General Chemistry I

Dr. Laila Mohammed Al-Harbi
Assistant professor

Web Site: http://lalhrbi.kau.edu.sa

Contact Info: Lalhrbi@kau.edu.sa

Technicalities

- **Locations:**
  - **Science tower 07 room 173 first floor**
  - phone 6400000 ext. 23024
  - email Lalhrbi@kau.edu.sa
  - web site: http://lalhrbi.kau.edu.sa
- **Exam schedule:**
- 1st exam : from lecture 1-11 (Chapters 1-4) **= 30 marks**
- 2nd exam: from lecture 12-24 (Chapters 5,7-9) **= 30 marks**
- Final exam: from lecture 1-33 **= 40 marks**
- (Chapters 1,2,3,4,5,7,8,914,15,24&25)

CHEM 110

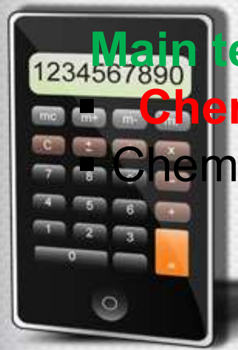
1	Course No.	Course Title	No. of Units			Pre-requisites
			Th.	Pr.	Credit	
	Chem 110	General Chemistry I	3	-	3	-

Course Objectives:

The course aims to introduce students to basic knowledge and principle in chemistry.

Course Description :

It provides an introduction to the general principles of chemistry for students planning a professional career in chemistry, a related science, the health professions, or engineering. By the end of this course the student will be able to understand the following: Significant figures, scientific notation and units, stoichiometry, atomic structure & periodic table, chemical bonding, gases, ionic equilibrium, basic principles of organic and basic principles of biochemistry .



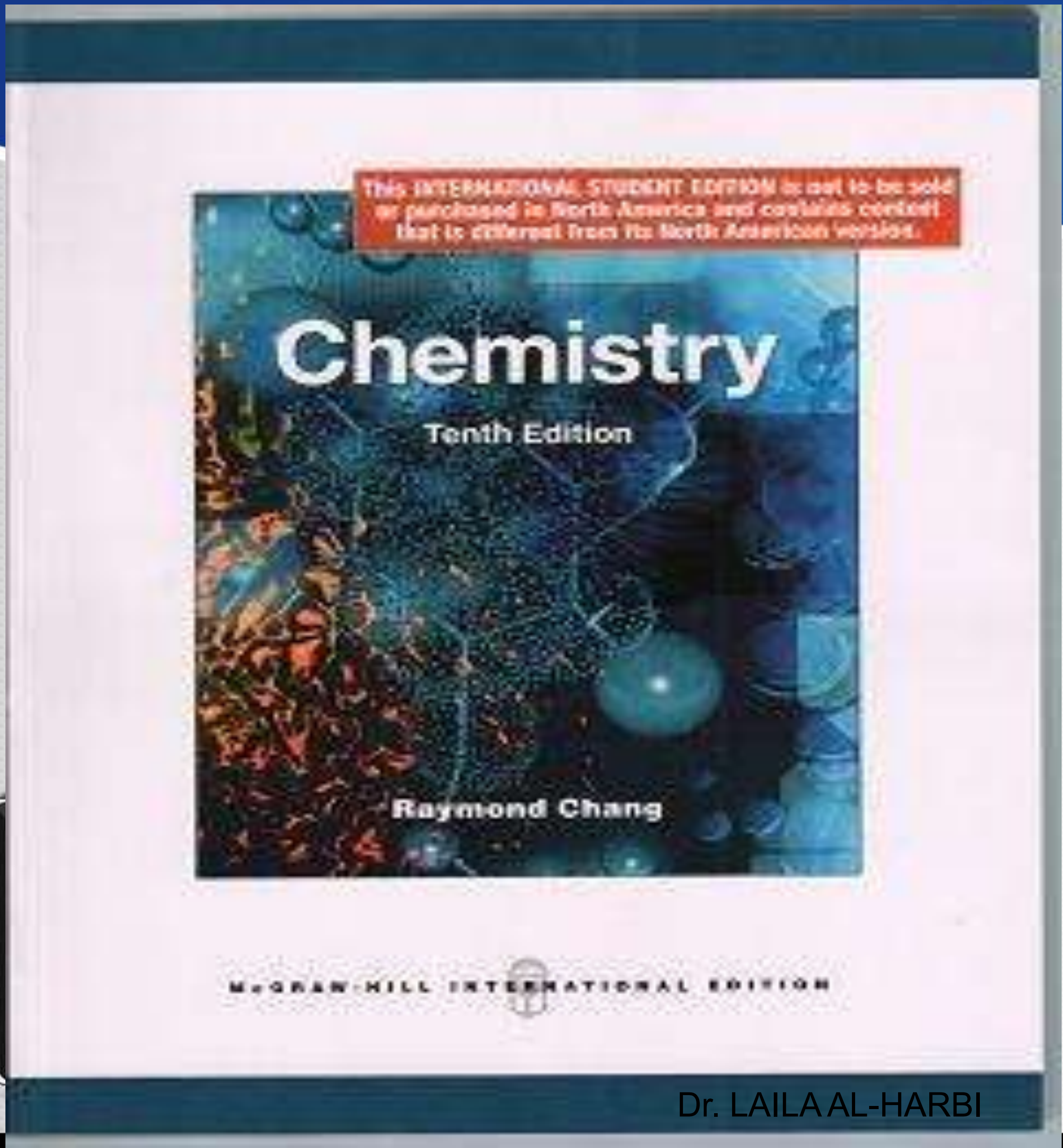
Main text books:

- **Chemistry, by Chang, 10th. ed., 2007, McGraw-Hill.**
- Chemistry, by Steven S. Zumdahl, 6th ed., Houghton Mifflin College Div.

Subsidiary books :

Chemistry, by Mortimer, 6th ed., Wadsworth Inc.

CHEM 110



Main text book :

- Chemistry, by Chang, 10^{th}. ed., 2007, McGraw-Hill.

نصائح للمذاكرة كيمياء 110

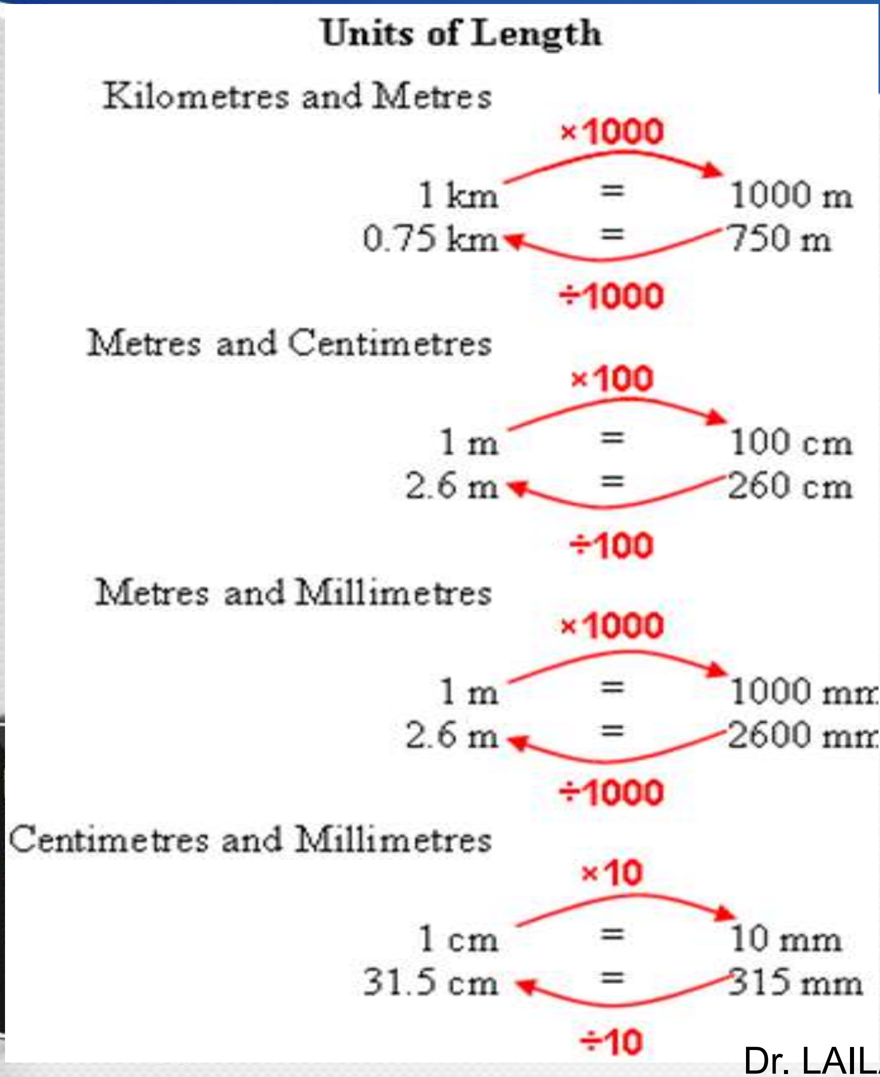
- المذاكرة أول باول
- عدم التغيب عن المحاضرات
- مذاكرة الكتاب الدراسي و المحاضرات
- حل بنك الاسئلة ... حيث يحوي على مجموعة مختلفة من الاسئلة بافكار متعدده ... تساعدك على التدرب
- التدرب على الاسئلة في الموقع التفاعلي حيث أن الطريقة شبيهة تماما بالاختبار الالكتروني
- http://prod.kau.edu.sa/faculties/science/website2/index.aspx

- و أخيرا ... التوفيق بيد الله سبحانه و تعالي


نصائح عامة للمذاكرة السليمة
وزارة التعليم
Ministry of Education
تعليم الكبار بجدة
توكل على الله وداوم على الأدعية
والأذكار ولا تؤخر صلواتك.
قم بتلخيص المعلومات
بشكلٍ مبسط وسهل
سواءً في عقلك أو كتابتها على الورق.
ركّز على النقاط المهمة في الدروس مع
المراجعة المستمرة لما قمت بحفظه.
لاتجهد نفسك كثيراً في المذاكرة،
ولاتسهر وخذ قسطاً كافياً من النوم.
لا تتخيل كثيراً صعوبة الأسئلة
بل حاول أن تركز على المذاكرة
وأعطي كل مادةٍ حقها في المذاكرة.
اجلس في مكان هادىء مع
إضاءه جيدة وأفتح نوافذ الغرفة.
حاول قراءة الموضوع وقم بتكرار
القراءة مع محاولة فهم المعاني فيها
وكتابتها أكثر من مرة لتثبيت الحفظ .
قم بوضع جدول منظم للمذاكرة
مع اعطاء فسحة وراحة لنفسك تقوم
من خلالها بنشاطات ممتعة كالرياضة.
تخيل نوع الأسئلة ولا تقف
عند معلومة لم تفهمها جيدا فضع
تحتها خطأً ثم أكمل المذاكرة وعد
إليها بعد الانتهاء من المذاكرة .
كرر الأدعية المؤثورة
من القرآن والسنة.
رؤية VISION
2030
ضمن برنامج كيف أذاكر عند الاختبار
مدرسة ابن سيرين المتوسطة والثانوية الليلية برابغ


1234567890

Chapter 1
The study of change



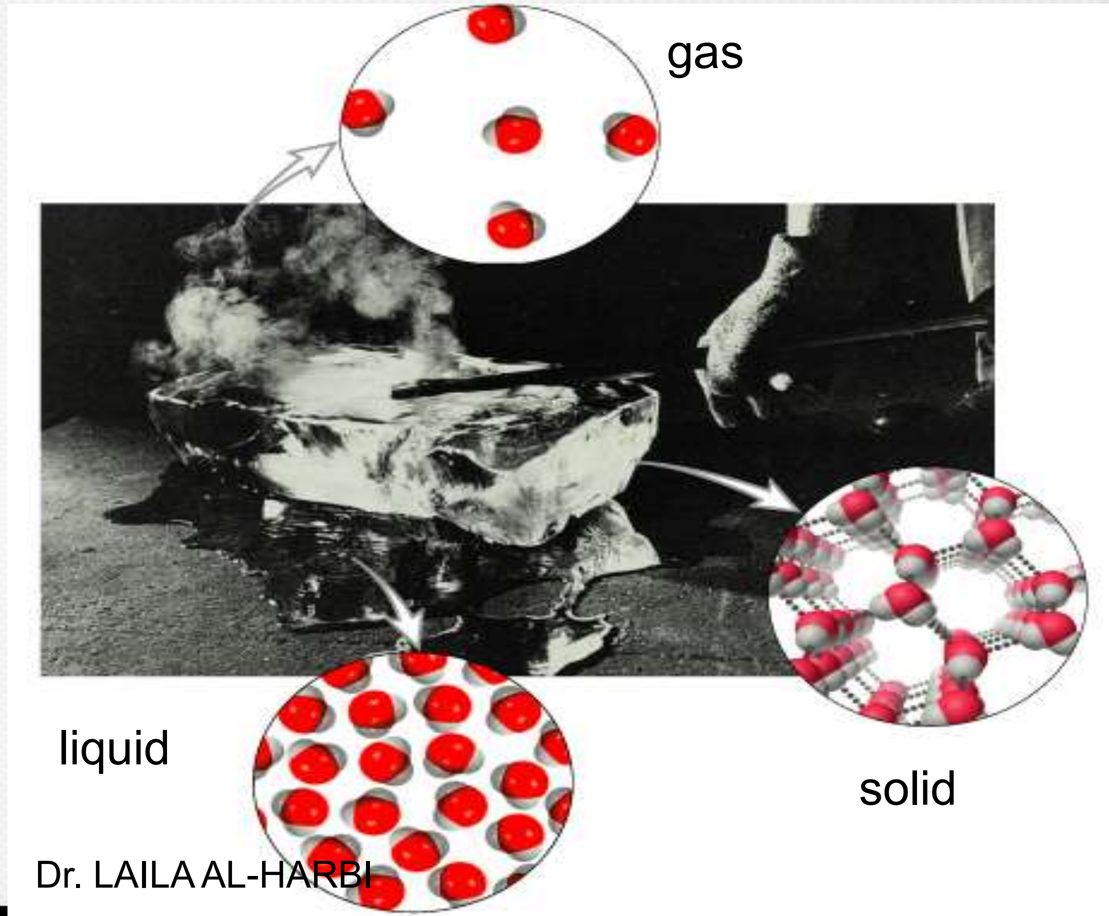
- Measurement
  - ❖ SI units
  - ❖ Mass and weight
  - ❖ Volume
  - ❖ Temperature scales

Objectives

- By the end of this chapter you should:
- Know the 7 SI basic units and their prefixes.
- Be able to convert from one unit to other.
- Know to derive units from the 7 SI basic units.
- Common units (L & mL)
- Know the temperature Scales

Introduction

- **Chemistry** is the study of matter and the changes it undergoes
- There are three states of matter



1.7 Measurement and Units
The International System of units (SI Units)

- used for commerce and science around

TABLE 1.2 SI Base Units

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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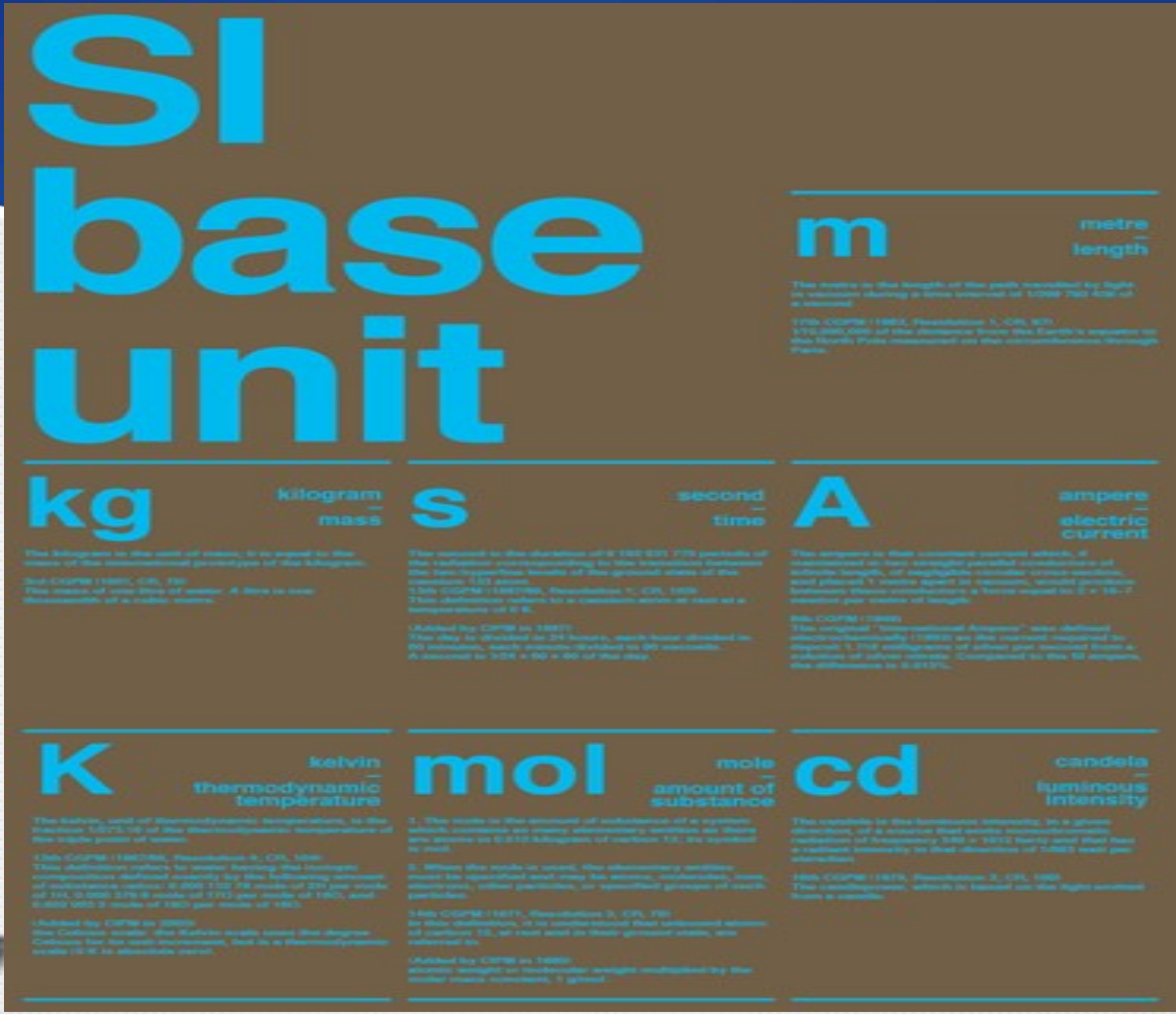

