King Khalid University

College of Science

Department of Chemistry

جامعة الملك خالد

KING KHALID UNIVERSITY

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كلية العلوم

قسم الكيمياء

Lecture notes

General Chemistry

1413Chem-4

For Engineering Students

These lecture notes are largely based on the following book:
Understanding General Chemistry
by Atef Korchef, Associate Professor at King Khalid University

CRC Press, Taylor and Francis Group, 1st Edition, 2022.
ISBN 9781032189147
https://www.routledge.com › Korchef › book

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Introduction to the course

General chemistry course 1413Chem-4 designed for engineering students delivers the fundamentals of general chemistry through a wide range of topics, relating the structure of atoms and molecules to the properties of matter. This course incorporates both lectures and laboratory experiments in developing and understanding chemical concepts and practices.

- Theory 3 hours/week
- Practice 2 hours/week

Chapters:

- Chapter 1. Matter and Measurements
- Chapter 2. Atoms, Molecules, Ions, and Moles
- Chapter 3. Calculations in Chemistry
- Chapter 4. Electron Configuration and Properties of Atoms
- Chapter 5. Molecular geometry
- Chapter 6. Intermolecular Forces and Properties of Matter.
- Chapter 7. Gases

Assessment details

- Quiz: 5 marks
- Homework: 5 marks
- Midterm exam: 30 marks
- Practical exam: 20 marks
- Final exam: 40 marks
- Total: 100 marks

Chapter 1

Matter and Measurements

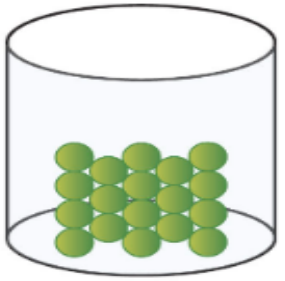
At the end of this chapter, the student will be able to

- define chemistry and matter and its states,
- differentiate between chemical change (chemical properties) and physical change (physical properties)
- differentiate between basic and derived quantities
- differentiate between intensive and extensive properties
- choose the appropriate units in the système internationale (SI) of measurement units
- define and use the basic methods and tools used in measurements such as significant figures and rounding off.
- define accuracy and precision

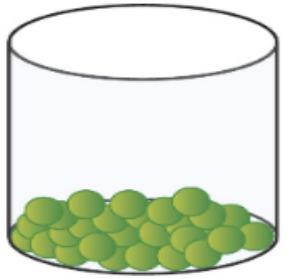
1.1. Matter

- **Chemistry** is the study of matter and the changes that matter undergoes
- **Matter** is defined as anything that **has a mass** and occupies **space** (volume) and is made up of particles.

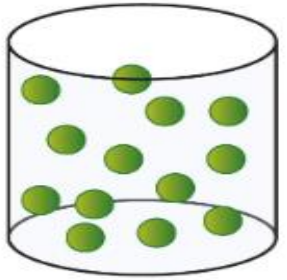
- **1.1.1. States of Matter**
  - **Solid**
    - particles close together in an orderly fashion
    - little freedom of motion
    - ***a solid has a fixed volume and shape***


Solid

  - **Liquid**
    - particles close together but not held rigidly in position
    - particles are free to move past one another
    - ***a liquid sample has a fixed volume but conforms to the shape of the part of the container that it fills***

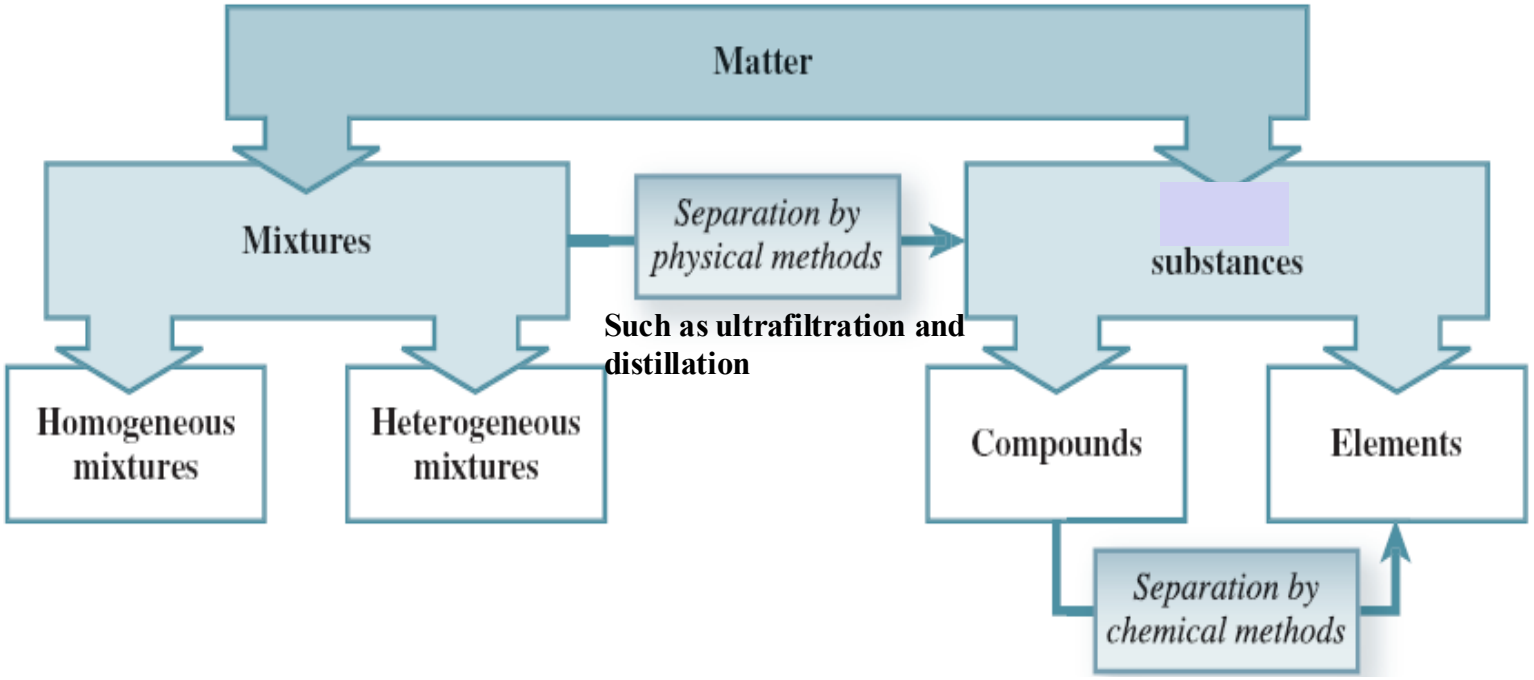

Liquid

  - **Gas**
    - particles randomly spread apart
    - particles have complete freedom of movement
    - ***a gas sample assumes both the shape and volume of container.***


Gas

1.1.2. Classification of Matter and methods of separation

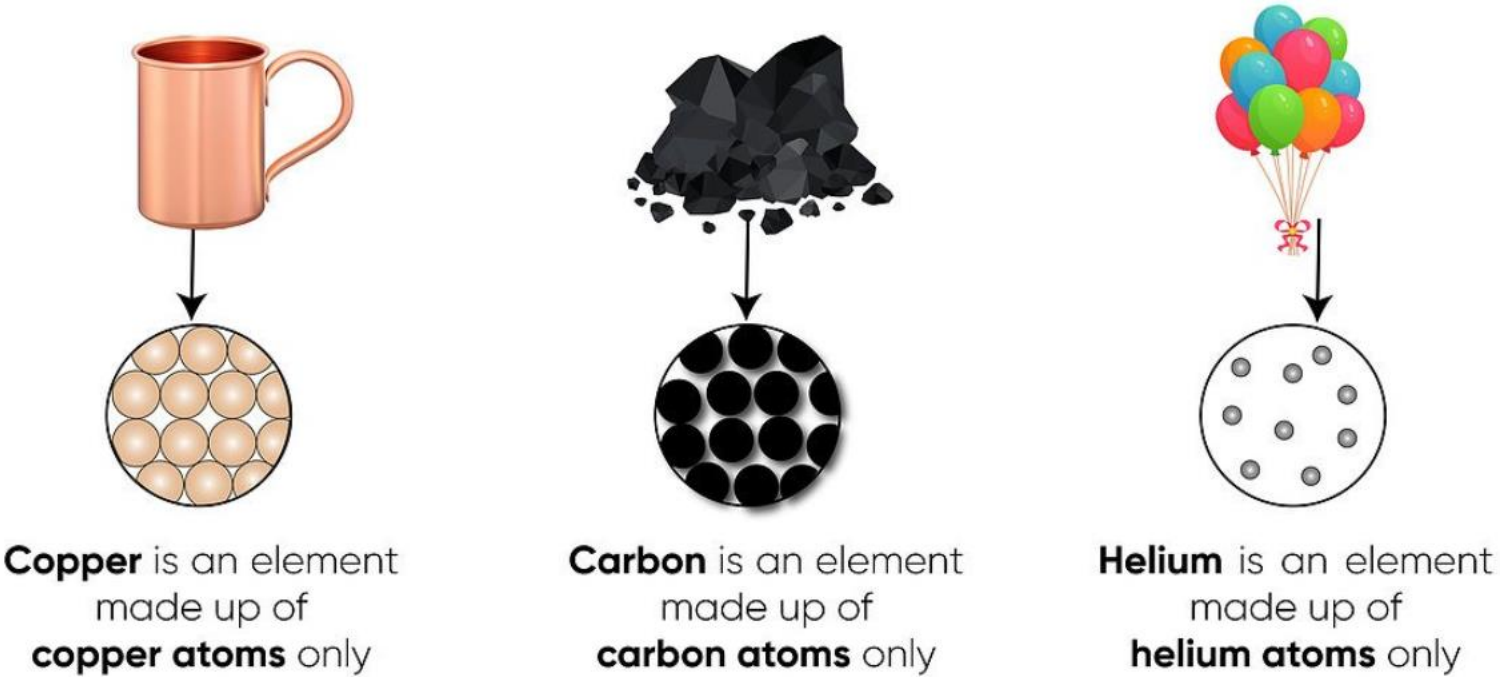
Matter is either classified as a **substance** or a **mixture** of substances. Substance can be either an **element** or a **compound.** A mixture can be either **homogeneous** or **heterogeneous.** Compounds could be separated to their elements by chemical methods such as **heat decomposition** and **electrolysis**. Mixtures could be separated to pure substances by physical methods such as **ultrafiltration** and **distillation**



Substances

- ***Element*** *is made up of* ***one kind of atom*** and cannot be separated into simpler substances by chemical means.

Examples 1.1. ***Carbon (C), Phosphorus (P), oxygen (O), hydrogen (H), Nitrogen (N), magnesium (Mg), Copper (Cu), Helium (He), iron (Fe),…***

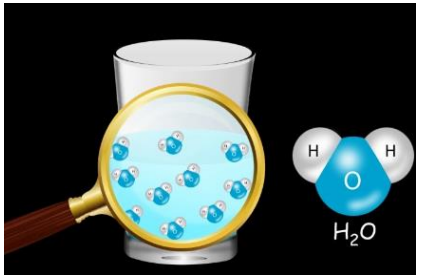


https://www.teachoo.com/12688/3446/Atom/category/Concepts/

Atom is the building-up unit of the element.

- ***A compound*** *is formed when* two or more elements are chemically combined in definite ratios.
- The compound is made up of **at least two kinds of atoms.**
- The properties of the compounds are different from the properties of their elements.

Examples 1.2: salt (NaCl), water H_2O, ethane (C_2H_6), carbon dioxide CO_2, HgO, $CaCO_3$, $CaSO_4 \cdot H_2O$, Na_2CO_3, $Mg(OH)_2$



Elements | Atoms | Molecules & Compounds | Tutway

- *Compounds can be separated into their elements by chemical means* such as:

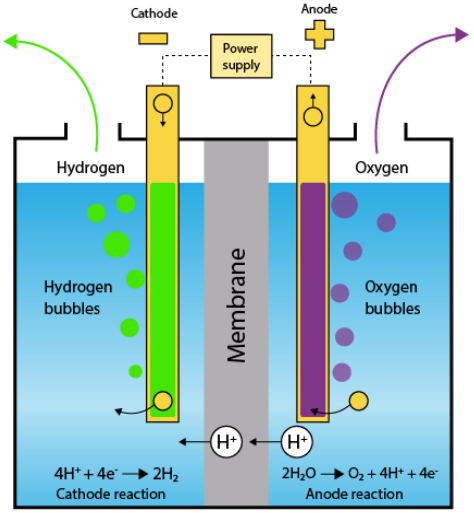
Heat decomposition of mercuric oxide (HgO)

$$HgO_{(sd)} \longrightarrow Hg_{(lq)} + \tfrac{1}{2} O_{2(g)}$$



Electrolysis of water :

$$H_2O_{(lq)} \longrightarrow H_{2(g)} + 1/2O_{2(g)}$$



Mixtures

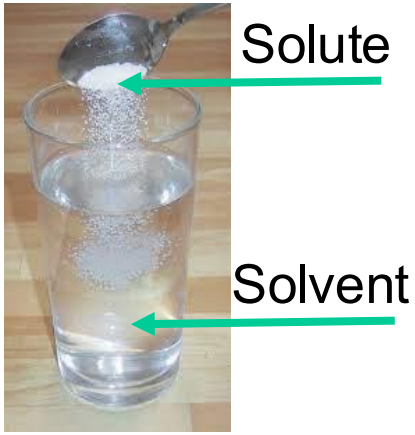
- ***Mixture:*** physical combination of two or more pure substances
  - Substances retain distinct identities
- **Types of Mixtures**
  - ***Homogeneous* Mixtures** : A homogeneous mixture is one in which the components are uniformly distributed. The composition of the mixture is *uniform* throughout
    - ***Examples 1.3.*** **sugar dissolved in water, air, seawater, NaOH solution (and all solutions), …**

❑ **Solution is a homogenous mixture** of two or more substances that are chemically **unreacted**.

A solution is made up of two components: the solute and the solvent:

• **Solute:** the substance present in the smaller amount and dissolved in the solvent.

• **Solvent:** the substance present in the larger amount and responsible for dissolving the solute.



Example 1.4: in NaCl solution, NaCl (sd) is the solute and water H_2O is the solvent.

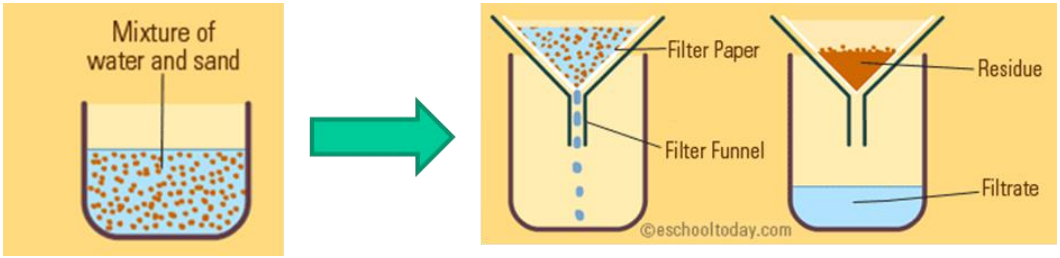
***Heterogeneous* Mixtures** : mixture is made of different substances that remain physically separate. Heterogeneous mixtures always have more than one phase (composition is *not* uniform throughout)

- ***Example* 1.5**: **sugar mixed with iron filings, sand and chalk powders in water, water and oil**

Mixtures can be separated to their components by different physical methods such as filtration and distillation:

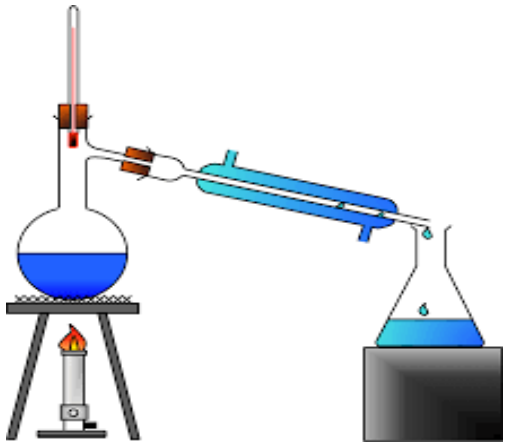
➢ **Filtration: used to separate a heterogeneous solid-liquid mixture**

Example 1.6: Mixture of water and sand



➢ **Distillation: used to resolve a homogenous solid-liquid mixture**

Example 1.7: Salt NaCl in water, seawater



1.1.3. Properties of matter

❑ **Chemical properties:** The ability of a substance to combine with or change into one or more other substances

Examples 1.8. Heat of combustion, enthalpy of formation, electronegativity

❑ **Physical properties:** Characteristics that can be observed or measured without changing the composition of the substance

Examples 1.9. Temperature, color, volume, mass, area, pressure, melting point, boiling point

- **Intensive properties** are properties which do not depend on the amount of matter

  Examples 1.10. **Temperature, density, boiling point, concentration, solubility, color**

- **Extensive properties** are properties which depend on the amount of matter

  Examples 1.11. **mass, volume, surface (area), amount of substance (number of moles)**

1.1.3.1. Chemical Change

- A chemical change involves making or breaking chemical bonds to create new substances. Chemical changes include the followings:

- **Oxidation reduction (Redox) reactions**: loss and gaining of electrons:

The oxidation of iron(II) to iron(III) by hydrogen peroxide in acidic medium

$Fe^{2+} \leftrightarrow Fe^{3+} + e^-$

$H_2O_2 + 2H^+ + 2e^- \leftrightarrow \mathbf{2H_2O}$

Overall equation: $\mathbf{2Fe^{2+} + H_2O_2 + 2H^+ \longrightarrow 2Fe^{3+} + 2H_2O}$ **(Redox reaction)**

- **Reaction of base and acid** (**neutralization reaction**):

$$HCl_{sol} + NaOH_{sol} \longrightarrow H_2O_{\ lq} + NaCl_{\ aq}$$

- **Heat decomposition**: $HgO_{(sd)} \longrightarrow Hg_{(lq)} + \frac{1}{2} O_{2\ (g)}$
- **Electrolysis of water:** $H_2O_{(Lq)} \longrightarrow H_{2(g)} + 1/2O_{2(g)}$
- **Iron rusting (corrosion):** $4Fe + 3O_2 \longrightarrow 2Fe_2O_3$
- **Combustion (**of ethane C_2H_6**):** $2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$
- **Burning** of wood

1.1.3.2. Physical Change

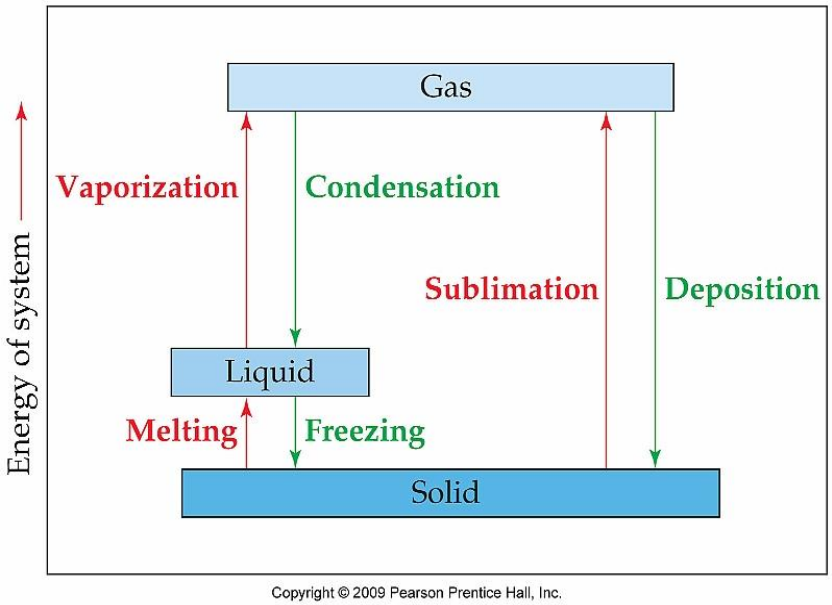
- A physical change alters a substance without changing its chemical identity.
- No new substance is created, no formation of new chemical bonds during a physical change.
- Physical changes occur when substances are mixed but do not chemically react

- **<u>Physical changes include the following:</u>**
  - Dissolving sugar and salts in water
  - Mixing sand with water or mixing oil with water
  - Crushing
  - Phase changes such as vaporization, condensation, freezing, sublimation, melting and deposition.

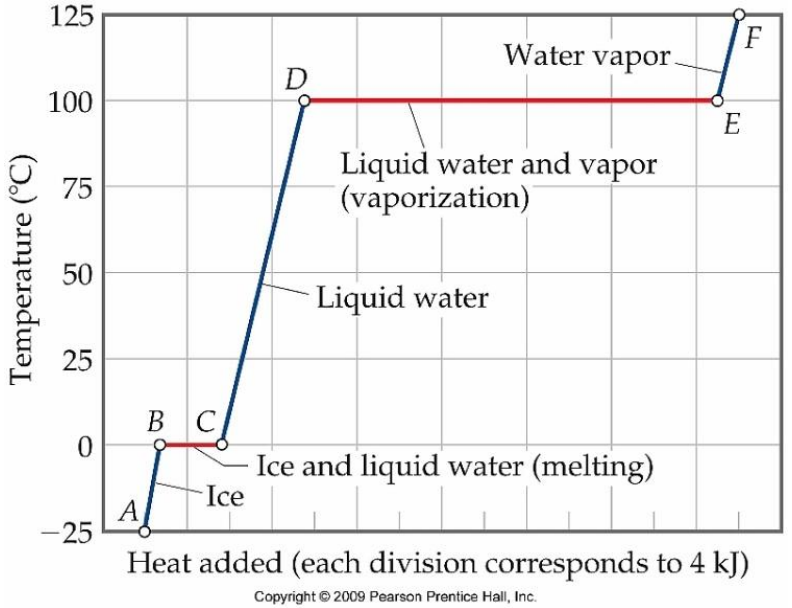
Example 1.12. $H_2O_{(lq)} \longrightarrow H_2O_{(g)}$

Phase Changes

- Phase changes are transformations from one phase to another. The energy required to go from one state to another state is called **enthalpy change (ΔH).**
- **melting, vaporization and sublimation** are ***endothermic*** *processes* (**need** energy, **ΔH>0**).
- **freezing, condensation and deposition** are ***exothermic*** *processes* (**release** energy, **ΔH<0**).
- **The temperature remains constant during phase changes** (energy is used to overcome attractive forces between molecules).



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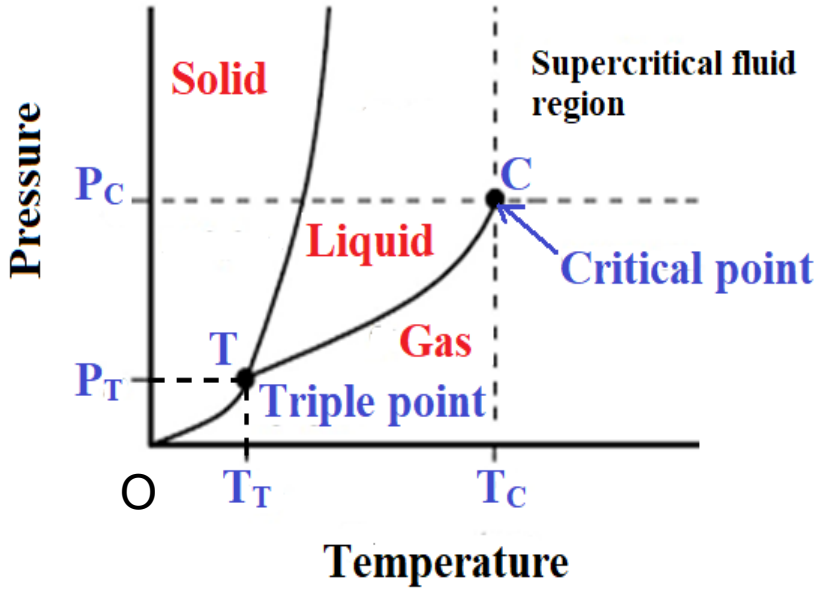


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Exercises Q1.1 to Q1.4, p.7-10

Phase Diagram

- A phase diagram is a graphic way to summarize the conditions under which equilibriums exist between the different states of matter.



- **The line *T* to *C*:** equilibrium between the liquid and the gas phases.

  The line *T* to *A*: equilibrium between the liquid and the solid phases.

  The line *O* to *T*: equilibrium between the solid and the gas phases.

- **The critical point *C***, is defined by the critical temperature (T_C) and the critical pressure (P_C). Beyond that point, the liquid and gas phases become indistinguishable, and the state is a supercritical fluid.
  - **Critical temperature** is the highest temperature at which a liquid can exist.
  - **Critical pressure** is the pressure required to create a liquid at this critical temperature.

- **Triple point T**: Liquid, solid and gas phases coexist in equilibrium.



1.2. PHYSICAL QUANTITIES
and the International System of units (SI units)

- ***Fundamental (basic) quantities*** cannot be defined in terms of other physical quantities

Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

- ***Derived quantities*** can be defined in terms of the fundamental physical quantities

quantity	***area***	***Volume***	***density***	**concentration**	**Pressure**
SI derived unit	**m^2**	**m^3**	**kg/m^3**	**mol/m^3**	**Pa (Pascal)**

$\rho = m/V$

m: mass and V: volume

$P = F/A$

F: force and A: area

SI Prefixes

Prefix	Symbol	Meaning	Example
Tera-	T	1×10^{12} (1,000,000,000,000)	1 teragram (Tg) = 1×10^{12} g
Giga-	G	1×10^{9} (1,000,000,000)	1 gigawatt (GW) = 1×10^{9} W
Mega-	M	1×10^{6} (1,000,000)	1 megahertz (MHz) = 1×10^{6} Hz
Kilo-	k	1×10^{3} (1,000)	1 kilometer (km) = 1×10^{3} m
Deci-	d	1×10^{-1} (0.1)	1 deciliter (dL) = 1×10^{-1} L
Centi-	c	1×10^{-2} (0.01)	1 centimeter (cm) = 1×10^{-2} m
Milli-	m	1×10^{-3} (0.001)	1 millimeter (mm) = 1×10^{-3} m
Micro-	μ	1×10^{-6} (0.000001)	1 microliter (μL) = 1×10^{-6} L
Nano-	n	1×10^{-9} (0.000000001)	1 nanosecond (ns) = 1×10^{-9} s
Pico-	p	1×10^{-12} (0.000000000001)	1 picogram (pg) = 1×10^{-12} g

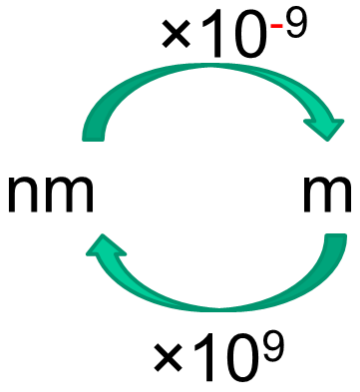
Practice 1.1. Convert the following into the required unit:

1 nm = ……. m

1m = ……nm

2 ng = ……mg

2.126 mg = ……..μg



Answer:

1 nm = 10^{-9} m

1m = 10^{9} nm

2 ng = 2×10^{-9} g = 2×10^{-9} ×10^{3} mg = 2×10^{-6} mg

2.126 mg = 2.126×10^{-3} g = 2.126×10^{-3}×10^{6} μg = 2.126×10^{3} μg = 2126 μg

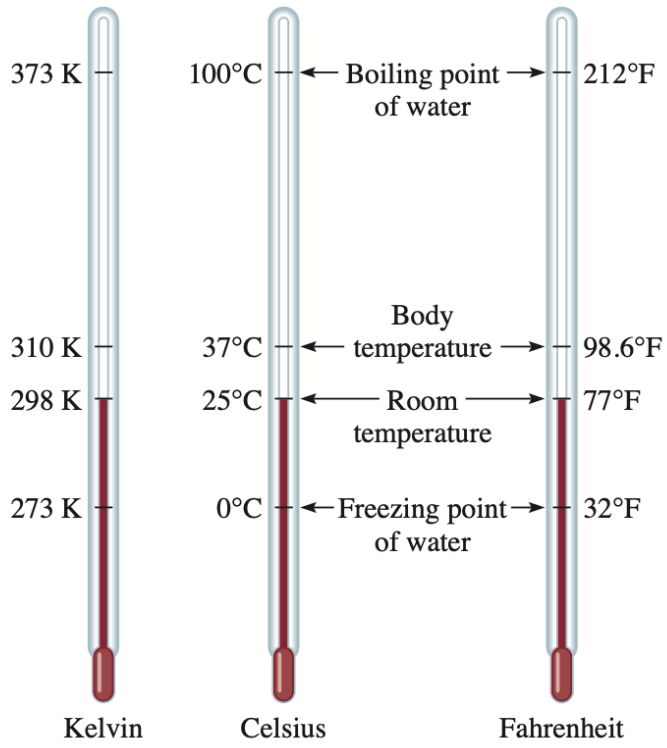
Temperature:

- **Celsius**
  - Represented by °C
  - Based on freezing point of water as 0°C and boiling point of water as 100°C
- **Kelvin**
  - Represented by K (no degree sign)
  - The *absolute* scale
- **Fahrenheit** (the English system) (°F)

- ***The SI unit of temperature is Kelvin (K)***

Equations for Temperature Conversions:

- $T\,(K) = T(°C) + 273$
- $T\,(°F) = 1.8 \times T(°C) + 32$



Practice on temperature conversions

Practice 1.2. **A clock on a local bank reported a temperature reading of 28 °C. What is this temperature in K and °F?**

Answer:

- $T\,(K) = T(°C) + 273 = 28\,°C + 273 = 301\,K$

28°C =(1.8×28)+32 °F= 82.4°F

Practice 1.3. **child has a temperature of 104 ° F. What is his temperature in (° C) and (K)?**

Answer: $T(°C) = \frac{[T(°F)-32]}{1.8} = \frac{[104-32]}{1.8} = 40\,°C$

$$T\,(K) = T(°C) + 273 = 40\,°C + 273 = 313\,K$$

Practice 1.4 **A temperature of 300 K, when expressed in °F, would be**

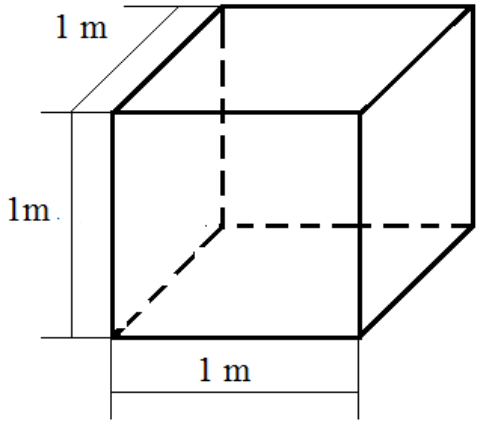
a) 392 °F. **b) 80.6 °F.** c) 27 °F. d) 0 °F.

