Name: ________________ Grade 11 Chemistry Summary C11_P.1

Unit 1: Matter, Chemical Trends, and Chemical Bonding (Page 1 of 2)

Rules for Determining Significant Digits

1. When 0's are between significant figures, 0's are always significant. ✓✓✓ 101
2. When the measurement is a whole number ending with 0's, the 0's are never significant. ✓✓✗ 340
3. When the measurement is less than a whole number, the 0's between the decimal and other significant numbers are never significant (they are place holders). ✗✗✗✓ 0.002
4. When the measurement is less than a whole number and the 0's fall after the other significant numbers, the 0's are always significant. ✗✗✗✓✓ 0.0020
5. When the measurement is less than a whole and there is a 0 to the left of the decimal, the 0 is not significant. ✗✗ 0.02
6. When the measurement is a whole number but ends with 0's to the right of the decimal, the 0's are significant. 3.0 ✓✓ 300. ✓✓✓

Rules for Reporting Significant Digits in Calculations

Rule 1: Multiplying and Dividing 2.0 x 6 = 1x10

The value with the fewest number of significant digits, going into the calculation, determines the number of significant digits that you should report in your answer.

Rule 2: Adding and Subtracting

The value with the fewest number of decimal places, going into the calculation, determines the number of decimal places that you should report in your answer

Rule 3: Rounding

To get the appropriate number of significant digits, you may need to round your answer.

If your answer ends in a number that is **greater than 5, round up**. For example, 3.456 can be rounded to 3.46, 2.3451 can be round to 2.35.

If your answer ends with a number that is **less than 5, round down**. for example, 1.234 can be rounded to 1.23.

If your answer **ends with exactly 5, round to even**. For example, 2.345 can be round to 2.34, but 2.375 is rounded to 2.38. ODD-UP RULE!

Unit 1: Matter, Chemical Trends, and Chemical Bonding (Page 2 of 2)

Standard atomic notation

P + N = Atomic mass

$${}^{A}_{Z}X$$

mass number → A (identity of element)

element → X

atomic number → Z

- The total number of protons and neutrons in the nucleus of an atom is the *mass number*, *A*.
- Each element has a characteristic number of protons in the nucleus. This is the *atomic number*, *Z*.
- *X* is the *symbol* of the element.

FACTORS

① NUCLEAR CHARGE nucleus (+) attracts e^- (−), ↑ p^+ ↑ attraction

ATOMIC RADIUS: distance from centre of nucleus to valence.

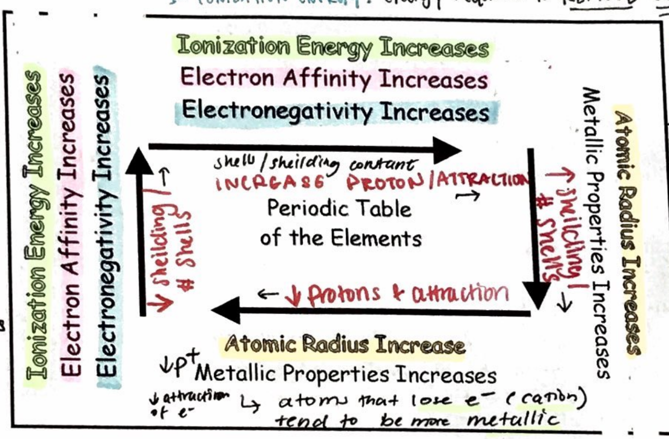
Electron Affinity: attractive force atom has for ADDING e^- to valence.

1st IONIZATION ENERGY: energy required to REMOVE ONE e^- from valence shell of neutral atom.

Periodic Trends

IONIC RADIUS

- Metal (CATIONS) lose e^-
  smaller ↓ increase ← increase
- Nonmetal (ANIONS) gain e^-
  LARGER ↓ INCREASE ← INCREASE



↳ atoms that lose e^- (cation) tend to be more metallic

↳ atoms that gain e^- (anion) tend to be nonmetallic

ELECTRONEGATIVITY: ability to HOLD on to e− in a covalent bond

WEIGHTED AVERAGE: all of an elements naturally existing isotopes & their relative abundance

RAM = Σ(% abundance × mass isotope)