SI base unit
m
metre — length
kg
kilogram — mass
s
second — time
A
ampere — electric current
K
kelvin — thermodynamic temperature
mol
mole — amount of substance
cd
candela — luminous intensity

Examples

The SI unit of mass is

(a) The pound

(b) The gram

(c) The kilogram

(d) The mole

The Kg is the SI unit of

(a) length

(b) mass

(c) temperature

(d) current

TABLE 1.3 Prefixes Used with SI Units

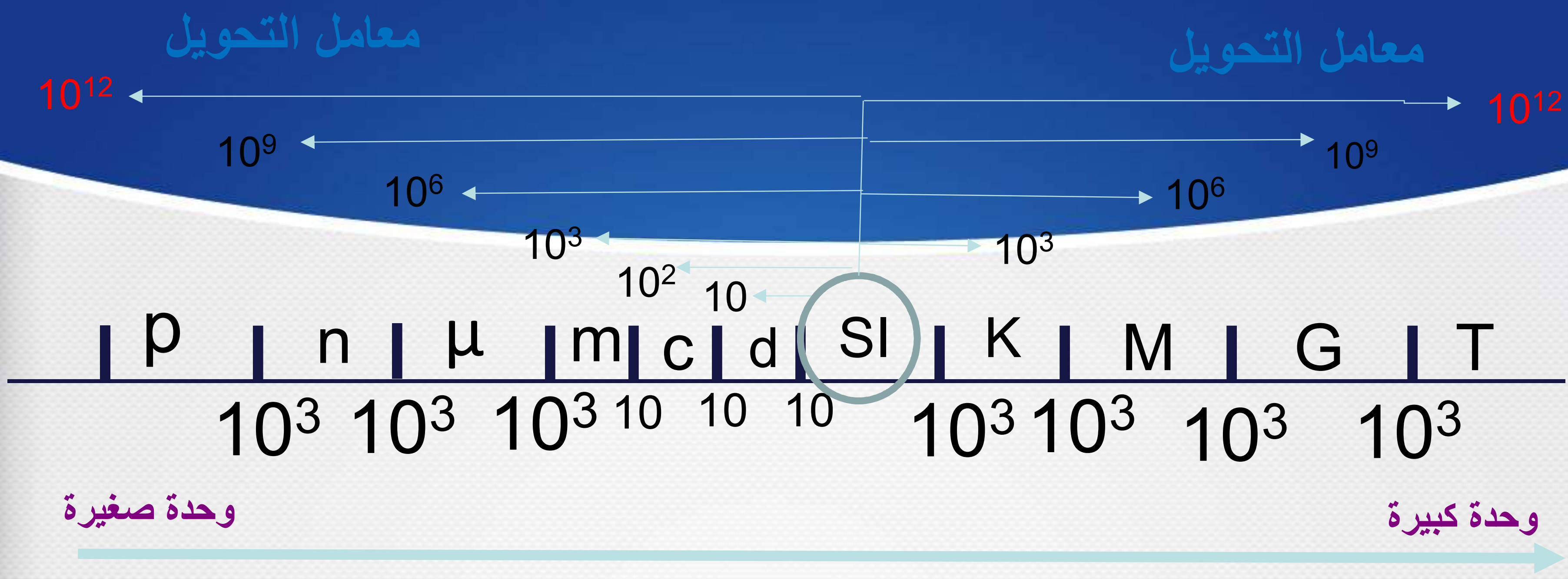
Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) = 1×10^{12} m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) = 1×10^{9} m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) = 1×10^{6} m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) = 1×10^{3} m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μm) = 1×10^{-6} m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) = 1×10^{-12} m

1m

To transfer between perifexs

- A) use numerical line خط الاعداد
- B) use transfer factor معامل التحويل

للتحويل من الوحده الاساسية و مضاعفتها أو المشتقات يمكننا استخدام خط الاعداد أو استخدام معامل التحويل (بضرب الطريفين في الوسطين) و لكن لا تخلطي بين الطريقتيين كما أن الطريقتين يعطوا نفس النتيجة اذا استخدموا بطريقة صحيحة.



بهذا الاتجاه تزيد الوحده

للتحويل من صغير الى كبير نقسم ÷ , للتحويل من كبير الى صغير نضرب ×

للتحويل من الوحده الاساسية و مضاعفتها أو المشتقات يمكننا استخدام خط الاعداد أو استخدام معامل التحويل (بضرب الطريفين في الوسطين) و لكن لا تخلطي بين الطريقتيين كما أن الطريقتين يعطوا نفس النتيجة اذا تم استخدامها بنفس بطريقة صحيحة.

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I. Convert 134 pm to m?

II. How many meters are in 134 pm?

الحل باستخدام خط الاعداد

الوحدة المطلوبة	الوحدة المعطاة
متر	بيكومتر

الناتج	المعامل	العملية الحسابية	القيمة
1.34×10^{-10}	10^{12}	÷	134

للتحويل من صغير الى كبير نقسم ÷

$$1\ pm = 1\times10^{-12}\ m$$

$$134\ pm = ??\ m$$

الحل باستخدام الجدول طرفين في وسطين

$$134\times1\times10^{-12}\ m = 1.34\times10^{-10}\ m$$

Examples

- The SI unit of mass is
- (a). The pound
- (b). The gram
- **(c). The kilogra**
- (d). The mole.

- The SI prefixes *giga* and *micro* represent, respectively:
- A. 10^{-9} and 10^{-6}.
- B. 10^{6} and 10^{-3}.
- C. 10^{3} and 10^{-3}.
- D. 10^{9} and 10^{-6}.

TABLE 1.3 Prefixes Used with SI Units

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mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) = 1×10^{6} m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) = 1×10^{3} m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μm) = 1×10^{-6} m

How many microseconds are in a second?

(a). 1×10^{-1}

(b). 1×10^{-6}

(c). 1×10^{-15}

(d). 1×10^{6}

1 µs= 1×10^{-6} s

x µs = 1 s

Explanation: Since the prefix *micro means* 1×10^{-6}, *there will be* $1 \times 10^{+6}$ microseconds in one second.

الوحدة المعطاة	الوحدة المطلوبة
s	µs

القيمة	العملية الحسابية	المعامل	الناتج
1	x	10^{6}	1×10^{6} µs

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Example

- Which of the following is the smallest distance?
- (a) 21 m → 21m
- (b) 2.1×10^2 cm → 2.1m
- (c) 21 mm → 0.021 m
- **(d) 2.1×10^4 pm → 2.1×10^{-8} m**

Put all of them in the same unit

Explanation: Even though 2.1×10^4 is the largest number in this question, the units of pm (picometers) are the smallest units here, making it the smallest distance.

p | n | µ | m | c | d | SI | K | M | G | T

10^3 10^3 10^3 10 10 10 10^3 10^3 10^3 10^3

$10^3 \times 10^3 \times 10^3 = 10^9$

- **A) 6.0 km is how many micrometers?**
- **Solution 1**

$1\ km = 10^3\ m$

$6\ km = x = \mathbf{6 \times 10^3\ m}$

$1 \mu m = 1 \times 10^{-6}\ m$

$x = 6 \times 10^3\ m$

$\mathbf{x = 6 \times 10^9}\mu m$

الوحدة المعطاة		الوحدة المطلوبة	
km		µm	
القيمة	العملية الحسابية	المعامل	الناتج
6	×	10^9	6×10^9

Explanation: convert first to meter then from meter to micro (two steps solution)

- **Example**
- The diameter of an atom is approximately 1×10^{-7} mm. What is this diameter when expressed in nanometers?
- A. 1×10^{-18} nm
- B. 1×10^{-15} nm
- C. 1×10^{-9} nm
- D. 1×10^{-1} nm

- $= 1 \times 10^{-7} \times 1 \times 10^{6} =$
- 1×10^{-1} nm = 0.1 nm

- **Example**
- Which of these quantities represents the largest mass?
- A. 2.0×10^{2} mg
- B. 0.0010 kg
- C. 1.0×10^{5} μg
- D. 2.0×10^{2} cg
- Put all of them in the same unit

A) 0.2 g
B)1 g
C) 0.1 g
D) 2 g

SI derived units

- are defined in terms of the seven base quantities via a system of quantity equations.
- The *SI derived units* for these derived quantities are obtained from these equations and the seven SI base units. For example
- Area **= width x length**
- **Unit of width = m**
- **Unit of length = m**
- **Unit of Area** = m× m = m^2

Treat units like numbers

Volume –

Volume = width × length × heights = m × m × m = m^3

SI derived unit for volume is cubic meter **(m^3)**

Common unit of volume is liter (L) and milliliter (ml)

The relation ship between liter (L) and ml (1L= 1000mL)

The relation ship between liter (L) and metric system

$$1\ L = 1\ dm^3$$

The relation ship between milliliter (ml) and metric system

$$1\ mL = 1\ cm^3$$

$$1\ dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$$

$$1\ cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$$

Example (1)

- How many liters are in 25 dm^3 ?
- since 1L = 1 dm^3
- 25 dm^3 = 25L

Example (2)

- **How many milliliters are in 32 cm^3 ?**
- **Since 1mL = 1 cm^3**
- **32 cm^3 = 32 mL**

Example (3)

- How many liters are in 250 cm^3 ?
- Since 1L = 1 dm^3
- and 1mL = 1 cm^3
- 250 cm^3 = 250 mL
- L → mL
- 250/1000 = 0.25 L

Examples

- The diameter of an atom is approximately 1×10^{-7} mm. What is this diameter when expressed in nanometers?

- A. 1×10^{-18} nm
- B. 1×10^{-15} nm
- C. 1×10^{-9} nm
- D. 1×10^{-1} nm

- How many cubic centimeters are there in exactly one cubic meter?

- A. 1×10^{-6} cm^3
- B. 1×10^{-3} cm^3
- C. 1×10^{-2} cm^3
- D. 1×10^{6} cm^3
- Solution
- $(1m\ )^3 = (cm\ )^3$
- $1m^3 = (1 \times 10^2)^3\ cm^3$
- $1m^3 = 1 \times 10^6\ cm^3$

Mass and Weight

- **Mass** is the measure of the amount of matter in an object.

  SI unit of mass is the **kilogram** (kg)

  $1\ kg = 1000\ g = 1 \times 10^3\ g$

- **Weight** is the measurement of the pull of gravity on an object.

  weight = *c* x mass

- The **Mass** of an object doesn't change when an object's location changes. **Weight**, on the other hand does change with location.
- Chemist are interested primarily in **mass**

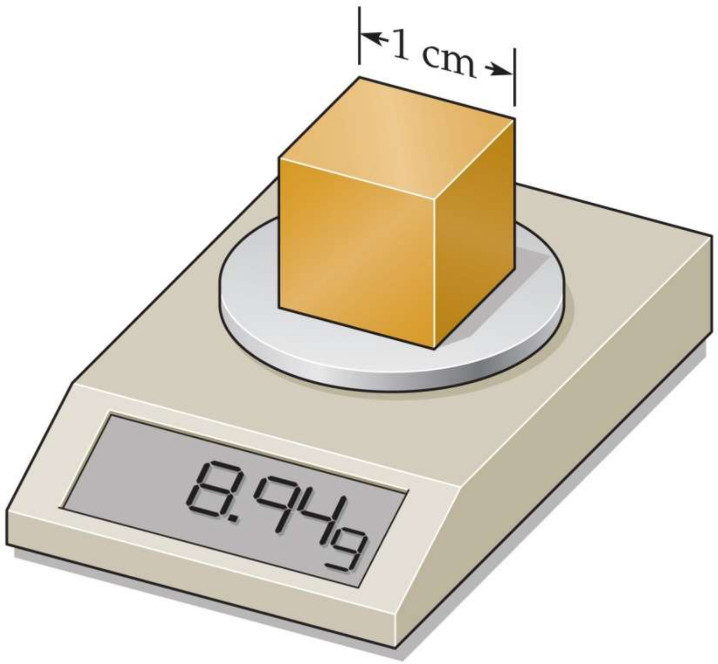
The weight of man on earth is 50 pounds is 8.25 pounds on moon

Density

Density is defined as the amount of matter in a given amount of space.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

- SI **derived** unit for density **is kg/m³**
- common units of density are g/mL , g/L
- Density decrease with temperature (g/ml)g/cm³ for liquid and solids g/L = 0.001g/ml for gases
- Because density of gases are very low



The density of copper is 8.94 g/cm³.

Questions in Density

Given	Given	Given
m & V	V & d	m & d
? d	? m	? V
d = m/V	m = d x V	V = m /d

Example 1.1

- **A piece of Gold metal has a volume of 15.6 cm³, with a mass of 301 g What is its density?**

$$d = \frac{m}{V}$$

301 g/ 15.6 cm^3
= 19.3 g/ cm^3

- **A piece of platinum metal with a density of 21.5 g/cm³ has a volume of 4.49 cm³. What is its mass?**

$$d = \frac{m}{V}$$

$$m = d \times V$$

= 21.5 g/$\cancel{cm^3}$ x 4.49 $\cancel{cm^3}$
= 96.5 g

Example 1.2

- **The density of mercury is 13.6 g/mL has a volume of 5.50 mL. What is its mass?**

$$d = \frac{m}{V}$$

$$m = d \times V$$

= 13.6 g/mL x 5.50 mL
= 74.8 g

- **The density of sulfuric acid is 1.41 g/mL has a volume of 242 mL. What is its mass?**

$$d = \frac{m}{V}$$

$$m = d \times V$$

= 1.41 g/mL x 242 mL
= 341.22 g

Temperature Scales

- Fahrenheit °F → °F = [(9/5) × °C] + 32
- Celsius °C → °C = (5/9) (°F - 32)
- Kelvin ° K → ° K = °C + 273.15

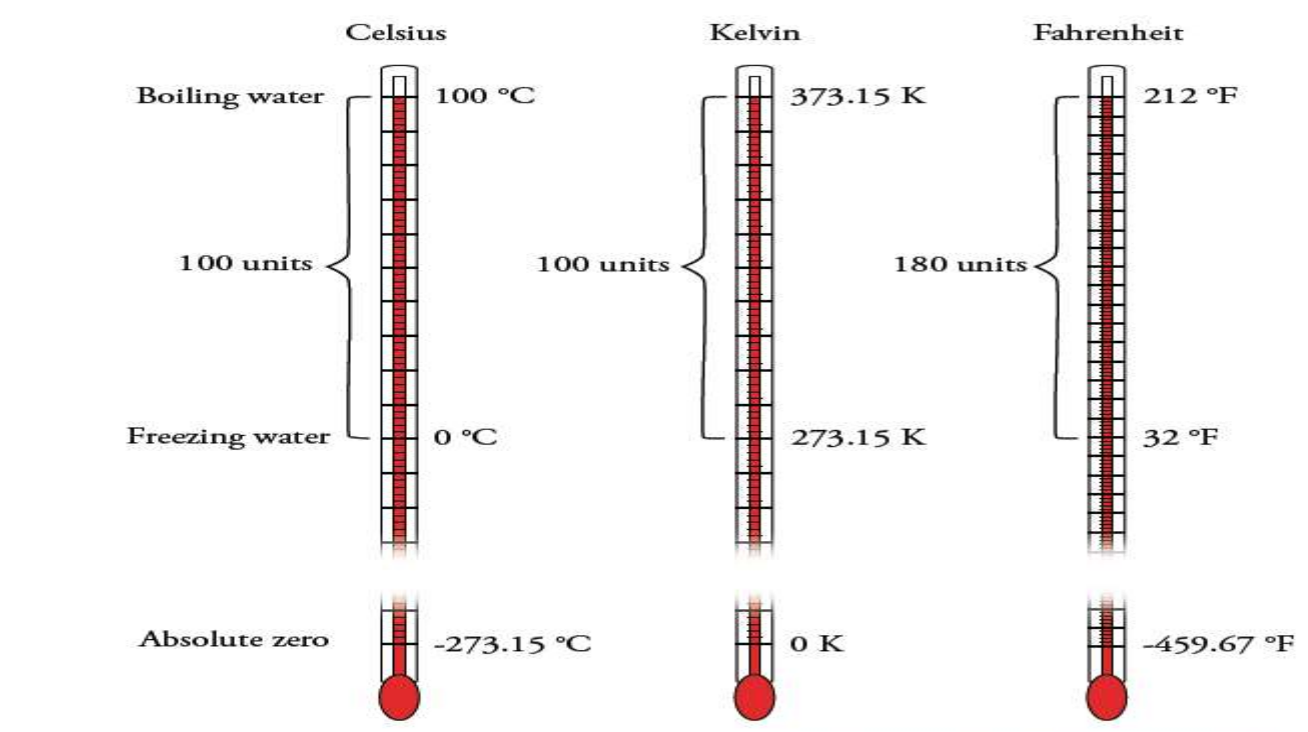


Figure 1.10 Comparing Temperature Scales

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Temperature Units Conversion