300 K = (300-273)°C =27°C =(1.8×27)+32 °F= 80.6°F

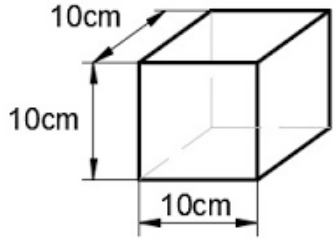
- **Volume** is an extensive property. Volume is a derived quantity.

  The **SI derived unit of volume is m^3 (**meter cubed)

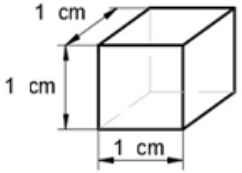
  but the unit liter (L) is more commonly used in the laboratory setting.



$1m^3 = 1000\ L$



$1dm^3 = 1L$



$1cm^3 = 1mL$

Practice 1.5. Convert the following to the desired unit:

a) $1m^3$ = …… mL.

b) 1 nL= …… cm^3

Answer:

a) $1m^3 = 1000\ L = 10^3\ L = 10^3 \times 10^3\ mL = 10^6\ mL$.

b) $1\ nL = 1 \times 10^{-9}\ L = 1 \times 10^{-9} \times 10^3\ mL = 10^{-6}\ mL = 10^{-6}\ cm^3$

- **Density**: Density is an intensive property (does not depend on the amount of matter). Density is a derived quantity.

$$\rho = m/V$$

ρ = density; *m* = mass and *V* = volume

  - **The SI unit for density is kg/m^3**

There are other common units for density such as (g/cm^3 or g/mL, g/L)
(*gas densities are usually expressed in g/L)

❑ **Practice 1.6.** A sample of mercury has a volume of 0.25 L and a mass of 3.4 kg. Calculate the density of mercury in g/cm^3?

Answer:

1 L = 1000 cm^3, Volume = 0.25 x 1000 = 250 cm^3

1 kg = 1000 g, Mass = 3.4 x 1000 = 3400 g

ρ = m/v = 3400 g/ 250 cm^3 = 13.6 g/cm^3

❑ **Practice 1.7.** **A cube of metal has a mass of 4.0 g and a length of 2.0 cm on each side. Calculate its density.**

a) 2.0 g/cm^3 **b) 0.50 g/cm^3** c) 0.20 g/cm^3 d) 5.0 g/cm^3

V =2×2×2 = 8 cm^3; ρ = m/ V = 4/8 = 0.5 g/cm^3

❑ **Solubility** is the maximum quantity (amount) of the solute that will dissolve in a given amount of the solvent, at a given temperature. Solubility depends on temperature.

The solubility may be stated in various units of concentration such as g/L, mol/L and g/100g of solvent.

Example 1.13: The solubility of sugar in water at 20 ^{o}C is 2040 g/L.

Solubility is an intensive property

Solubility is a derived quantity

Practice 1.8. The solubility of potassium nitrate (KNO_3) in water is 270 g/L at 20 ^{o}C. **Calculate the mass** of KNO_3 dissolved in 17.8 mL of water at 20 ^{o}C.

Answer: at 20 ^{o}C,

270g KNO_3 → 1000 mL water (=1L)

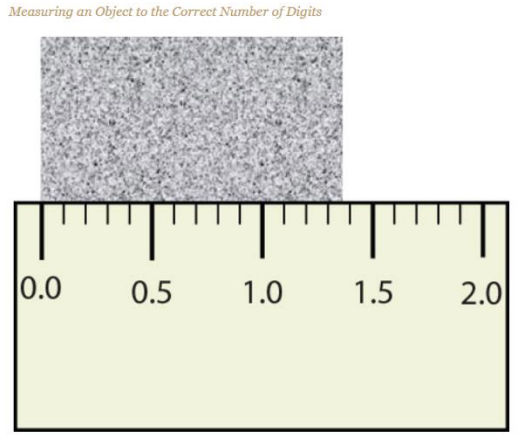
Mass KNO_3 → 17.8 mL water

Mass KNO_3 = (17.8 × 270) / 1000 = **4.8 g**

Exercises Q1.3 p.7 and Q2.6 p.36

1.3. Measurements

To measure the width of this object, you cannot report the object's width any more exactly than to the nearest millimeter. All you can do is estimate the next decimal place in the measurement. *It makes no sense to report any digits after the first uncertain one.*



The concept of significant figures takes this limitation into account. The **significant figures** of a measured quantity are defined as all the digits known with certainty and the first uncertain, or estimated, digit.

The length of the object:

1.38 cm

Certain digit

Certain digit

Uncertain digit or Estimated digit

1.3.1. Rules for significant figures

- **Any non-zero digit is significant**

3455 km 4 significant figures (4 SF)
28.7 s 3 significant figures (3 SF)

- **Zeros between non-zero digits are significant**

12051 m 5 significant figures
25400032 nL 8 significant figures
208.0025 cm 7 significant figures

- **Zeros to the left of the first non-zero digit are not significant**

0.0058 mol 2 significant figures

- **Zeros to the right of the last non-zero digit are significant if a decimal is present**

25.00 mL 4 significant figures
30.0 g 3 significant figures

- **Are zeros to the right of the last non-zero digit significant if the decimal is not present?**

7500 m

It is an ambiguous case:
Do we mean 7500 m, measured to the nearest meter? Nearest 10 meters?

The best approach here is to use exponential notation (scientific notation):

$$m \times 10^n$$

Where m is a number written with only one digit before the decimal point. The coefficient m establishes the number of significant figures, and the power of ten locates the decimal point:

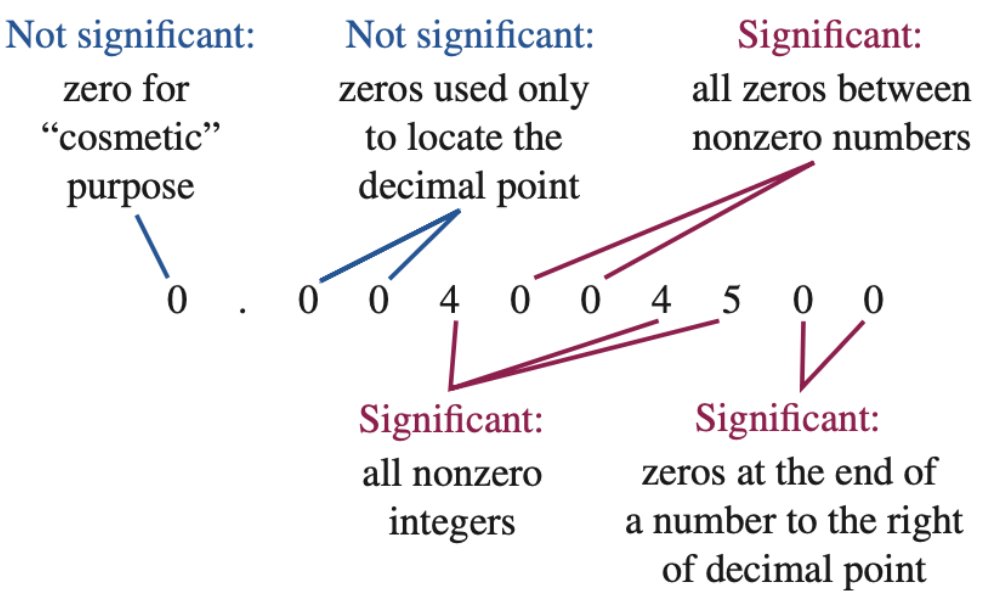
2 significant figures	*3 significant figures*	*4 significant figures*
7.5×10^3 m	7.50×10^3 m	7.500×10^3 m

Practice 1.9. How many significant figures are there in the measurements 0.004004500 km, **and** 1.03×10^{-3} m?

Answer

0.004004500 km

7 significant figures



1.03×10^{-3} m (Scientific notation)

3 significant figures

1.3.2. Simplified Rounding rules

- If the number is less than 5 round "down".

0.242313 → (Rounded off to 2 SF) → *0.24*

3.674 → (Rounded off to 3 SF) → *3.67*

- If the number is greater than 5 round "up".

0.0679 → (Rounded off to 1 SF) → *0.07*

1.5396 → (Rounded off to 4 SF) → *1.540*

- When the first digit dropped is 5, and 5 is the last digit or followed by zeros, make the preceding digit **even**

2.0845 → (Rounded off to 4 SF) → *2.084*

2.132755 → (Rounded off to 5 SF) → *2.1328*

3.305 → (Rounded off to 3 SF) → 3.30 **Note that zero is an even number**

0.074500 → (Rounded off to 2 SF) → *0.074* *(5 is followed by zeros)*

0.07457 → (Rounded off to 2 SF) → *0.075* *Note that 5 is not the last digit*

1.3.3. Calculations with measured numbers

❑ *Addition and subtraction*

Answer cannot have more digits to the right of the decimal than any of original numbers

- ***Example 1.12.*** **102.50 + 0.231 – 12.1 = 90.6**

  2 digits after decimal point | 3 digits after decimal point | 1 digit after decimal point | 1 digit after decimal point

  Calculator answer: 90.631 rounded off to 90.6.

❑ *Multiplication and division*

Final answer contains **the smallest number of significant figures**

- ***Example 1.13.*** **1.4 x 8.011 / 5.12 = 2.2**

  2SF 4 SF 3 SF 2SF

(Limited by 1.4 to 2 significant figures in answer)

(Calculator answer: 2.1905078125 round to 2.2)

Note that exact numbers do not limit answer because exact numbers have an infinite number of significant figures

Example 1.14: A penny minted has a mass of 2.5 g. If we have 3 pennies, the total mass is

3 x 2.5 g = 7.5 g

In this case, 3 is an exact number and does not limit the number of significant figures in the result.

PRACTICE 1.10. **Calculate the following to the correct number of significant figures:**

(2.20 - 0.0111)/0.1165 = ?

- **(2.20 - 0.0111) = ?** **(correct answer: 2.19 (2 decimals))**

  (2 decimals) (4 decimals) (2 decimals)

  – Calculator answer: **2.1889** → 2.19 (with 2 decimals)

- **2.19/0.1165 = ?**

  3SF 4SF 3SF

  – Calculator answer: **18.7982832618** → **18.8 (3 SF)**

⇨ **(2.20 - 0.0111)/0.1165 = 18.8**

PRACTICE 1.11. **Calculate the following to the correct number of significant figures:**

(8.4 × 46) = …….

(2SF) (2SF) (2SF)

Calculator answer: **386.4** (4SF)

$386.4 = \underline{3.86}4 \times 10^2$ → 3.9×10^2 **(scientific notation)**

rounded off to 2 SF

(8.4 × 46) = $\underline{3.9} \times 10^2$

(2SF) (2SF) (2SF)

Exercise Q2.7 p.36

1.3.4. Accuracy and precision

- **Accuracy** is how close a measurement is to the accepted value (or true value).
- **Precision** is how closely measurements of the same thing are to one another

Practice 1.14. The true value of a neutralization volume is 12.370 mL. Describe accuracy and precision for each set of measurements:

- *Set A* (12.345 mL, 12.346 mL, 12.344 mL) ⟹ Precise, not accurate
- *Set B* (12.357 mL, 12.337 mL, 12.393 mL) ⟹ Not precise, not accurate
- *Set C* (12.369 mL, 12.370 mL, 12.371 mL) ⟹ Precise and accurate

Explanation:

- The values in set A are close to one another but are not close to the true value so set A is precise but not accurate.
- The values in set B are not close to one another and are not close to the true value so set B is neither precise nor accurate
- The values in set C are close to one another and close to the true value so set C is both precise and accurate.

Exercise Q2.5 p.33

Practice on chapter 1

Q1. $CaSO_4 \cdot 1/2H_2O$ is

a) an element b) a compound c) a heterogeneous mixture d) a homogeneous mixture

Q2. The density of water at 20°C is 0.9982 g/mL. Which of the following sets of density measurements is precise but not accurate?

a) (0.9989, 0.9993, 0.9991) b) (0.9981, 0.9980, 0.9982)

c) (0.9881, 0.9880, 0.9882) d) (0.8989, 0.8993, 0.8991)

Q3. Which of the following is a derived quantity?

a) Mass b) Volume c) Temperature d) Amount of substance

Q4. The correct answer to express the arithmetic operation (2.00×0.009)/(9.410-8.0) is

a) 0.014 b) 0.01 c) 0.0142 d) 0.0143

Q5. Electrolysis of water is

a) a physical change b) a chemical change c) a neutralization process d) none of these

Q6. The number of significant figures (SF) in 0.0100 is

a) 1 SF. b) 3 SF. c) 4 SF. d) 5 SF.

Q7. A temperature of 173 K, when converted to °C equals

a) 273 °C b) 446 °C c) -100 °C d) 100 °C

Q8. Rounding off 3.1055 to four significant figures results in

a) 3.105 b) 3.106 c) 3.1057 d) 3.100

Q9. 600 nm equals

a) 6×10^{-3} m b) 0.6 µm c) 6×10^{-2} µm d) 6×10^{3} m

Q10. Which of the following is an intensive property?

a) Volume b) Mass c) Amount of substance d) Density

Q11. Which of the following is the smallest measure of the mass of a substance?

a) 2 Mg b) 2 mg c) 2 ng d) 2 µg

Q12. Which of the following is an endothermic process?

a) Sublimation. b) Freezing. c) Condensation d) Deposition.

Q13. In a phase diagram, the triple point represents the equilibrium between

a) Liquid and solid. b) liquid and gas. c) liquid, solid and gas. d) solid and gas.

Chapter 2

Atoms, Molecules, Ions and Moles

At the end of this chapter, the student will be able to

- ❑ describe the atomic structure
- ❑ differentiate between the properties of the atom's components
- ❑ differentiate between the atom, the molecule, the ion, and the mole
- ❑ Calculate the number of protons, neutrons, and electrons in an atom or an ion
- ❑ recognize the isotopes of an element
- ❑ differentiate between the chemical formula, the empirical formula, and the structural formula
- ❑ define the ionic compounds

2.1. The Atom Theory

Democritus expressed the belief that all matter consists of very small, indivisible particles, which he named *atomos* (meaning uncuttable or indivisible).

Dalton's atomic theory:

1. Elements are composed of tiny particles called atoms.

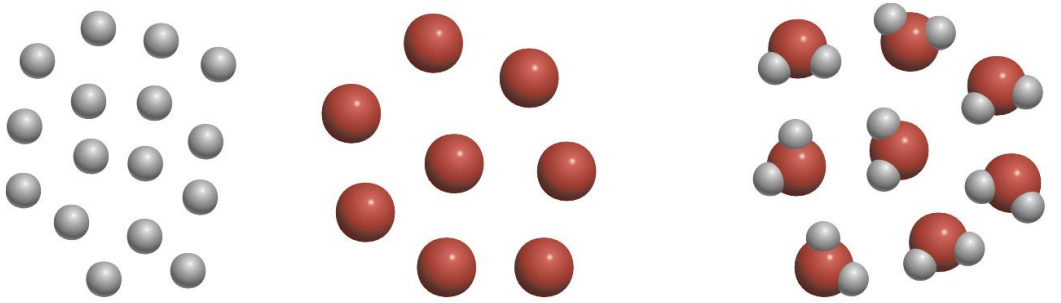
2. All atoms of a given element are identical, having the same size, mass, and chemical properties. Atoms of different elements show different properties.

3. Compounds are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.

4. A chemical reaction involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.

Dalton made no attempt to describe the structure of atoms.

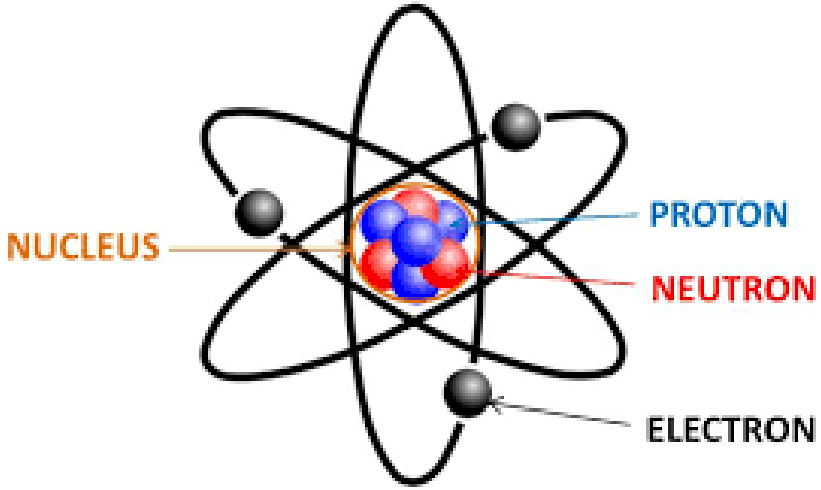
(a) According to Dalton's atomic theory, atoms of the same element are identical, but atoms of one element are different from atoms of other elements. (b) Compound formed from atoms of elements X and Y. In this case, the ratio of the atoms of element X to the atoms of element Y is 2:1.



2.2. The atomic structure

Atom consists of three subatomic particles:
Protons (p) : have a positive charge (+)
Neutrons (n): are neutral and have a mass slightly greater than that of the protons
Electrons (e): have a negative charge (-) and a tiny mass

Most of the atom is an empty space

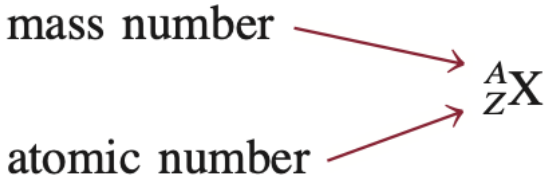


		Charge	
Particle	Mass (g)	Coulomb	Charge Unit
Electron*	9.10938×10^{-28}	-1.6022×10^{-19}	-1
Proton	1.67262×10^{-24}	$+1.6022 \times 10^{-19}$	$+1$
Neutron	1.67493×10^{-24}	0	0

Mass of proton ≈ mass of neutron >> mass of electron

Atom is the smallest particle (building up unit) of an element, having the same chemical properties as the bulk element.

2.3. Symbol of an element



$$Mass\ number\ \boldsymbol{A} = Number\ of\ protons\ \boldsymbol{Z} + Number\ of\ neutrons\ \boldsymbol{N}$$

$$A = Z + N$$

$$The\ number\ of\ neutrons\ N = A - Z$$

For a neutral atom: $The\ number\ of\ electrons\ e- = The\ number\ of\ protons\ Z$

Practice 2.1. How many protons, neutrons and electrons are there in $^{12}_{6}C$?

Answer: For the carbon atom $^{12}_{6}C$: $A = 12\ and\ Z = 6$

$$The\ number\ of\ protons\ Z = 6$$

$$The\ number\ of\ neutrons\ N = A - Z = 12 - 6 = 6\ neutrons$$

$^{12}_{6}C\ is\ a\ neutral\ atom\ (since\ there\ is\ no\ sign\ assigned\ to\ the\ symbol)$:

$$The\ number\ of\ electrons\ = The\ number\ of\ protons\ = 6\ e-$$

Practice 2.2. **How many protons, neutrons and electrons are there in $^{197}_{79}Au$?**

Answer: For the carbon atom $^{197}_{79}Au$: $A = 197\ and\ Z = 79$

$$The\ number\ of\ protons\ Z = 79$$

$$The\ number\ of\ neutrons\ N = A - Z = 197 - 79 = 118\ neutrons$$

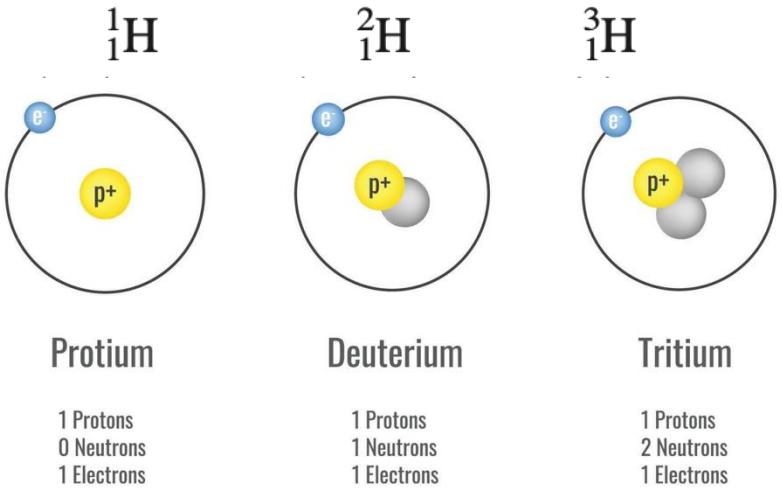
$^{197}_{79}Au\ is\ a\ neutral\ atom\ (since\ there\ is\ no\ sign\ assigned\ to\ the\ symbol)$:

$$The\ number\ of\ electrons\ = The\ number\ of\ protons\ = 79\ e-$$

2.4. Isotopes

Isotopes are atoms of a given element that **differ in the number of neutrons** (**same number of protons**)

Example 2.1. Isotopes of hydrogen



Example 2.2. Isotopes of carbon

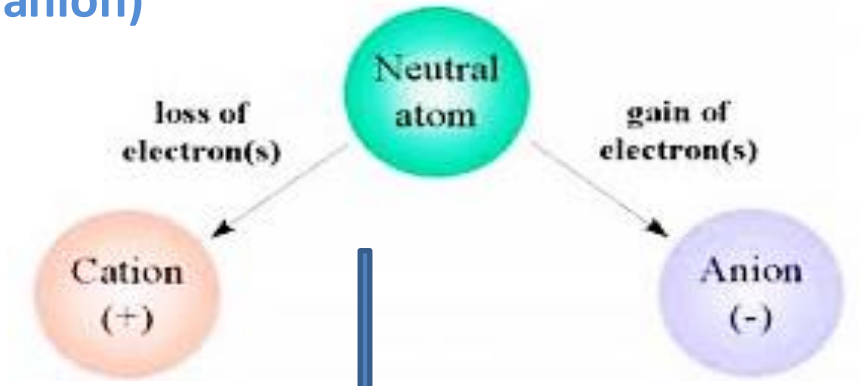
$^{12}_{6}C$ Number of Neutrons = 12 – 6 = 6

$^{13}_{6}C$ Number of Neutrons = 13 – 6 = 7

$^{14}_{6}C$ Number of Neutrons = 14 – 6 = 8

2.5. Ions: Anions and Cations

A neutral atom (or a group of atoms) becomes **an ion** either by **losing electrons (cation)** or by **gaining electrons (anion)**



$Na \longrightarrow Na^{+} + 1e^{-}$

$Mg \longrightarrow Mg^{2+} + 2e^{-}$

$Fe \longrightarrow Fe^{3+} + 3e^{-}$

$Cl + 1e^{-} \longrightarrow Cl^{-}$

$O + 2e^{-} \longrightarrow O^{2-}$

$^{9}_{4}Be^{2+}$

no. of (P) = 4
no. of (e) = 4 - 2 = 2
no. of (n) = 9 - 4 = 5

$^{17}_{8}O^{2-}$

no. of (P) = 8
no. of (e) = 8 + 2 = 10
no. of (n) = 17 - 8 = 9

Symbol	Charge	(P)	(n)	(e)
$^{75}_{33}As^{5+}$	+ 5	33	75-33 = 42	33 - 5 = 28
$^{14}_{7}N^{3-}$	- 3	7	14-7 = 7	7 + 3 = 10

Exercise Q3.4, p.51

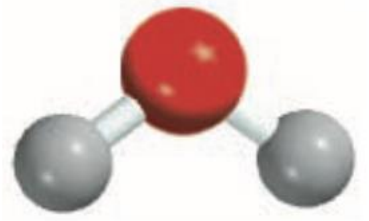
2.6. Molecules

A **molecule** is a neutral **group of two or more atoms** bonded together that may or may not be the same element.

- **Homonuclear molecules** formed by one type of atoms
  Examples 2.3. O_2 , O_3 , H_2 , N_2 , Cl_2 , Br_2 , I_2 , F_2, …



- **Heteronuclear molecules** formed by more than one type of atoms.
  Examples 2.4. H_2O, NH_3 , $NaCl$, $NaOH$, $CaSO_4 \cdot 2H_2O$, Na_2CO_3



Heteronuclear molecule is the smallest particle (building up unit) of a compound.

❑ **Molecular Formula** represents the **actual number and types** of atoms in a molecule

Examples 2.5. H_2 , Cl_2 , H_2O_2 , C_6H_6 , $C_6H_{12}O_6$, $CaCO_3$, Fe_2O_3 , $CaSO_4 \cdot 2H_2O$, ….

❑ **Empirical or Simplest Formula** is the formula that gives only the **relative number of atoms** of each type in a molecule (gives rise to the smallest set of whole numbers of atoms)

Molecular formula = D × (simplest formula)

D means how many times simplest formula repeated to give molecular formula.

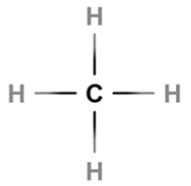
D = 1,2,3…

Examples 2.6.

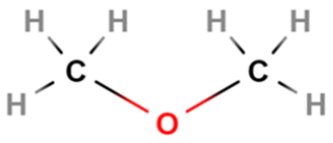
H_2O_2	D=2	Empirical formula: HO
$C_6H_{12}O_6$	D= 6	Empirical formula: CH_2O
H_2O	D=1	Empirical formula: H_2O
$C_3H_6O_3$	D= 3	Empirical formula: CH_2O

❑ **Structural Formula** is the formula that indicates **the attachment of atoms**

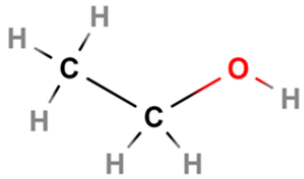
Examples 2.7.



methane CH_4



Ether C_2H_6O



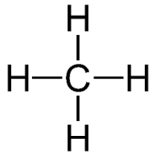
Ethanol C_2H_6O

PRACTICE 2.3

1. Which of the following is the simplest formula of C_6H_{14}?

a) C_6H_{14} b) CH c) C_3H_7 d) CHO

2. Which of the following is the structural formula of methane CH_4?

a) CH_4 b) CH c) C_2H_8 d) 

3. The molecular formula is the formula that

a) gives only the relative number of atoms of each type in a molecule
b) represents the actual number and types of atoms in a molecule
c) indicates the attachment of atoms in a molecule
d) represents a group of atoms of the same element containing different number of protons

4. Which of the following is the building up unit of a compound?

a) Atom b) Homonuclear molecule c) Heteronuclear molecule d) Electron

5. Which of the following is not an ion?

a) H^+ b) NH_4^+ c) O_3 d) SO_4^{2-}

2.7. The mole

Mole - the amount of a substance that contains the same number of entities as there are atoms in exactly 12g of carbon-12.