↑ sum of

* mass of isotope must be rounded

ISOTOPE: same element w/ diff atomic mass (diff n°)

e.g. carbon-12
carbon-13

RADIOISOTOPE: unstable nucleus + emits radiation energy → stable form

② ELECTRON SHELLS / SHIELDING

→ more shell → more distance b/w e^- + nucleus
→ less attraction

→ inner e^- shells block attractive force b/w valence e^- + nucleus

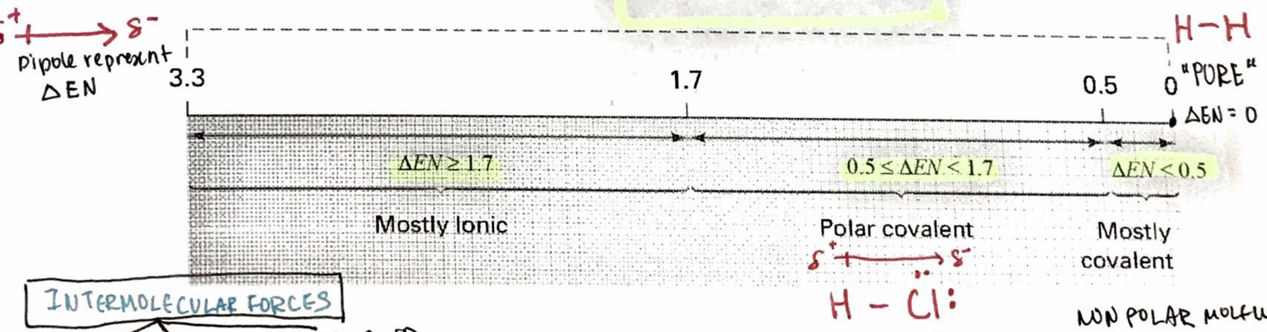
REACTIVITY:

- Metal → lose e^- (IE)
- Nonmetal → gain e^- (EA)

Chemical bonds range in character from mostly ionic to mostly covalent (shair pair of e^-)

* visit lewis structure sheet

$$\Delta EN = |EN_1 - EN_2|$$

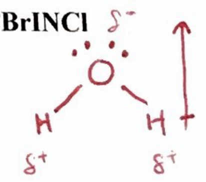


INTERMOLECULAR FORCES

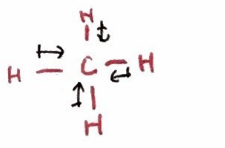
- L.D.F
  - nonpolar molecule
  - cause: instantaneous induced dipole
- D-D INT.
  - polar
  - δ− end attract to δ+ end
    ↳ permanent dipole
- H-BOND
  - polar containing F, O, N, H
  - $H^{\delta+}$ attract to $F^{\delta-}$, $O^{\delta-}$, or $N^{\delta-}$

weak force ⟶ STRONGEST FORCE!

Diatomic Elements: **HOFBrINCl**



NON POLAR MOLECULE



- symmetry
- dipoles cancel out!

Name: ______

Unit 2: Chemical Rea

1. Synthesis Reacti
$4Fe_{(s)} + 3O_{2(g)}$
$2H_2(g) + O_{2(g)}$
$2Na_{(s)} + Cl_{2(g)}$
* $SO_{3(g)} + H_2O_{(l)}$
* $CaO_{(s)} + H_2O_{(l}$

2. Decomposition
$2MgO_{(s)} \rightarrow 2N$
$2NaNO_{3(s)} \rightarrow$
$MgCO_{3(s)} \rightarrow$
$CaSO_{4(s)} \rightarrow$
$Cu(OH)_{2(s)}$
$2KClO_{3(s)}$

3. Combustion
Complete C
◆ C
Incomplete
◆ T
◆ C

4. Single Dis
One met
1. C
C
2.
3.

Single
√C

Unit 2: Chemical Reactions (Page 1 of 2)

1. Synthesis Reactions (AKA combination or formation reactions)

$4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$ $2H_2(g) + O_{2(g)} \rightarrow 2H_2O_{(g)}$	When a metal or a non-metal element reacts with oxygen, the product is an oxide. ← When react w̄ oxygen
$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$	A metal and a non-metal reacts to form a binary compound.
* $SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$	When a non-metallic oxide reacts with water, the product is an acid.
* $CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)}$	When a metallic oxide reacts with water, the product is a base.

