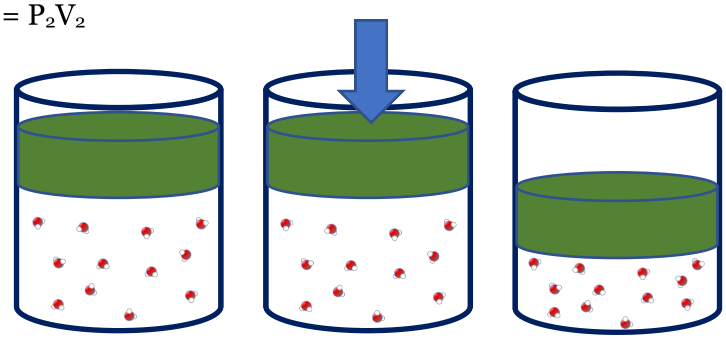


Kinetic Molecular Theory of Gases

1. Gases are composed of very tiny particles that are very far apart relative to the size of the particles.
2. Collisions between gas particles are perfectly elastic (collisions can transfer heat but the total energy of the system remains constant)
3. Gas particles move in continuous, rapid, random, straight-line motion.
4. The particles of a gas exert no forces of attraction or repulsion.
5. $KE = \frac{1}{2} mu^2$

Boyle's Law: at constant temperature, the volume of a gas is inversely proportional to the pressure

$PV = k$ or $P_1V_1 = P_2V_2$



If $V_2 = \frac{1}{2} V_1$ then $P_2 = 2P_1$

Charles' Law: at constant pressure, the volume of a gas is directly proportional to the absolute temperature

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

If $T_2 = 2\ T_1$ then $V_2 = 2V_1$

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Combined Gas Law:

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

If $T_1 = T_2$ then $P_1V_1 = P_2V_2$ or Boyle's Law results

If $P_1 = P_2$ then $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ or Charles' Law results

In gas law problems, the temperature *must always* be expressed in Kelvins

Example: At a temperature of 0°C, a 400. mL gas sample has a pressure of 101.3 kPa. If the pressure is changed to 76.5 kPa, at what temperature will the volume be 551 mL?

$T_1 = 0°\text{C} = 273\ \text{K}$
$V_1 = 400.\ \text{mL}$
$P_1 = 101.3\ \text{kPa}$
$T_2 = ?$
$V_2 = 551\ \text{mL}$
$P_2 = 76.5\ \text{kPa}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \therefore T_2 = \frac{P_2V_2T_1}{P_1V_1}$$

$$= \frac{(76.5\ kPa)(551\ mL)(273\ K)}{(101.3\ kPa)(400.mL)}$$

$= 283.992090$ K (calculator answer)

$= 284$ K (correct sig figs)

$= 11°$C (temperature in °C)

Deviations from Gas Laws

Remember that the KMT of gases states that

1. the particle size is small compared to the distance between the particles
2. there are no forces of attraction between the particles in a gas

But if the particles are forced very close together

- the volume of the particles does become significant
- the forces of attraction between the particles also becomes significant

Looking at the combined gas law, there are two ways to force the volume to become small

$$\frac{PV}{T} = K$$

1. make the pressure very high (as P goes up, V becomes small to maintain constant, K)
2. make the temperature very low (as T decreases, V decreases to maintain constant, K)

Example: What changes would cause gases to deviate from ideal behavior?

Answer: Increase the pressure and decrease the temperature.

At which temperature and pressure will a sample of neon gas behave most like an ideal gas?

(1) 300. K and 2.0 atm
(2) 300. K and 4.0 atm
[✓ 3] 500. K and 2.0 atm
(4) 500. K and 4.0 atm

Avogadro's Hypothesis: at constant temperature, and pressure, equal volumes of two gases will contain the same number of particles

$\frac{n}{V} = \text{k}$ or $\frac{n_1}{V_1} = \frac{n_2}{V_2}$

Example: Compare the total number of gas molecules in cylinder A to the total number of gas molecules in cylinder B.

Answer: They are equal.