This amount is 6.022×10^{23}. The number is called Avogadro's number and is abbreviated as N_A. ($N_A = 6.022 \times 10^{23}$)

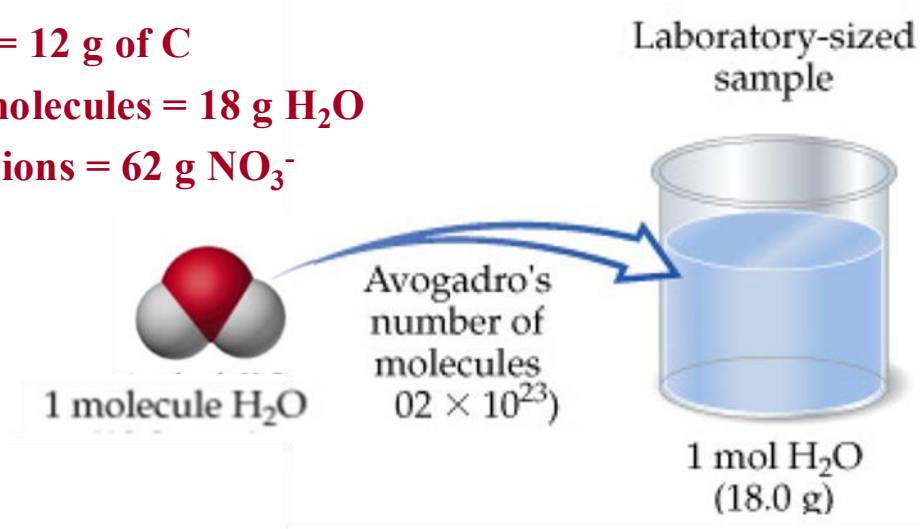
One mole (1 mol) contains $N_A = 6.022 \times 10^{23}$ entities

entities: atoms (ex: H, C, O, Cl, Ca, Fe, …); molecules (ex: H_2, O_2, $CaCl_2$, NaOH, …);

ions (ex: H^+, Cl^-, OH^-, SO_4^{2-}, NH_4^+, …)

Examples 2.8.

- **1 mole of C = 6.022×10^{23} of C atoms = 12 g of C**
- **1 mole of H_2O = 6.022×10^{23} of H_2O molecules = 18 g H_2O**
- **1 mole of NO_3^- = 6.022×10^{23} of NO_3^- ions = 62 g NO_3^-**



2.8. Periodic Table of Elements

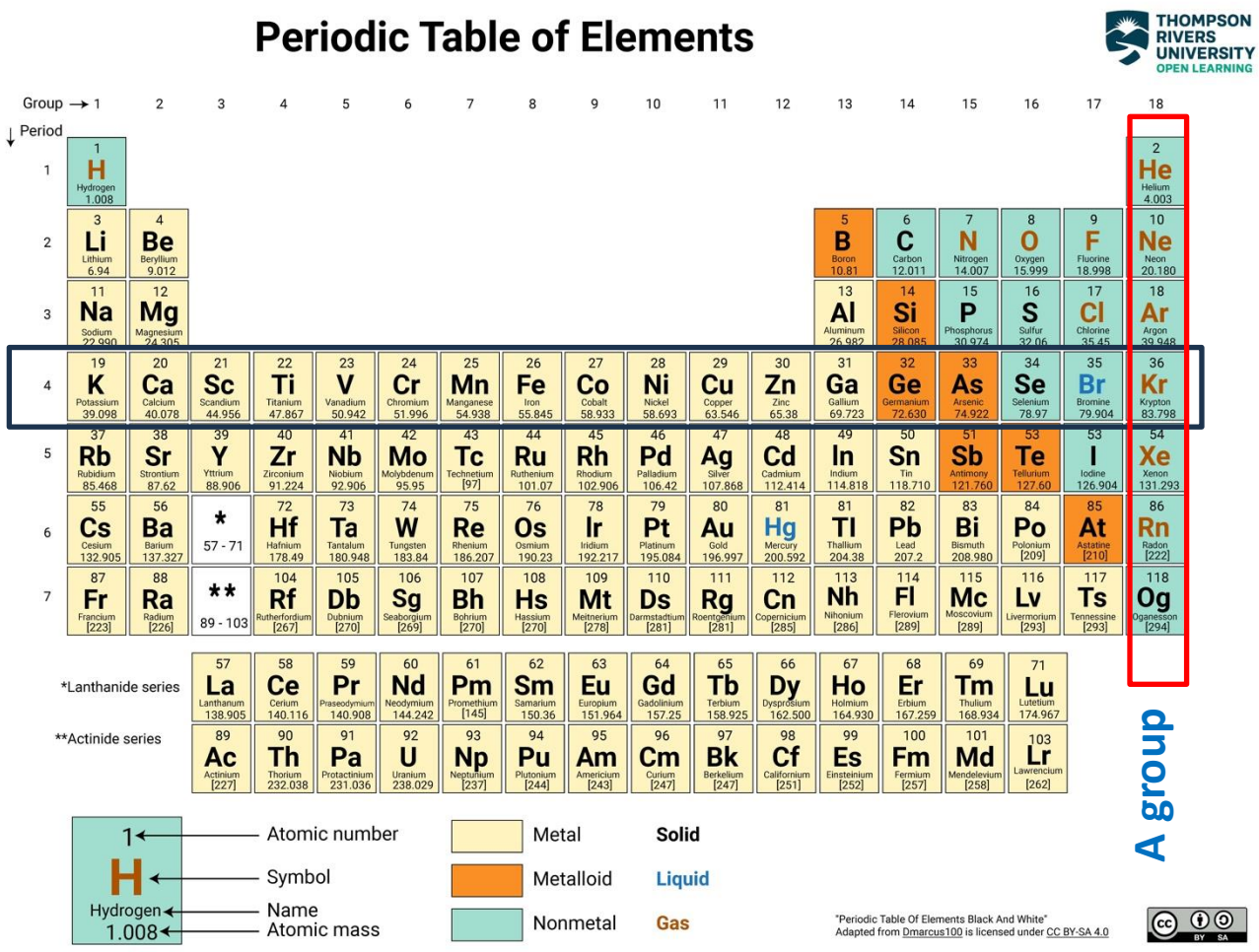
- **The arrangement of elements** is in order of **increasing atomic number (Z)**
- **The horizontal rows** of the table (numbered from 1 to 7) are called **periods**.
- **The vertical columns** of the table (numbered from 1 to 18) are called **groups or families**. Element in the same group have similar but not identical characteristics.

Example 2.9.

Group 1: very reactive
Group 18: noble gases, very stable

A period

- Most of the elements are metals



2.9. Ionic compounds

- Most elements are metals. ***Metals*** **easily lose electrons.**
- ***Nonmetals*** **tend to gain electrons**
- *Metals and nonmetals form compounds with each other. These compounds are called* ***ionic compounds****.*

metal ion + nonmetal ion → *ionic compounds*

❑ Formula of ionic compounds

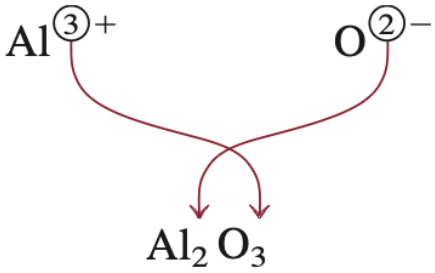
➢ **If the charge of cation = charge of anion, the ionic compound will be 1:1 [cation : anion]**

Examples 2.10. The formula of the ionic compound formed between

Na^+ **and** Cl^- → **NaCl**

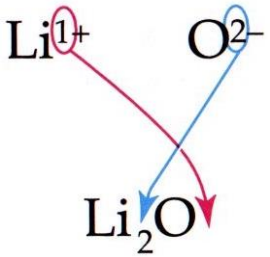
Mg^{2+} **and** S^{2-} → **MgS**

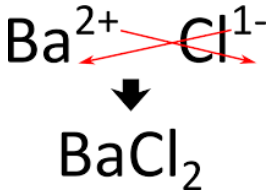
Al^{3+} **and** N^{3-} → **AlN**

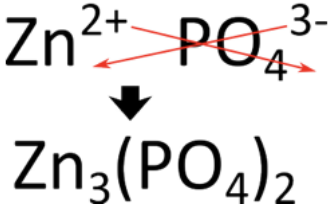


➢**If the charge of the cation and the charge of the anion are different, use the cross over method to determine the formula of the ionic compound.**

Examples 2.11.







Practice on chapter 2

Q1. ^{12}C, ^{13}C and ^{14}C are called

a) molecules b) compounds c) ions d) isotopes

Q2. The formula of the ionic compound formed between Fe^{2+} and SO_4^{2-} ions is

a) $FeSO_4$ b) $Fe_3(SO_4)_2$ c) $Fe_2(SO_4)_3$ d) $Fe_3(SO_4)$

Q3. The simplest (empirical) formula of glucose $C_6H_{12}O_6$ is

a) $C_6H_{12}O_6$ b) $C_3H_4O_3$ c) CHO d) CH_2O

Q4. Cations form when

a) neutrons are added to or removed from an atom b) electrons are removed from an atom

c) electrons are added to an atom d) all of these

Q5. The formula that indicates the attachment of atoms in a molecule is the

a) structural formula b) ionic formula c) molecular formula d) simplest formula

Q6. Which of the following is a homonuclear molecule?

a) H_2O b) CaF_2 c) KNO_3 d) Cl_2

Q7. The atomic number of an element is the number of

a) protons and electrons b) neutrons and electrons c) protons d) protons and neutrons

Q8. The numbers of neutrons (n) and electrons (e-) in radium ion $^{226}_{88}Ra^{2+}$ are

a) 226 n and 88e- b) 138 n and 88 e- c) 138 n and 86 e- d) 226 n and 86 e-

Q9. The horizontal rows in the periodic table are called

a) groups or families b) isotopes c) periods d) ionic compounds

Q10. The atomic mass of iron ($_{26}Fe$) is 56. The atomic mass of the ferric ion (Fe^{3+}) is

a) 26 b) 56 c) 29 d) 59

Q11. When the cation Ti^{3+} gains an electron, it becomes

a) Ti^{+} b) Ti^{4+} c) Ti^{2+} d) Ti^{-}

Q12. The ionic compound is formed between

a) two metals b) two nonmetals

c) a metal and a nonmetal d) two cations

Q13. One mole of H_2O contains

a) N_A molecules of H_2O b) N_A atoms of hydrogen

c) 3 molecules of H2O d) $2\times N_A$ molecules of oxygen

Chapter 3

Calculations in Chemistry

At the end of this chapter, the student will be able to

- ❑ explain the relationship between the molar mass and the atomic mass or molecular mass
- ❑ calculate the number of moles, and the mass percentage of an element in a compound
- ❑ Calculate the number of atoms or molecules knowing the number of moles
- ❑ Determine the molecular formula of a compound using chemical calculations
- ❑ Interpret a balanced chemical equation
- ❑ show the relationship between the number of moles (or mass) of the reactants and the number of moles (or mass) of the products in a balanced chemical equation

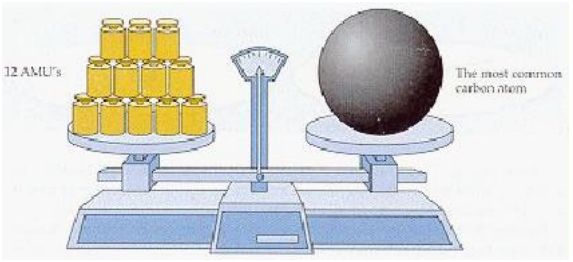
3.1. Atomic mass

Atomic mass (the mass of one atom) is the average mass of all the isotopes of an atom.

- It is commonly expressed in unified atomic mass units (amu)

•1 atomic mass unit is defined as 1/12 of the mass of a single carbon-12 atom:
$1amu = (1/12) \times 12/(6.022 \times 10^{23}) = 1/(6.022 \times 10^{23}) = 1.660 \times 10^{-24}$ g

$$1amu = (1/N_A)\ g = 1.660 \times 10^{-24}\ g$$



- **Hydrogen (H)** atom has a mass of **1 amu**
- **Carbon (C)** atom has a mass of **12 amu**
- **Oxygen (O)** atom has a mass of **16 amu**

The atomic masses of the elements are given in the periodic table of elements.

3.2. Molecular mass

Molecular Mass (mass of one molecule) =
Sum of atomic masses of all the atoms in the molecule

Practice 3.1. Determine the molecular mass of H_2SO_4 and $C_6H_{12}O_6$

Atomic masses: H=1, C=12, O= 16, S= 32

- One molecule of H_2SO_4 contains 2 atoms of H, 1atom of S and 4 atoms of O:

Molecular mass of H_2SO_4 $= 2\times$(atomic mass of H)$+1\times$(atomic mass of S)
$+ 4\times$(atomic mass of O)
$= (2 \times 1) + (1 \times 32) + (4 \times 16) = 98$ amu

- One molecule of $C_6H_{12}O_6$ contains 6 atoms of C, 12 atoms of H and 6 atoms of O:

Molecular mass of glucose $C_6H_{12}O_6$
$= 6 \times 12 + 12 \times 1 + 6 \times 16 = 180$ amu

Practice 3.2. **Calculate in grams:**

a) the mass of one atom of selenium (Se)

b) the mass of one molecule of H_2O

Atomic masses: (H = 1, O = 16, Se = 78.96)

***N_A* = 6.022 × 10^{23}**

Answer:

a) The atomic mass of Se equals 78.96 amu

The mass of one atom of Se in grams is

$78.96/N_A = 78.96/(6.022 \times 10^{23}) = 1.311 \times 10^{-22}$ g

b) The molecular mass of H_2O is $(2 \times 1 + 1 \times 16) = 18$ amu

The mass of one molecule of H_2O in grams is

$18/N_A = 18/(6.022 \times 10^{23}) = 2.989 \times 10^{-23}$ g

Exercises Q4.1 p.75

3.3. Molar Mass

❑ **For monatomic elements, the molar mass is the numerical value on the periodic table expressed in g/mol**

EXAMPLES 3.1.

Element	Atomic mass	Molar mass (M)
H	1 amu	1 g/mol
Na	23 amu	23 g/mol
Fe	56 amu	56 g/mol

❑ **For molecules, the molar mass is the sum of the molar masses of each of the atoms in the molecular formula.**

***EXAMPLES 3.2.* (Atomic masses: H=1, C= 12, 0=16 and Ca=40)**

➢ **One molecule of H_2O contains 2 atoms H and 1 atom O:**

The molecular mass of H_2O is equal to $2\times1+1\times16 = 18$ amu

The molar mass of H_2O is 18 g/mol

➢ **One molecule of $CaCO_3$ contains 1 atom Ca, 1 atom C and 3 atoms O:**

The molecular mass of $CaCO_3$ is equal to $1\times40+ 1\times12 + 3\times16 = 100$ amu

The molar mass of $CaCO_3$ is 100 g/mol

3.4. Number of moles:

$$n = \frac{m}{M}$$

n : number of moles of the substance
m : mass of the substance in grams
M: molar mass of the substance (g/mol)

Practice 3.3: **Calculate the number of moles of calcium carbonate $CaCO_3$ in 14.8 g of calcium carbonate.**
(atomic masses: C = 12, O = 16, Ca = 40)

***Answer*:**

Molar mass of $CaCO_3$ M = (1 x 40) + (1 x 12) + (3×16) = 100 g/mol
Number of moles: n = m / M = 14.8 / 100 = 0.148 mol

Practice 3.4. **What is the mass in grams of 0.287 mol of Acetylsalicylic acid $C_9H_8O_4$? (Atomic masses: H = 1, C = 12, O = 16)**

Answer:

Molar mass of $C_9H_8O_4$: M = (9×12) + (8×1) + (4×16) = 180 g/mol
n= m/M ⟹ m = n × M = 0.287 × 180 = 51.7 g

3.5. The number of atoms or molecules: $N = n \times N_A$

Practice 3.5. **Calculate the number of Se atoms in 1 g of Se.**
(atomic mass: Se = *78.96*)

Answer:

The molar mass of Se = 78.96 g/mol

Number of Se moles $n = m / M = 1 / 78.96 = 1.266 \times 10^{-2}$ mol

Number of Se atoms $N = n \times N_A = 1.266 \times 10^{-2} \times 6.022 \times 10^{23}$

$= 7.624 \times 10^{21}$ atoms

Practice 3.6. **Calculate the number of molecules of calcium carbonate $CaCO_3$ in 14.8 g of calcium carbonate.**
(atomic masses: C = 12, O = 16, Ca = 40)

Answer:

Molar mass of $CaCO_3 = (1 \times 40) + (1 \times 12) + (3 \times 16) = 100$ g/mol

Number of moles $n = m / M = 14.8 / 100 = 0.148$ mol

Number of molecules $N = n \times N_A = 0.148 \times 6.022 \times 10^{23}$

$= 8.91 \times 10^{22}$ molecules

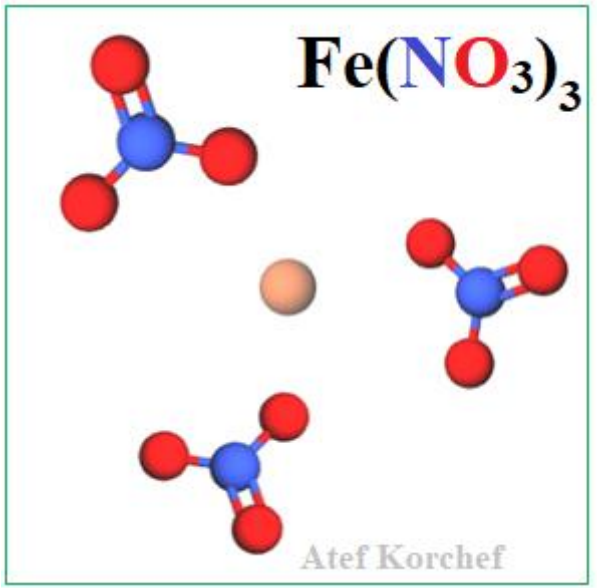
Exercises Q4.3 to Q4.7 p.78

3.6. Mass percent of an element in a compound:

$$\text{mass \% (element)} = \left(\frac{n \times \text{molar mass of element}}{\text{molar mass of compound}}\right) \times 100\%$$

n is the number of moles of the element in 1 mole of the compound.
n is determined from the molecular formula of the compound.

Practice 3.7. **Calculate the mass percent of iron Fe, nitrogen N and oxygen O in $Fe(NO_3)_3$.**
(Atomic masses: N = 14, O = 16, Fe = 56)



Answer

Molar mass M of $Fe(NO_3)_3$

$M = M(Fe) + 3 \times M(NO_3)$
$= M(Fe) + 3 \times [M(N) + 3 \times M(O)]$
$= 56 + 3 \times [14 + 3 \times 16]$
$= 242$ g/mol

mass % (Fe) = [(1 × M(Fe)) / M($Fe(NO_3)_3$] × 100
= [(1 ×56)/242]× 100
= 23.14 %

mass % (N) = [(3 × M(N)) / M($Fe(NO_3)_3$] × 100
= [(3 ×14)/242]× 100
= 17.36 %

mass % (O) = [(9 × M(O)) / M($Fe(NO_3)_3$] × 100
= [(9 ×16)/242]× 100
= 59.50 %

23.14 % + 17.36% + 59.50 % = 100.00%

Exercises Q4.8 to Q4.10 p.78-79

3.7. Empirical and Molecular Formulas Calculation

Molecular Formula = D× (Empirical Formula); D means how many times simplest formula repeated to give molecular formula.

$$D = \frac{\text{Molar mass of molecular formula}}{\text{Molar mass of empirical formula}}$$

Practice 3.8. **The simplest formula of vitamin C is $C_3H_4O_3$ and its molar mass is 176 g/mol. What is the molecular formula of vitamin C? (*Atomic masses: H = 1, C = 12, O = 16*)**

***Answer*:**

Molar mass of empirical formula = 3 xM(C) + 4 x M(H) + 3 x M(O)

= (3 x 12) + (4 x 1) + (3 x 16) = 88 g/mol

Molar mass of the molecular formula = 176 g/mol

D = 176 / 88 = 2

The molecular formula of vitamin C is $C_6H_8O_6$

Practice 3.10. **Elemental analysis of lactic acid (M = 90 g/mol) shows that 100 g of this compound contains 40.0 g (C), 6.7 g (H) and 53.3 g (O). Find the empirical formula and the molecular formula of lactic acid.**

Atomic masses: (H = 1, C = 12, O = 16)

Answer:

- The number of moles of C: n(C) = m(C)/M(C) = 40/12 = 3.33 mol
- The number of moles of H: n(H) = m(H)/M(H) = 6.7/1 = 6.7 mol ⇒ $C_{3.33}H_{6.7}O_{3.33}$
- The number of moles of O: n(O) = m(O)/M(O) = 53.3/16 = 3.33 mol

Divide each mole number by the smallest number of moles (3.33):

Relative mole ratio of (C) = 3.33 / 3.33 = 1

Relative mole ratio of (H) = 6.7 / 3.33 = 2 ⇒ The empirical formula is CH_2O

Relative mole ratio of (O) = 3.33 / 3.33 = 1

The molar mass of the empirical formula CH_2O is

$M(CH_2O) = 12 + 2\times1 + 16 = 30$ g/mol

Or the molar mass of the molecular formula is M(lactic acid) = 90 g/mol

So, **D = M(lactic acid)/M(CH_2O) = 90/30 = 3**.

The molecular formula is obtained by multiplying the subscripts of the empirical formula by 3. ⇒ The molecular formula of lactic acid is $C_3H_6O_3$.

Exercises Q4.12 to Q4.14 p.79

3.8. Balancing Chemical Equations: Determining the *coefficients* that provide equal numbers of each type of atom on each side of the equation:

1. Write the **correct** formula(s) for the reactants on the left side and the **correct** formula(s) for the product(s) on the right side of the equation.

Ethane reacts with oxygen to form carbon dioxide and water

$$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$$

2. Change the numbers in front of the formulas (***coefficients***) to make the number of atoms of each element the same on both sides of the equation. **Do not change the subscripts**.

$2C_2H_6$ **NOT** C_4H_{12}

3. **Start by balancing those elements that appear in only one reactant and one product.**

$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$ **start with C or H <u>but not</u> with O**

$C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$

$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$

4. Balance those elements that appear in two or more reactants or products.

$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$ multiply O_2 by $\frac{7}{2}$

$C_2H_6 + \frac{7}{2} O_2 \longrightarrow 2CO_2 + 3H_2O$ remove fraction; multiply both sides by 2

$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$

5. Check to make sure that you have the same number of each type of atom on both sides of the equation.

$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$

Reactants	Products
4 C	4 C
12 H	12 H
14 O	14 O

Practice 3.11. Which of the following is a balanced chemical equation?

a) $C_2H_6 + 7\ O_2 \rightarrow 2CO_2 + 3\ H_2O$

b) $2HgO \rightarrow 2Hg + 2O_2$

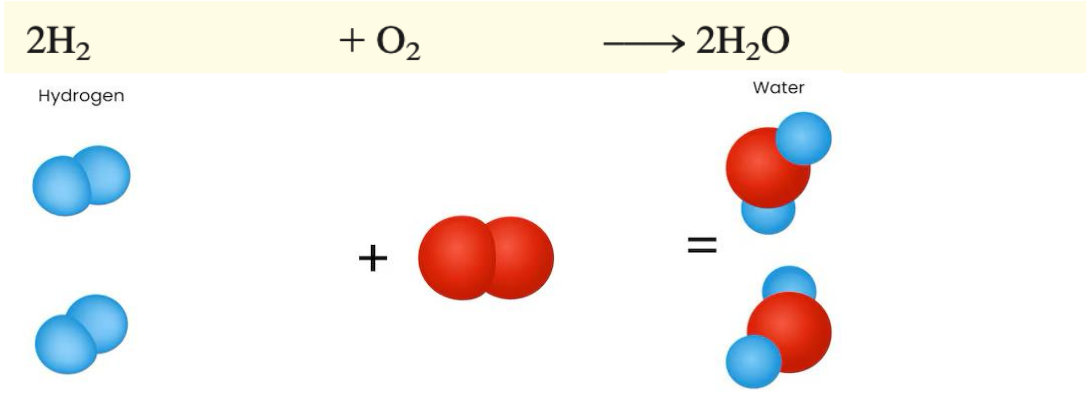
c) $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

d) $2H_2O \rightarrow 2H_2 + 1/2O_2$

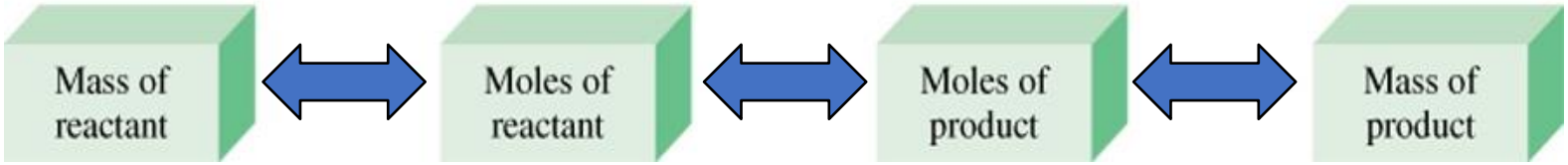
3.9. Interpretation of a balanced chemical equations

The law of conservation of mass states that atoms cannot be created or destroyed in a chemical reaction.

During a chemical reaction, the total mass of the **products** must be equal to the total mass of the **reactants**



3.10. Stoichiometry: **Study of mass relationships between reactants and products in a chemical reaction:** Reactants ⟶ Products



Exercises Q4.15 to Q4.18 p.79-80

Practice 3.12. Butane C_4H_{10} reacts with oxygen to form carbon dioxide and water. Determine the mass of carbon dioxide obtained when 5.8 g of butane reacts.

Atomic masses (H= 1, C=12, O=16)

Answer: $2C_4H_{10\,(g)} + 13\,O_{2\,(g)} \rightarrow 8CO_2 + 10\,H_2O_{\,(g)}$

2 mol C_4H_{10} ⟶ 8 mol CO_2

$2\times(4\times12+10\times1) = 116$ g ⟶ $8\times(1\times12+2\times16) = 352$ g

5.8 g ⟶ $m(CO_2) = ?$ $\quad m(CO_2) = (5.8\times352)/116 = 17.6$ g

Practice 3.13. Titanium is prepared by the reaction of titanium(IV) chloride $TiCl_4$ with molten magnesium at high temperature:

$$TiCl_4(g) + 2Mg(l) \rightarrow Ti(s) + 2MgCl_2(l)$$

What masses (in kg) of $TiCl_4$ and Mg are required to produce 250 kg of Ti?

Atomic masses (Mg = 24.3, Cl = 35.5, Ti = 47.9)

Answer:

1 mol $TiCl_4$ ⟶ 1 mol Ti

$1\times(1\times47.9+4\times35.5) = 189.9$ g ⟶ 47.9 g

$m(TiCl_4) = ?$ g ⟶ 250 kg = 2.5×10^5 g

$m(TiCl_4) = (2.5\times10^5\times189.9)/47.9 = 991100$ g = 991.1 kg

1 mol $TiCl_4$ ⟶ 2 mol Mg

$1\times(1\times47.9+4\times35.5) = 189.9$ g ⟶ $2\times24.3 = 48.6$ g

99.1×10^3 g ⟶ $m(Mg) = ?$

$m(Mg) = (991.1\times10^3\times48.6)/189.9 = 253.6\times10^3$ g = 253.6 kg

Practice on chapter 3

Atomic masses: **H = 1, C = 12, N = 14, O = 16, Na = 23, Mg = 24, S = 32, Cl = 35.5, Ca = 40, Fe = 56, Cu = 63.5.**

Avogadro's number: $N_A = 6.022 \times 10^{23}$

Q1. What are the values of (x,y) that make the following reaction balanced ?

$$2\ C_xH_y + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$$

a) (8,9) b) (8,18) c) (4,18) d) (8,25)

Q2. The number of moles in 275 g of calcium carbonate ($CaCO_3$) is

a) 17.5 moles b) 27.5 moles c) 2.75 moles d) 100 moles

Q3. The molar mass of $(NH_4)_2SO_4$ is:

a) 114 g/mol b) 100 g/mol c) 132 g/mol d) 13.20 g/mol

Q4. The mass percent of sulfur (mass% S) in copper sulfate ($CuSO_4$) is

a) 20 % b) 75 % c) 55% d) 40 %

Q5. One mole of H_2O contains

a) 1 mole of O b) $N_A = 6.022 \times 10^{23}$ molecules of H_2O

c) $2 \times 6.022 \times 10^{23}$ atoms of H d) All of these

Q6. From the balanced equation for the combustion of methane (CH_4)

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$$

The mass of oxygen (O_2) needed for the combustion of 16 g of methane is

a) 16 g b) 36 g c) 32 g d) 64 g

Q7. The simplest formula of a compound is CH_2O and its molar mass is 180 g/mol. The molecular formula of the compound is

a) C_2H_4O b) C_2H_6O c) $C_6H_{12}O_6$ d) $C_2H_4O_2$

Q8. How many molecules are there in 9.8 g of sulfuric acid (H_2SO_4)?

a) 602×10^{22} b) 6.022×10^{22} c) 6.022×10^{23} d) 6.022×10^{24}

Q11. The mass of one atom of oxygen (O) is

a) 16 amu b) 16 g/mol c) $16/N_A$ g d) both of a) and c) are correct

Q12. 1 amu equals

a) 1 g b) 1 mol c) $1/N_A$ g d) $12/N_A$ g

Q13. Which of the following is a balanced chemical equation?

a) $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$ b) $5H_2 + N_2 \rightarrow 5NH_3$

c) $2HgO \rightarrow 2Hg + 1/2O_2$ d) $Ca^{2+} + 2HCO_3^- \rightarrow CaCO_3 + 2CO_2 + H_2O$

Homework - CALCULATIONS

Atomic masses: H = 1, C = 12, N = 14, O = 16, Na = 23, Mg = 24, S = 32, Ca = 40, Fe = 56, Zn = 65.
Avogadro's number: $N_A = 6.022 \times 10^{23}$

Question 1

1. Calculate the molar mass of $CaSO_4 \cdot 2H_2O$.
2. Calculate the number of moles of H_2SO_4 in 49 g of this compound.
3. Calculate the mass of 0.04 mol of $CaCO_3$.
4. Determine the mass percent of nitrogen (N) in zinc nitrate $Zn(NO_3)_2$.
5. Determine the number of molecules of H_2O_2 in 17 g of this compound.
6. Determine the mass of one molecule of C_2H_6 in grams.
7. Determine the mass of one atom of magnesium (Mg) in grams.
8. The empirical formula of a compound is CH_2O and its molar mass is 60 g/mol. What is the molecular formula of this compound?