2. Decomposition Reactions

$2MgO_{(s)} \rightarrow 2Mg_{(s)} + O_{2(g)}$	A binary compound decomposing into its elements.
$2NaNO_{3(s)} \rightarrow 2NaNO_{2(s)} + O_{2(g)}$	A metal nitrate decomposing into a metal nitrite and oxygen gas.
$MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)}$	A metal carbonate decomposing into a metal oxide and carbon dioxide.
$CaSO_{4(s)} \rightarrow CaO_{(s)} + SO_{3(g)}$	A metal sulfate decomposing into a metal oxide and sulfur trioxide.
$Cu(OH)_{2(s)} \rightarrow CuO_{(s)} + H_2O_{(l)}$	A metal hydroxide decomposing into a metal oxide and water.
$2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$	A metal chlorate decomposing into a metal chloride and oxygen gas.

3. Combustion Reactions

Complete Combustion $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$

- Complete combustion involves: Fuel (hydrocarbon) + oxygen → carbon dioxide + water

Incomplete Combustion $CH_4 + \frac{3}{2}O_{2(g)} \rightarrow CO + 2H_2O + CO_2 + C$

- Too little oxygen leads to an incomplete combustion reaction.
- Carbon monoxide (and other products such as CO_2 and C) is produced.

4. Single Displacement Reactions

One metal displacing another metal from a compound: (Three kinds)

1. One metal replaces another metal in an ionic compound.
   $Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$
2. A metal displaces hydrogen from an acid, creating hydrogen gas.
   $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
3. A metal displaces hydrogen from a water molecule, creating hydrogen gas.
   $2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$

$Na + HOH \rightarrow NaOH + H_{2}$ (g)

Single Displacement Reactions Involving Halogens

$Cl_{2(g)} + CaBr_{2(aq)} \rightarrow CaCl_{2(aq)} + Br_{2(l)}$

Halogen Reactivity: F > Cl > Br > I

Metal Activity Series →

Li
Cs
Rb
K
Ba
Sr
Ca
Na
Mg
Be
Al
Mn
Zn
Cr
Fe
Cd
Co
Ni
Sn
Pb
H
Sb
As
Bi
Cu
Ag
Pd
Hg
Pt
Au

ex: As will replace Cu will not

S.D. of metal will only occur when the single starting metal is HIGHER on series

Unit 2: Chemical Reactions (Page 2 of 2)

5. Double Displacement Reactions

$$NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

Double Displacement Reactions that Form a Precipitate

$$MgCl_{2(aq)} + 2KOH_{(aq)} \rightarrow Mg(OH)_{2(s)} + 2KCl_{(aq)}$$

A solubility table can tell which compounds is insoluble and becomes a precipitate

Double Displacement Reactions that Produce a Gas

$$Na_2CO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + \boxed{H_2CO_{3(aq)}}$$

Acid + Carbonate Compound

But **$H_2CO_{3(aq)}$ is unstable** and decomposes to carbon dioxide and water as shown.

$$H_2CO_{3(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}$$

Overall, we can write this two-step reaction as follows:

$$Na_2CO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$

$$CaSO_{3(aq)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + \boxed{H_2SO_{3(aq)}}$$

Acid + Sulfite Compound

But **$H_2SO_{3(aq)}$ is unstable** and decomposes to sulfur dioxide and water as shown.

$$H_2SO_{3(aq)} \rightarrow H_2O_{(l)} + SO_{2(g)}$$

The name of $H_2SO_{3(aq)}$ is sulfurous acid

Overall, we can write this two-step reaction as follows:

$$CaCO_{3(aq)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + SO_{2(g)}$$

$$NH_4Cl_{(aq)} + NaOH_{(aq)} \rightarrow \boxed{NH_4OH_{(aq)}} + NaCl_{(aq)}$$

Base + Ammonium Compound

But **$NH_4OH_{(aq)}$ is unstable** and decomposes to ammonia and water as shown.

$$NH_4OH_{(aq)} \rightarrow H_2O_{(l)} + NH_{3(g)}$$

Overall, we can write this two-step reaction as follows:

$$NH_4Cl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + NH_{3(g)}$$

The Formation of Water in a Neutralization Reaction

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$$

$$HNO_{3(aq)} + NaOH \rightarrow H_2O_{(l)} + NaNO_{3(aq)}$$

When an acid reacts with a base, water and a salt and/or gas is produced.