1. Degrees Celsius ^0C: Scale 0 → 100 Thus: 100 divisions or 100 degrees

2. Kelvin K: Scale 273 → 373

Thus: 100 divisions or 100 degrees

1K = 1 C

3. Degrees Fahrenheit ^0F : Scale from 32→ 212

Thus: 180 divisions or 180 degrees

Thus: the size of degree in ^0F scale is only 100/180 or

5/9 of a degree on the ^0C scale $1^0F = (5/9)\ 1^0C$

Example 1.3

- Convert 224 ^{0}C to degrees Fahrenheit?
- °F = (9 ^{0}F /5 ^{0}C) × °C + 32
- [°F = (9 ^{0}F /5 ^{0}C) × 224 °C] + 32 ^{0}C = 435 ^{0}F

- Convert -452 ^{0}F to degrees Celsius.
- °C = (5 ^{0}C /9 ^{0}F) (°F - 32 ^{0}F)
- °C = (5 ^{0}C /9 ^{0}F) (-452 °F - 32 ^{0}F) = -269 ^{0}C

- Convert -38.9 ^{0}C to degrees Kelvin..
- ° K = [-38.9 °C + 273.15 °C] × 1 K/ 1 ^{0}C = 234.3 K

- Ammonia boils at -33.4°C. What temperature is this in °F?
- A. -60.1°F
- B. -92.1°F
- C. -28.1°F
- D. +13.5°F

$F = (9\ ^0F / 5\ ^0C) \times °C + 32$

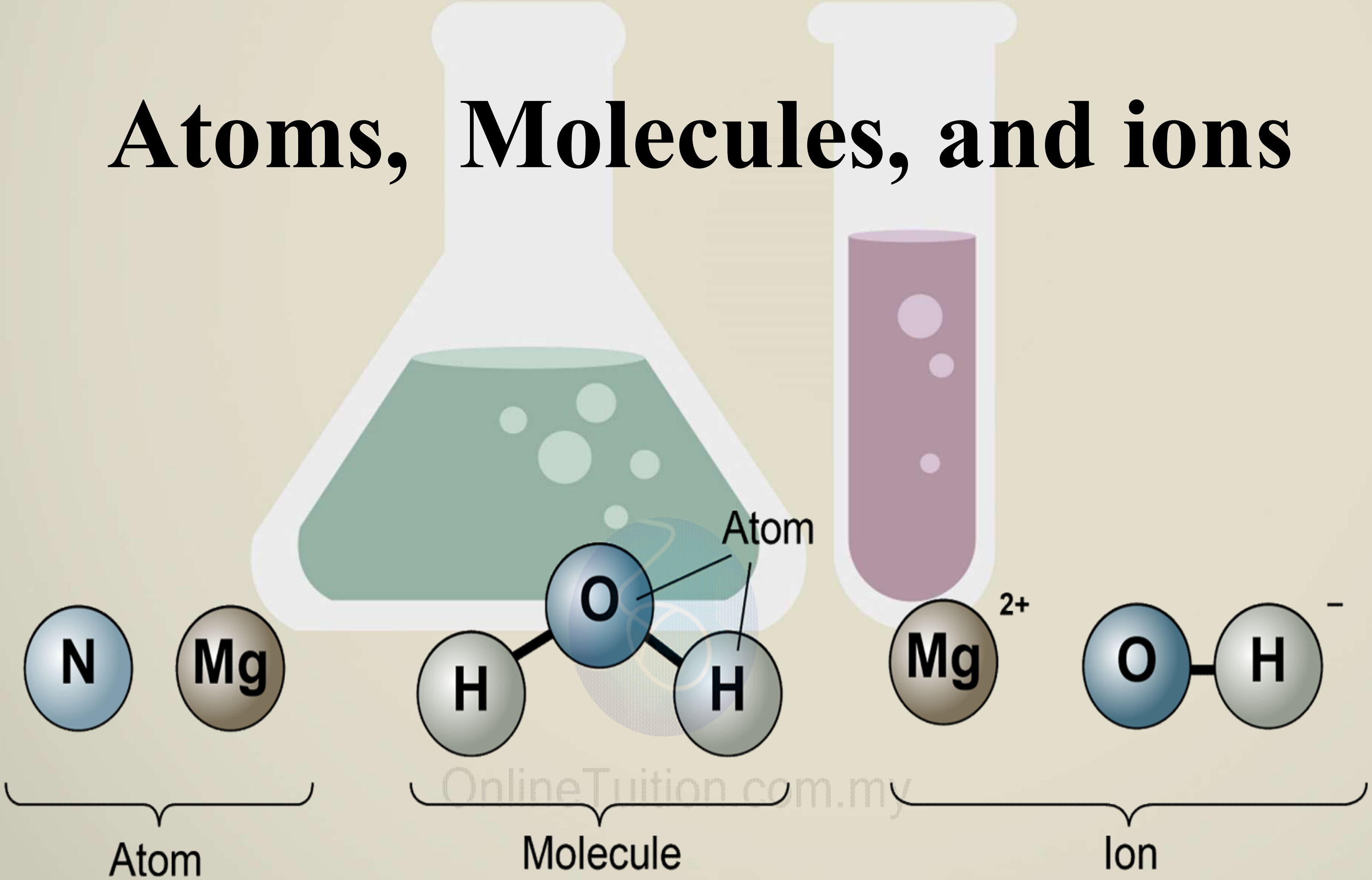
$[°F = (9\ ^0F / 5\ ^0C) \times -33.4\ °C] + 32\ ^0C = -28.1\ ^0F$

Useful sites

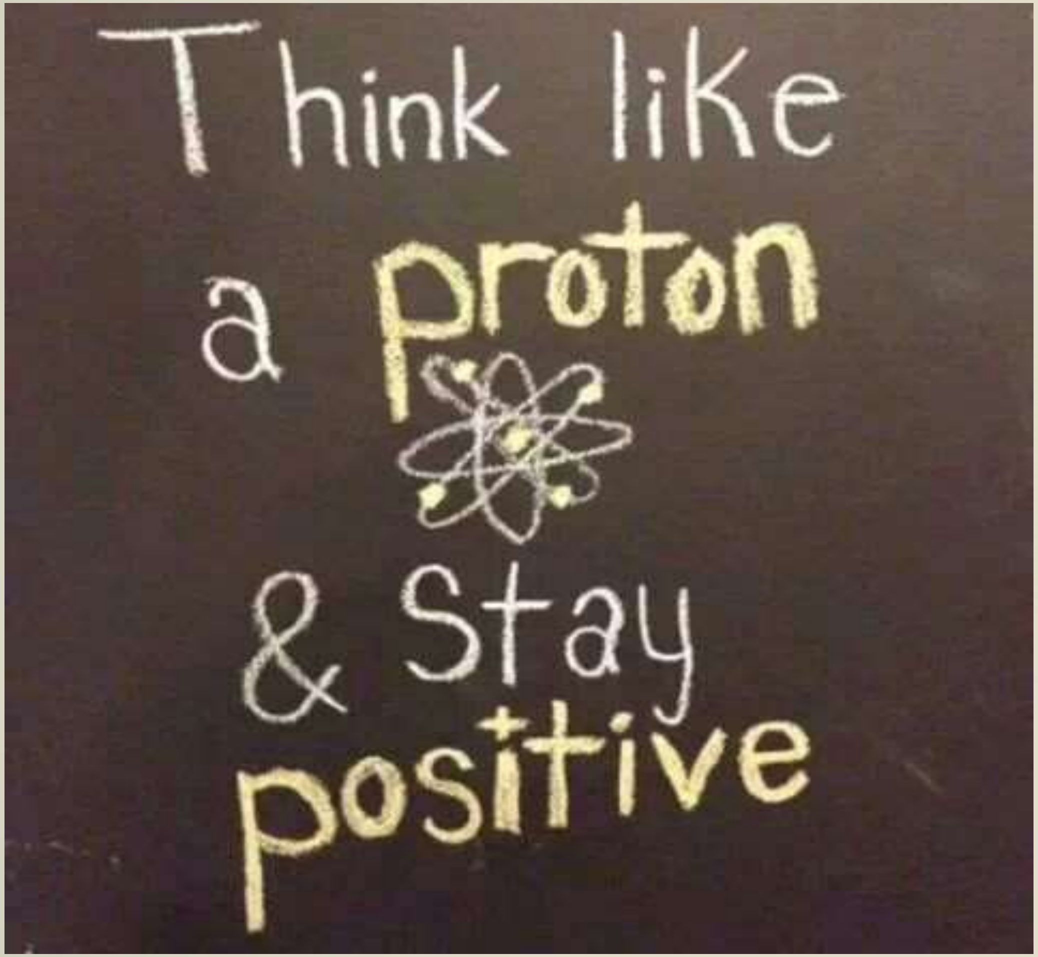
- https://www.youtube.com/watch?v=djTNUp4XIRo
- http://www.convertunits.com/

Chapter 2

Atoms, Molecules, and ions




العادات السبع للنجاح
1
كن مبادرا
2
ليكن هدفك واضحا
منذ البداية
3
رتب اولوياتك الاهم
مقدم على المهم
4
اعمل مع المجموعة
5
افهم جيدا ما يقال ثم
افهم جيدا ما تقول
6
نمي طاقتك الشخصية
الكامنة
7
تعامل بمنطق الكسب
المتبادل


Think like
a proton
& Stay
positive

Chapter 2
Atoms. Molecules, and ions

- 2.3 atomic number, mass number, and isotopes
- 2.4 the periodic table
- 2.5 molecules and ions
- 2.6 chemical formulas
- ❖ Molecular formula
- ❖ Molecular models
- ❖ Ionic formulas
- ➢ 2.7 naming compounds
- ❖ Ionic compounds
- ❖ Molecular compounds

Objectives

- By the end of this chapter you should:
- Know atomic number, mass number, and isotopes
- Be able to distinguish between molecules (diatomic & polyatomic) and ions (cation & anions) .
- Know different Chemical formulas
- Know how to Name Ionic & covalent compounds

2.3 atomic number, mass number, and isotopes

- Protons and electrons are the only particles that have a charge.
- Protons and neutrons have essentially the same mass.
- The mass of an electron is so small we ignore it.
- **Atomic number (Z)** = number of protons in nucleus)
- ***Mass number*** (A) = number of protons + number of neutrons
- = atomic number (Z) + number of neutrons

Particle	Charge	Mass (amu)
Proton	Positive (1+)	1.0073
Neutron	None (neutral)	1.0087
Electron	Negative (1−)	5.486×10^{-4}

Note that th № of P= № of e^-

Symbols of Elements

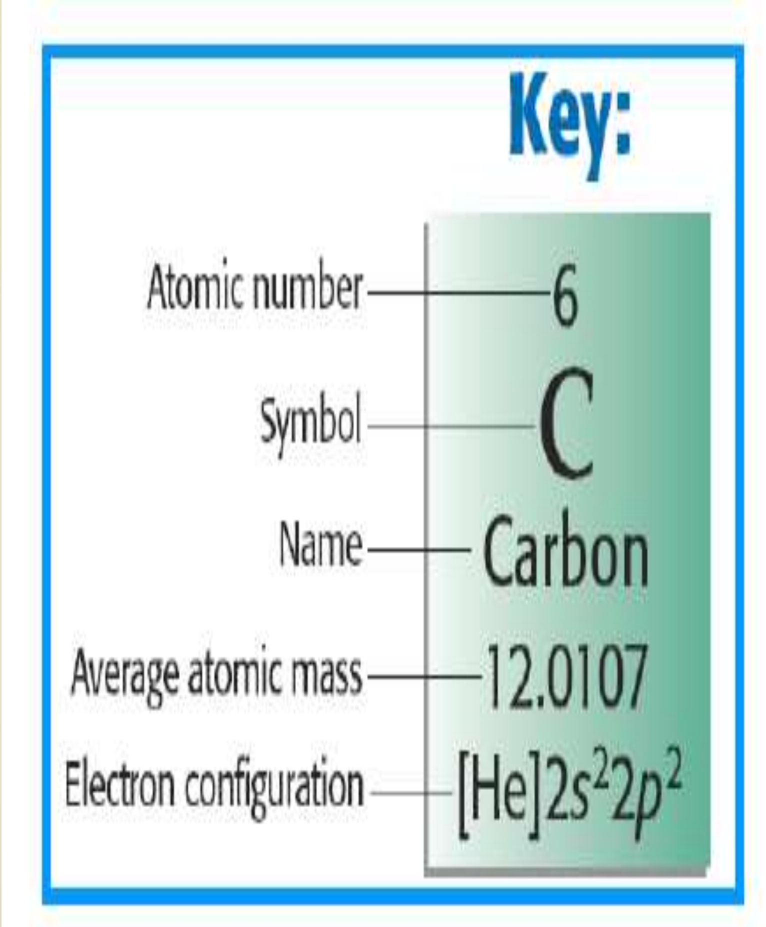
Mass Number → A
Atomic Number → Z

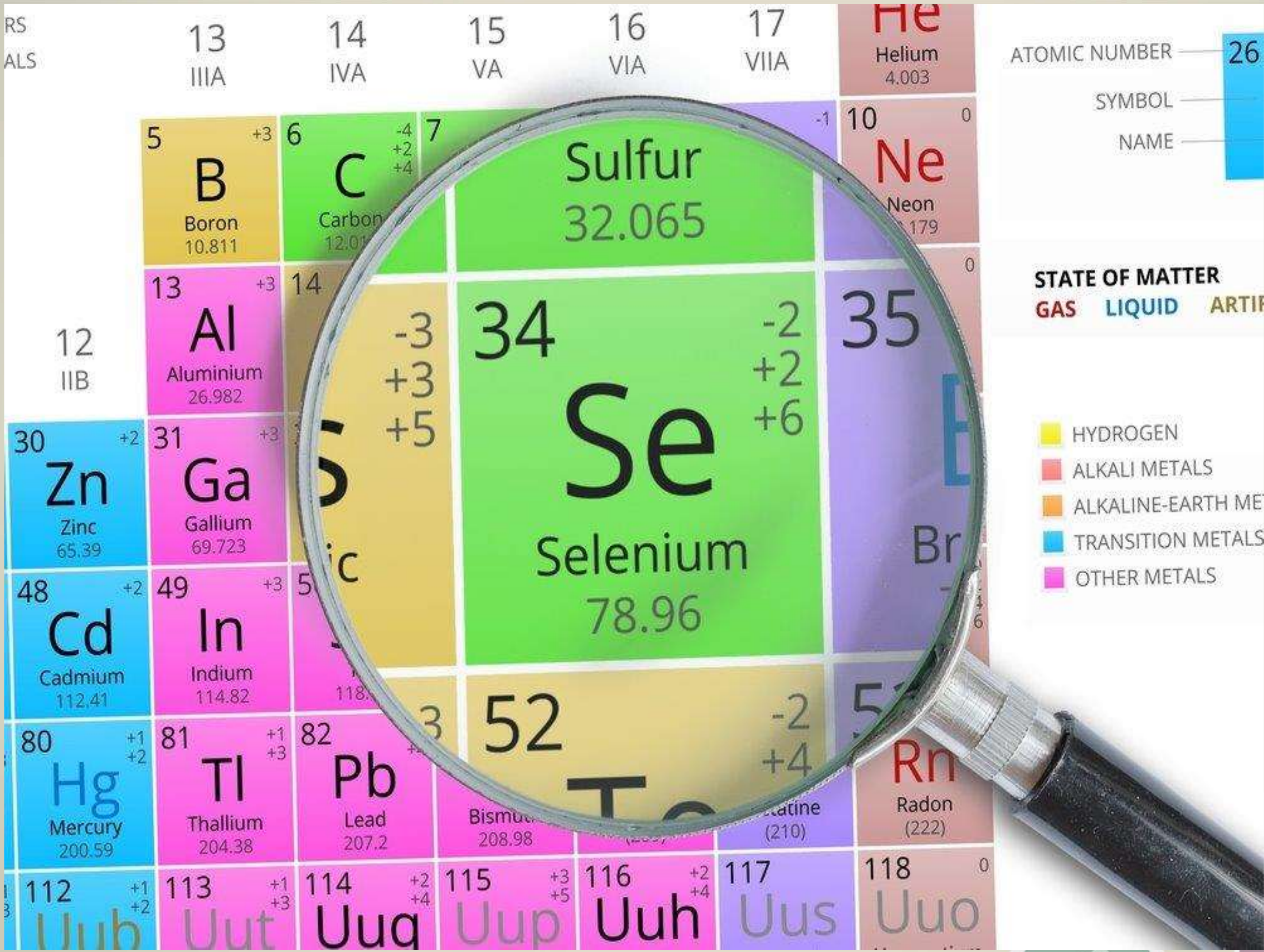
$^{A}_{Z}X$ ← Element Symbol

- Symbols of Elements

All atoms of the same element have the same number of protons: The atomic number (Z)

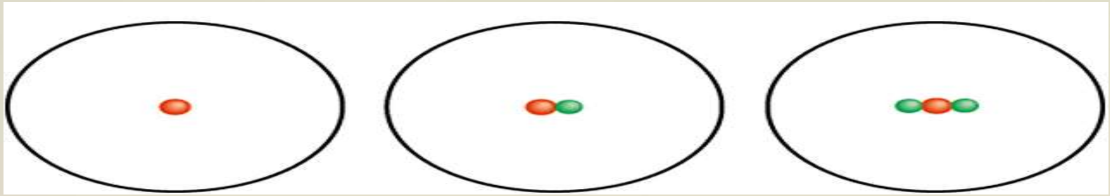
- The mass of an atom in atomic mass units (amu) is the total number of protons and neutrons in the atom.