Question	1	2	3	4	5	6	7	8
Answer	172 g/mol	0.5 mol	4 g	14.8%	3.011×10^{23} molecules	$30/N_A$ g	$24/N_A$ g	$C_2H_4O_2$

Question 2

Octane (C_8H_{18}), one of many components of gasoline, mixes with oxygen from the air and burns to form carbon dioxide and water vapor.

1. Write the balanced chemical equation for this reaction.
2. Calculate the mass of oxygen (O_2) needed for the combustion of 220 g of octane.

ANSWER

1) $2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$
2) 771.93 g

Chapter 4

Electron Configuration and Properties of Atoms

At the end of this chapter, the student will be able to

- ❑ Recognize the wave-particle duality of light and matter
- ❑ Describe the Bohr's model of the hydrogen atom
- ❑ Calculate the wavelength of the line corresponding to the transition of an electron from one energy level to another.
- ❑ define the quantum numbers
- ❑ define the different rules for the electron configuration
- ❑ write the electron configuration of atoms
- ❑ determine the valence shell and valence electrons and the number of unpaired electrons in an atom
- ❑ describe the periodic trends in atoms' properties using the periodic table of elements.

4.1. Electromagnetic Radiation (light) - Wave like

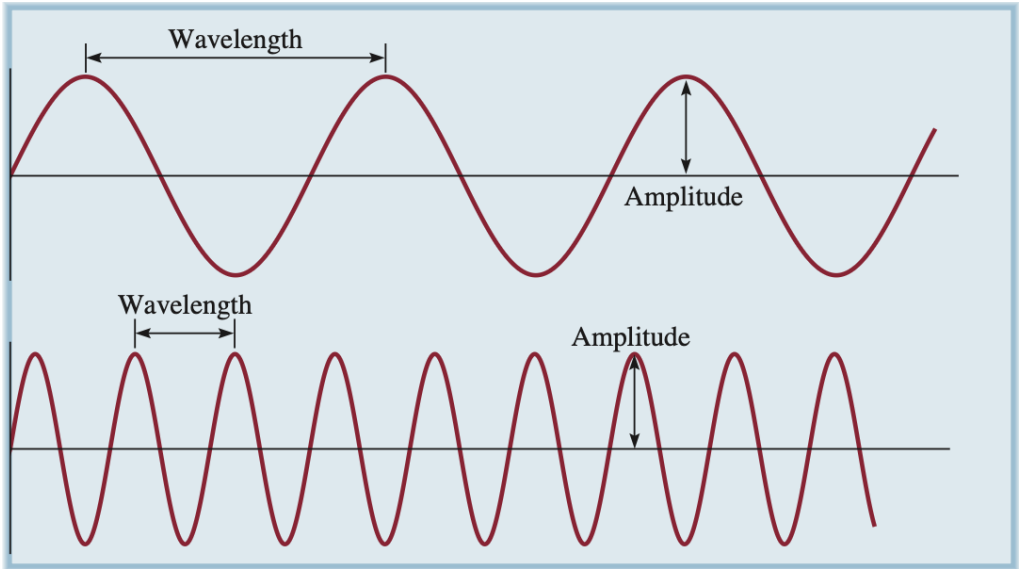
Wavelength, λ, is the distance from one crest to the next in the wave. Measured in units of distance.

Frequency, ν, is the number of complete cycles per sec., expressed in s^{-1} or Hz

Speed of Light, C, same for all electromagnetic radiation.

C= 3 x 10^8 m/s in vacuum

$$\nu = c/\lambda$$



Practice 4.1. **The yellow light given by the sodium lamp has λ = 589 nm.**
What is the frequency of this radiation?
C = 3 × 10^8 m/s

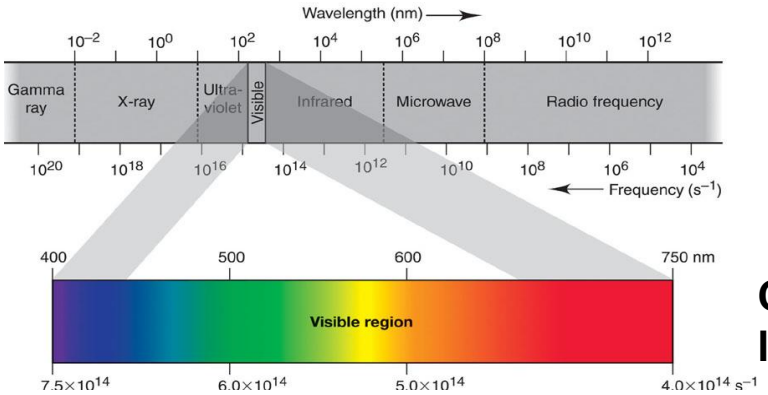
Answer:

λ = 589 nm = 589 × 10^{-9} m

υ = C/λ = 3 × 10^8 / (589 x 10^{-9}) = 5.09 x 10^{14} s^{-1} = 5.09 x 10^{14} Hz

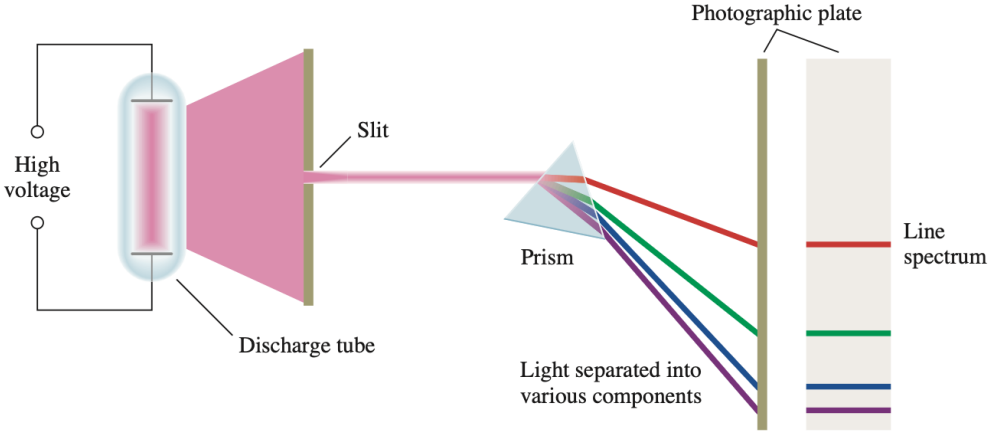
The continuous spectrum of light

Visible light: 400 – 700 nm
Ultraviolet: <400 nm
Infrared: >700 nm

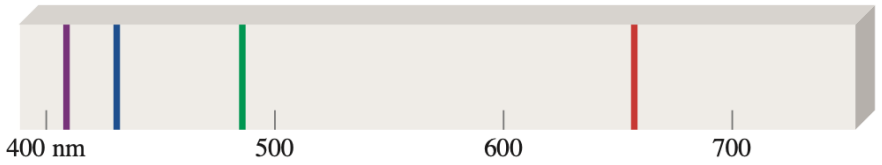


Continuous spectrum of light in the visible region

The line spectrum of hydrogen



The spectrum of hydrogen element is a discontinuous spectrum (line spectrum) ⟶ **Why?**



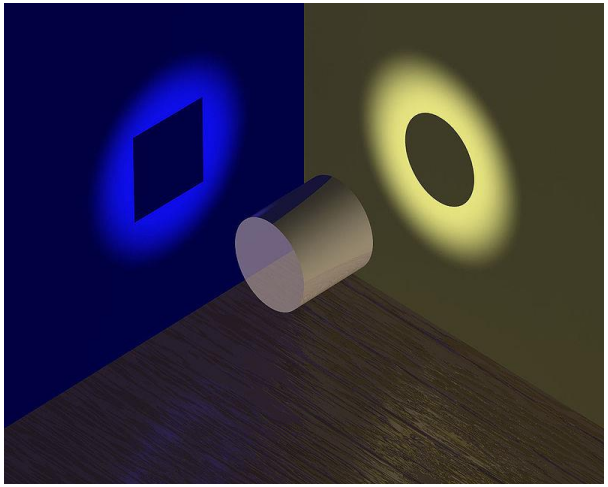
The wave-particle duality of light

❑ **Planck**: Atoms can gain or lost only certain quantities of energy, $\Delta E = nh\nu$ (where n is a positive integer, 1, 2, 3, etc) by absorbing or emitting radiation. This means energy is quantized.

❑ **The wave-particle duality of light** (proposed by **Einstein)**: **Light** consisted of packets of energy called **photons** (particle-like) that had wave-like properties as well.

$$\mathbf{E_{photon} = h\nu = hc/\lambda = \Delta E_{atom}}$$

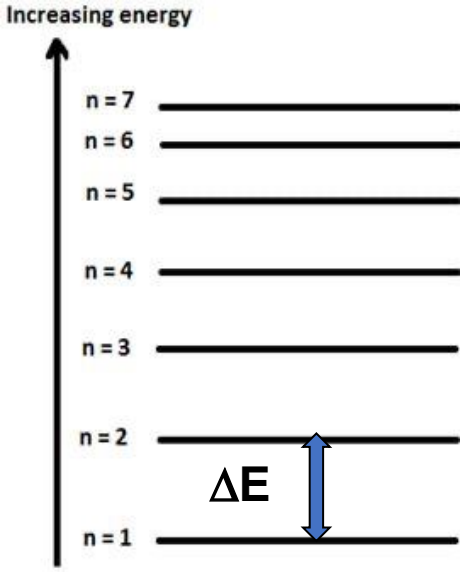
Planck's constant **h = 6.63 x 10^{-34} Js**



4.2. Bohr's Model of the Hydrogen Atom

❑ Hydrogen atom has only certain allowable energy levels for the electron orbits. (quantized energy)

- Ground state: the lowest possible energy level an electron can be at (n=1).
- Excited state: an energy level higher than the ground state (n>1).



The energy level

$$E_n = - R_E / n^2 = - 2.18 \times 10^{-18} / n^2$$

where n = 1, 2, 3,....∞

The energy constant $R_E = 2.18 \times 10^{-18}$ J

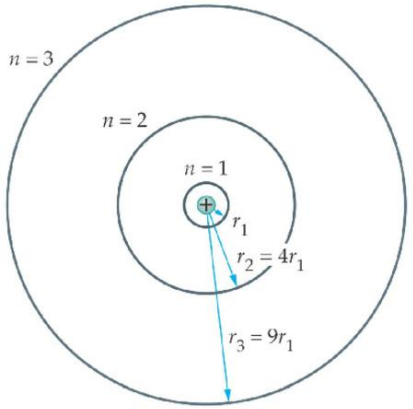
Ground state (lowest energy)

❑The electron traveled in circular orbits.

❑ Each orbit has certain radius r_n

$$r_n = 0.5 \times n^2 \text{ Å}$$

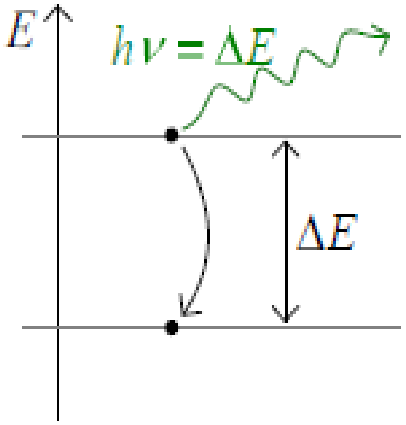
Example 4.1. For n=3, $r_n = 0.5 \times 3^2 = 4.5$ Å



❑When the electron moves from one orbit to another, it absorbs or emits a photon whose energy equals the difference in energy between the two orbits.

Electron moves from a higher energy level orbit to a lower energy level orbit

Emits a photon

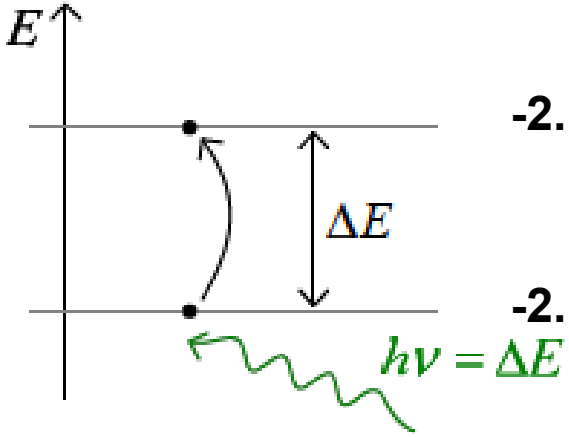


$-2.18 \times 10^{-18} / (n_{initial})^2$ (In **Joule**)

$-2.18 \times 10^{-18} / (n_{final})^2$

Electron moves from a lower energy level orbit to a higher energy level orbit

absorbs a photon



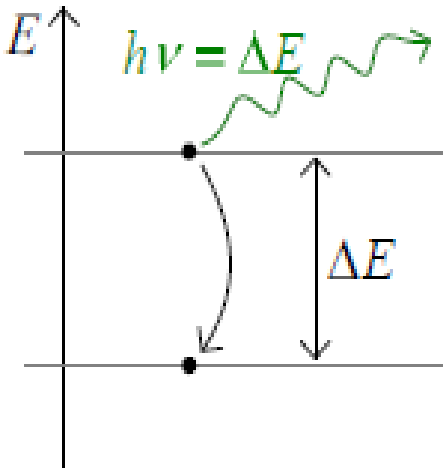
$-2.18 \times 10^{-18} / (n_{final})^2$

-2.18×10^{-18} J $/ (n_{initial})^2$

$$\Delta E = E_{final} - E_{initial} = -2.18 \times 10^{-18} \times \left(\frac{1}{n_{final}^2} - \frac{1}{n_{initial}^2}\right)$$

Practice 4.2. **Calculate the wavelength (λ) of the line corresponding to the transition of an electron from n = 4 to n = 2 in the hydrogen atom.**

$R_E = 2.18 \times 10^{-18}$ J, C= 3 x 10^8 m/s and h = 6.63 x 10^{-34} Js



$n_i = 4$ $\quad E_4 = -2.18 \times 10^{-18} / (4^2)$

$n_f = 2$ $\quad E_2 = -2.18 \times 10^{-18} / (2^2)$

$\Delta E = E_{final} - E_{initial} = E_2 - E_4 = -2.18 \times 10^{-18} [(1/2^2) - (1/4^2)]$

$= -2.18 \times 10^{-18} [(1/4) - (1/16)]$

$= -4.1 \times 10^{-19}$ J .

The sign (-) indicates a lose of energy since light is emitted from an atom when the electron moves from a higher level to a lower level ($n_{final} < n_{initial}$)

$\Delta E = h\, C/\lambda \longrightarrow \lambda = h\, C/\Delta E,$

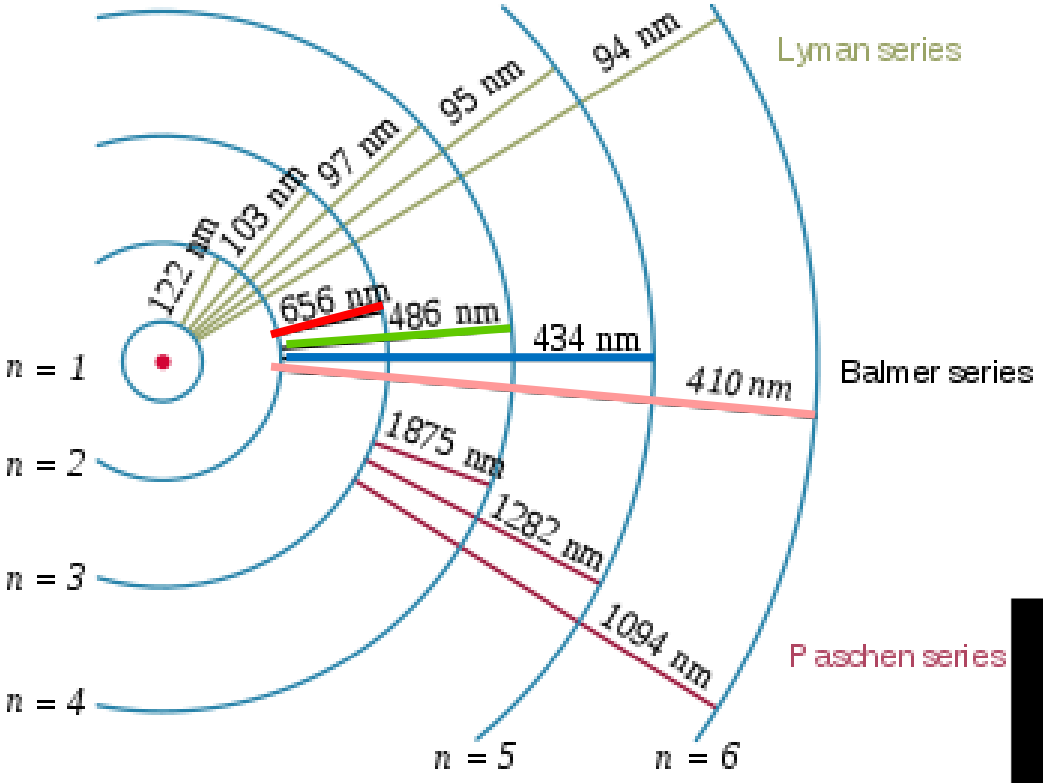
The wavelength λ of the light emitted is

$\lambda = (6.63 \times 10^{-34} \times 3 \times 10^8) / 4.1 \times 10^{-19} = 4.86 \times 10^{-7}$ m = 486 nm (green light)

Practice 6.3 p.126 and Q6.3 p.142

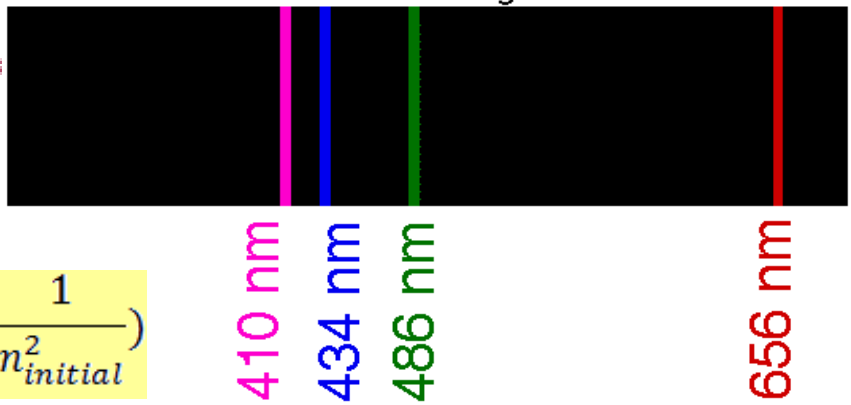
Explanation of the line spectrum of hydrogen

The lines of the hydrogen spectrum correspond to the transition of electrons from a higher energy level to a lower energy level:



for the visible series, $n_f = 2$ and $n_i = 3, 4, 5, \ldots$

Hydrogen emission spectrum in the visible region



$$\Delta E = E_{final} - E_{initial} = -2.18 \times 10^{-18} \times \left(\frac{1}{n_{final}^2} - \frac{1}{n_{initial}^2}\right)$$

And $\lambda = h\,C/\,\Delta E$

4.3. de Broglie hypothesis: The Wave-Particle Duality of Matter

- de Broglie hypothesizes **particles could have a wave behavior.**
- **The electrons have wave-like properties**
- **The de Broglie relation:**

$$\lambda = h/mv = h/p$$

h: Planck's constant: h= 6.62×10^{-34} Js
m: mass of the particle
v: velocity of the particle
p: momentum of the particle (**p = mv**)

- Every particle has a wave nature as well but it is only truly evident when a particle is very light such as an electron (m = 9.11×10^{-28} g).
- However, we cannot observe both the wave nature and the particle nature of the electron at the same time.
- According to **the Heisenberg Uncertainty Principle,** we **cannot know both** the position and the speed of the electron at a given time.

Practices 6.4 and 6.5 p.129

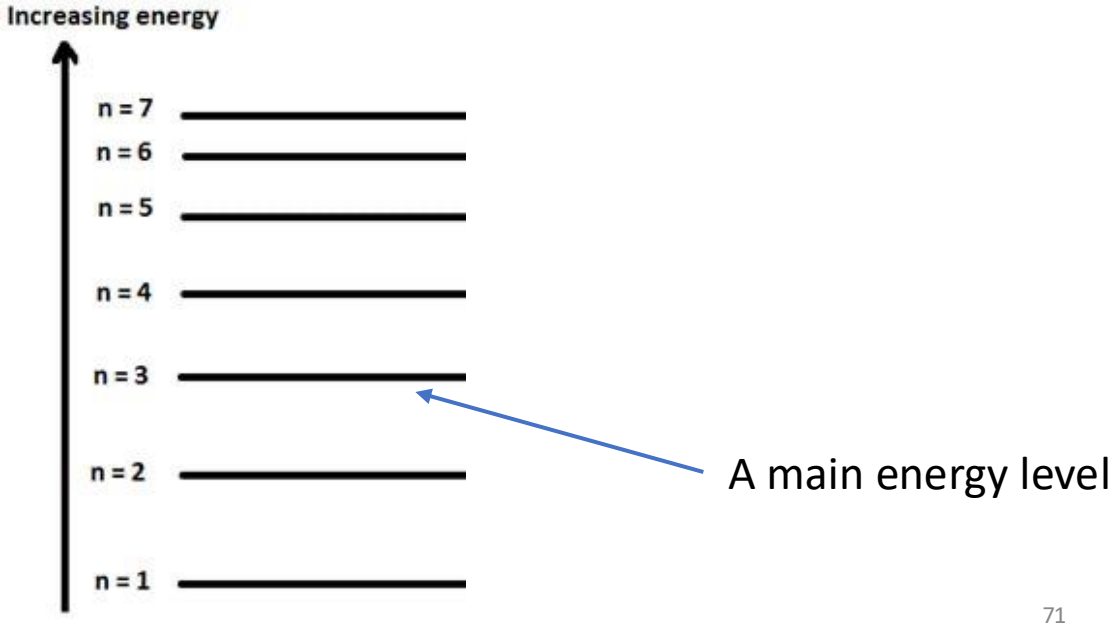
4.4. Quantum Numbers and Atomic Orbitals

Atomic Orbital: **A region** in space in which there is **high probability of finding an electron**.

Quantum numbers specify the properties of atomic orbitals and their electrons, ***i.e., an atomic orbital is specified by three quantum numbers (n, l, m_l). The spin of the electron is specified by the quantum number*** m_s *(or* ***s****)*

Principal Quantum Number, n

- Indicates main energy levels: **n = 1, 2, 3, 4…**
- Each main energy level has sub-levels



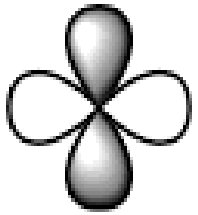
Orbital Quantum Number, ℓ

(Angular Momentum Quantum Number/ Azimuthal quantum number)

- ℓ indicates the shape of orbital **sublevels**
- **ℓ = 0, 1, …., (n-1)**

Example 4.3. When n=4, ℓ = 0, 1, 2 and 3.

ℓ=0	ℓ=1	ℓ=2	ℓ=3
s	p	d	f



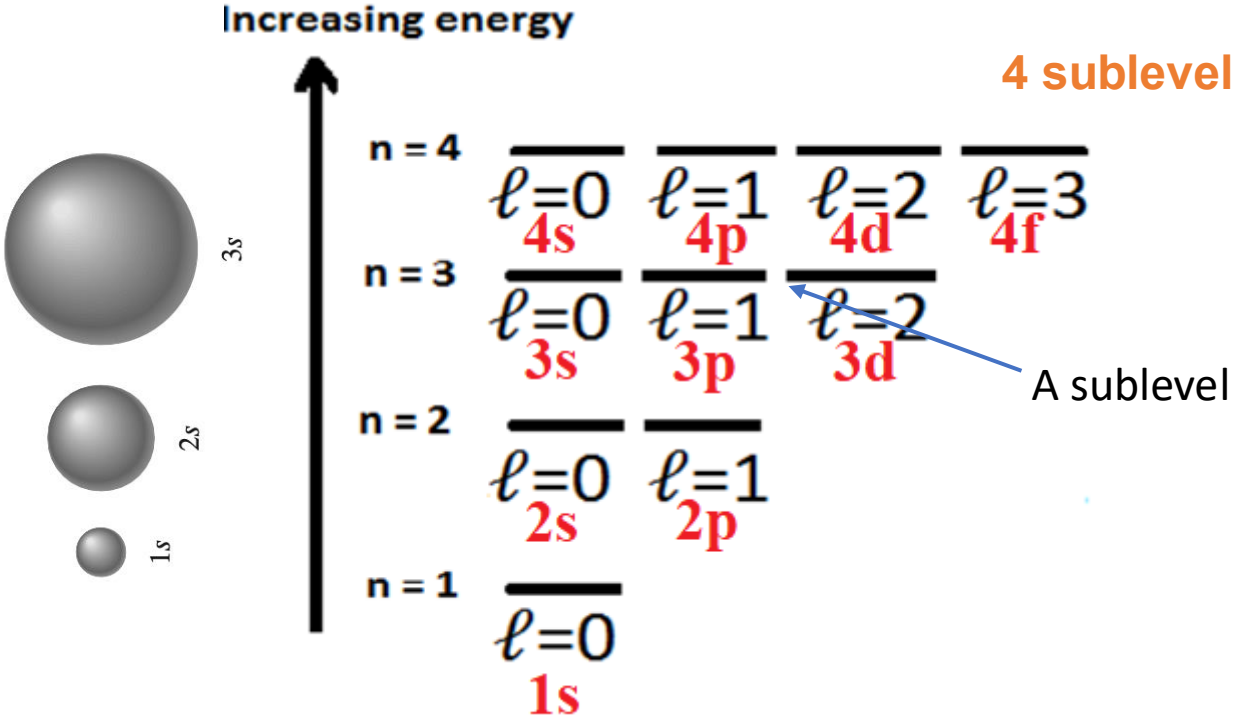
d



p orbital



s orbital

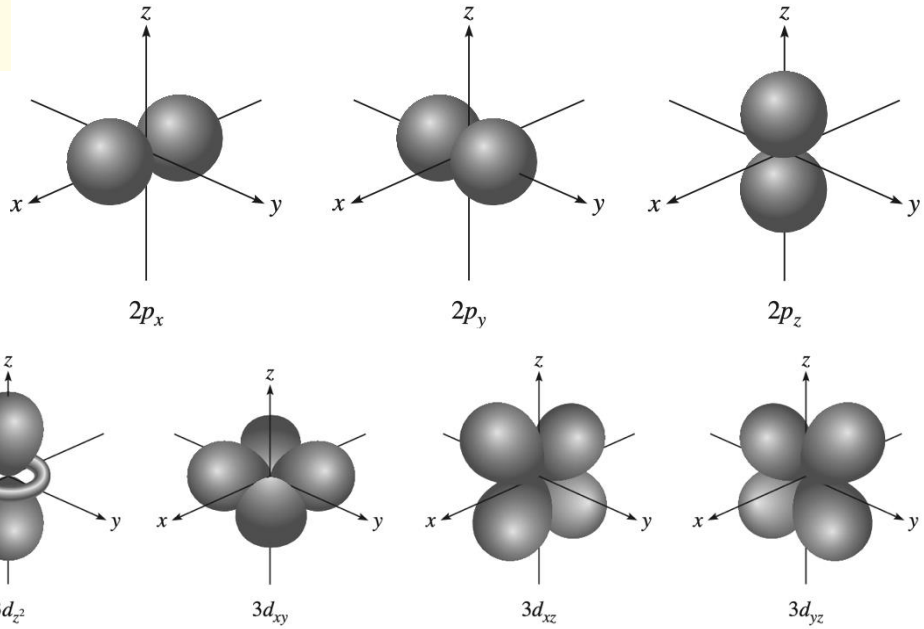


Magnetic Quantum Number, m_ℓ

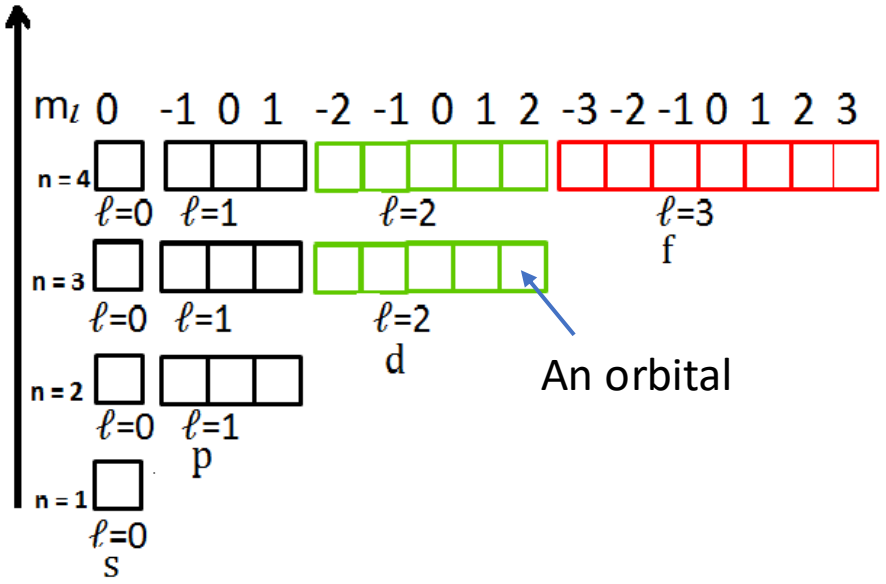
- Indicates the orientation of the orbital in space.
- Equals to integer values, including zero ranging from - ℓ to +ℓ

n	ℓ	m_ℓ	Number of Orbitals
1	0	0	1
2	0	0	1
	1	−1, 0, 1	3
3	0	0	1
	1	−1, 0, 1	3
	2	−2, −1, 0, 1, 2	5

- For each main energy level n, the number of m_ℓ values represents the number of orbitals.
- The sublevel s contains one orbital, p contains 3 orbitals, d contains 5 orbitals and f contains 7 orbitals.