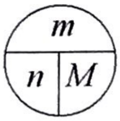
Particles involved in different types of Radioactive Decay

${}^{4}_{2}He$:alpha particle ${}^{1}_{0}n$: neutron ${}^{0}_{-1}e$: beta particle ${}^{1}_{2}p$ or ${}^{1}_{2}H$: proton

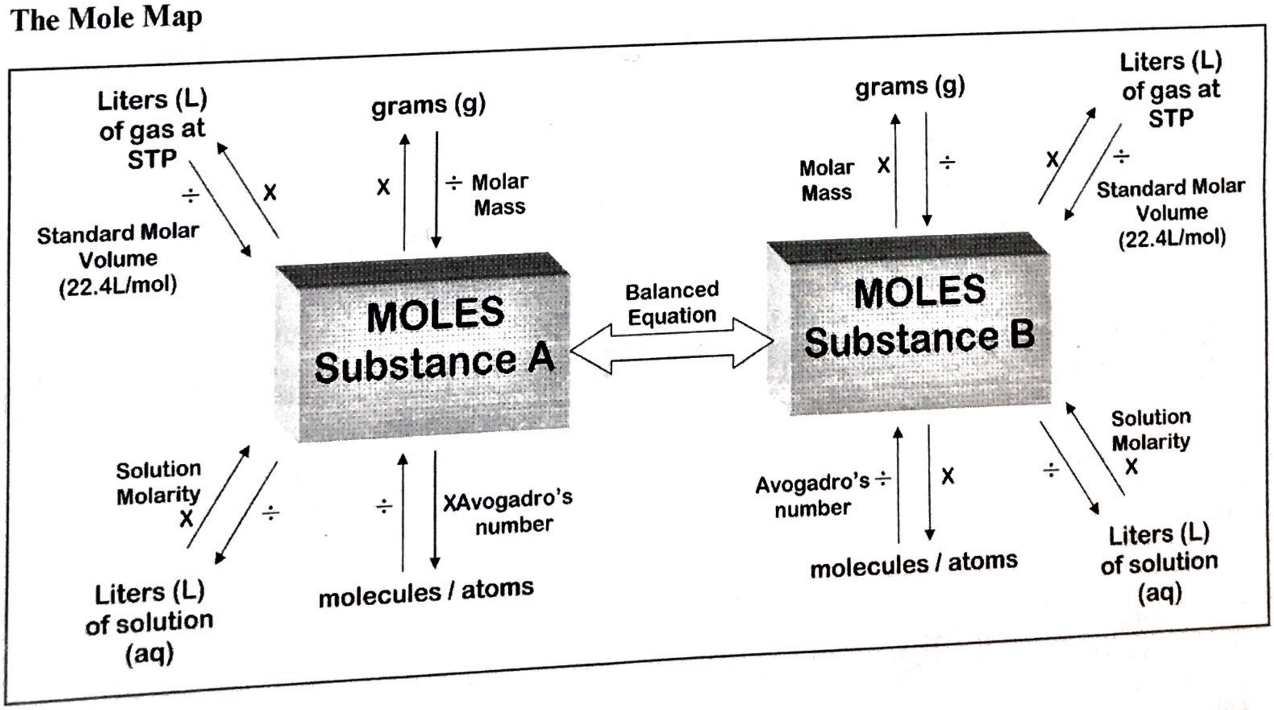
Unit 3: Quantities in Chemical Reactions (Page 1 of 1)

smallest ratio → $\frac{M.F.}{E.F.} = \frac{MM_{MF}}{MM_{EF}} = X$

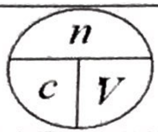
↑ rounded

Equation	Variables	Name
$m_{av} = f_1 m_1 + f_2 m_2 + \ldots$	m_{av} = average atomic mass u f_1 = relative abundance of isotope 1 m_1 = atomic mass of isotope 1 u f_2 = relative abundance of isotope 2 m_2 = atomic mass of isotope 2 u	Average atomic mass
$N = n \cdot N_A$ 	N = number of representative particles (atoms, molecules, formula units) n = number of moles N_A = Avogadro constant (6.022×10^{23})	Avogadro constant
$m = n \cdot M$ 	m = mass of sample g n = amount (# of moles) mol M = molar mass g/mol	amount and molar mass
$\text{mass \% of } X = \frac{\text{mass}_X}{\text{mass}_{\text{compound}}} \times 100\%$	mass_X = mass of element X g $\text{mass}_{\text{compound}}$ = mass of compound g	percent composition
$\text{\% yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$	actual yield g theoretical yield g	percentage yield