13
IIIA
14
IVA
15
VA
16
VIA
17
VIIA
Helium
4.003
ATOMIC NUMBER
26
SYMBOL
NAME
5 +3
B
Boron
10.811
6
C
Carbon
Sulfur
32.065
10 0
Ne
Neon
STATE OF MATTER
GAS
LIQUID
13 +3
Al
Aluminium
26.982
34
-2
+2
+6
Se
Selenium
78.96
-3
+3
+5
35
Br
12
IIB
30 +2
Zn
Zinc
65.39
31
Ga
Gallium
69.723
HYDROGEN
ALKALI METALS
TRANSITION METALS
OTHER METALS
48 +2
Cd
Cadmium
112.41
49 +3
In
Indium
114.82
52
-2
+4
80
Hg
Mercury
200.59
81
Tl
Thallium
204.38
82
Pb
Lead
207.2
208.98
(210)
Rn
Radon
(222)
112
113
114
Uuq
115
116
Uuh
117
Uus
118 0
Uuo

Isotopes

- Isotopes are atoms of the same element with different masses.
- Isotopes have different numbers of neutrons.



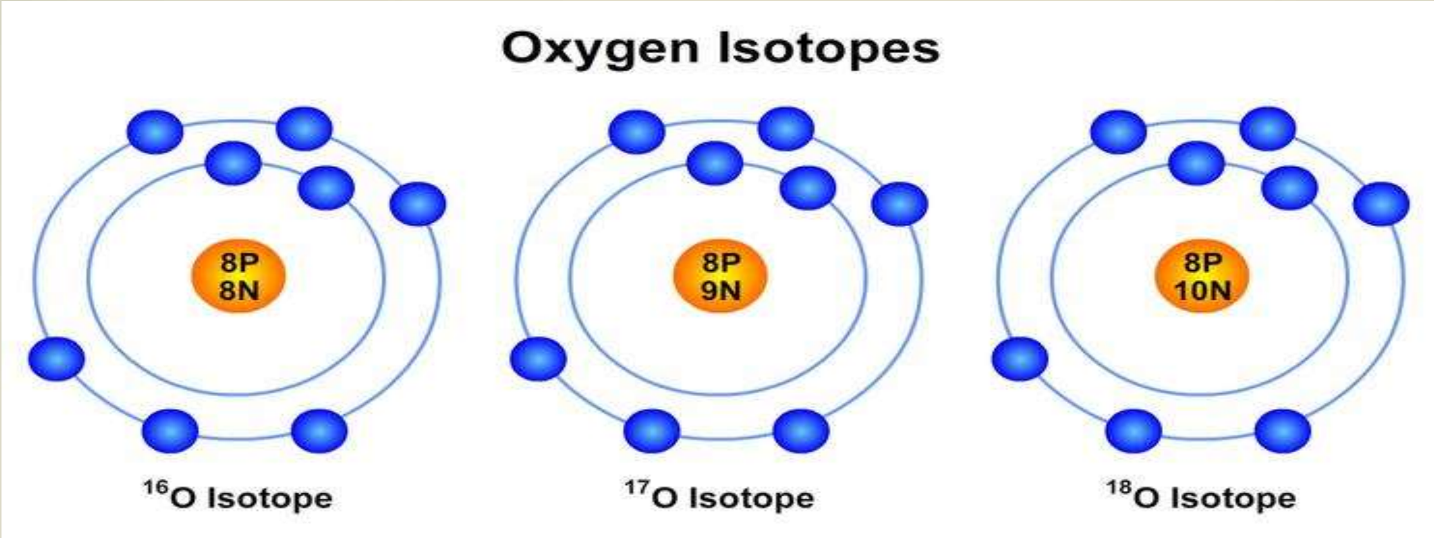
Hydrogen $^{1}_{1}H$

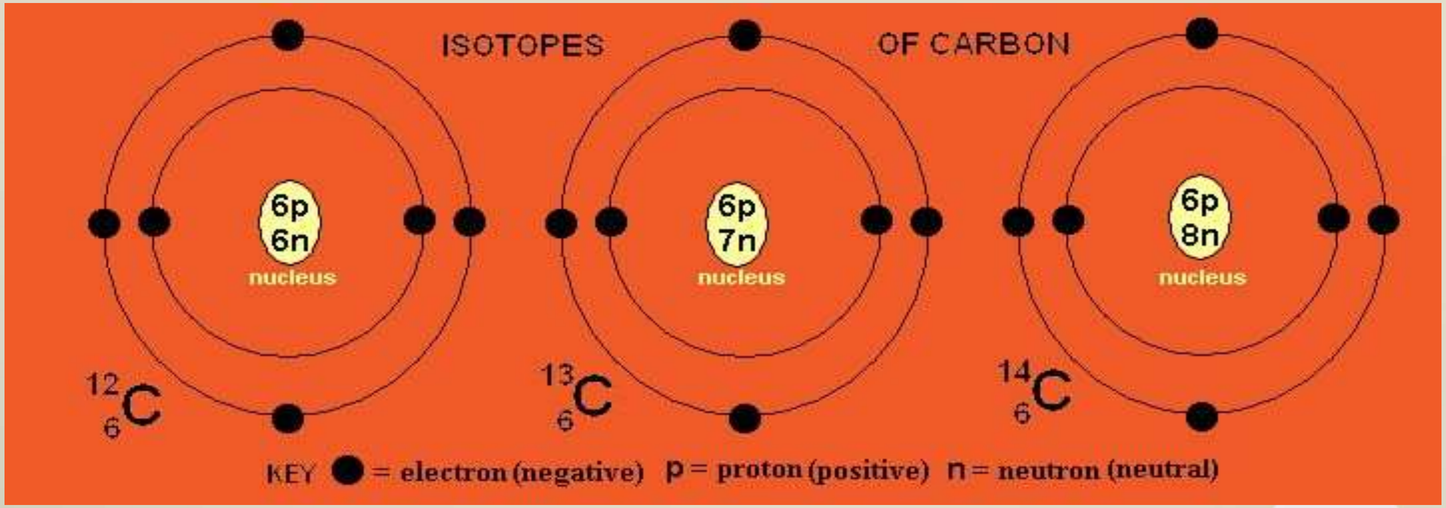
Deuterium $^{2}_{1}H$ (D)

Tritium $^{3}_{1}H$ (T)

$^{235}_{92}U$ $^{238}_{92}U$

Chemically, isotopes are not very different from each other.


Oxygen Isotopes
8P
8N
8P
9N
8P
10N
^{16}O Isotope
^{17}O Isotope
^{18}O Isotope


ISOTOPES OF CARBON
6p
6n
nucleus
6p
7n
nucleus
6p
8n
nucleus
$^{12}_{6}C$
$^{13}_{6}C$
$^{14}_{6}C$
KEY ● = electron (negative) p = proton (positive) n = neutron (neutral)

Isotopes

- **Isotopes :** Not all atoms of the same element have the same mass due to different numbers of neutrons in those atoms. (Same Z, different A)
- There are, for example, three naturally occurring isotopes of uranium:
  - Uranium-234   Uranium-236   Uranium-238

- isobaric: nuclear transformation in which nuclei have the same (A) but different (Z).
- ^{58}Fe on ^{58}Ni / ^{64}Ni on ^{64}Zn / ^{48}Ca on ^{48}Ti.

- Isotones (Same N, different A)
- $^{39}_{18}Ar$ & $^{40}_{19}K$ (N = 21)

Example 2.1

- Give the number of protons, neutrons and electrons in each of the following species
- $^{20}_{11}\text{Na}$ $^{22}_{11}\text{Na}$ $^{17}_{8}\text{O}$ $^{14}_{6}\text{C}$

	$^{20}_{11}\text{Na}$	$^{22}_{11}\text{Na}$	$^{17}_{8}\text{O}$	$^{14}_{6}\text{C}$
Mass Number	20	22	17	14
Atomic Number	11	11	8	6
Number of electrons	11	11	8	6
Number of protons	11	11	8	6
Number of neutrons	20-11 = 9	22-11=11	17-8=9	14-6=8

- The nucleus of an atom contains:

a. protons and neutrons.
b. protons and electrons.
c. electrons and neutrons.
d. air.

- Atoms with identical atomic numbers but different mass numbers are called:

a. mutants.
b. isomers.
c. Isotopes.
d. symbiots.

Example 7

- Consider the following nuclei:
- ^{14}C; ^{14}N; ^{12}C; ^{13}N
- Which are isotopes? Isotones? Isobars?
- ^{14}C and ^{12}C are isotopes of C
- ^{13}N and ^{14}N are isotopes of N
- ^{14}C and ^{14}N are isobars (A =14)
- ^{12}C and ^{13}N are isotones (N = 6).

An ***ion*** is an atom, or group of atoms, that has a net positive or negative charge.

cation – ion with a positive charge

If a neutral atom **loses** one or more electrons it becomes a cation.

Na — 11 protons, 11 electrons

Na^{+} — 11 protons, 10 electrons

anion – ion with a negative charge

If a neutral atom **gains** one or more electrons it becomes an anion.

Cl — 17 protons, 17 electrons

Cl^{-} — 17 protons, 18 electrons

Ion (№ of e ≠ p)

(lose or gain electrons_)

Cation: (№ of e < p)

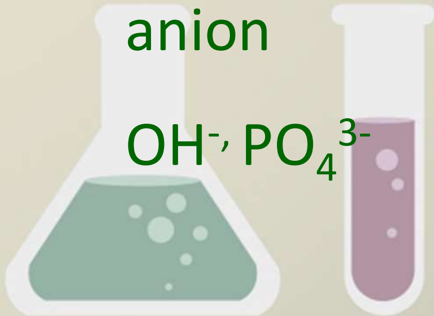
an ion with a +ve charge (lose electron/s)

- Monoatomic cation: Na^+, K^+, Mg^{2+}
- Polyatomic cation: NH_4^+

Anion: (№ of e > p)

an ion with a -ve charge (Gain electron/s)

- Monoatomic anion: Cl^-, Br^-,S^{2-}
- Polyatomic anion: OH^-, PO_4^{3-}



A magnesium ion, $_{12}Mg^{2+}$, has

A. 12 protons and 13 electrons.

B. 24 protons and 26 electrons.

C. 12 protons and 10 electrons.

D. 24 protons and 22 electrons.

E. 12 protons and 14 electrons.

A sulfide ion, $_{16}S^{2-}$, has:

A. 16 protons and 16 electrons

B. 32 protons and 16 electrons

C. 16 protons and 14 electrons

D. 16 protons and 18 electrons

E. 32 protons and 18 electrons

How many protons and electrons are in

$$^{27}_{13}Al^{3+}$$

13 protons, 10 (13 – 3) electrons

How many protons and electrons are in

$$^{78}_{34}Se^{2-}$$

34 protons, 36 (34 + 2) electrons

Use the following table and choose which of the species are **neutral**?

Atom or ion element	I	II	III	IV	V	VI
Atom or ion electrons (e)	6	10	18	10	28	7
Atom or ion protons (p)	6	8	17	11	30	7
Atom or ion neutrons (n)	6	8	18	11	36	6

A. III and V
B. IV and V
C. II and III
D. I and VI

Use the following table and choose which of the species are **negatively charged**?

Atom or ion element	I	II	III	IV	V	VI
Atom or ion electrons (e)	6	10	18	10	28	7
Atom or ion protons (p)	6	8	17	11	30	7
Atom or ion neutrons (n)	6	8	18	11	36	6

A. III and V
B. IV and V
C. II and III
D. I and VI

Atoms with the **same number of electrons and number of protons** are called…

A. Ions
B. isotopes
C. neutral atoms
D. different atoms

Periodic Table of elements

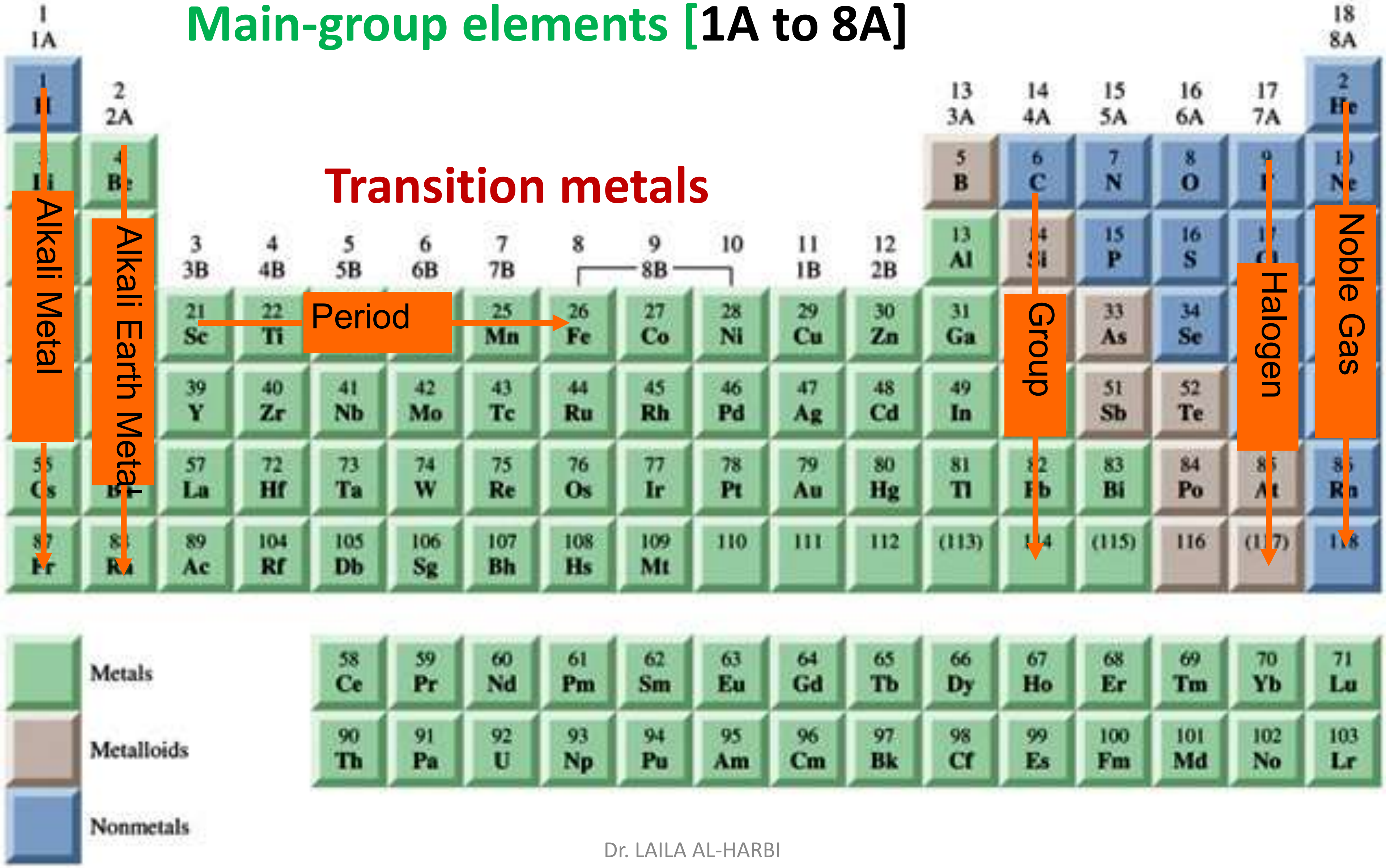
- Elements arranged in order of **increasing atomic number**.
- Horizontal Rows in periodic table are called **periods.**
- Vertical Columns are **groups or families; elements have similar properties.**
- **representative elements: A Group; transition elements: B Group**