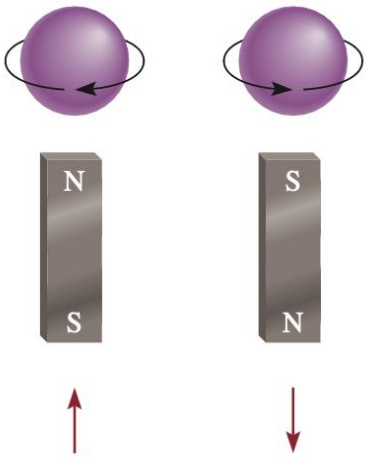


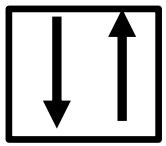
- In a main energy level n, there are n^2 orbitals, i.e., for n = 4 there are $4^2 = 16$ orbitals.



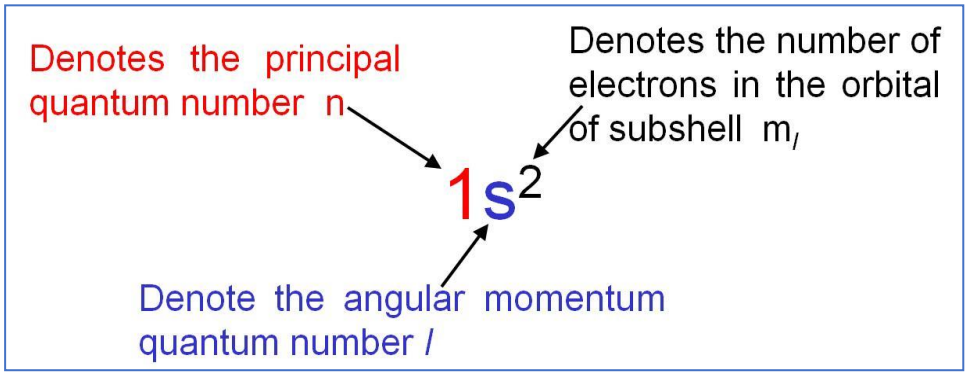
Electron Spin Quantum Number, (m_s or s)

- Indicates the direction in which an electron spins in an orbital (clockwise or counterclockwise).
- **m_s = -1/2, +1/2**





Practice 6.8 p.132



PRACTICE 4.3.

Q1. When the principle quantum number n=4, the orbital quantum number ℓ can take the values

a) 0,1,2,3,4 b) -4,-3,-2,-1, 0, 1,2,3,4 c) 0,1,2,3 d) -1/2, 1/2

Q2. The name of the sub-level with n= 3 and l = 2 is

a) 2d b) 3d c) 3p d) 2s

Q3. When the orbital quantum number ℓ = 1, the magnetic quantum number $m_ℓ$ can take the values

a) 0,1 b) -1, 0, 1 c) 0 d) -1/2, 1/2

Q4. The quantum number that can take the values -1/2 and 1/2 is the magnetic quantum number ($m_ℓ$).

a) True b) False

Q6.4 p.143

4.5. Electronic configurations

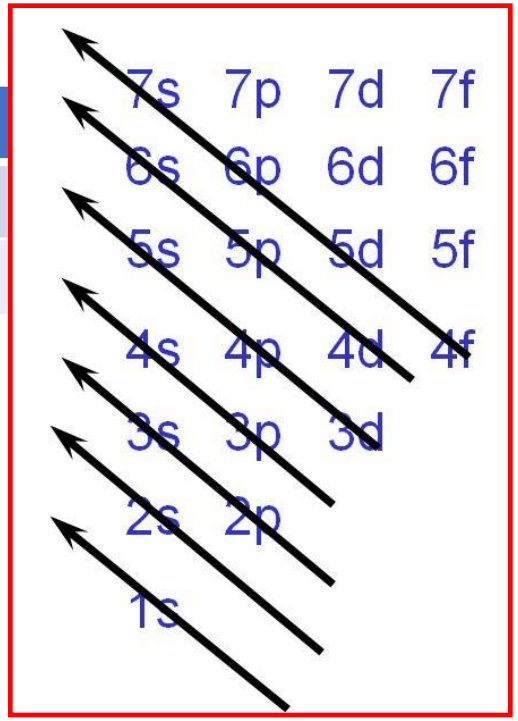
Electronic configuration is how the electrons are distributed among the various atomic orbitals in an atom. There are 3 rules to build up electronic configurations:

4.5.1. Aufbau (building up) Principle

Electrons enter the orbitals in order of ascending energy.

- ✓**Increased n+ ℓ, increased energy**
- ✓**Same value of n+ ℓ, lowest energy for lowest n**

orbital	1s	2s	2p	3s	3p	4s	3d
n +ℓ	1+0	2+0	2+1	3+0	3+1	4+0	3+2
Energy	↗	↗	↗	↗	↗	↗	↗



2p and 3s have the same value of n+ ℓ, 2p has a lower energy than 3p

General pattern for filling the sublevels

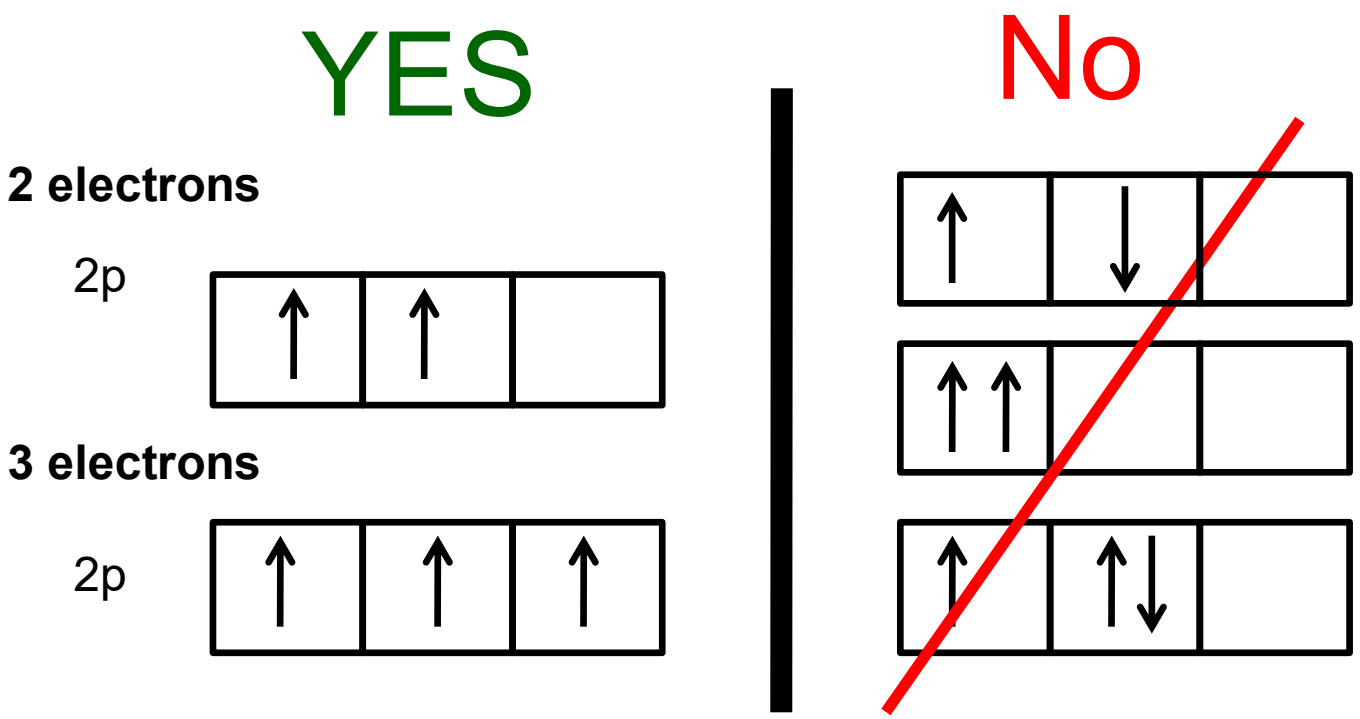
1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s …

Practice 6.9 p.133

4.5.2. Hund's Rule :

When in orbitals of equal energy, electrons will try to remain unpaired. The electrons occupy degenerate orbitals singly to the maximum extent possible and with their spins parallel.

Placing two electrons in one orbital means that, as they are both negatively charged, there will be some electrostatic repulsion between them. Placing each electron in a separate orbital reduces the repulsion and the system is more stable.



4.5.3. Pauli's exclusion principle :

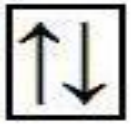
No two electrons in the same atom have the same set of the four quantum numbers (n , ℓ, $m_ℓ$, m_s).

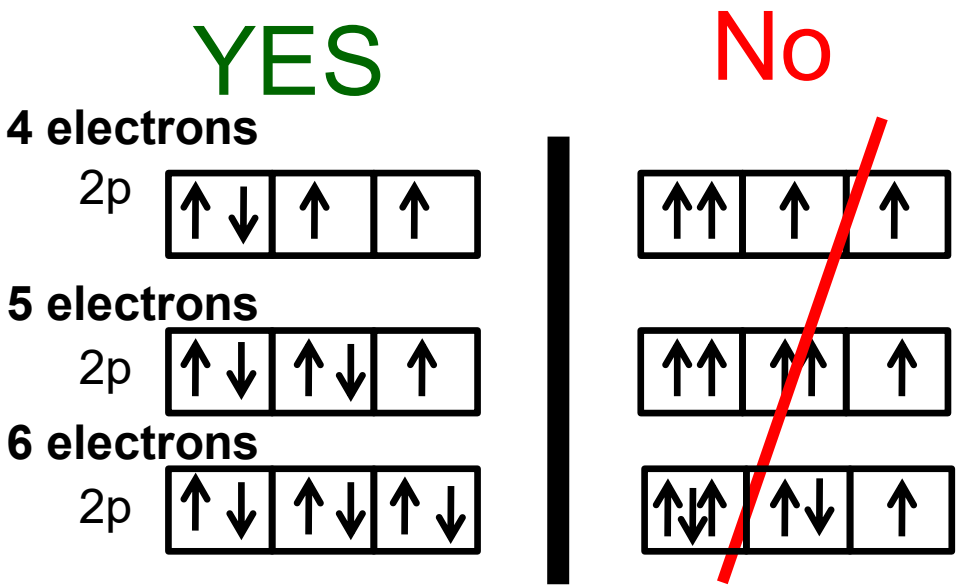
This means that:

Orbitals of the same energy must be occupied singly and with the same spin before pairing up of electrons occurs.

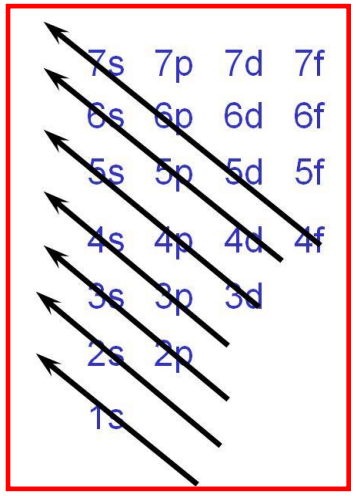
Electrons occupying the same orbital must have opposite spins.

An orbital is occupied au maximum by 2 electrons. An orbital containing paired electrons is presented as :





Practice 4.4. Determine the electronic configurations of $_{4}Be$, $_{17}Cl$, $_{26}Fe$, and $_{35}Br$.



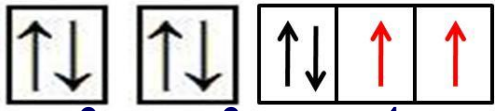
$_{4}Be$: $1s^2 2s^2$

$_{17}Cl$: $1s^2 2s^2 2p^6 3s^2 3p^5$

$_{26}Fe$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

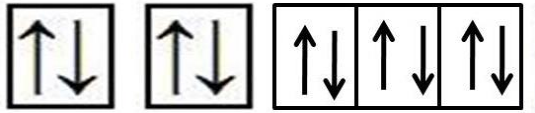
$_{35}Br$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

Practice 4.5. How many **unpaired electrons** are there in $_{8}O$, $_{10}Ne$, and $_{15}P$?



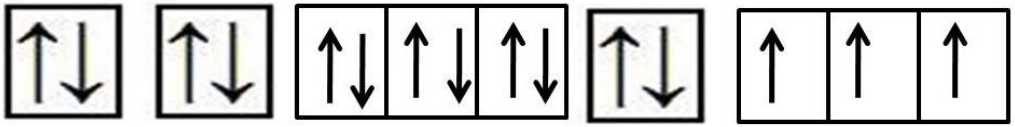
$_{8}O$: $1s^2$ $2s^2$ $2p^4$

2 unpaired electrons



$_{10}Ne$: $1s^2$ $2s^2$ $2p^6$

zero unpaired electrons



$_{15}P$: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^3$

3 unpaired electrons

4.5.4. Exceptions in some transition metals: Cr and Cu

❑ **Chromium $_{24}$Cr:**

1 electron

Expected electronic configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4$

✓ **Real electronic configuration**: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

❑ **Copper $_{29}$Cu:**

1 electron

Expected electronic configuration: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$

✓ **Real electronic configuration**: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$

The reason for these anomalies is the greater stablity of d subshells that are either half-filled (d^5) or completely filled (d^{10})



4.5.5. Valence-Shell and Valence Electrons

- The most outer sublevels **(for the highest n) are called valence-shell**
- Electrons in the most outer sublevels are called **valence electrons**
- **Transition metal**: metal whose atom has an **incomplete d subshell** or which can give rise to cations with an incomplete d subshell. For transition metals, the valence shell is **ns (n-1)d**

Examples 4.4.

Elements	Valence shell	Valence electrons
$_{9}F$: $1s^2$ $2s^2$ $2p^5$	2s 2p	7
$_{11}Na$: $1s^2$ $2s^2$ $2p^6$ $3s^1$	3s	1
$_{18}Ar$: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$	3s 3p	8
$_{35}Br$: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^{10}$ $4p^5$	4s 4p	7
Transition metals	**Valence shell**	**Valence électrons**
$_{26}Fe$: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^6$	4s 3d	8
$_{28}Ni$: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^8$	4s 3d	10

4.5.6. ELECTRONIC CONFIGURATION OF IONS

Electrons are removed first from the highest occupied orbitals (except for transition metals).

SODIUM $_{11}Na$ $1s^2\ 2s^2\ 2p^6\ 3s^1$ 1 electron removed from the 3s orbital

$_{11}Na^+$ $1s^2\ 2s^2\ 2p^6$



CHLORINE $_{17}Cl$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ 1 electron added to the 3p orbital

$_{17}Cl^-$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$



Isoelectronic configuration

Atoms or ions having the same electronic configuration

Example 4.5. $_8O^{2-}$, $_{11}Na^+$, $_9F^-$, Ne

$_8O^{2-}$: $1s^2\ 2s^2\ 2p^6$

$_{11}Na^+$: $1s^2\ 2s^2\ 2p^6$

$_9F^-$: $1s^2\ 2s^2\ 2p^6$

$_{10}Ne$: $1s^2\ 2s^2\ 2p^6$

Q6.5 p.144

4.5.7. Electronic configuration of transition metal ions

Electrons in the **ns** orbital are removed **before** any electrons in the **(n-1)d** orbitals.

Examples 4.6.

TITANIUM $_{22}Ti$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$

Ti^+: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^2$ (1 electron is removed from 4s)

Ti^{2+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2$ (2 electrons are removed from 4s)

Ti^{3+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1$ (1 electron is removed from 3d)

Ti^{4+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ (2 electrons are removed from 3d)

Practice 4.5. Determine the electronic configurations of Fe, Fe^{2+} and Fe^{3+}

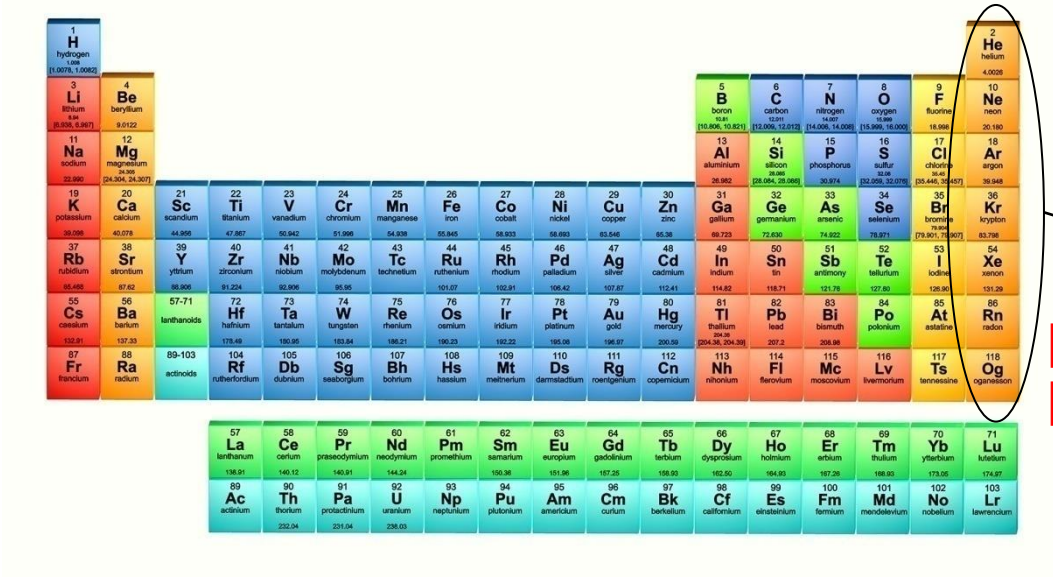
Answer:

Iron $_{26}Fe$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

Fe^{2+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$ (2 electron are removed from 4s)

Fe^{3+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$ (1 electron is removed from 3d)

4.5.8. Short cut electronic configuration



Noble gases:

$[_2He]$, $[_{10}Ne]$, $[_{18}Ar]$, $[_{36}Kr]$, $[_{54}Xe]$ $[_{86}Rn]$, $[_{118}Og]$

Short cut electronic configuration

$_{10}Ne: 1s^2\ 2s^2\ 2p^6$

$_{17}Cl: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ → $_{17}Cl: [Ne]\ 3s^2\ 3p^5$

$_{18}Ar: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

$_{28}Ni: 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$ → $_{28}Ni: [Ar]\ 3d^8\ 4s^2$

Practice 4.6. **Which of the following is the short cut electronic configuration of $_{40}Zr$? Atomic numbers: $_{18}Ar$, $_{36}Kr$ and $_{54}Xe$**

a) $[Ar]\ 4d^2 5s^2$ b) $[Xe]\ 4d^2 5s^2$ c) $[Kr]\ 4d^1 5s^2$ d) $[Kr]\ 4d^2 5s^2$

4.6. Periodic table of elements and properties of atoms

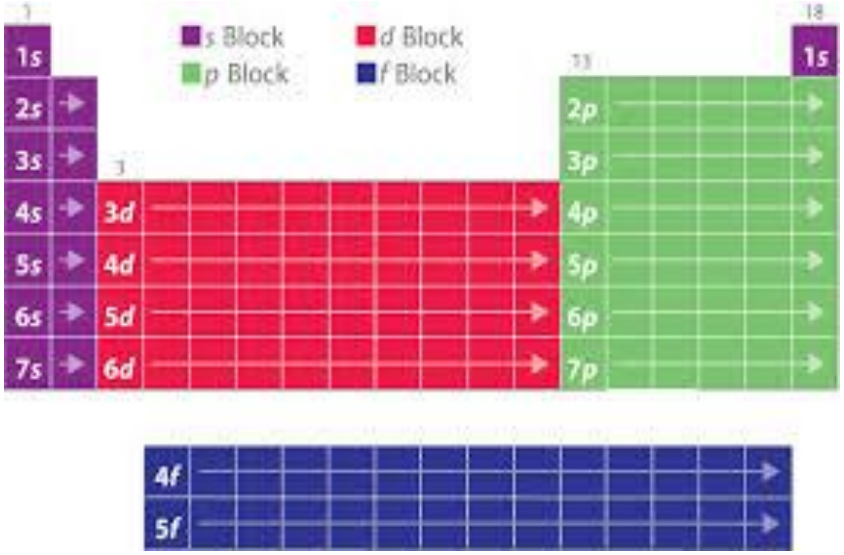
- The **vertical columns** of the table (Numbered from 1 to 18) are called **groups** or **families**. Element in the same group have similar but not identical characteristics.
- ***Elements in the same group have the same outer shell electron configuration (same number of outer shell electrons), and hence similar chemical properties***.
- The **horizontal rows** of the table (Numbered from 1 to 7) are called **periods**. Each **period** contains elements with electrons in the **same outer shell**.

Element	Configuration	Element	Configuration	Element	Configuration
H Z=1	$1s^1$	B Z=5	$1s^2 2s^2 2p^1$	He Z=2	$1s^2$
Li Z=3	$1s^2 2s^1$	C Z=6	$1s^2 2s^2 2p^2$	Ne Z=10	$1s^2 2s^2 2p^6$ (2p is now full)
Na Z=11	$1s^2 2s^2 2p^6 3s^1$	N Z=7	$1s^2 2s^2 2p^3$	Ar Z=18	$1s^2 2s^2 2p^6 3s^2 3p^6$
K Z=19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	O Z=8	$1s^2 2s^2 2p^4$	Kr Z=36	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
Rb Z=37	$1s^2 2s^2 2p^6 3s^2 3p^6 4 s^2 3d^{10} 4p^6 5s^1$	F Z=9	$1s^2 2s^2 2p^5$	Xe Z=54	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$

Elements in the same group have the same outer shell electron configuration; They behave the same chemically.

Each period contains elements with electrons in the same outer shell.

noble gases (in the last group) contain completely filled outer shell.

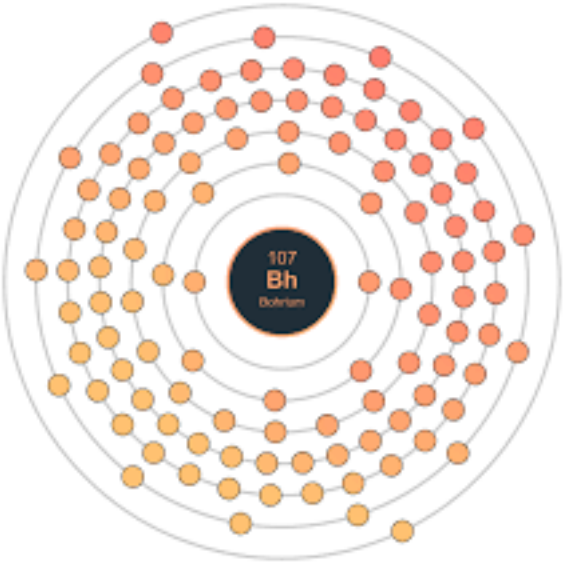


4.6.1. Effective nuclear charge (Z_{eff})

- The inner electrons have a repelling effect on the valence electrons. This effect, called shielding effect, reduces the electrostatic attraction between the positively charged protons in the nucleus and the outer electrons.

- Moreover, the repulsive forces between electrons in a many-electron atom further offset the attractive force exerted by the nucleus.

- The effective nuclear charge (Z_{eff}) is the net positive charge experienced (felt) by an electron in a multi-electron atom



$$Z_{eff} = Z - S$$

Z_{eff}: The effective nuclear charge,
Z: Atomic number of the atom,
S: Number of inner electrons
(in the core, also called shielding e-)

Practice: Calculate Z_{eff} for $_{11}Na$

Answer: $_{11}Na$: $1s^2 2s^2 2p^6 3s^1$ Z=11 and S= 10 so Z_{eff} = 11-10 = 1

Cation is smaller than the atom from which it is formed.

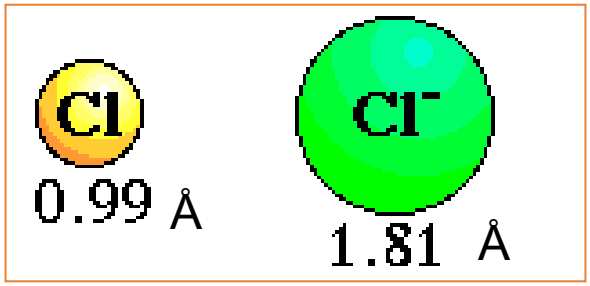
Excess of protons in the ion draws the outer electrons closest to nucleus.



Outer electrons of Na^+ feel **stronger attraction**

Anion is larger than the atom from which it is formed.

More repulsion between electrons and the ionic radius increases.

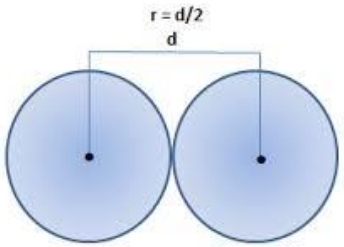


ELEMENTS OF THE SAME PERIOD			ELEMENTS OF THE SAME GROUP		
Element	S	$Z_{eff}=Z-S$	Element	S	$Z_{eff}=Z-S$
$_{11}Na:1s^2 2s^2 2p^6 3s^1$	10	11-10=1	$_{3}Li:1s^2 2s^1$	2	3-2=1
$_{13}Al:1s^2 2s^2 2p^6 3s^2 3p^1$	10	13-10=3	$_{11}Na:1s^2 2s^2 2p^6 3s^1$	10	11-10=1
$_{15}P:1s^2 2s^2 2p^6 3s^2 3p^3$	10	15-10=5	$_{19}K:1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	18	19-18=1
$_{17}Cl:1s^2 2s^2 2p^6 3s^2 3p^5$	10	17-10=7	$_{37}Rb:1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$	36	37-36=1

- **When we move from the left to the right across a period, Z_{eff} increases**
- **When we move in a group, Z_{eff} remains constant**

4.6.2. Atomic size or Radius

The atomic size (or radius, r) is the half distance of closest approach between two identical atoms.



When we move from the left to the right across a period, the atomic size decreases

When we move from the top to the bottom in a group, the atomic size increases



Atomic radii (in picometers) of representative elements according to their positions in the periodic table. Note that there is no general agreement on the size of atomic radii. We focus only on the trends in atomic radii, not on their precise values.