The Mole Map



Unit 4: Solutions and Solubility (Page 1 of 2)

Equation	Variables	Name
$c = \frac{n}{V}$ 	c = molar concentration mol/L or M n = amount of solute mol V = volume of solution L	Molar concentration (AKA Molarity)
molality $= \frac{n}{m}$	molality m n = amount of solute mol m = mass of solution kg	Molality
$\%(\text{m/v}) = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100\%$		Mass/volume percent
$\%(\text{m/m}) = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 100\%$		Mass/mass percent % solution (by mass)
$\%(\text{v/v}) = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100\%$		Volume/volume percent % solution (by volume)
$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$ ppm or $\frac{\text{mass of solute}}{\text{mass of solution}} = \frac{x\text{ g}}{10^6\text{ g of solution}}$		Parts per Million (ppm)
$\text{ppb} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^9$ ppb or $\frac{\text{mass of solute}}{\text{mass of solution}} = \frac{x\text{ g}}{10^9\text{ g of solution}}$		Parts per Billion (ppb)
$c = \frac{(\text{density of solution})(\text{percentage of solute by mass})}{\text{Molar mass of solute}}$		Concentration and density of solution
$c_1V_1 = c_2V_2$		Diluting a solution

Density of water: 1 kg / L = 1 g / mL

$$\text{Percent(age) Error} = \frac{|\text{Experimental Value} - \text{Actual Value}|}{\text{Actual Value}} \times 100\%$$

Compounds of ions with **small charges** tend to be **soluble**.
Compounds with **large ions** tend to be more **soluble** than compounds with small ions.

Table The Colour of Some Common Ions in Aqueous Solution

Table The Flame Colour of Selected Metallic Ions

Name: ______________

Unit 4: Solutions and Solubility (Page 2 of 2)

The Arrhenius Theory of Acids and Bases

- An acid is a substance that dissociates in water to produce one or more hydrogen ions, H^+. $HCl_{(g)} \rightarrow \boxed{H^+}_{(aq)} + Cl^-$
- A base is a substance that dissociates in water to form one or more hydroxide ions, OH^-.

$NaOH_{(s)} \rightarrow Na^+_{(aq)} + \boxed{OH^-}_{(aq)}$

The Brønsted-Lowey Theory of Acids and Bases

- An acid is a substance from which a proton (H^+ ion) can be removed.
- A base is a substance that can remove a proton (H^+ ion) from an acid. (gain H^+)

$$\overset{\text{ACID}}{HCl} + \overset{\text{BASE}}{H_2O} \xrightarrow{\text{gains } H^+} \overset{\text{conjugate acid}}{H_3O^+} + \underset{\text{conjugate base}}{Cl^-}$$

lose H^+

Table Comparing the Arrhenius Theory and the Brønsted-Lowry Theory

Theory	Arrhenius	Brønsted-Lowry
Acid	any substance that dissociates to form H^+ in aqueous solution	any substance that provides a proton to another substance
Base	any substance that dissociates to form OH^- in aqueous solution	any substance that receives a proton from an acid
Example	$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$	$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

The seven **strong acids**:

1.	hydrochloric acid,	HCl ✓
2.	hydrobromic acid,	HBr ✓
3.	hydroiodic acid,	HI ✓
4.	nitric acid,	HNO_3 ✓
5.	chloric acid,	$HClO_3$
6.	perchloric acid,	$HClO_4$
7.	sulfuric acid,	H_2SO_4 ✓

Common **strong bases**:

1.	sodium hydroxide,	$NaOH$ ✓
2.	potassium hydroxide,	KOH ✓
3.	calcium hydroxide,	$Ca(OH)_2$ ✓
4.	strontium hydroxide,	$Sr(OH)_2$ ✓
5.	barium hydroxide,	$Ba(OH)_2$ ✓
	Lithium Hydroxide	$LiOH$

$$pH = -\log[H^+_{(aq)}] \qquad [H^+_{(aq)}] = 10^{-pH}$$

$$pOH = -\log[OH^-_{(aq)}] \qquad [OH^-_{(aq)}] = 10^{-pOH}$$

$$pH + pOH = 14$$

weak acids

Acetic acid	CH_3COOH
aq Carbon dioxide	$CO_{2(aq)}$
aq Sulfur dioxide	$SO_{2(aq)}$
Hydrofluoric acid	$HF_{(aq)}$
Hydrocyanic acid	HCN

weak bases

AMMONIA	NH_3
AMINE	$R-NH_2$

Al, ZN, Ag not multivalent

Unit 5: Gases and Atmospheric Chemistry (Page 1 of 1)