Group	Name	Elements
1A	Alkali metals	Li, Na, K, Rb, Cs, Fr
2A	Alkaline earth metals	Be, Mg, Ca, Sr, Ba, Ra
6A	Chalcogens	O, S, Se, Te, Po
7A	Halogens	F, Cl, Br, I, At
8A	Noble gases (or rare gases)	He, Ne, Ar, Kr, Xe, Rn

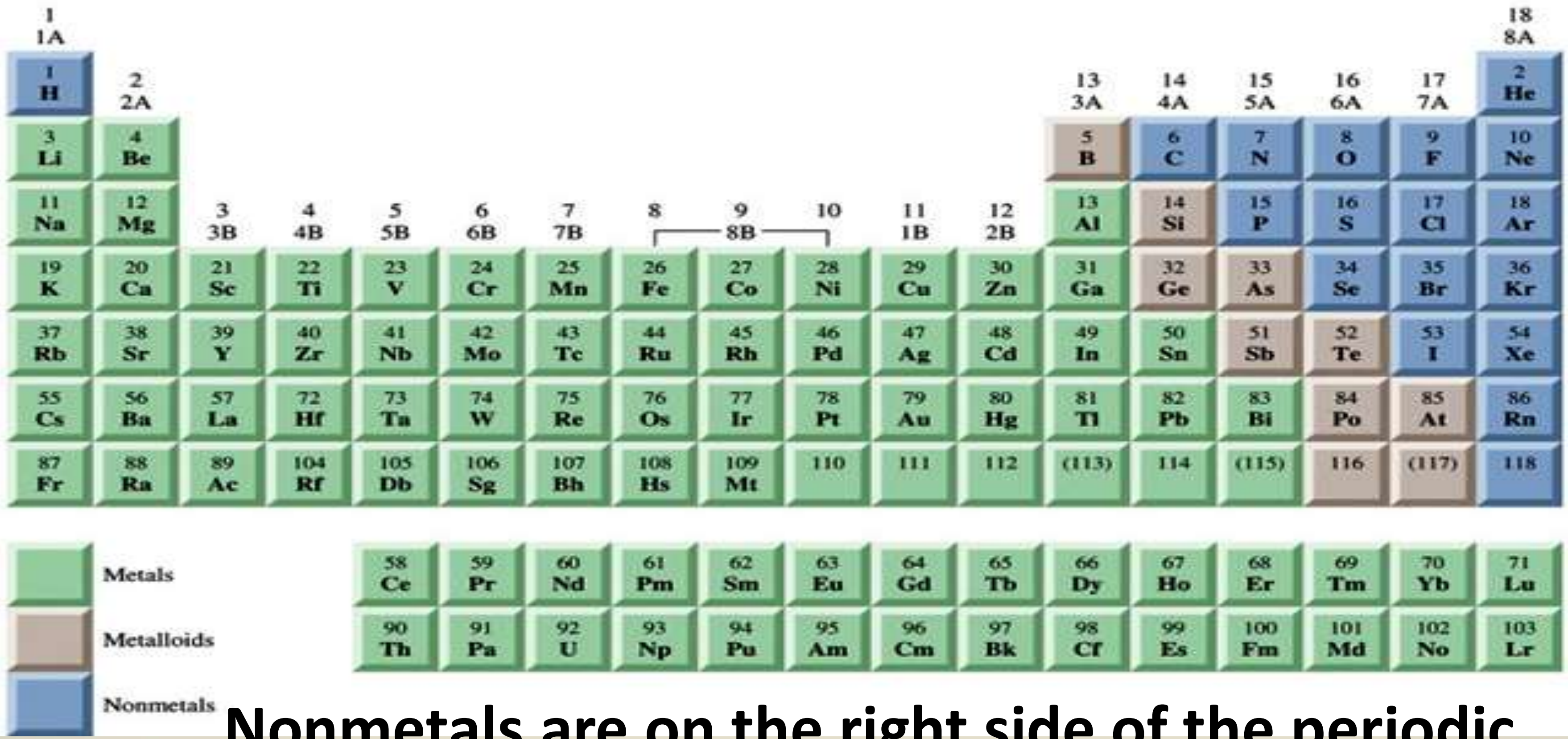
These five groups are known by their names

Modern Periodic Table

Main-group elements [1A to 8A]



- Elements in the periodic table are divided into three categories:
  - **Metal:** (in green colour, Most elements) is a good conductor of heat and electricity
  - **Nonmetal:** (in blue colour, 17 elements) is a poor conductor of heat and electricity
  - **Metalloid:** (in brown colour, 8 elements) has properties that are intermediate between those of metals and nonmetals



Nonmetals are on the right side of the periodic table (with the exception of H) (blue).

Metalloids border the stair-step line (with the exception of Al, Po, and At).

Metals are on the left side of the chart (green color)

- Positive ions are called:

a. positrons.

b. anions.

c. cations.

d. nucleons.

- The elements located in group 7A of the periodic table are called:

a. alkali metals.

b. noble gases.

c. chalcogens.

d. halogens.

What are the ions present in the compound $(NH_4)_2SO_4$?

- NH_3, H_2, and SO_2
- N^{3-}, H^+, S^{2-}, O^{2-}
- NH_4^{2+} and SO_4^-
- NH_4^+ and SO_4^{2-} (3 IONS)
- $(NH_4^+)_2$ and SO_4^{2-}

1- Selenium ($_{34}Se$) element is:

(a) a nonmetal

(b) found in group 6A

(c) both a and b

2- Gallium (Ga) element is found in the periodic table in

(a) period 3, group 1B

(b) period 3A, group 4

(c) period 4, group 1A

(d) period 4, group 3A

3- Which of the following sets of elements is expected to have similar chemical properties?

a) Sulfur and phosphorous

b) Sulfur and oxygen

c) Sulfur and argon

4- Which of these elements is most likely to be a good conductor of electricity?

Which of the following is metal?

A. N

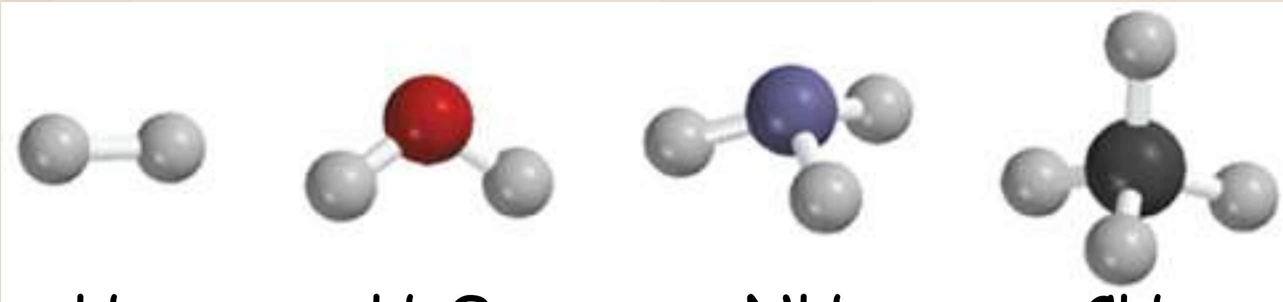
B. S

C. He

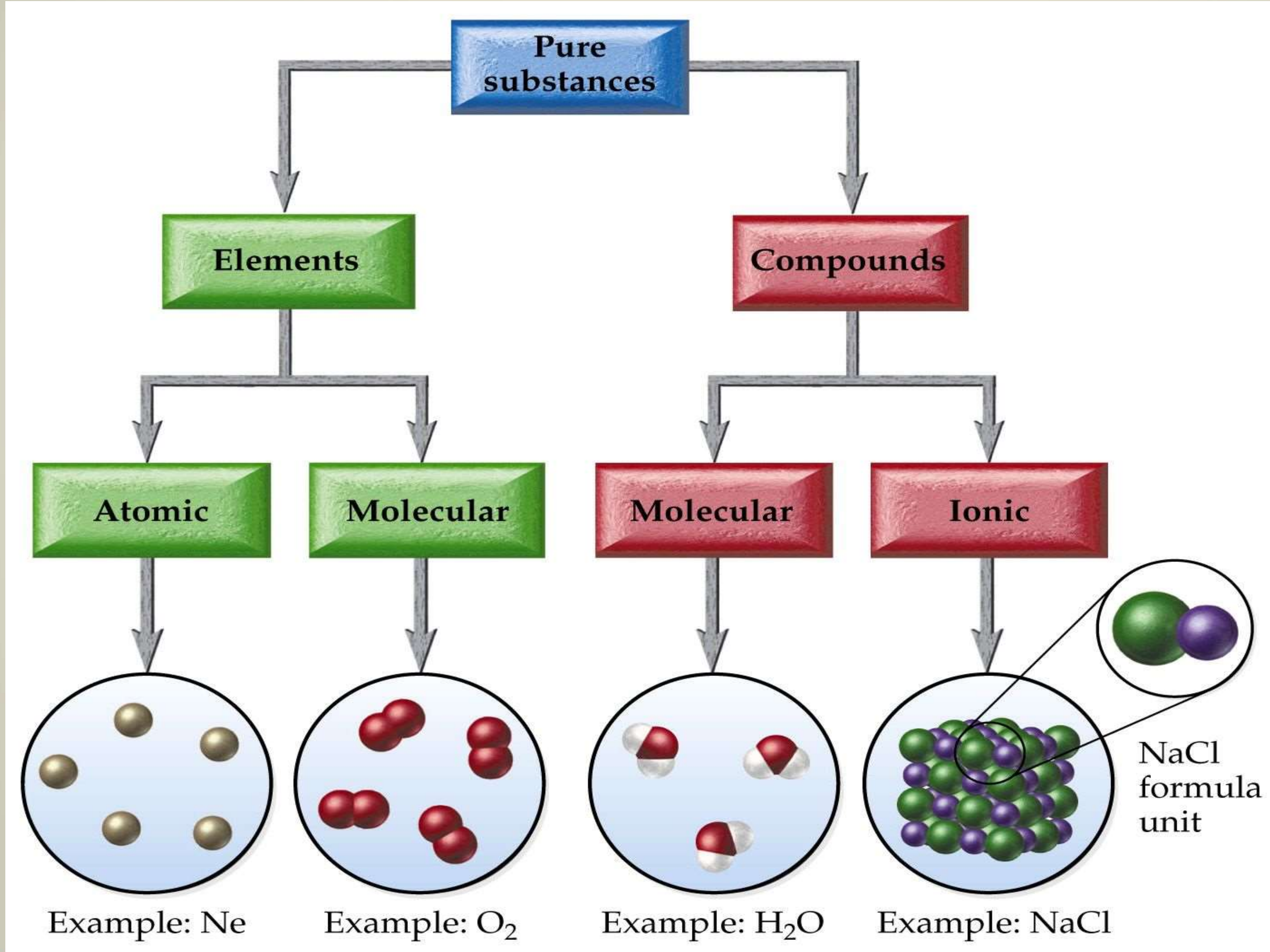
D. Fe

2.5 Molecules and Ions

A molecule: is an aggregate of two or more atoms in a definite arrangement held together by chemical forces, a molecule may contain atoms of the same element or atoms of two or more elements.



H_2 H_2O NH_3 CH_4


Pure substances
Elements
Compounds
Atomic
Molecular
Molecular
Ionic
NaCl formula unit
Example: Ne
Example: O_2
Example: H_2O
Example: NaCl

2.5 Molecules and Ions

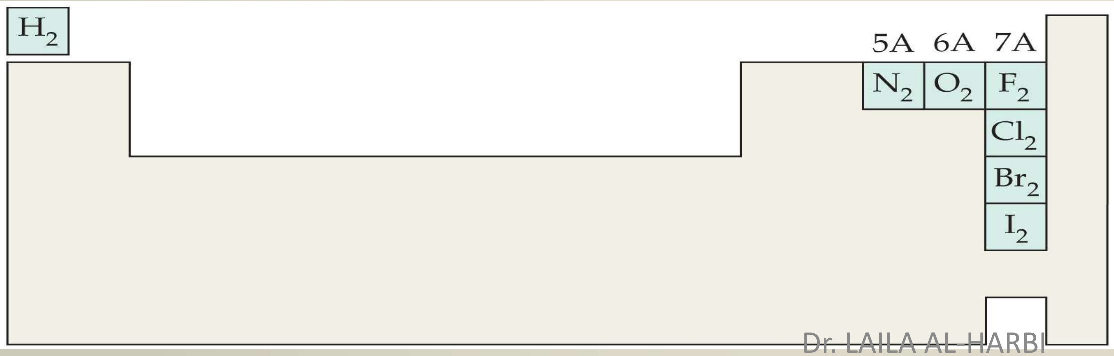
Molecule

Diatomic molecule: contains only two atoms

H_2, N_2, O_2, Br_2, HCl, CO

Polyatomic molecule: contains more than two atoms

O_3, H_2O, NH_3, CH_4



An ion: is an atom or a group of atoms that has a net positive (+ve) or negative (-ve) Charge.

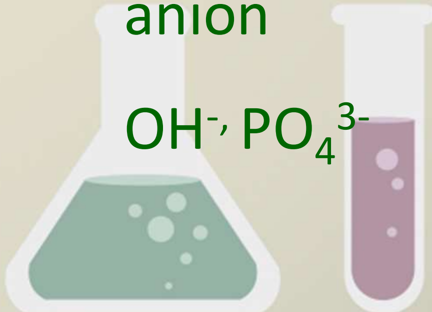
Ion

Cation: an ion with a +ve charge (lose electron/s)

- Monoatomic cation: Na^+, K^+, Mg^{2+}
- Polyatomic cation: NH_4^+

Anion: an ion with a -ve charge (Gain electron/s)

- Monoatomic anion: Cl^-, Br^-, S^{2-}
- Polyatomic anion: OH^-, PO_4^{3-}



- Which of the following is an example of polyatomic cation?

- A) Mg^{+2}
- **B) NH_4^{+1}**
- C)O^{-2}
- D) SO_4^{-2}

- Which of the following is an example of monatomic anion?

- A) Mg^{+2}
- B) NH_4^{+1}
- **C)O^{-2}**
- D) SO_4^{-2}

2.6 Chemical formulas

- **Molecular formulas** give the actual numbers and types of atoms in a molecule.

E.g. CH_4, H_2O_2, C_2H_4, $C_6H_{12}O_6$

- **Empirical formulas** give the smallest whole number ratio of atoms in a molecule. The empirical formula of many compounds is the same as the molecular formula

E.g. CH_4, HO, CH_2, CH_2O

- **Ionic formulas:** the number of electrons lost & gained must be equal, so + and -charge cance out.

An *empirical formula* shows the simplest whole-number ratio of the atoms in a substance

molecular		**empirical**	
H_2O	1 →	H_2O	
$C_6H_{12}O_6$	6 →	CH_2O	✓
	3 →	$C_2H_6O_2$	✗
O_3	3 →	O	
N_2H_4	2 →	NH_2	

EXAMPLE 2.3

- Write the Empirical formulas for the following molecules

✓ Acetylene C_2H_2 divided by 2 CH

✓ Glucose $C_6H_{12}O_6$ divided by 6 CH_2O

✓ Nitrous oxide N_2O , the Empirical formulas is same as molecular formula N_2O

✓ Caffeine $C_8H_{10}N_4O_2$ divided by 2 $C_4H_5N_2O$

Examples

- Which of the following is empirical formula

A. O_3 >>>>>> O

B. H_2SO_4

C. S_8 >>>>>>> S

D. $C_6H_{12}O_6$ >>>> CH_2O

المطلوب الصيغة التي لايمكن تبسيطها

الاجابة الثانية لايمكن تبسيطها هي الاجابة الصحيحة

- Which of the following is molecular formula

A. CO_2

B. H_2SO_4

C. S_8

D. CH_4O

المطلوب الصيغة التي يمكن تبسيطها

الاجابة الثالثة هي الوحيده التي يمكن تبسيطها هي الاجابة الصحيحة

Rules for writing ionic formula:

- 1) Write down formulas of ions
- 2) Combine the smallest number of ions to give the charge sum equal to 0; if the charges are not equal, find the lowest common multiple

- E.g. Predict the formula for the compound formed from the following elements:
- Potassium bromide K^+ Cl^- = KCl
- Zinc iodide Zn^{+2} I^- = ZnI_2
- Aluminum oxide Al^{+3} O^{-2} = Al_2O_3

$2 \times +3 = +6$ $3 \times -2 = -6$

Al_2O_3

Al^{3+} O^{2-}

Examples 2.4

- Write the formula of magnesium nitride?