4.6.3. Ionization energy (I)

- **The ionization energy is the amount of energy it takes to detach one electron from a gaseous atom**

- If it is easy to detach an electron, it has low ionization energy.
- If it is hard to detach an electron, it has a high ionization energy
- ***The larger the atom the easier it is to detach an electron.*** *The smaller the atom the harder to detach an electron*

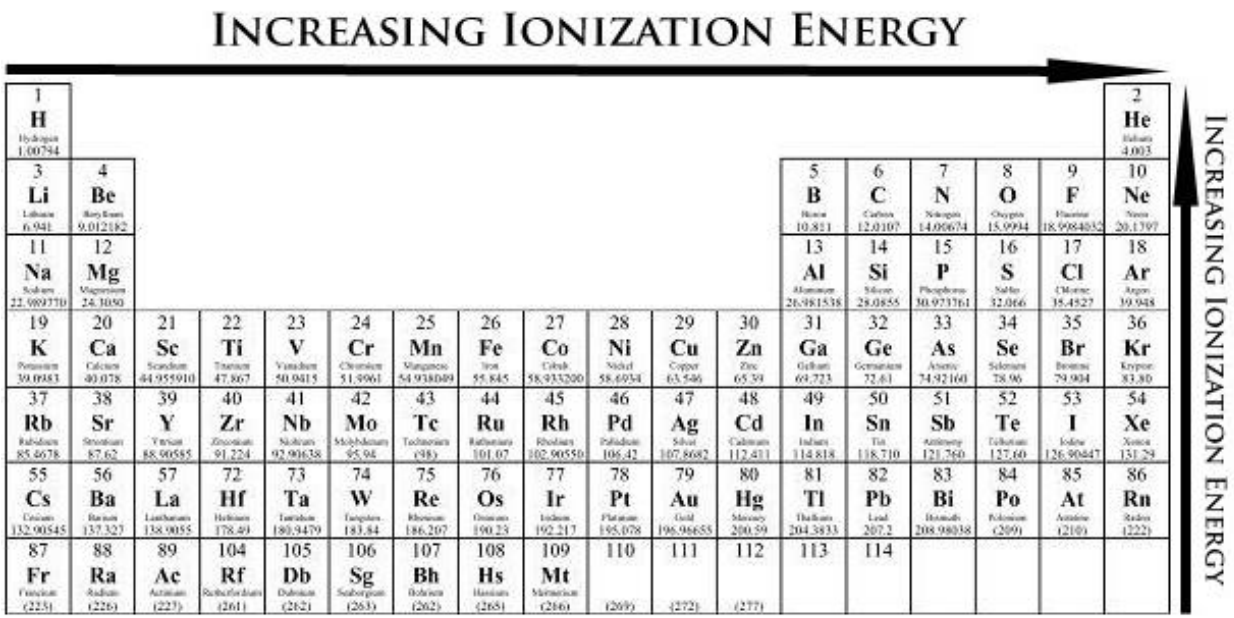
$$\text{energy} + X(g) \longrightarrow X^+(g) + e^- \quad \text{First ionization energy } I_1$$

$$\text{energy} + X^+(g) \longrightarrow X^{2+}(g) + e^- \quad \text{second ionization } I_2$$

$$\text{energy} + X^{2+}(g) \longrightarrow X^{3+}(g) + e^- \quad \text{third ionization } I_3$$

$$I_1 < I_2 < I_3$$

The first ionization energy increases from left to right across a period and from bottom to top in a group.

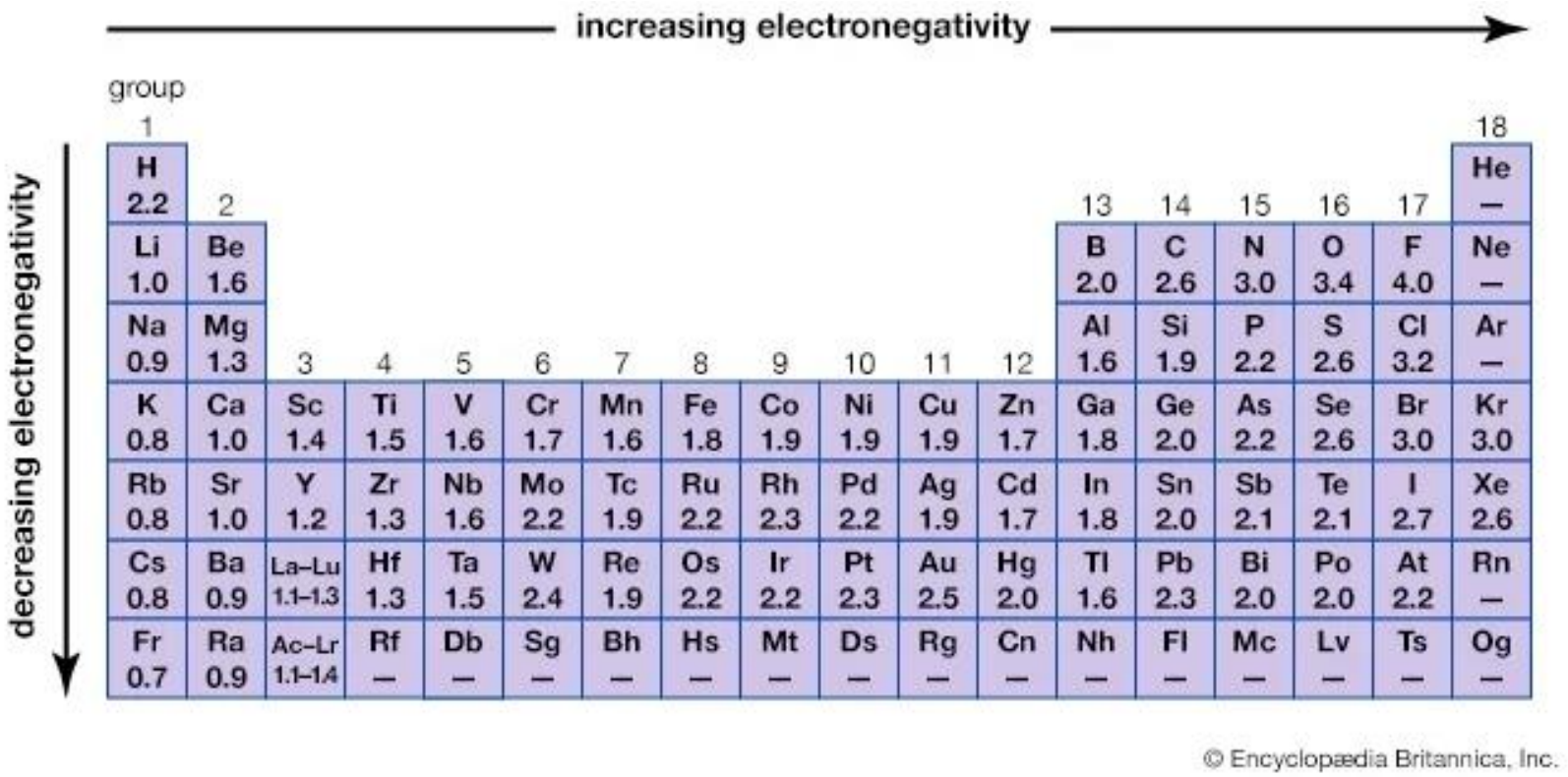


4.6.4. Electronegativity (E_N)

- **Electronegativity** is the ability of an atom to attract a pair of electrons towards itself in a chemical bond.

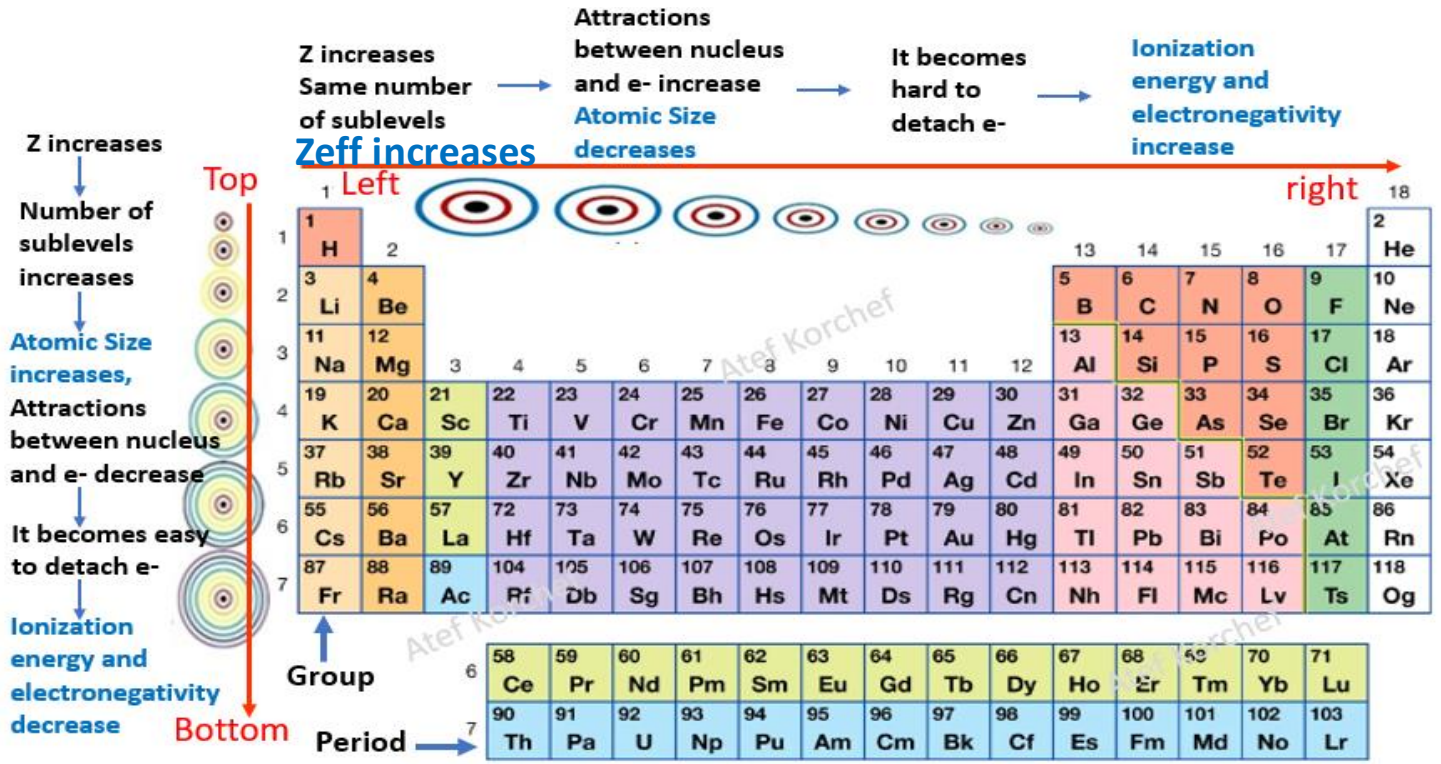
  Same electronegativity ⟹ A : A

  B is more electronegative than A ⟹ A :B



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Q7.1 to Q7.6 p.170 -172



Practice 4.6.

Arrange the following elements $_{12}$Mg, $_{16}$S, $_{13}$Al and $_{17}$Cl from the highest to the lowest

1. **atomic radius**
2. **ionization energy**
3. **Electronegativity**
4. **Effective nuclear charge (Z_{eff})**

Answer: **1.** decreasing atomic radius: Mg > Al > S >Cl

2. decreasing ionization energy: Cl > S > Al > Mg

3. decreasing electronegativity: Cl > S > Al > Mg

4. decreasing Z_{eff}: Cl > S > Al > Mg

Practice on chapter 4

Q1. The wave-particle duality of light means that

a)light consists of particles called photons that have wave-like properties b) electrons behave as a wave

b) light shining on certain metal plates caused a flow of electrons. d) all of these

Q2. The wavelength (λ) of the line corresponding to the transition of an electron from n = 3 to n = 2 is

($R_E = 2.18 \times 10^{-18}$ J, C= 3 x 10^8 m/s and h = 6.63 x 10^{-34} Js)

a) 486 nm b) 656 nm c) 1250 nm d) 32 nm

Q3. Which of the following sets of quantum numbers (n, l, m_l, m_s) describes an electron in an orbital

a)(2, 3, 1, 1/2) b) (3, 2, 3, 1/2) c) (2, 1, 1, -1/2) d) (3, 1, 0, -1/4)

Q4. Choose the correct order of the following sublevels from the lowest to the highest energy.

a) 3p < 4p < 3d < 4s b) 3p < 3d < 4s < 4p c) 3d < 3p < 4s < 4p d) 3p < 4s < 3d < 4p

Q5. Which one of the following about d orbitals is correct?

a) They are found in all principal energy levels. b) There are 5 types of d orbitals

c) They are spherical in shape d) Each d orbital can hold up to 3 electrons

Q6. What is the number of sublevels associated with n = 2

a)1 b) 2 c) 4 d) 9

Q7. What is the number of orbitals associated with n = 2

a)1 b) 2 c) 4 d) 9

Q8. The electronic configuration of copper $_{29}Cu$ is

a)$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$ b) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$ d) $s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5 4p^3$

Q9. The electronic configuration of $_{17}Cl$ is

a) $1s^2 2s^2 2p^6 3s^2 3p^6$ b) $1s^2 2s^2 2p^6 3s^2 3p^5$ c) $1s^2 2s^2 2p^6 3p^5 3s^2$ d) $1s^2 2p^6 2s^2 3p^5 3s^2$

Q10. Iron (Fe) is a transition metal. The electronic configuration of $_{26}Fe^{2+}$ is

a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ b) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$ c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ d) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$

Q11. How many unpaired electrons are there in $_{20}Ca^{2+}$?

a)0 b) 2 c) 10 d) 20

Q12. The ions $_8O^{2-}$, $_{11}Na^+$ and $_9F^-$

•are isotopes b) have isoelectronic configurations c) are compounds d) are called sublevels

a)Q13. The effective nuclear charge (Z_{eff}) of oxygen ($_8O$) is

b)a) 2 b) 4 c) 6 d) 8

Q14. Arrange the elements $_3Li$, $_{11}Na$, $_{19}K$ and $_{37}Rb$, of the same group, from the highest to the lowest electronegativity

a) $_3Li >_{11}Na >_{19}K >_{37}Rb$ b) $_{37}Rb >_{19}K >_{11}Na >_3Li$ c) $_3Li >_{19}K >_{11}Na >_{37}Rb$ d) $_{37}Rb >_{11}Na >_{19}K >_3Li$

Q15. In the periodic table, when we move from the top to the bottom in a group, the atomic size

a) increases b) decreases c) remains constant d) both increases and decreases

Q16. The energy of light is

a) inversely proportional to frequency
b) directly proportional to frequency
c) constant
d) equal to frequency

Q15. Two states differ in energy by 3.03 x 10^{-19} J. Calculate the frequency (υ) of the corresponding radiation.

a) 656 Hz b) 3.03x 10^{14} Hz c) 4.57 x 10^{14} Hz d) 12 Hz

Chapter 5

Molecular Geometry

At the end of this chapter, the student will be able to

- ❑ draw the Lewis dot representation of an atom or an ion.
- ❑ Recognize the different types of chemical bonding
- ❑ recognize the different molecular geometries.
- ❑ calculate the formal charge of an atom in a molecule or an ion.
- ❑ Define the resonance structures.
- ❑ recognize the VSEPR theory.
- ❑ determine the geometry of molecules.

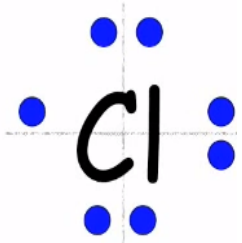
5.1. Lewis dot representation of atoms

Lewis dot representations or electron dot diagrams are a simplistic way of showing the valence électrons of an atom.

Example 5.1.

$_{17}Cl$: $1s^2 2s^2 2p^6 3s^2 3p^5$

- Write out the chemical symbol
- Find the number of valence electrons
- Draw the valence electrons in the following configuration



Valence Electrons = 7

Elements in the same group have the same Lewis dot representations because they have the same valence shell electron configuration.

Group	1	2	13	14	15	16	17	18
Element	Li	Be	B	C	N	O	F	Ne
Electron Configuration	$[He]2s^1$	$[He]2s^2$	$[He]2s^2 2p^1$	$[He]2s^2 2p^2$	$[He]2s^2 2p^3$	$[He]2s^2 2p^4$	$[He]2s^2 2p^5$	$[He]2s^2 2p^6$
Lewis Symbol	Li·	·Be·	·Ḃ·	·Ċ·	·Ṅ:	:Ȯ:	·Ḟ:	:N̈e:
	Na	Mg	Al	Si	P	S	Cl	Ar
	$[Ne]3s^1$	$[Ne]3s^2$	$[Ne]3s^2 3p^1$	$[Ne]3s^2 3p^2$	$[Ne]3s^2 3p^3$	$[Ne]3s^2 3p^4$	$[Ne]3s^2 3p^5$	$[Ne]3s^2 3p^6$
	Na·	·Mg·	·Ȧl·	·Ṡi·	·Ṗ:	:S̈:	·C̈l:	:Ar̈:

When we draw electron dot diagrams for ions, we show the original valence shell of the atom and not the valence shell of the ion.

Example 5.2.

$_{11}Na$: $1s^2 2s^2 2p^6 3s^1$



$_{11}Na^+$: $1s^2 2s^2 2p^6$

Na^+

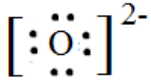
(no dots around the Na)

Example 5.3.

$_{8}O$: $1s^2 2s^2 2p^4$



$_{8}O^{2-}$: $1s^2 2s^2 2p^6$



Q8.2 and Q8.3 p.200

5.2. Chemical bonding

Ionic bond results from the transfer of electrons from a **metal to a nonmetal**. It involves the electrostatic attraction between oppositely charged ions (cation and anion)

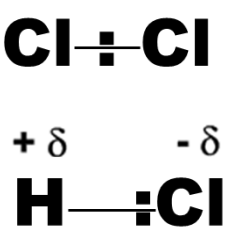


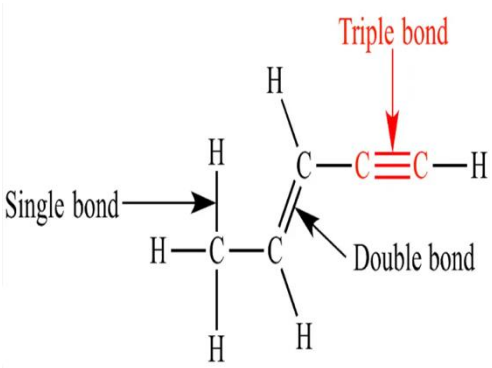
$$\cdot Li + :\ddot{F}\cdot \longrightarrow Li^+ \; :\ddot{F}:^- \quad \text{(or LiF)}$$

$$1s^2 2s^1 \qquad 1s^2 2s^2 2p^5 \qquad 1s^2 \qquad 1s^2 2s^2 2p^6$$

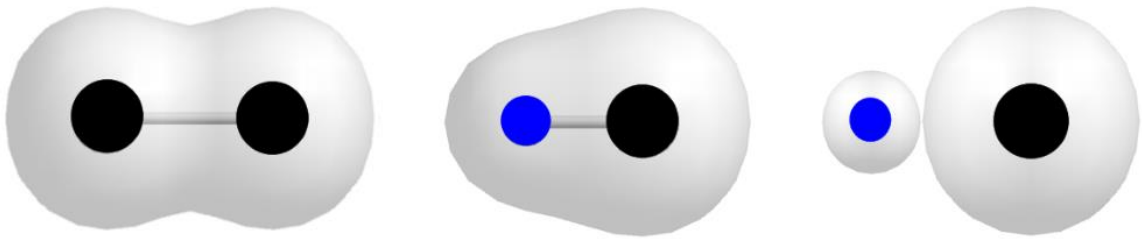
Covalent bond results from sharing electrons between **nonmetals.**

- Nonpolar covalent bond is when nonmetals equally share electrons between them.
- Polar covalent bond occurs between two nonmetals in which electrons are shared unequally.





Covalent bonds include simple, double, and triple bonds.



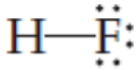
https://www.angelo.edu/faculty/kboudrea/general/shapes/00_lewis.htm

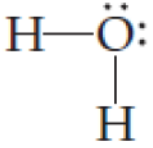
Q8.1 p.200

5.3. The octet rule

- **Octet rule:** In chemical bonding, atoms gain, lose, or share electrons until they are surrounded by **8 valence electrons** (4 electron pairs).
- Hydrogen requires only 2 electrons

Examples 5.5.





- **Note that there are many exceptions to the octet rule:**

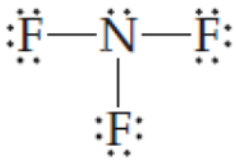
  ➢Atoms from the 3rd period (such as Na, Mg, Al, Si, P, S, Cl) and Xe can accommodate more than an octet (expanded octet).

  Examples 5.6. PCl_5, ClF_3, XeF_2, SF_6

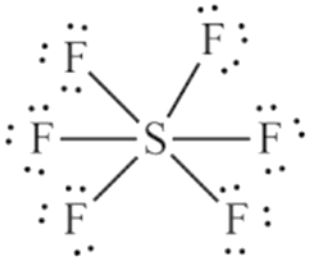
  ➢Be, B and N can accommodate less than an octet to form covalent compounds.

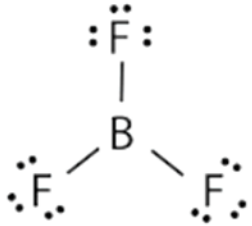
  Examples 5.7. $BeCl_2$, BF_3

Example 5.4.

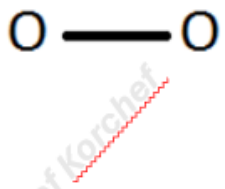
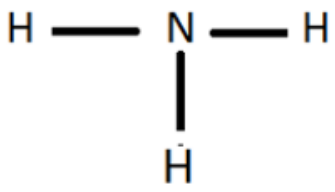
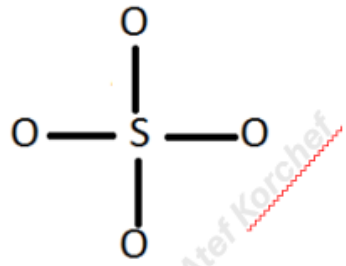


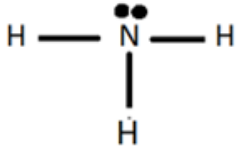
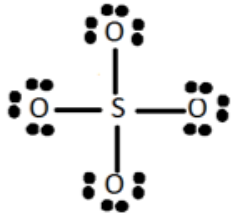
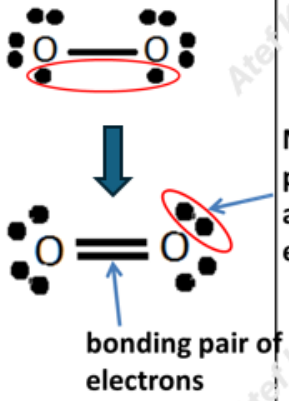
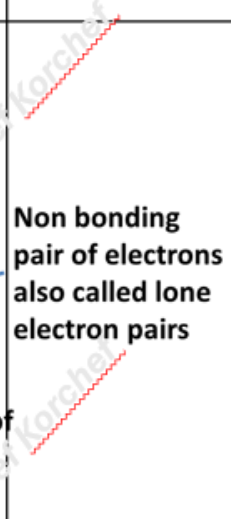
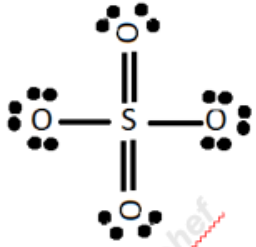
Example 5.6.





5.4. Rules for drawing Lewis structures

Rules	Example 1. O_2	Example 2. NH_3	Example 3. SO_4^{2-}
1. Identify metal/non metal and Add up the number of valence electrons for each of the atoms in the molecule. If the molecule is an ion, then add one electron for each negative charge; substract one electron for each positive charge	Oxygen is a nonmetal Nonmetal/nonmetal Covalent bond $_8O$: $1s^2 2s^2 2p^4$ VE= 6 e- Total valence electrons 2×6 = 12 e-	Non metal/non metal Covalent bonding NH_3 1 N 1 × 5 e- 3 H 3× 1 e- =8 e-	Non metal/non metal Covalent bonding SO_4^{2-} 1 S 1 × 6 e- 4 (O) 4× 6 e- + 2 e- =32 e-
2. Draw a trial structure. The trial structure has single bonds connecting the atoms in your molecule. Double or triple bonds are not part of this step. *The central atom is the atom which has the least electronegativity. Hydrogen (H) always goes on the outside*	 2 e- are involved in the trial structure of O_2	 2×3 = 6 e- are involved in the trial structure of NH_3	 2×4 = 8 e- are involved in the trial structure of SO_4^{2-}

Rules	Example 1.	Example 2.	Example 3.
3. Substract the number of electrons in your trail structure from the number of electrons in step 1. These are the electrons that need to be in your trial structure, but are not. It is possible to have this step be equal to zero	12e- - 2 e- = 10 e-	8 e- - 6 e- = 2 e-	32e- - 8 e- = 24 e-
4. Take the electrons from step 3 and put this on your trial structure. Your goal is to give every atom (except hydrogen) an octet. Whatever you do, don't put down more electrons than you are allowed.			
5. If you cannot give every atom an octet is step 4, then rearrange the electrons that you have to form double or triple bonds. This step is a last resort. •Note that beyond the third period, the d-orbitals are low enough in energy to participate in bonding and accept the extra electron density. Sulfur for example can accommodate more than one octet by using its empty valence d orbital.			A second possible structure of SO_4^{2-} 

Atef Korchef

5.5. Formal charge

Formal charge (FC) is the charge that an atom would have if electronegativity differences are ignored:

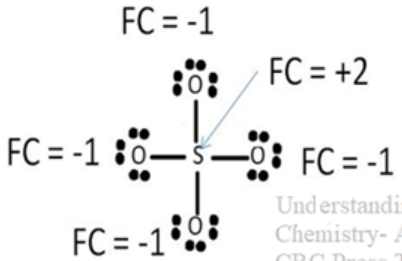
$$FC = n(Ve^-) - \left[n(NBe^-) + \frac{n(Be^-)}{2}\right]$$

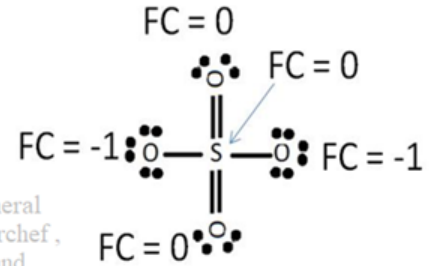
where n(Ve-) is the number of valence electrons,
n(NBe-) is the number of nonbonding electrons and n(Be-) is the number of bonding electrons

The sum of all formal charges must equal the overall charge of the molecule/ion

Best Lewis structure has the fewest number of nonzero formal charges

Structure 1		**Structure 2**	
atom	FC	atom	FC
S	6-0-(8/2) = 2	S	6-0-(12/2) = 0
O (4 atoms)	6-6-(2/2) = -1	O (2 atoms)	6-4-(4/2) = 0
		O (2 atoms)	6-6-(2/2) = -1





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The sum of all formal charges equals the overall charge of SO_4^{2-} ion (= -2).

The best Lewis structure of SO_4^{2-} is structure 2 because it has the fewest number of nonzero formal charges.

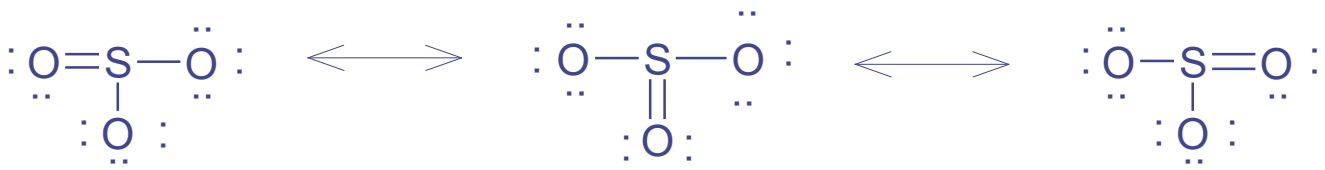
5.6. Resonance structures

Resonance is a manner of describing delocalized electrons within certain molecules or polyatomic ions where the bonding cannot be expressed by a single Lewis formula.

A molecule or an ion with such delocalized electrons is represented by two or more contributing structures to show bonding called resonance structures.

Double-headed arrows are used to indicate the resonance formulas.

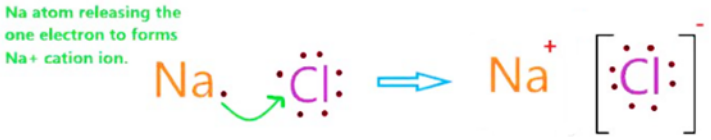
Example 5.8. Three possible structures for SO_3



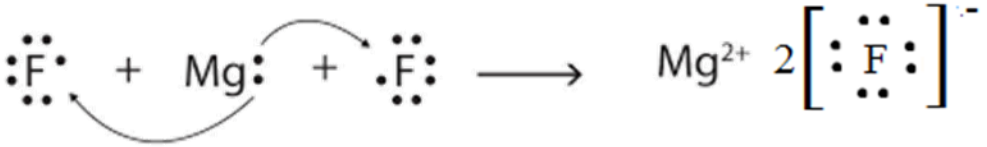
5.7. Lewis structures of ionic compounds

Examples 5.7.