The **kinetic molecular theory of gases** makes the following assumptions:

gas particles are much smaller than spaces b/w them

- The volume of an individual gas molecule is negligible compared to the volume of the container holding the gas. This means that individual gas molecules, with virtually no volume of their own, are extremely far apart and most of the container is "empty" space.
- There are neither attractive nor repulsive forces between gas molecules.
- Gas molecules have high translational energy. They move randomly in all directions, in straight lines.
- When gas molecules collide with each other or with a container wall, the collisions are perfectly elastic. This means that when gas molecules collide, somewhat like billiard balls, there is no loss of kinetic energy.

  T Avg kinectic energy of gas
- The average kinetic energy of gas molecules is directly related to the temperature. The greater the temperature, the greater the average motion of the molecules and the greater their average kinetic energy.

  ↑T ↑ kinetic ~~motion~~ energy

  ↳ speed of gas particle changes directly w kelvin temp.

$V \propto \frac{1}{P}$

Equation	**Variables**	**Name**
$P_1V_1 = P_2V_2$	P = Pressure of gas (either kPa or atm) V = Volume of gas (L) T = Temperature in Kelvin (K) starts at absolute 0 direct relate to avg kinetic energy (temp) ∴ 0K = particles don't move. (minimal) n = amount (mol) R = gas constant = 8.314 kPa·L·mol^{-1}·K^{-1}, or = 0.08206 atm·L·mol^{-1}·K^{-1} $PV=nRT$ $R = \frac{PV}{nT} = \frac{kPa \cdot L}{mol \cdot K}$ not constant M = molar mass m = mass (g) d = density (g/L)	Boyle's law (volume & pressure)
$\frac{V_1}{T_1} = \frac{V_2}{T_2}$		Charles's law (volume & Temperature) $T \propto V$
$\frac{P_1}{T_1} = \frac{P_2}{T_2}$		Gay-Lussac's law (pressure & volume) $P \propto T$
$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$		combined gas law
$\frac{V_1}{n_1} = \frac{V_2}{n_2}$		Avogadro's Law (volume & moles)
$PV = nRT$ $R = \frac{PV}{nT}$ $PVM = mRT$ $PM = dRT$ * $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$		ideal gas law
$d = \frac{M}{V_m}$, $V_m = \frac{V}{n}$	d = density M = molar mass V_m = molar volume $d = \frac{PM_{gas}}{RT}$	density
$P_{total} = P_A + P_B + \ldots$	$P_A = P_{total} \times \frac{n_A}{n_{total}} = P_{total} \times \frac{V_A}{V_{total}}$ $\frac{P_{gas}}{P_{Total}} = \frac{n_{gas}}{n_{total}} = \frac{V_{gas}}{V_{total}}$	Dalton's law of partial pressure

* [K] = [°C] + 273.15 [K]: Kelvin, [°C]: degree Celsius ; 1 L = 1000 cm^3

Memorized! * 1 standard atmosphere = 1 atm = 101.3 kPa = 760 mm Hg = 760 torr = 14.7 lb·in^{-2}

$D = 4.00\ kPa \times \frac{1\ atm}{101.3\ kPa} =$

* **STP**: P = 101.3 kPa, T = 273K. At STP, volume of 1 mole of gas = 22.4 L

SATP: P = 100 kPa, T = 298K. At SATP, volume of 1 mole of gas = 24.8 L

maybe w dAP

1 mole of gas contains 6.02 × 10^{23} molecules of the gas.

$P_{absolute} = P_{relative} + P_{atmospheric}$

$P_{gas} = P_{atm} + \Delta h$ if $P_{gas} > P_{atm}$

$P_{gas} = P_{atm} - \Delta h$ if $P_{gas} < P_{atm}$

pressure: force per unit area.
measure of the frequency & force of collisions of gas particles w the walls of their container