$3 \times +2 = +6$ $\quad$ $2 \times -3 = -6$

Mg_3N_2

Mg^{2+} $\quad$ N^{3-}

- Write the formula of
a)chromium sulfate ,
b)titanium oxide?

$2 \times +3 = +6$ $\quad$ $3 \times -2 = -6$

$Cr_2(SO_4)_3$

Cr^{3+} $\quad$ SO_4^{2-}

$2 \times +4 = +8$ $\quad$ $4 \times -2 = -8$

Ti_2O_4

Ti^{4+} $\quad$ O^{2-}

2.7 naming compounds

- ❖ Ionic compounds
- ❖ Molecular compounds
- **Ionic compounds**

consist of metals (positive ions (cations) and negative ions (anions).

A. Naming Cations

1. **Fixed charge metals:**

Cations have same name as the metal element. (Groups1A, 2A, 3A , transition metals) have specific charge.

Ag^{+} silver ion Zn^{2+} zinc ion , Al^{3+} Aluminum ion

Li^{+} lithium ion Ca^{2+} calcium ion

Naming Ionic compounds

Metals
Positive ion
Cation

(1A , 2A ,3A & transition metals)

Fixed charge metals

Cations have same name as the metal element

Ag^{+} silver ion , Zn^{2+} zinc ion , Al^{3+} Aluminum ion , Li^{+} lithium ion , Ca^{2+} calcium ion

Variable charge metals

(STOCK SYSTEM)

If the metal can form more than 1 cation, the charge is indicated by a Roman numeral in parenthesis after the metal name.

Fe^{2+} iron(II) , Au^{+} gold(I), Cu^{+} copper(I) ,
Fe^{3+} iron(III) Au3^{+} gold(III) Cu^{2+} copper (II)
$Hg_2{}^{+2}$ mercury(I) Hg^{2+} mercury (II)

Non metals
Negative ion
Anion

monoatomic anions

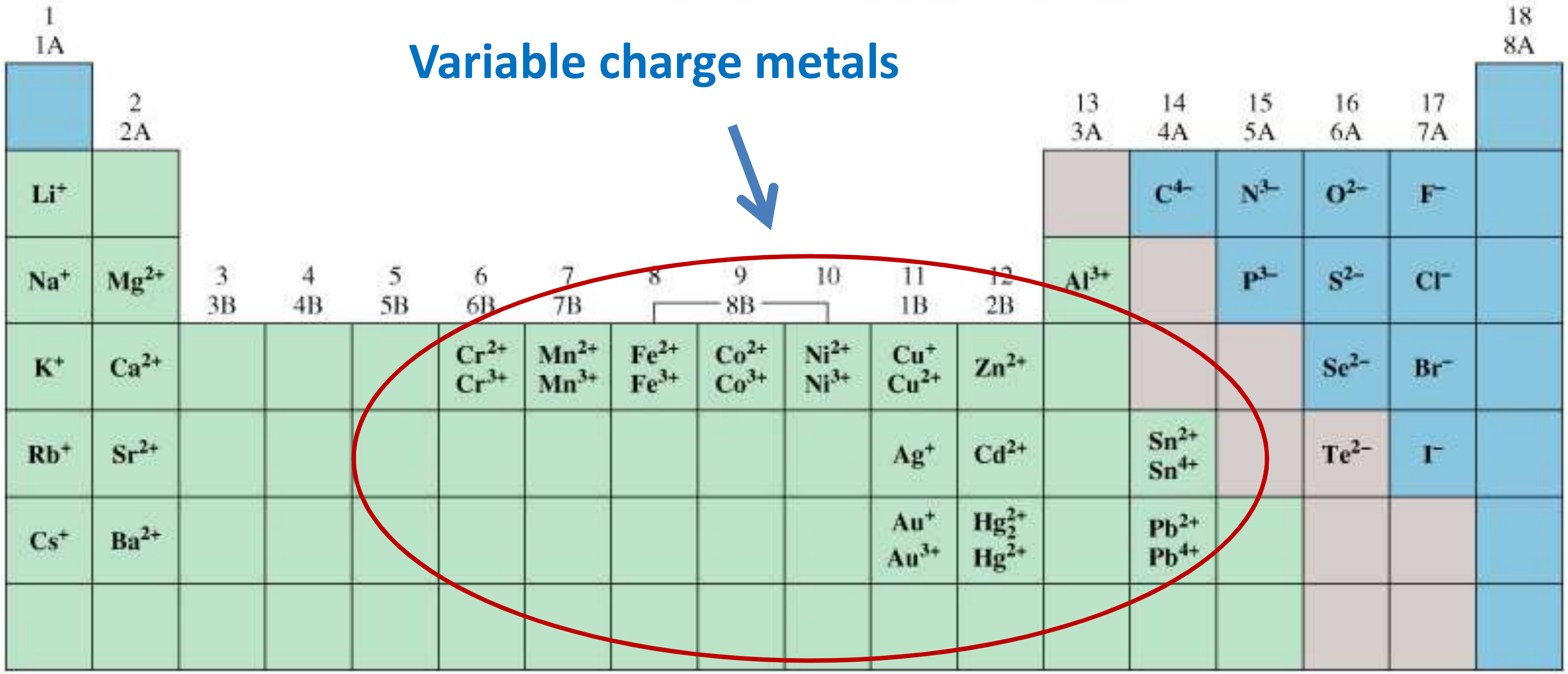
Polyatomic anions

Dr. LAILA AL-HARBI

- 2. **Variable charge metals:**
- If the metal has more than one oxidation state, the charge is indicated by a Roman numeral in parenthesis after the metal name.
- Most of the transition metals are variable charge metals.
- E.g. Common metals which exist in more than one positive state:
- Fe^{2+} iron(II) ,Au^{+} gold(I), Cu^{+} copper(I) ,
  Fe^{3+} iron(III) Au3^{+} gold(III) Cu^{2+} copper (II)
  $Hg_2{}^{+2}$mercury(I) Hg^{2+} mercury (II)
- 3. **Polyatomic cations**: consist of nonmetals:

  H_3O^{+} hydronium ion $NH_4{}^{+}$ ammonium

Variable charge metals



NOTE

- Some Cations of variable charge have name for each oxidation state
- Example

Fe^{2+} iron(II) ferrous , Fe^{3+} iron(III) ferric

Cu^{+} copper(I) cuprous , Cu^{2+} Copper (II) cupric

$Hg_2{}^{+2}$mercury(I) **mercurous**

Hg^{2+} mercury (II) **mercuric**

Mercury (Hg) is the only metal has this formula when it form cation with only one positive charge :

Hg_2^{2+} NOT Hg^{+}

The cation of two positive charges has the formula Hg_2^{+}

B. Naming Anions

- 1. monoatomic anions: change ending to -ide
- E.g. Oxygen→ Oxide   Sulfur → Sulfide

Hydrogen →Hydride   chlorine → Chloride

Florine → Floride   Bromine → Bromide

- Polyatomic anions: most end in -ate or -ite; usually contain O (oxy)
- Know polyatomic anions on handout.
- Carbonate CO_3^{-2} , Nitrate NO_3^{-}, Sulfate SO_4^{-2} ,
- Phosphate PO_4^{-3}

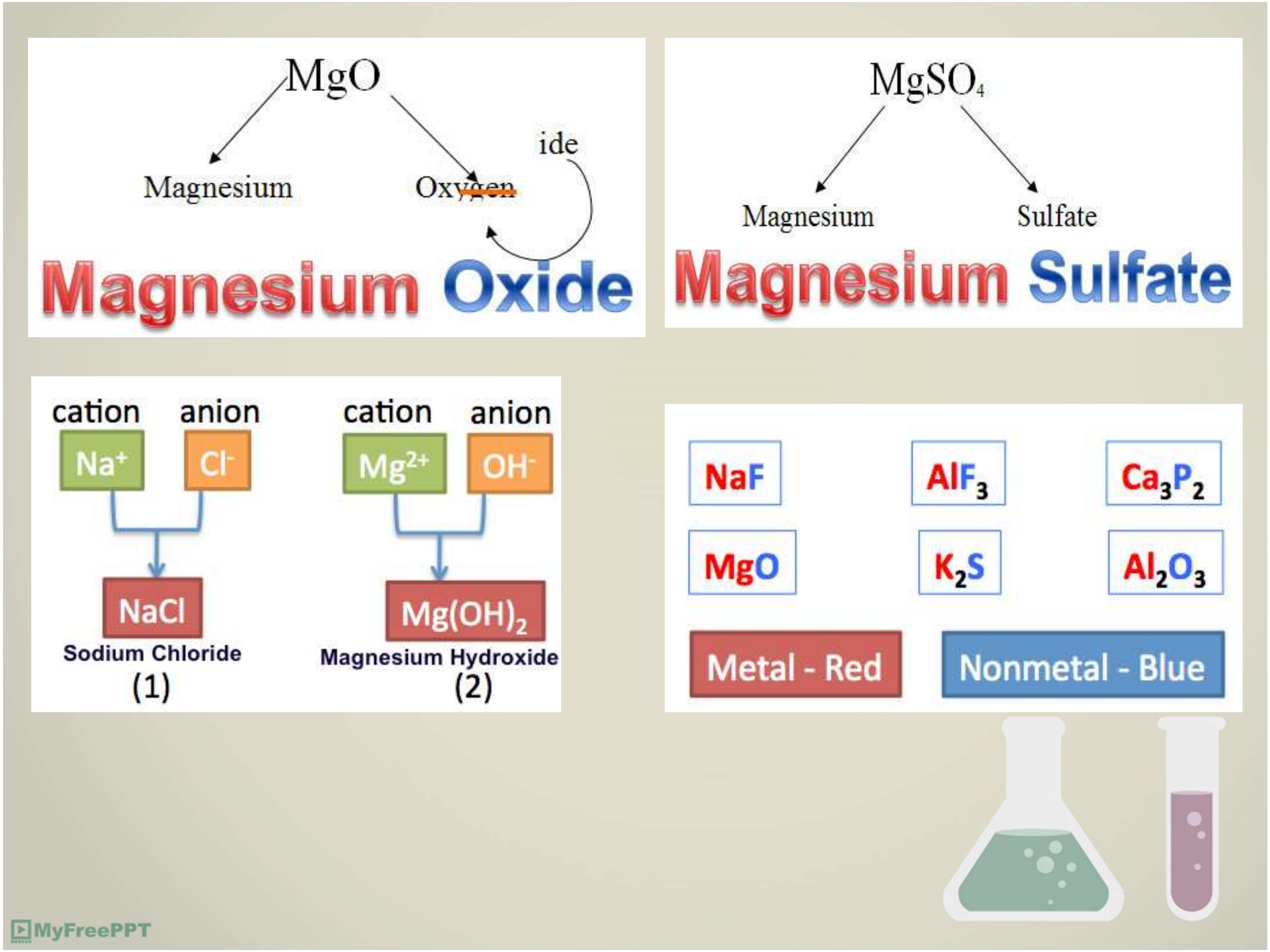
Cyanide CN^- , Hydroxide OH^- , Oxide O_2^{-2}

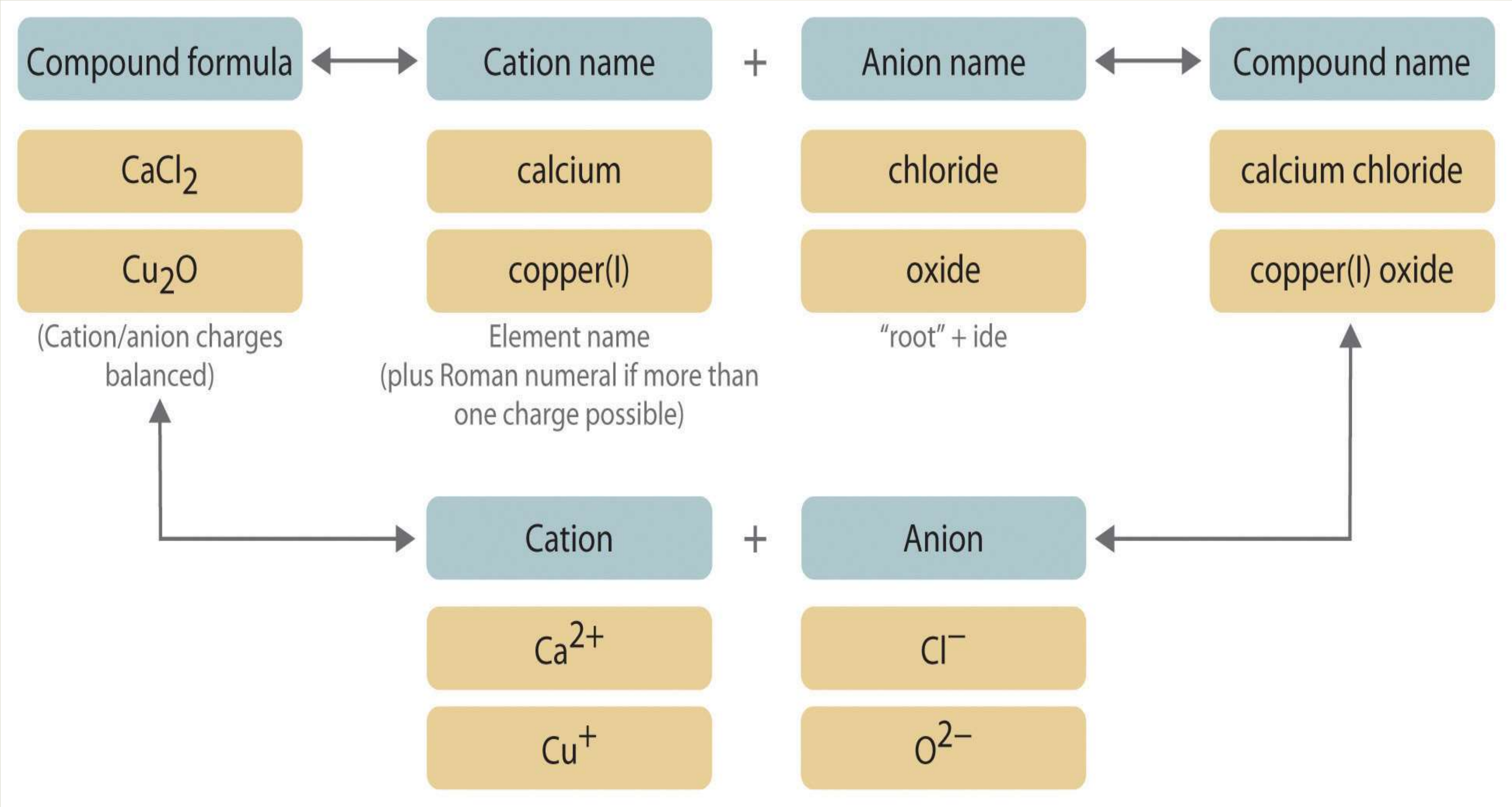
See table 2.3

- Ionic compounds names start with the positive ion (metal) (include Roman numeral in parenthesis ONLY IF metal has variable charge) followed by the negative ion (nonmetal).
- $NaCl$ Sodium Chloride
- $BaCl_2$ Barium Chloride
- K_2O Potassium oxide
- KNO_3 Potassium Nitrate
- Na_2CO_3 Sodium Carbonate
- $FeCl_2$ Iron(II) Chloride → ferrous Chloride
- $FeCl_3$ Iron(III) Chloride → ferric Chloride
- Cr_2S_3 Chromium(III) Sulfide
- $(NH_4)_3PO_4$ Ammonium Phosphate
- **$Cu(NO_3)_2$ Cupper(II)nitrate**
- **PbO Lead(II) oxide**

TABLE 2.3 Names and Formulas of Some Common Inorganic Cations and Anions

Cation	Anion
aluminum (Al^{3+})	bromide (Br^-)
ammonium (NH_4^+)	carbonate (CO_3^{2-})
barium (Ba^{2+})	chlorate (ClO_3^-)
cadmium (Cd^{2+})	chloride (Cl^-)
calcium (Ca^{2+})	chromate (CrO_4^{2-})
cesium (Cs^+)	cyanide (CN^-)
chromium(III) or chromic (Cr^{3+})	dichromate ($Cr_2O_7^{2-}$)
cobalt(II) or cobaltous (Co^{2+})	dihydrogen phosphate ($H_2PO_4^-$)
copper(I) or cuprous (Cu^+)	fluoride (F^-)
copper(II) or cupric (Cu^{2+})	hydride (H^-)
hydrogen (H^+)	hydrogen carbonate or bicarbonate (HCO_3^-)
iron(II) or ferrous (Fe^{2+})	hydrogen phosphate (HPO_4^{2-})
iron(III) or ferric (Fe^{3+})	hydrogen sulfate or bisulfate (HSO_4^-)
lead(II) or plumbous (Pb^{2+})	hydroxide (OH^-)
lithium (Li^+)	iodide (I^-)
magnesium (Mg^{2+})	nitrate (NO_3^-)
manganese(II) or manganous (Mn^{2+})	nitride (N^{3-})
mercury(I) or mercurous (Hg_2^{2+})*	nitrite (NO_2^-)
mercury(II) or mercuric (Hg^{2+})	oxide (O^{2-})
potassium (K^+)	permanganate (MnO_4^-)
rubidium (Rb^+)	peroxide (O_2^{2-})
silver (Ag^+)	phosphate (PO_4^{3-})
sodium (Na^+)	sulfate (SO_4^{2-})
strontium (Sr^{2+})	sulfide (S^{2-})
tin(II) or stannous (Sn^{2+})	sulfite (SO_3^{2-})
zinc (Zn^{2+})	thiocyanate (SCN^-)