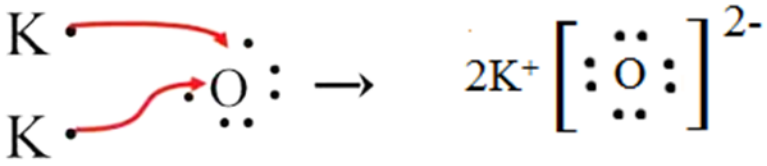
Electrons are transferred from a metal to a nonmetal to form an ionic compound. In the Lewis structure of the ionic compounds the nonmetal (anion) is put in brackets.



Lewis structure of NaCl



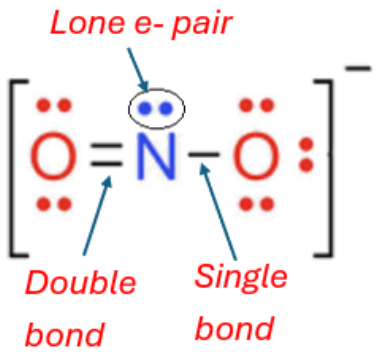
Lewis structure of MgF_2



Lewis structure of K_2O

5.8. Electron Groups and Molecule Notation

On a **central atom**
1.Each lone pair of electrons constitutes one electron group.
2.Each bond constitutes one electron group regardless of whether it is a single, a double or a triple bond.



There are **3 electron groups** on the central atom N of NO_2^-
1 lone e- pair
1 single bond
1 double bond

AX_2E

For the notation of a molecule:
The letter **A** is used to represent the central atom,
Xm is used to represent the number of atoms (m) bonded to the central atom
En is used to represent the number of lone electron pairs (n) on the central atom.

Practice: Determine the electron groups and the notation of the molecule HCN.

Atomic numbers: $_1H$, $_6C$, $_7N$,

Answer:

Lewis structure



There are **2 electron groups** on the central atom C of HCN
1 single bond
1 triple bond

AX_2

5.9. Valence Shell Electron Pair Repulsion (VSEPR) model

regions of high electron density around the central atom go as far apart as possible to minimize repulsions.

- Lone pairs of electrons (unshared pairs) require more volume than shared pairs

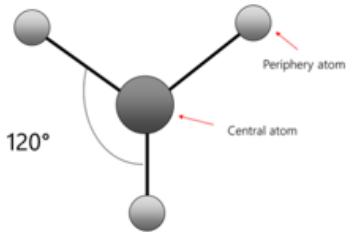
- There is an ordering of repulsions of electrons around central atom

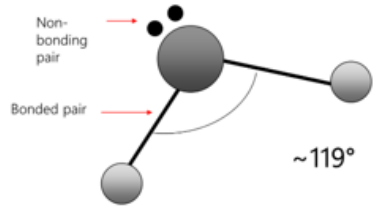
  - lone pair to lone pair repulsion is strongest
  - lone pair to bonding pair repulsion is intermediate
  - bonding pair to bonding pair repulsion is weakest

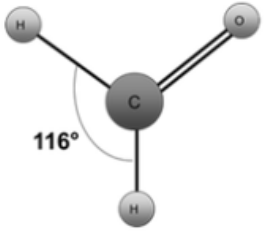
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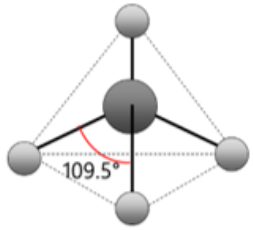
- Electron domains for multiple bonds exert a greater repulsive force on adjacent electron domains than do electron domains for single bonds (i.e., CH_2O)

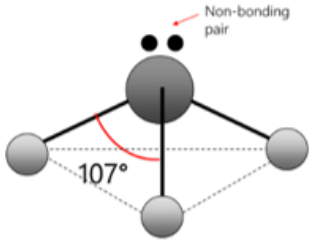
Example 5.8.

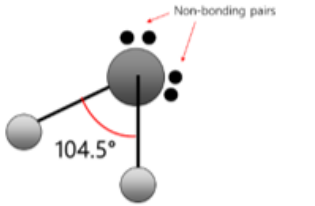










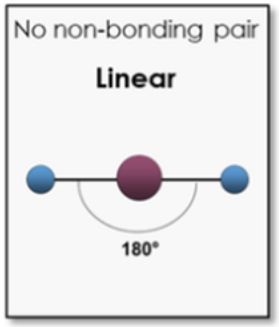


lone pair to lone pair repulsion is why bond angles in water are less than 109.50

5.10. Molecular Shapes

Basic and **derivative** shapes

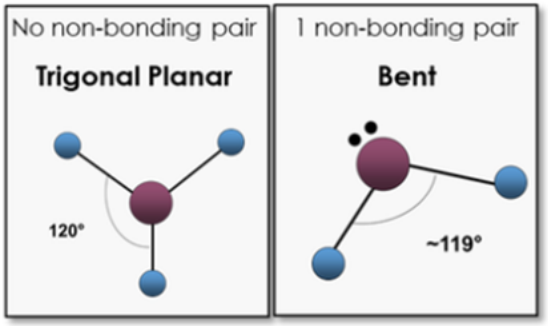
(Examples 5.9-20.)



AX_2

Examples:

$HgCl_2$, ZnI_2 , CS_2 , CO_2



AX_3

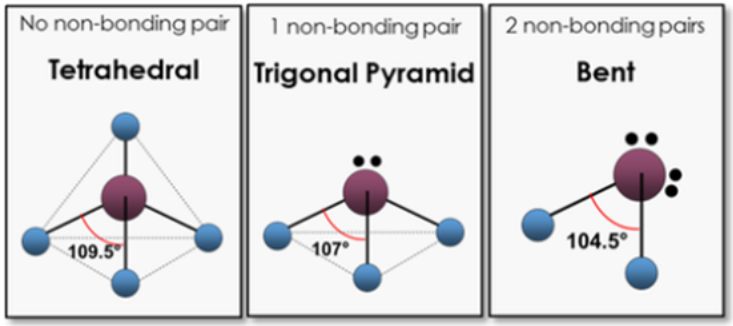
Examples:

BF_3 , GaI_3

AX_2E

Examples:

SO_2 , $SnCl_2$



AX_4

Examples:

CCl_4 , CH_4 , BF_4^-

AX_3E

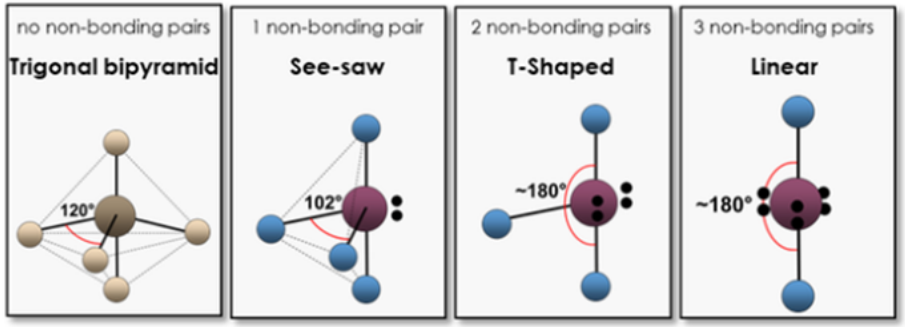
Examples:

NH_3 , OH_3^-

AX_2E_2

Examples:

H_2O , $SeCl_2$



AX_5

PCl_5 , PF_5

AX_4E

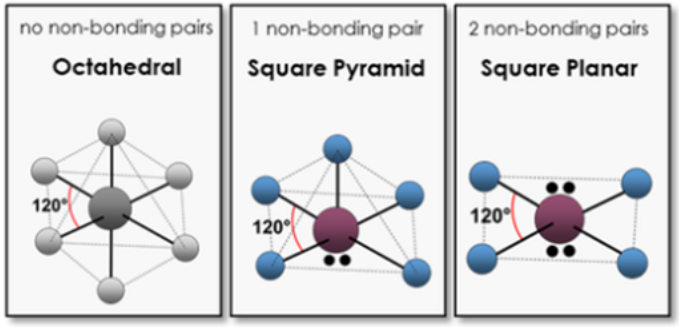
$TeCl_4$, SF_4

AX_3E_2

ClF_3 , BrF_3

AX_2E_3

I_3^- , ICl_2^-



AX_6

SF_6 , PF_6^-

AX_5E

IF_5 , BrF_5

AX_4E_2

ICl_4^- , BrF_4^-

5.11. How do we determine the shape of a molecule?

- Draw the Lewis structure
- How many things are there around the central atom? (count the total number of electron pairs around the central atom ; count multiple bonds as one bonding pair)
- What structure is this molecule based on? (arrange the electron pairs in one of the five basic geometries to minimize electron- electron repulsion)
- Which atoms are replaced by lone electron pairs?
- What is the final structure?

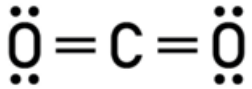
- N.B. The arrangement of electron groups (or electron domains) around the central atom is called **electron-domain geometry** (which corresponds to the basic geometry), i.e., the electron domain geometry of SO_2 is trigonal planar and its molecular geometry is bent.

Practice 5.1

Application: Determine the shape of CO_2 and SO_2 molecules.

Atomic numbers: $_6C$, $_8O$, $_{16}S$

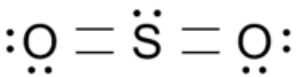
Answer:



2 groupes (double bonds)

AX_2

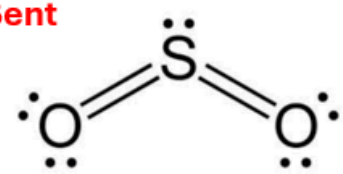
Linear

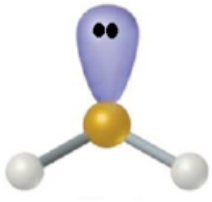


3 groupes (2 double bonds and 1 lone e- pair)

AX_2E

Bent





molecular geometry: Bent

Electron-domain geometry: Trigonal planar

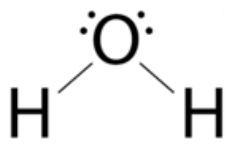
Atef Korchef

Practice 5.2

Practice: Determine the shape of H_2O, HCO_3^-, O_3, and IF_5 molecules.

Atomic numbers: ${}_1H$, ${}_6C$, ${}_8O$, ${}_9F$, ${}_{53}I$

Answer:



4 groupes (2 single bonds and 2 lone e- pair)

AX_2E_2

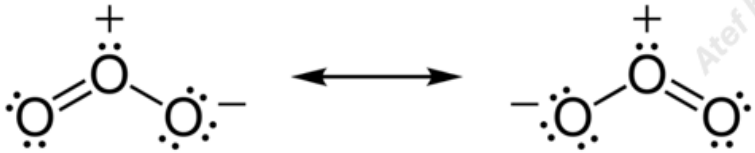
Bent



3 groupes (2 single bonds and 1 double bond)

AX_3

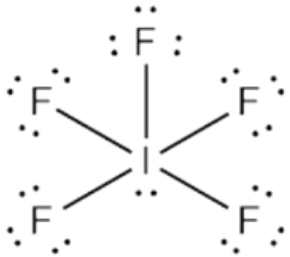
Trigonal planar

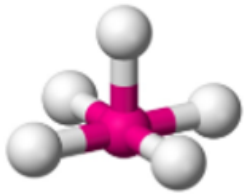


3 groupes (1 single bond, 1 double bond and 1 lone e- pair)

AX_2E

Bent





6 groupes (5 single bonds, and 1 lone e- pair)

AX_5E

Square pyramid

Q8.3 to Q8.6 p.200-201

Practice on chapter 5

1. How many groups does the central atom in water (H_2O) have?

A) 1
B) 2
C) 3
D) 4

2. What does VSEPR stand for?

A) Valence Shell Electron Pair Repulsion
B) Valence Shell Electron Proton Repulsion
C) Variable Shell Electron Pair Repulsion
D) Valence Shell Energy Pair Reorganization

3. According to VSEPR theory, which of the following shapes corresponds to a molecule with four bonding pairs and no lone pairs?

A) Linear
B) Trigonal Planar
C) Tetrahedral
D) Octahedral

4. Which molecule has a bent molecular geometry due to two lone pairs of electrons on the central atom?

A) CO_2
B) H_2O
C) NH_3
D) CH_4

5. What is the molecular shape of SF_6 according to VSEPR theory?

A) Octahedral
B) Tetrahedral
C) Pyramidal
D) Linear

6. In the nitrate ion (NO_3^-), which atom carries a formal charge of -1?

A) Nitrogen
B) One of the oxygen atoms
C) All oxygen atoms
D) The entire ion carries the negative charge

7. How many lone pairs of electrons are present in the central atom of NH_3?

A) 1
B) 2
C) 0
D) 3

8. Which molecule exhibits a trigonal planar geometry?

A) BF_3
B) H_2O
C) NH_3
D) CH_4

9. Which of the following has a linear shape?

A) H_2O
B) CO_2
C) NH_3
D) SF_4

10. What is the molecular geometry for a molecule with five bonding pairs and one lone pair?

A) Octahedral
B) Trigonal Bipyramidal
C) Square Pyramidal
D) T-shaped

Chapter 6

Intermolecular Forces and Properties of Matter

At the end of this chapter, the student will be able to

- ❑ Differentiate between polar and nonpolar molecules
- ❑ Identify and differentiate between the different types of intramolecular forces and intermolecular forces
- ❑ interpret properties of matter based on the intermolecular forces,
- ❑ define the thermodynamic equilibrium between the different states of matter
- ❑ recognize the exothermic and endothermic phase changes
- ❑ analyze a phase diagram.

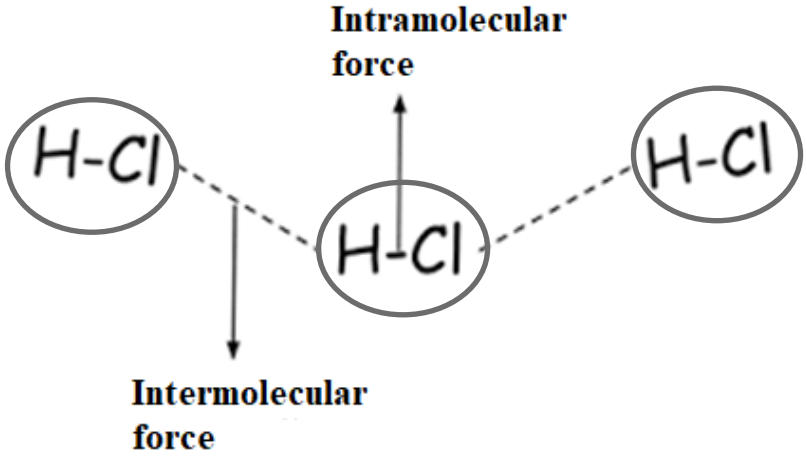
6.1. Introduction

***Inter*molecular forces** occur **between** molecules**:**

- ***van der Waals interactions***
  - **Keesom forces (between polar molecules)**
  - **Debye forces (between a nonpolar molecule and a polar molecule)**
  - **London dispersion forces (between nonpolar molecules)**

- Ion-dipole forces

- **Hydrogen bond (**Hydrogen on one molecule attached to (**N, O, or F**) atoms and either (**N, O, or F**) on another molecule.



In a substance, intramolecular forces (chemical bonds) are <u>stronger</u> than intermolecular forces.

6.2. Examples of Polar and nonpolar molecules

Reminder: ***Electronegativity*** is a measure of the tendency (ability) of an atom to attract a bonding pair of electrons towards itself.

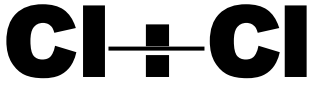
Nonpolar molecules occur when there is an equal sharing of electrons between their atoms.

Examples (6.1) of nonpolar molecules:

- O_2, H_2, Cl_2 , I_2, Br_2
- CO_2



- Hydrocarbons: CH_4 , C_2H_6 , C_3H_8 ,….



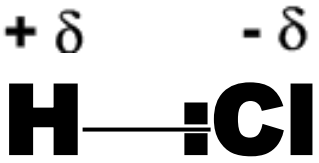
There is an equal sharing of electrons.
Cl_2 is a nonpolar molecule.

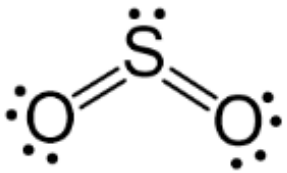
Polar molecules occur when atoms do not share electrons equally.

A dipole *forms, with part of the molecule carrying a slight positive charge and the other part carrying a slight negative charge. This happens when there is a difference between the electronegativities of the atoms or because of the molecular geometry.*

Examples (6.2) of polar molecules:

- HBr, HCl, HI, H_2O, HF, NH_3
- SO_2 , O_3





Cl is more electronegative than H. The two atoms did not share electrons equally. **HCl is a polar molecule. HCl molecule forms a permanent dipole.**

6.3. Intermolecular forces:

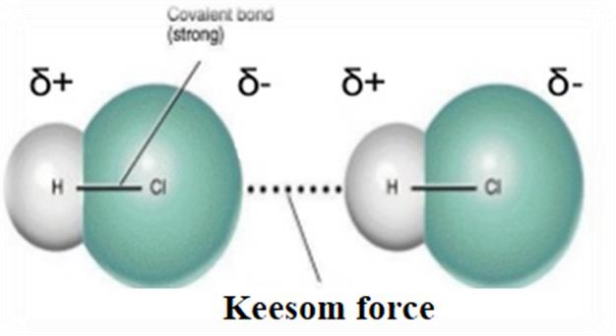
6.3.1. Keesom Forces (Dipole-dipole interactions)

- **Act *between polar molecules*.**
- Oppositely charged ends attract and like ends repel.
- Moderate strength



Interaction between permanent dipoles.

Examples 6.3. HCl, HBr, HI



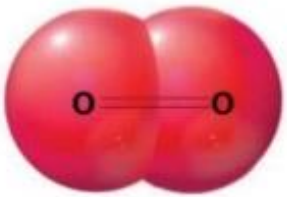
Keesom forces become stronger when

- ❑ the difference of **electronegativity** between the atoms forming the molecule increases.
- ❑ the temperature decreases.
- ❑ the distance between molecules decreases.

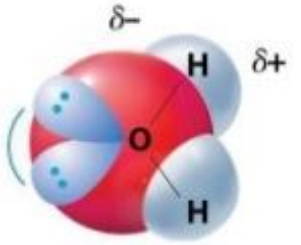
6.3.2. Debye Forces

- Occur between **a nonpolar** molecule and a **polar** molecule.
- The permanent dipole in the polar molecule induces an electric dipole in the nonpolar molecule when it comes extremely close.

Examples 6.4. Debye forces between O_2 and H_2O

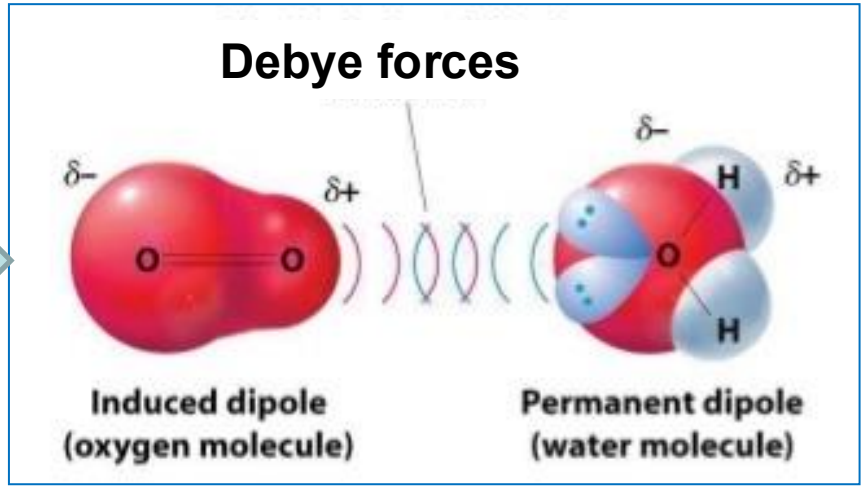


Oxygen molecule
(nonpolar)



Water molecule
(polar; permanent dipole)

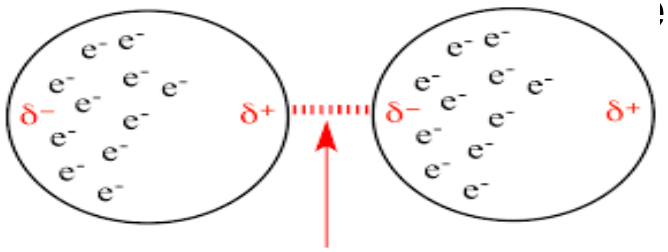
When O_2 comes extremely close to H_2O



6.3.3. London Dispersion Forces

❑ **Polarizability** is the ease with which the electron distribution around an atom or molecule can be distorted ('How easy is it to form an induced dipole').

❑ Instantaneous dipole, that occurs accidentally to electrons movement, induces a similar dipole neighboring atom or molecule.



London dispersion force

❑ London forces occur in **nonpolar molecules**.

❑ **Weak** forces

Examples 6.5.

Homonuclear diatomic molecules: H_2, Cl_2, I_2, O_2, F_2, Br_2, …

Hydrocarbons: CH_4, C_3H_8, C_6H_{14} , ….

Carbon dioxide CO_2

Noble gases : He, Ne, Ar, ….

London dispersion forces become stronger when

❑ the polarizability increases. Polarizability increases when

- ✓ The size of atoms or molecules increases.
- ✓ The number of electrons increases.

❑ the distance between molecules decreases.

❑ the chain of carbon atoms in hydrocarbons becomes longer.

PRACTICE

Practice 6.1. **Which of the followings shows the highest London dispersion forces?**

a) C_3H_8 **b) C_4H_{10}** **c) C_5H_{12}** **d) C_6H_{14}**

Answer : **C_6H_{14} shows the highest London dispersion forces**

$$C_3H_8 < C_4H_{10} < C_5H_{12} < C_6H_{14}$$

Increasing London forces

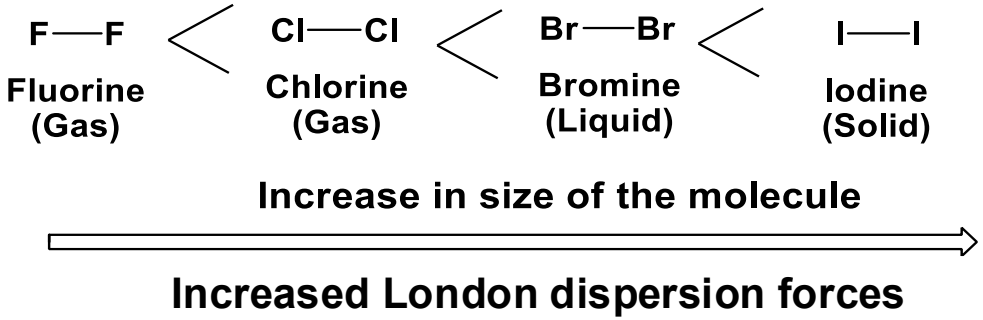
Practice 6.2. **According to the following order from the lowest to the highest atomic size:**

F < Cl < Br < I

explain why F_2 and Cl_2 are gases, Br_2 is a liquid and I_2 is a solid at room temperature.

Answer :

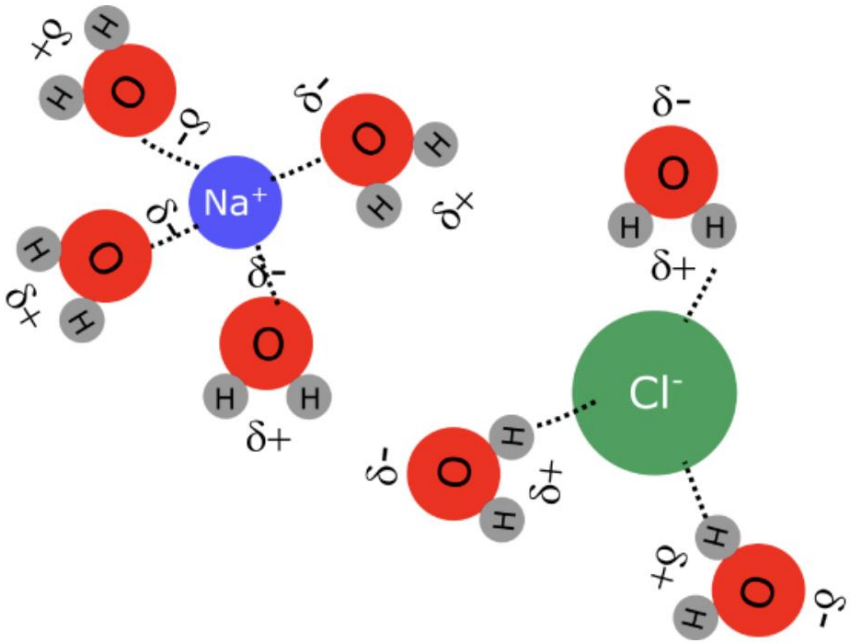
F_2, Cl_2, Br_2 and I_2 are nonpolar molecules, and they show London dispersion forces. As the size increased, London forces increased:



6.3.4. Ion-dipole Forces

- Occur between an ion (cation or anion) and a polar molecule (a dipole)
- Commonly found in solution where ionic compounds are dissolved in polar liquids

Example 6.6. Dissolution of NaCl in water



https://chem.libretexts.org/

Factors affecting **ion–dipole forces** include:

1. Charge of the ion

- Higher ionic charge produces stronger attraction.
- Example 6.7: Mg^{2+} attracts water molecules more strongly than Na^{+}.



https://chem.libretexts.org/

2. Size of the ion

- Smaller ions have higher charge density, so they create stronger ion–dipole forces.
- Example 6.8: Li^{+} forms stronger interactions with water than K^{+}.

3. Polarity of the molecule

- Molecules with a larger dipole moment ($\mu = \delta \times d$) interact more strongly with ions.
- Example 6.9: Water shows stronger ion–dipole attraction than less polar molecules.

4. Distance between ion and dipole

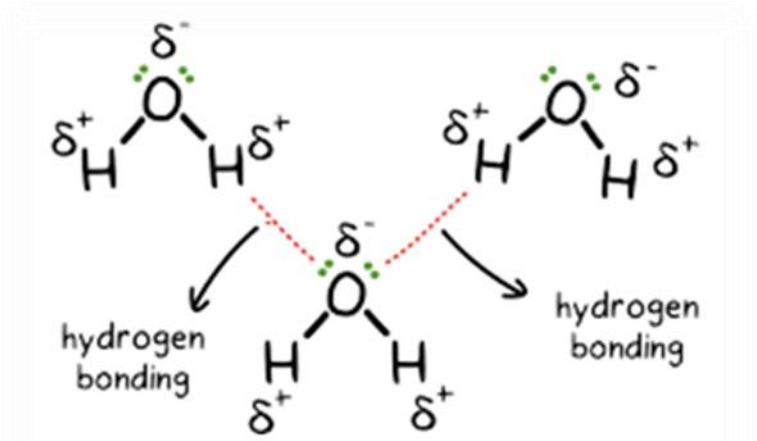
- Shorter distance between the ion and the polar molecule increases the force of attraction.

6.3.5. Hydrogen Bonding

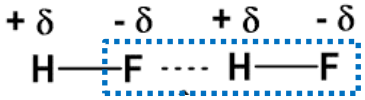
- Hydrogen on one molecule attached to a highly electronegative atom (**N, O, or F**) and either (**N, O, or F**) on another molecule.
- **Strong** dipole-dipole force.
- Occurs between Polar molecules.

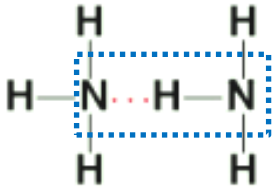
Types of Hydrogen Bonds

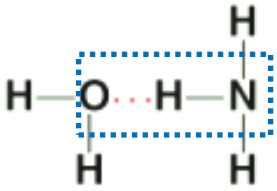
N—H···N—	O—H···N—	F—H···N—
N—H···O—	O—H···O—	F—H···O—
N—H···F—	O—H···F—	F—H···F—

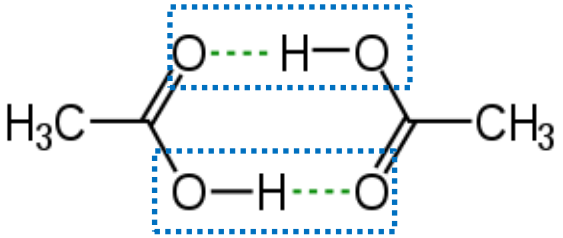


Examples 6.10. H_2O, HF, NH_3, CH_3COOH







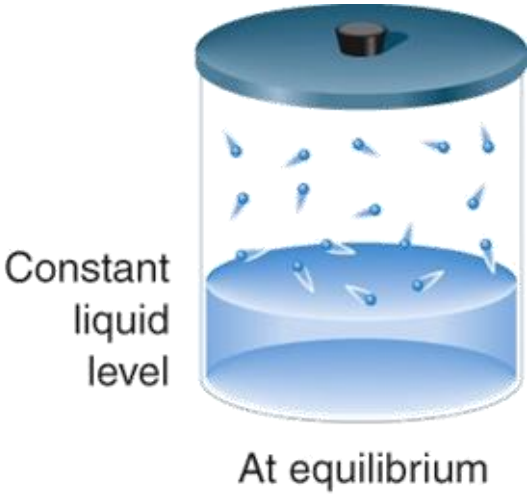


Q9.1 to Q9.3 p.230 -231

6.4. Properties of Matter

6.4.1. Vapor pressure

The pressure exerted by the vapor on the surface of the liquid **at equilibrium** is called the **vapor pressure**.