MgO
Magnesium
Oxygen
ide
Magnesium Oxide
MgSO4
Magnesium
Sulfate
Magnesium Sulfate
cation
anion
Na+
Cl-
NaCl
Sodium Chloride
(1)
cation
anion
Mg2+
OH-
Mg(OH)2
Magnesium Hydroxide
(2)
NaF
AlF3
Ca3P2
MgO
K2S
Al2O3
Metal - Red
Nonmetal - Blue
MyFreePPT


Compound formula
Cation name
+
Anion name
Compound name
$CaCl_2$
calcium
chloride
calcium chloride
Cu_2O
copper(I)
oxide
copper(I) oxide
(Cation/anion charges balanced)
Element name (plus Roman numeral if more than one charge possible)
"root" + ide
Cation
+
Anion
Ca^{2+}
Cl^-
Cu^+
O^{2-}

Example 2.6

- Write the chemical formula for the following compounds
- Mercury(I)nitrite

$Hg_2(NO_2)_2$

- Cesium sulfide

Ce_2S

- Calcium phosphate
- $Ca_3(PO_4)_2$

PRACTIES EXERICISE 2.6

- Write the chemical formula for the following compounds
- Rubidium sulfate
- Rb_2SO_4
- Barium hydride
- BaH_2

Example 2.5 p61:

Name the following compounds:

(a) $Cu(NO_3)_2$

1. **Cation:** Copper cation (can form two types of cation →Stock system) → Copper (II)
2. **Anion:** NO_3^- anion has a common name Nitrate

Thus: the name of the compound is:

Copper (II) nitrate

(b) KH_2PO_4

1. **Cation:** K form only one type of cation → Potassium Note: <u>not potassium (I)</u>
2. **Anion:** $H_2PO_4^-$ has a common name dihydrogen phosphate

Thus: the name of the compound is:

Potasium dihydroen phosphate

(c) NH_4ClO_3

1. **Cation:** NH_4^+ has a common name ammonium
2. **Anion:** ClO_3^- has a common name chlorate

Thus: the name of the compound is:

Ammonium chlorate

Example 2.6 p62:

Write chemical formulas for the following compounds:

(a) Mercury (I) nitrite

Roman number (I) shows that mercury has +1 charge $\rightarrow Hg_2^{2+}$

Nitrite is a common name of NO_2^-

Thus: the chemical formula is:

$Hg_2(NO_2)_2$

H.W. Solve the practice exercise p62

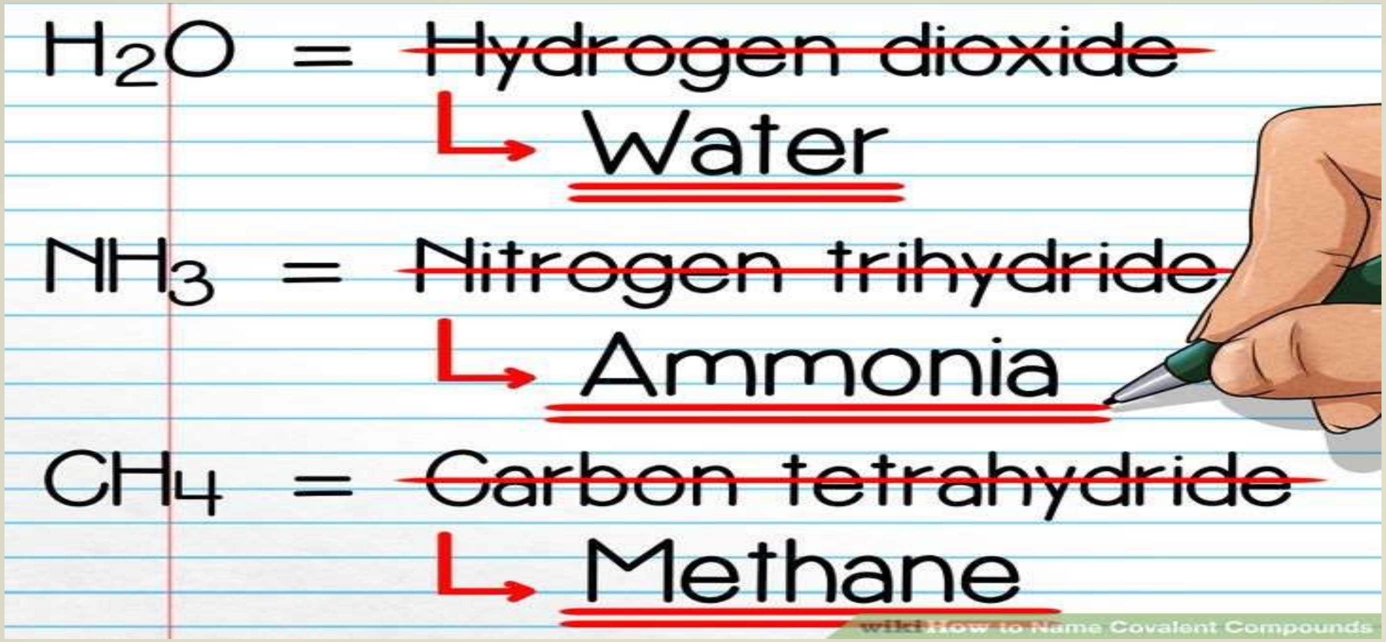
Molecular compounds

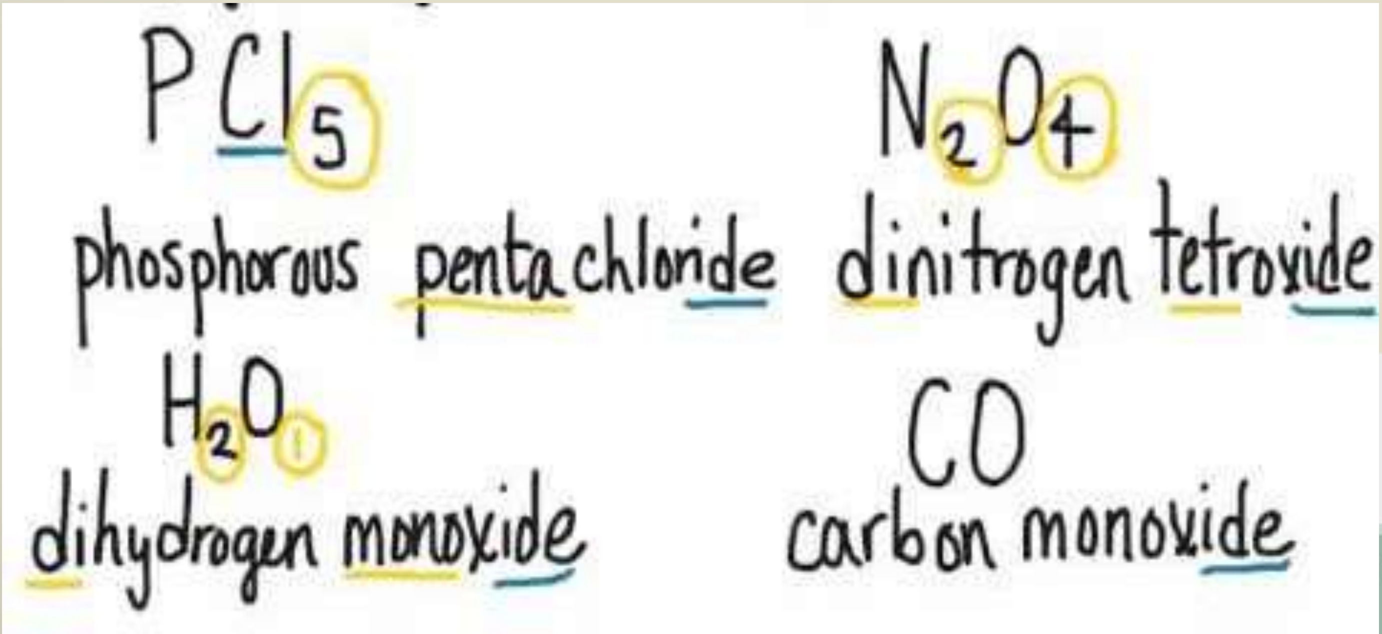
- nonmetals or nonmetals + metalloids
- common names
  - H_2O water
  - NH_3 ammonia
  - CH_4 methane
  - H_2S hydrogen sulfide
  - SiH_4 silane
  - B_2H_6 diborane
- 1) Name 1st element & use a prefix (table 2.4) to indicate the number of atoms.
- .
- 2)Name 2nd element & include prefix for number of atoms (see table 2.4).
- 3) Change ending of 2nd element to –ide.

TABLE 2.4

Greek Prefixes Used in Naming Molecular Compounds

Prefix	Meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10


H2O = Hydrogen dioxide
Water
NH3 = Nitrogen trihydride
Ammonia
CH4 = Carbon tetrahydride
Methane
wikiHow to Name Covalent Compounds


PCl5
phosphorous pentachloride
N2O4
dinitrogen tetroxide
H2O
dihydrogen monoxide
CO
carbon monoxide

Note

- **Note that mono- is never used for the first element**

- **For oxides, the ending “a” in the prefix is omitted.**
- N_2O_4 dinitrogen tetroxide **not** (dinitrogen tetraoxide)

- **For oxides, the ending “o” in the prefix is omitted.**

- N_2O dinitrogen monoxide **not** (dinitrogen monooxide)

Molecular Compounds

HI	hydrogen iodide
NF_3	nitrogen trifluoride
Br_2O_7	Dibromine heptoxide
SO_2	sulfur dioxide
N_2Cl_4	dinitrogen tetrachloride
NO_2	nitrogen dioxide
N_2O	dinitrogen monoxide
ICl_3	Iodine trifchloride

IONIC COMPOUNDS

- Iron (III) sulfide $\rightarrow Fe_2S_3$
- Silver dichromate $\rightarrow Ag_2Cr_2O_7$
- Sodium phosphide $\rightarrow Na_3P$
- Cobalt (III) nitrite $\rightarrow Co(NO_2)_3$
- Tin(IV) chloride $\rightarrow SnCl_4$
- Chromium(III) thiocyanate $\rightarrow Cr(SCN)_3$
- Lead(IV) oxide $\rightarrow PbO_2$
- Calcium phosphite $\rightarrow Ca_3(PO_3)_2$
- Arsenic(V) sulfide $\rightarrow As_2S_5$
- manganese(VII) oxide $\rightarrow Mn_2O_7$

MOLECULAR COMPOUND

- Tetrasulfur octoxide $\rightarrow S_4O_8$
- Aluminum hydride $\rightarrow AlH_3$
- Diphosphorus pentasulfide $\rightarrow P_2S_5$
- Sulfur hexafloride SF_6
- Dinetrogen pentoxide P_2O_5
- Disulfur pentafluoride S_2F_{10}

Which of these pairs of elements would be most likely to form an ionic compound?

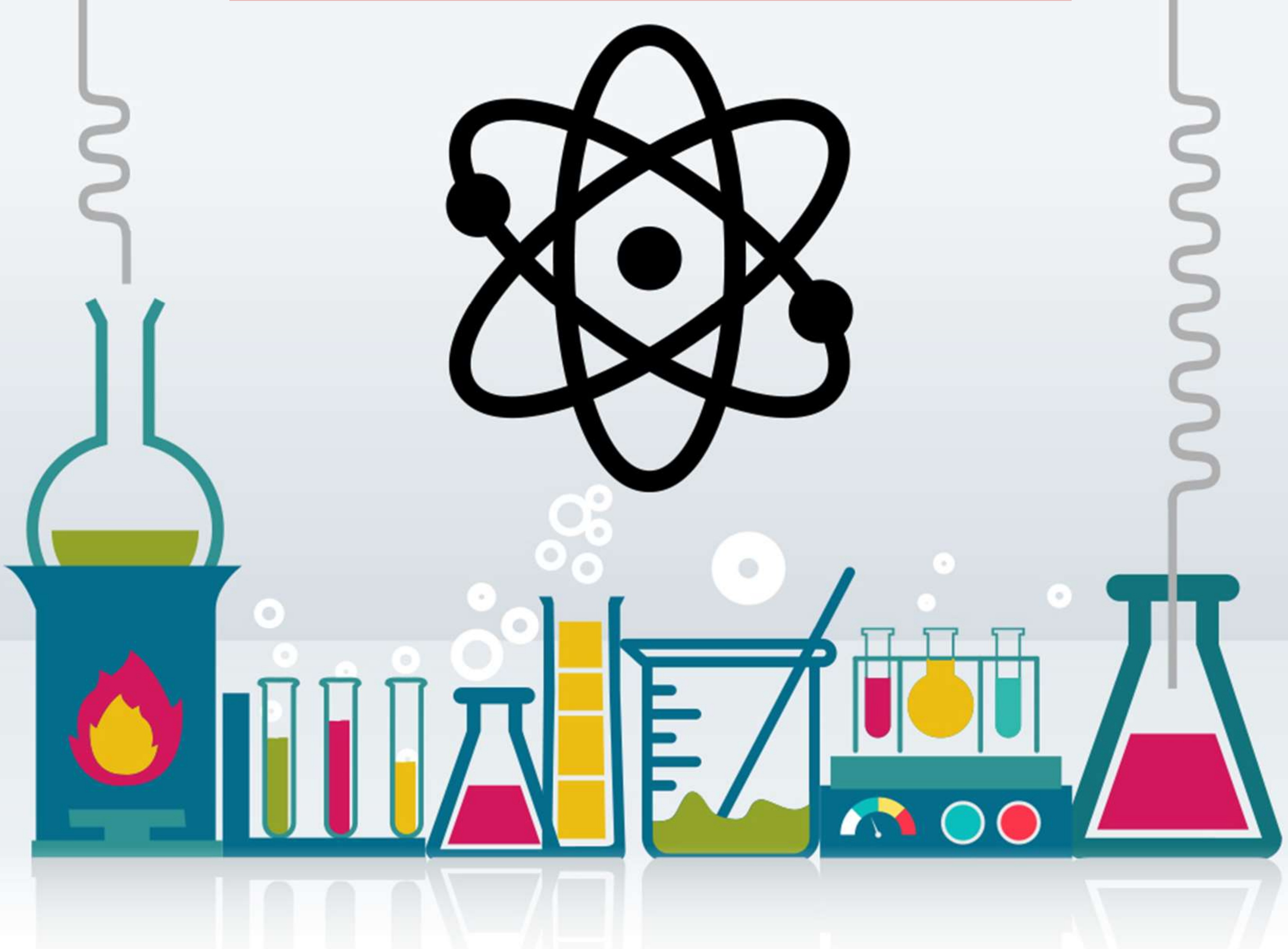
(a) P and Br
(b) Cu and K
(c) C and O
(d) O and Zn
(e) Al and Rb

Which of these pairs of elements would be most likely to form a molecular compound?