In a **closed system**, when the rate at which the liquid is entering the gas phase equals the rate at which the vapor is returning to the liquid phase, the system is at equilibrium. After this time, the liquid level will remain constant. The pressure exerted by the vapor on the surface of the liquid **at equilibrium** is called the **vapor pressure**.

The vapor pressure of a liquid increases when

- ✓ the intermolecluar forces decreases.
- ✓ the temperature increases.

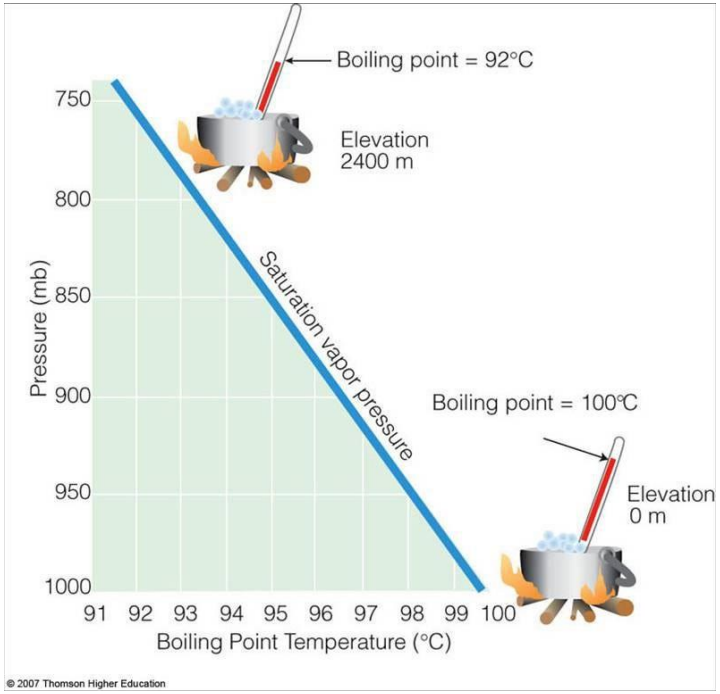
6.4.2. Boiling Point

The temperature at which the vapor pressure equals the external pressure (atmospheric pressure) is called the **boiling point**. This is the point where bubbles of vapor form within the liquid.



- The boiling point of a liquid at **1 atm** pressure is the **normal boiling point**.

- The boiling point **increases** with **increasing external pressure**.
- The molecules that have **strong intermolecular** forces have **high boiling point**.



PRACTICE 6.3.

Which of the compounds hexane (C_6H_{14}) and propane (C_3H_8) has the highest boiling point ? Which of these two compounds has the lowest vapor pressure ? Explain your answer.

Answer: Hexane C_6H_{14} *has a longer chain than propane* C_3H_8. *Therefore,* C_6H_{14} *is held by higher attractive molecular forces via London forces than* C_3H_8.

Hexane C_6H_{14} has lower vapor pressure than propane C_3H_8.
Hexane C_6H_{14} has higher boiling point than propane C_3H_8.

The substance having the highest boiling point has the lowest vapor pressure and conversely, the substance having the lowest boiling point has the highest vapor pressure.

6.4.3. Viscosity of a liquid

❑ **Viscosity** is **the resistance to flow** of a liquid.

When the viscosity of a liquid increases, the liquid flows more slowly.



The viscosity of honey is higher than the viscosity of water. Honey flows more slowly than water.

Viscosity of a liquid increases when

✓the **intermolecular forces** increases

✓the **temperature** decreases

Q9.4 p.233

Practice on chapter 6

Q1. Keesom forces are present between

a) nonpolar molecules. b) CO_2 molecules. c) polar molecules. d) hydrocarbons molecules.

Q2. Which of the following changes would increase the vapor pressure of a liquid?

a) An increase in temperature. b) An increase in the intermolecular forces in the liquid.

c) An increase in the viscosity of the liquid. d) An increase in the size of the open vessel containing the liquid.

Q3. Ion-dipole forces are present between

a) Na+ and Cl^-. b) H_2O and Cl^-.

c) H_2O and O_2. d) CO_2 and C_2H_6

Q4. The intermolecular forces in HBr molecules are

a) Keesom forces b) polar covalent bonds c) London forces d) hydrogen bonds.

Q5. Which one of the following compounds shows hydrogen bonds

a) HCl b) NH_3 c) C_2H_6 d) C_5H_{12}

Q6. Which of the following compounds shows Keesom forces

a) HCl b) CH_4 c) O_2 d) C_5H_{12}

Q7. Which of the following shows London dispersion forces

a) HCl b) O_3 c) NH_3 d) C_6H_{14}

Q8. Which one of the following compounds shows the strongest London forces?

a) CH_4 b) C_3H_8 c) C_4H_{10} d) C_5H_{12}

Q9. When honey is heated, its viscosity

a) increases b) decreases c) remains constant d) All of these

Q10. The boiling point

a) is the pressure exerted by the vapor on the liquid surface at equilibrium.

b) is the resistance of a liquid to flow.

c) is the temperature at which the vapor pressure of a liquid equals the external pressure.

d) decreases when intermolecular forces become stronger.

Chapter 7

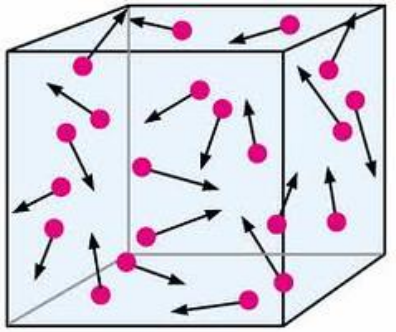
Gases

At the end of the present chapter, the student would be able to:

- ❑ Interpret the temperature and the pressure of a gas.
- ❑ recognize the different gas laws and derive the ideal gas law.
- ❑ calculate the gas density from the general law of ideal gas
- ❑ define the partial pressure of a gas in a mixture of gases and the relationship between the partial pressure and the total pressure.
- ❑ define diffusion and effusion of gases
- ❑ differentiate between an ideal gas and a real gas

7.1. What is an ideal gas?

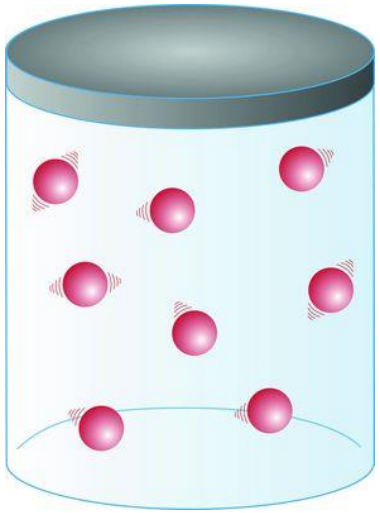
Kinetic Theory of Gases



- The gas particles occupy no volume. That is, they are points.
- Gas particles move in straight lines, being separated from each other and they only change directions while colliding against each other or against the walls of the container.
- These collisions are elastic. That is, no energy is gained or lost during collisions.
- There is no attractive or repulsive forces between the gas particles.
- The average kinetic energy K_E of these particles is directly proportional to the absolute temperature T (K). This means that hotter the gas is, faster the gas particles move:

A gas can be described by the following variables:

- **Temperature (T):** [T (K) = T(°C) + 273]
- **Number of moles (n):** [n = m/ M]
- **Volume (V)**
- **Pressure (P):** force per unit area (P = F/A)



(P, V, n, T)

SI unit of pressure is Pascal (Pa) but the units bar, atmosphere (atm) and millimeters of mercury (mmHg) are commonly used

$$1 \text{ atm} = 1.013 \times 10^5 \text{ Pa} = 760 \text{ mmHg}$$

Practice 7.1. A gas has a pressure of P=120 kPa. What is the pressure in atm and mmHg?

Answer:

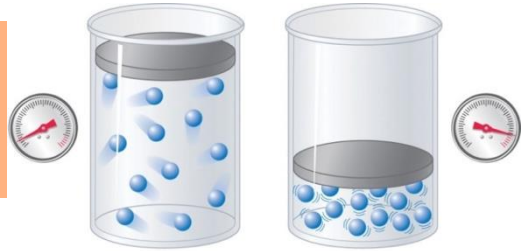
P = 120 kPa x 1000/(1.013 x 10^5) = 1.185 atm

P = 1.185 atm = 1.185×760 mmHg = 900.6 mmHg

7.2. Boyle's Law

At constant temperature and gas amount, the volume of the ideal gas is inversely proportional to its pressure

Boyle's Law: $V = \dfrac{\text{constant}}{P}$ **n and T are constant**



PV = constant $\longrightarrow$ $P_1V_1 = P_2V_2$

PRACTICE 7.2. **A sample of chlorine gas occupies a volume of 945 mL at a pressure of 750 mmHg. What is the pressure of the gas (in mmHg) if the volume is reduced to 315 mL at a constant temperature?**

Answer:

$P_1 = 750$ mmHg $\qquad P_2 = ?$

$V_1 = 945$ mL $\qquad V_2 = 315$ mL

At constant gas amount and temperature: $P_1 \times V_1 = P_2 \times V_2$

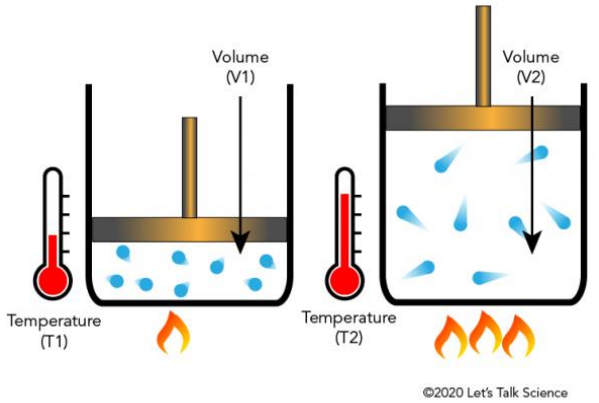
$$P_2 = \frac{P_1 \times V_1}{V_2} = \frac{750 \text{ mmHg} \times 945 \text{ mL}}{315 \text{ mL}} = 2250 \text{ mmHg}$$

7.3. Charles's Law

At constant pressure and gas amount, the volume of the ideal gas is directly proportional to its absolute temperature

Charles's Law: $V = \text{constant} \times T$ **n and P are constant**

$$\frac{V}{T} = \text{constant} \longrightarrow \frac{V_1}{T_1} = \frac{V_2}{T_2}$$



PRACTICE 7.3. **A sample of carbon monoxide gas (CO) occupies 3.2 L at 125°C. At what temperature will the gas occupy a volume of 1.6 L if the pressure remains constant?**

Answer:

$V_1 = 3.20$ L $\qquad V_2 = 1.60$ L

$T_1 = 125 + 273 = 398$ K $\qquad T_2 = ?$

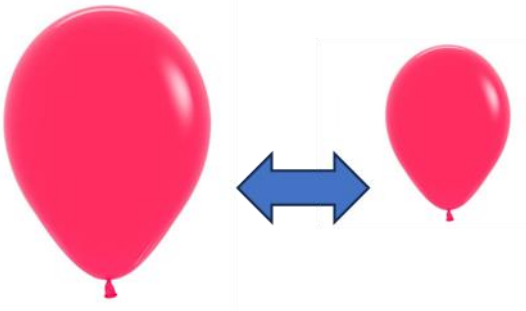
At constant pressure and gas amount: $V_1/T_1 = V_2/T_2$

$$T_2 = \frac{V_2 \times T_1}{V_1} = \frac{1.6\text{ L} \times 398\text{ K}}{3.2\text{ L}} = 199\text{ K}$$

7.4. Avogadro's Law

At constant temperature and pressure, the volume of the ideal gas is directly proportional to the amount of substance

Avogadro's Law: $V = \text{constant} \times n$ **P and T are constant**



$$\frac{V}{n} = \text{constant} \longrightarrow \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

PRACTICE 7.4. **A balloon contains 1.1 mol of helium (He) occupying 26.2 L. Calculate the volume will the gas occupy when the balloon contains 2.31 mol of He if the pressure and temperature remain constant?**

Answer: $n_1 = 1.1$ mol $\quad n_2 = 2.31$ mol)

$V_1 = 26.2$ L $\quad V_2 = ?$

At constant pressure and temperature: $V_1/n_1 = V_2/n_2$

$$V_2 = \frac{V_1 \times n_2}{n_1} = \frac{26.2 \text{ L} \times 2.31 \text{ mol}}{1.1 \text{ mol}} = \mathbf{55\ L}$$

7.5. Gay- Lussac's Law

At constant volume and amount of substance, the pressure of the ideal gas is directly proportional to the absolute temperature.

Gay-Lussac's Law : P = constant × T (V and n are constant)

$$\frac{P}{T} = \text{constant} \longrightarrow \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

PRACTICE 7.5. **The temperature of a sample of O_2 gas is 27 °C at 1.1 atm. What would the pressure of the gas be if the gas was heated to 70 °C at constant volume?**

Answer: $P_1 = 1.1$ atm, $P_2 = ?$

$T_1 = 273+27 = 300$ K, $T_2 = 273+70 = 343$ K

At constant volume and number of moles: $P_1/T_1 = P_2/T_2$

$$P_2 = \frac{P_1 \times T_2}{T_1} = \frac{1.1 \times 343}{300} = 1.25 \text{ atm}$$

7.6. THE IDEAL GAS LAW

$$PV = nRT$$

Gas constant R = 0.082 atmLmol^{-1}K^{-1}
(used when P is given in atm and V in L)

PRACTICE 7.6. **For an ideal gas, calculate the pressure of the gas if 0.215 mol occupies 338 mL at 32 °C.**

R = 0.082 L atm/Kmol

Answer:

$n = 0.215$ mol

$V = 338$ mL $= 338 \times 10^{-3}$ L $= 0.338$ L

$T = 32 + 273 = 305$ K

$P = ?$

PV = nRT ⟶ P = nRT/V = (0.215×0.082× 305) / 0.338 =15.9 atm

PRACTICE 7.7. **Octane (C_8H_{18}) is burned to give CO_2 and H_2O. How many liters of O_2 measured at 0.974 atm and 24°C are required to burn 1 g of octane?**
(Atomic masses: H = 1, C = 12, O = 16); **R = 0.082 Latm/molK**

Answer:

The molar mass of of C_8H_{18} : $M(C_8H_{18}) = (8\times12) + (18\times1) = 114$ g/mol
The molar mass of O_2 : $M(O_2) = 2 \times 16 = 32$ g/mol

$$2\ C_8H_{18\ (g)} + 25\ O_{2\ (g)} \rightarrow 16\ CO_{2\ (g)} + 18\ H_2O_{\ (lq)}$$

2 moles of C_8H_{18} ⟶ 25 moles of O_2
(2 x 114)=228 g ⟹ (25 x 32) = 800 g
1 g ⟹ $m(O_2)$ = ?
$m(O_2)$ = (800×1)/228 = 3.51 g

Number of moles of O_2 : $n(O_2) = m(O_2) / M(O_2) = 3.51 / 32 = 0.109$ mol

O_2
- **P = 0.974 atm,**
- **T = 24 + 273 = 297 K**
- **n = 0.109 mol**
- **V = ?**

PV = nRT ⟹ V= nRT/P
V = (0.109 × 0.082 × 297)/ 0.974 = 2.73 L

7.7. Standard Temperature and Pressure (STP):

P = 1 atm and T = 0 °C = 273 K

PRACTICE 7.8. **A quantity of He gas occupies a volume of 16.5 L at 78° C and 45.6 atm. What is its volume at STP?**
R = 0.082 L atm/Kmol

Answer:

V = 16.5 L
T= 78 +273 = 351 K
P = 45.6 atm
n = ?

PV = nRT ⟶ n= PV/RT
n = 45.6×16.5 / (0.082 × 351)
= 26 mol

⟶ STP

V = ?
T= 273 K
P = 1 atm
n = ? (26 mol)

PV = nRT ⟶ V= nRT/P
V = (26 ×0.082× 273) / 1
= 582 L

PRACTICE 7.9. **Calculate the volume of one mole of a gas at STP?**

Answer:

PV = nRT V= nRT/P
n = (1×0.082×273) / 1
= 22.4 L

7.8. Calculating density (d) and Molar Mass (*M*) of a Gas

PV = nRT or n = m / *M* *then* $PV = \frac{m}{M}RT$

$P = \frac{m}{VM}RT$ or *d* = *m*/*V* *then* $P = \frac{d}{M}RT$

$d = \frac{PM}{RT}$ and $M = \frac{dRT}{p}$

PRACTICE 7.10. **What is the density of carbon dioxide (CO_2) gas at STP?**

(Atomic masses: C = 12, O = 16), **R = 0.082 L atm/Kmol**

Answer:

STP ⟹ T = 273 K and P = 1atm

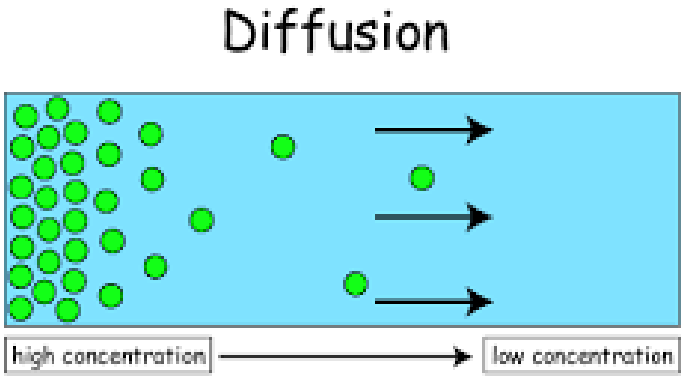
$M(CO_2) = (1 \times 12) + (2 \times 16) = 44$ g/mol

$d = [P \times M]/[R \times T] = [1 \times 44] / [0.082 \times 273] = 1.96$ g/L

PRACTICE 7.11. **An unknown gas has a density of 0.18 g/L at 1.1 atm and 300 K. Calculate its molar mass.**

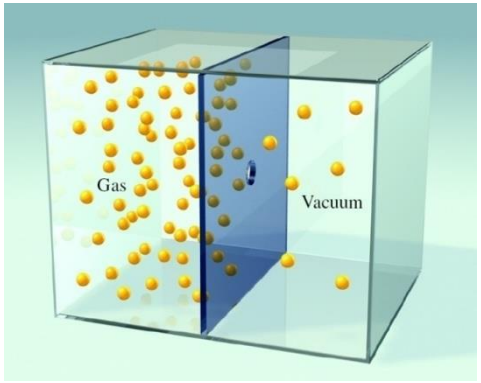
Answer: $M = dRT/P = (0.18 \times 0.082 \times 300)/1.1 = 4$ g/mol

7.9. Diffusion and Effusion



❑ **7.9.1. Diffusion** is defined as the movement of gaseous molecules from an area of high concentration to an area of low concentration.

❑ **7.9.2. Effusion** is the movement of gaseous molecules through a small **hole.**



7.9.3. Graham's Law of Effusion

The rate of effusion of a gas is inversely related to the square root of its molar mass.

$$\text{Rate of effusion} \ \alpha \ 1/(M)^{½}$$

$$\frac{r_1}{r_2} = \frac{t_2}{t_1} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

r_1: rate of effusion of gas 1
r_2: rate of effusion of gas 2
t_1 : time flow of gas 1
t_2 : time flow of gas 2
M_1: molar mass of gas 1
M_2: molar mass of gas 2

PRACTICE 7.12. **Calculate the ratio of the effusion rates of helium (He) and methane (CH_4). Atomic masses: H= 1, He = 4, C =12**

Answer: The molar mass of CH_4 is $M(CH_4) = 12+4\times1=16$ g/mol

The molar mass of He is M(He) = 4 g/mol

$$\frac{r(He)}{r(CH_4)} = \left[\frac{M(CH_4)}{M(He)}\right]^{\frac{1}{2}} = \left(\frac{16}{4}\right)^{\frac{1}{2}} = 2$$

PRACTICE 7.13. **Which one of the followings about the effusion rates of gases carbon dioxide (CO_2) and oxygen (O_2) is correct?**

a) (CO_2) has a higher effusion rate than (O_2)
b) (CO_2) has a lower effusion rate than (O_2)
c) (CO_2) and (O_2) have equal effusion rates
d) none of these

PRACTICE 7.14. **An unknown gas (X) effuses 1.4 times faster than oxygen (O_2). What is the molar mass of the unknown gas? Atomic mass: O = 16**

Answer: $$\frac{r(X)}{r(O_2)} = 1.4 = \left[\frac{M(O_2)}{M(X)}\right]^{\frac{1}{2}} \Rightarrow 1.4^2 = \left[\frac{M(O_2)}{M(X)}\right] \Rightarrow M(X) = \left[\frac{M(O_2)}{1.4^2}\right]$$

$$M(X) = \left[\frac{32}{1.4^2}\right] = 16 \text{ g/mol}$$

7.10. Partial Pressure of a Gas in a Mixture of Gases

- Each gas in a mixture behaves as if it is the only gas present
- its pressure, called the Partial Pressure, is calculated from PV = nRT with n equal to the number of moles of that gas.

Dalton's Law of Partial Pressures

The gas pressure in a container is the sum of the partial pressures of all the gases present:

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

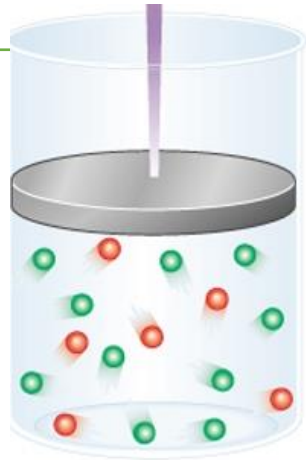
In a mixture of gases:

$$P_{total}\, V = n_{total}\, RT$$

The partial pressure of a gas (i) in a mixture of gases is

$$P_i = X_i P_{total}$$

where $X_i = n_i/n_{total}$ **is the mole fraction of gas (i)**



$n_{total} = n_1 + n_2$

PRACTICE 7.15. **What pressure in atm is exerted by a mixture of 2 g of H_2 and 8 g of N_2 at 273 K in a 10 L container?**

Calculate the partial pressures of N_2.

Atomic masses (H = 1, N= 14) and R = 0.082 L atm/Kmol

Answer:

Number of moles of H_2: $n_{H2} = m/M = 2/2 = 1$ mol

Number of moles of N_2: $n_{N2} = 8/28 = 0.286$ mol

$$P_{total} = n_{total} \frac{RT}{V} = (n_{H2} + n_{N2}) \frac{RT}{V}$$

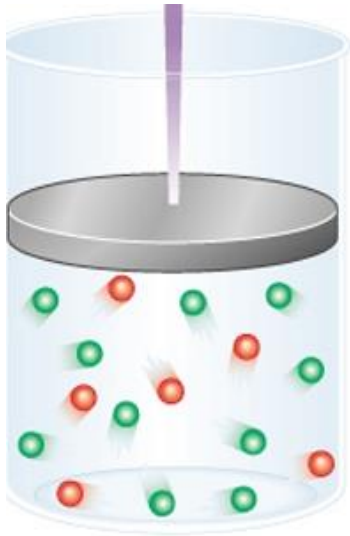
$P_{total} = [(1 + 0.286) \times 0.082 \times 273]/10 =$ **2.88 atm**

The mole fraction of N_2:

$X_{N2} = n_{N2}/(n_{H2} + n_{N2}) = 0.286/(1+0.286) = 0.22$

The partial pressure of N_2:

$P_{N2} = X_{N2} \times P_{total} = 0.22 \times 2.88 =$ **0.63 atm**



A mixture of H_2 and N_2 gases

7.11. Real Gases

- There are attraction forces between the particles
- The volume of gas is not neglected with respect to the volume of the container.

General Gas Law for Real Gases (Van Der Waals Equation):

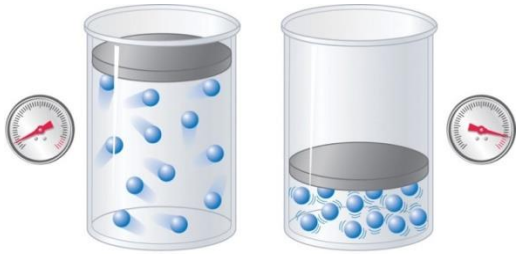
$$(P + a\frac{n^2}{V^2}) \times (V - nb) = nRT$$

P: Pressure of real gas
a, b, R: constants
n: Number of moles of real gas
V: Volume of real gas
T: Absolute temperature

A gas deviates from the ideal behavior (it is a real gas)

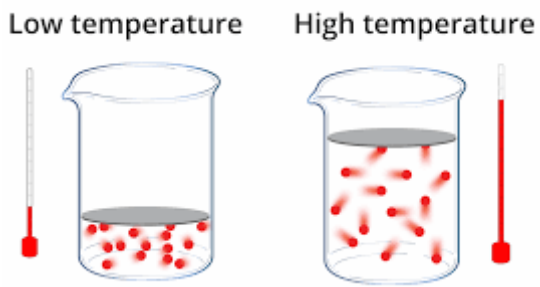
➢ at high pressure

- The molecules of a gas possess **finite volumes.**
- At short distances of approach, gas molecules exert **attractive force** upon one another.



➢ at low temperature

- The kinetic energy of the gaseous molecules decreases, and the molecules can no longer overcome the intermolecular attractive forces.
- The gas liquefies (condenses to a liquid).



Q10.1 to Q10.10 p.251 -254

Practice on chapter 7

***(Atomic masses:* H = 1, He = 4, C = 12, O = 16); $R = 0.082$ Latm/molK**

Q1. Which one of the followings about the effusion rates of gases helium (He) and methane (CH_4) is correct?

a)(He) has a higher effusion rate than (CH_4) b) (He) has a lower effusion rate than (CH_4) c) (He) and (CH_4) have equal effusion rates d) none of these

Q2. When a balloon is filled with oxygen gas at constant pressure and temperature, its volume increases with the increase of the amount of oxygen. This corresponds to

a) Boyle's law b) Charles's law c) Avogadro's law d) Dalton's law

Q3. An unknown gas has a density of 1.56 g/L at 0.98 atm and 65 °C. Its molar mass is

a) 0.02 g/mol b) 1.56 g/mol c) 44 g/mol d) 18 g/mol

Q4. STP means that

a) T = 273 K and P = 1 Pa. b) T = 0 °F and P = 1 atm.

c) T = 0 K and P = 1 atm. d) T = 0 °C and P = 1 atm.

Q5. The equation PV = nRT represents

a) the ideal gas law b) the general gas law for real gases.

c) Boyle's law. d) Dalton's law of partial pressures.

Q6. In the equation ($P_i = X_i \times P_{total}$), which relates the partial pressure of a gas (i) to the total pressure in the container, X_i represents the

a) mole fraction of gas (i). b) number of moles of gas (i).

c) molar mass of gas (i). d) density of gas (i).

Q7. Diffusion is defined as

a) the proximity of several measurements to the true (accepted) value.

b) the movement of molecules through a small hole into an empty container.

c) the sum of the partial pressures of all the gases present in a container.

d) the movement of gaseous molecules from an area of high concentration to an area of low concentration.

Q8. An ideal gas occupies a volume of 8.2 L at STP. What is its number of moles?

a) 2.7 mol b) 0.37 mol c) 3.7 mol d) 0.27 mol

Q9. Calculate the volume occupied by 1 mol of an ideal gas at STP.

a) 22.4 L b) 1 L c) 273 L d) 8.4 L

Q10. Dalton's law of partial pressure states that

a) the volume of the ideal gas is inversely proportional to its pressure at a constant temperature

b) the total gas pressure in a container is the sum of the partial pressures of all the gases present.

c) the pressure of the ideal gas is proportional to the temperature at a constant volume

d) the pressure of the ideal gas is proportional to the amount of substance

Q11. Boyle's law, which relates the pressure and volume at constant amount of substance and temperature, states that the volume of an ideal gas is

a) directly proportional to its pressure b) inversely proportional to its pressure

c) inversely proportional to the amount of substance d) proportional to the gas constant R

Q12. An ideal gas is composed of particles that

a) are far away from each other b) there is no attraction forces between them

c) their volume is neglected with respect to the volume of the container d) all of the above.

Q13. According to Charles's law, when the volume of an ideal gas increases at constant pressure and gas amount, the temperature

a) increases b) decreases

c) remains constant d) both increases and decreases depending on the volume

Q14. A gas deviates from ideal behavior at

a) high pressure and temperature. b) low pressure and temperature.

c) high pressure and low temperature. d) low pressure and high temperature.

Q14. The gas pressure in a vessel containing 2 moles of H_2 and 4 moles of CO_2 is 2 atm. Calculate the partial pressure of CO_2 gas.

a) 0.66 atm b) 1.32 atm. c) 0.68 atm d) 2 atm