(a) Na and Br
(b) C and O
(c) Ca and O
(d) Zn and O
(e) Mg and Cl

Chapter 3

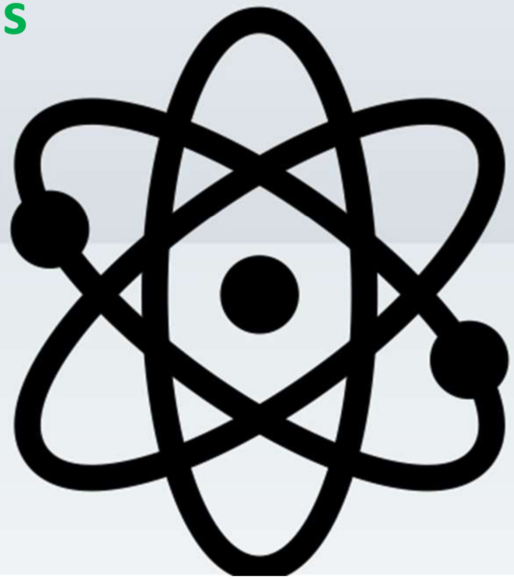
Mass relationships in chemical reactions



المعلومات الأساسية
وكتابتها على شكل
خريطة ذهنية
وتحديد
مواعيد
للمذاكرة
المذاكرة. فإذا كان الطالب
بطيئا يجب أن يبدأ مبكرا
النوم لفترة
كافية
استخدام الألوان
لتحديد المعلومات
المهمة
الجلوس
في مكان
مريح
وهادئ
تناول الأطعمة
الصحية
تلخيص
أهم العناصر
استخدام
الكلمات
المفتاحية
الابتعاد
عن الأجهزة
الالكترونية
المذاكرة الجماعية
وتبادل الملخصات
نصائح عامة
للمذاكرة
نصائح علمية لمذاكرة

Chapter 3
Mass relationships in chemical reactions

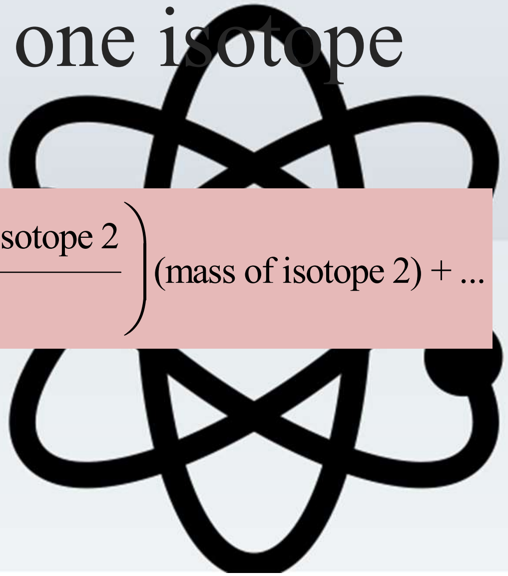
- **3.1 atomic mass**
- **3.2 Avogadro's number and molar mass of an element**
- **3.3 molecular mass**
- **3.5 percent composition of compounds**
- **3.6 experimental determination of empirical formula**
- **3.7 Chemical Reactions and Chemical Equations**
- **3.9 limiting reagents**
- **3.10 reaction yield**



3.1 atomic mass

- Each atom have more than one isotope with different abundance
- **Average atomic Mass**: the average mass of all of the isotopes of an element, each one weighted by its proportionate abundance
- Science each atom have more than one isotope with different abundance

$$\text{Average Atomic Mass} = \left(\frac{\%\ \text{abundance of isotope 1}}{100}\right)(\text{mass of isotope 1}) + \left(\frac{\%\ \text{abundance of isotope 2}}{100}\right)(\text{mass of isotope 2}) + \ldots$$



Average atomic mass

Average atomic mass of Lithium

- Natural lithium is:
- 7.42% ^{6}Li (6.015 amu)
- 92.58% ^{7}Li (7.016 amu

$$\frac{(7.42\% \times 6.015) + (92.58\% \times 7.016)}{100}$$

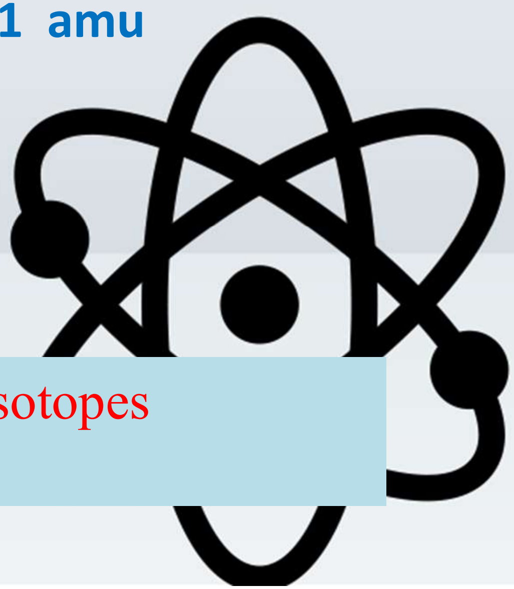
= 6.941 amu

Average atomic mass of carbon

- Natural Carbon is:
- 1.18% ^{13}C (13 amu)
- 98.9% ^{12}C (12 amu)

$$\frac{(98.9\ \% \times 12) + (1.18\% \times 13)}{100}$$

= 12.01 amu



The average atomic mass is between the atomic masses of the isotopes
And near the value of the highest abundance

EXAMPLE 3.1

- ^{65}Cu (30.91percent)

Atomic mass 64.9278

- ^{63}Cu (69.091percent)

Atomic mass 62.93

$$\frac{(30.91\% \times 64.9278) + 69.091\% \times 62.93)}{100}$$

= 63.55 amu

PRACTIES EXERICISE 3.1

- ^{10}B (19.78 percent)

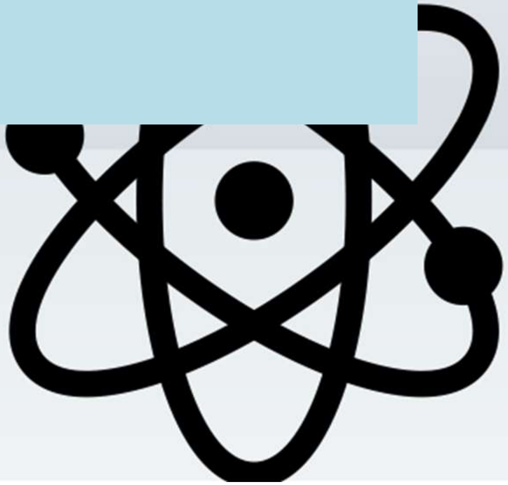
Atomic mass 10.0129

- ^{11}B (80.78percent)

Atomic mass 11.0093

$$\frac{(19.78\ \% \times 10.0129) + 80.78\%\ 11.0093)}{100}$$

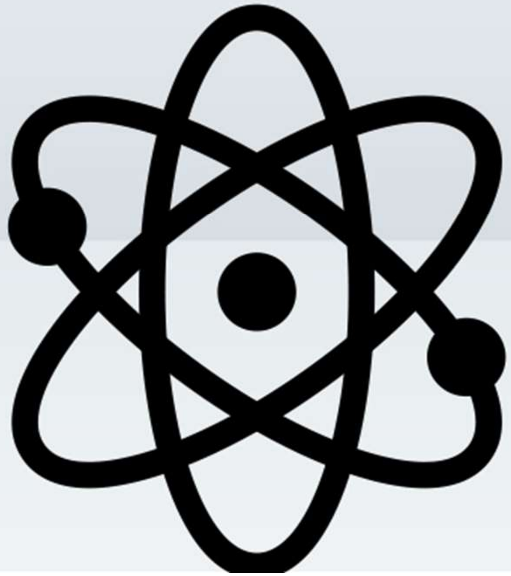
=10.81amu



Iodine has two isotopes ^{126}I and ^{127}I, with the equal abundance.
Calculate the average atomic mass of Iodine ($_{53}I$).

(a) 126.5 amu
(b) 35.45 amu
(c) 1.265 amu
(d) 71.92 amu

equal abundance MEAN each atom has abundance 50% .

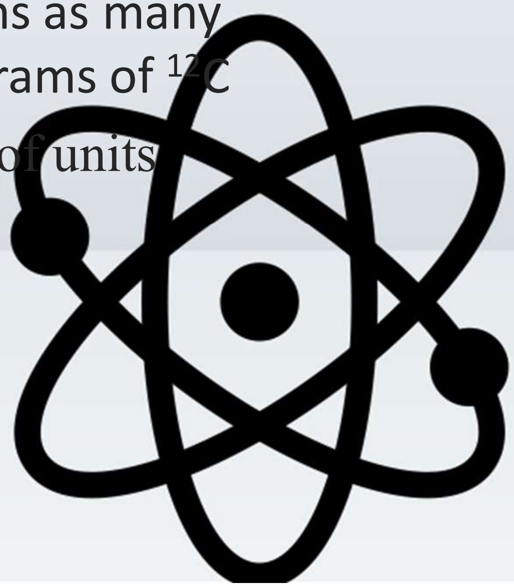


3.2 Avogadro's number and molar mass of an element

- **Atomic mass** is the mass of an atom in atomic mass units (amu)

By definition: 1 atom ^{12}C "weighs" 12 amu

- On this scale ^{1}H = 1.008 amu ^{16}O = 16.00 amu
- **Avogadro's Number , Is the number of atoms in exactly 12 grams of carbon-12 ($N_A = 6.022 \times 10^{23}$)**
- The **mole (mol)** is the amount of a substance that contains as many elementary entities as there are atoms in exactly 12.00 grams of ^{12}C
- One mole of a substance contains an Avogadro's Number of units



THUS:

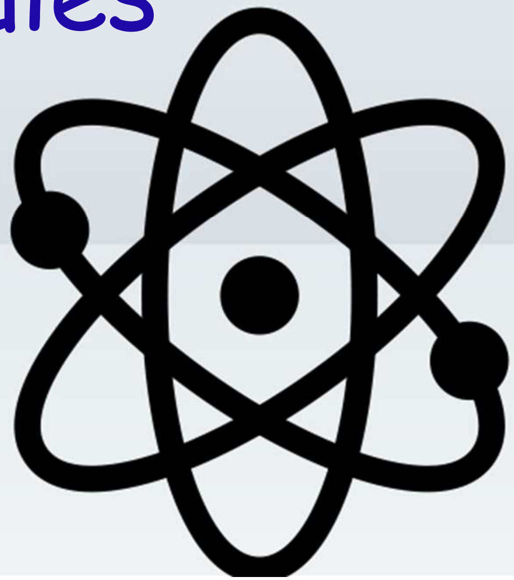
one mole of H atoms has

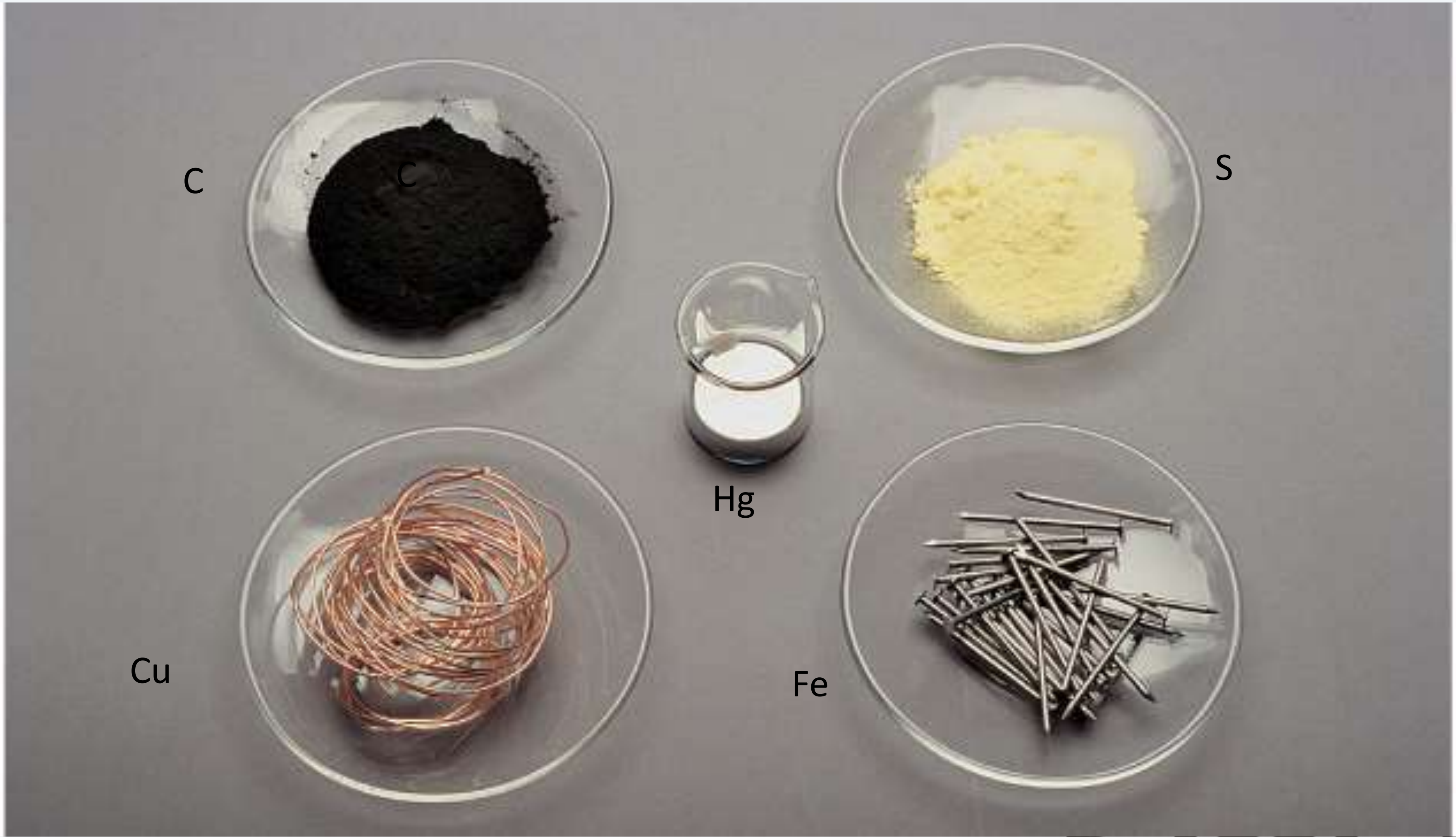
6.022×10^{23} atoms

&

One mole of H_2 molecules has

6.022×10^{23} molecules





One mole of these substances contain = 6.022 x 10^{23} atoms

but is not equal because they have different molar masses

Molar Mass

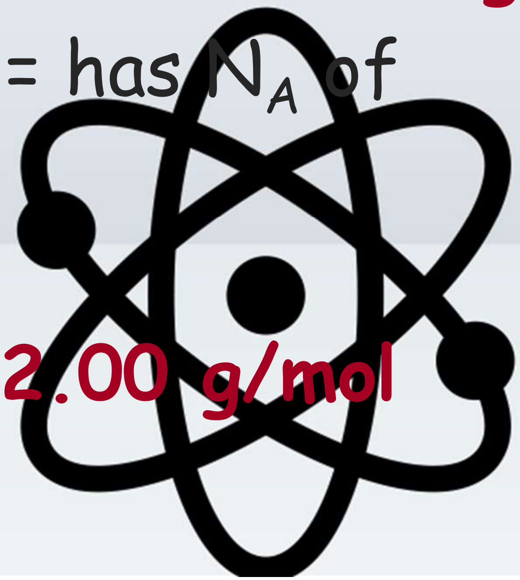
- **Molar mass (M):** the mass (in g or kg) of one mole of a substance;

M = mass/mol = g/mol

For ONE MOLE: 1 amu = 1 g

- **The atomic mass of ^{12}C is 12.00 amu = 12.00 g**
- 1 mole of ^{12}C = 12.00 amu = 12.00 g = has N_A of atoms = has 6.022 x 10^{23} atoms
- Thus:

The Molar Mass (M) of ^{12}C = 12.00 g/mol



Molar Mass (g/mol)

=

Atomic Mass (amu)

Examples:

1. The atomic mass of Na = 22.99 amu
   The molar mass of Na = 22.99 g/mol

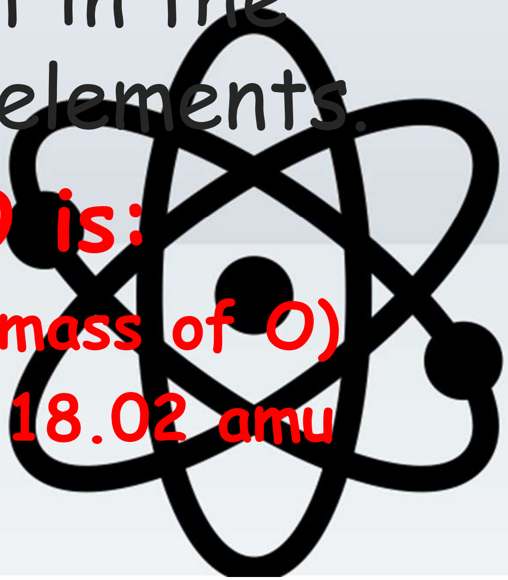
2. The atomic mass of P = 30.97 amu
   The molar mass of P = 30.97 g/mol

Molecular Mass

- **Molecular Mass (molecular weight):** is the sum of the atomic masses (in amu) in the molecule. **(MOLECULE)**
- **Molecular Mass:** multiply the atomic mass of each element by the number of atoms of that element present in the molecule and sum over all the elements.
- **e.g. Molecular Mass of H_2O is:**

(2 x atomic mass of H) + (1x atomic mass of O)

(2 x 1.008 amu) + (1x 16.00 amu) = 18.02 amu



Example

- What is the molar mass of the following compound ?
- NH_3 , CH_3COOH , Na_2SO_4 , $C_6H_{12}O_6$
- $NH_3 = (1\times14)+(3\times1) = 17$ g/mol
- $C_2H_4O_2 = (2\times12)+(4\times1)\ )+(2\times16) = 60$ g/mol
- $Na_2SO_4 = (2\times23)+(1\times32)\ )+(4\times16) = 142$ g/mol
- $C_6H_{12}O_6 = (6\times12)+(12\times1)\ )+(6\times16) = 180$ g/mol

EXAMPLE 3.5

- **Calculate the molecular masses (in amu) of the following compounds ?**
- Sulfur dioxide $\mathbf{SO_2}$ = 32.07+2 (16) = 64.07 amu

- Caffeine $\mathbf{C_8H_{10}N_4O_2}$

$$= \mathbf{8}(12.01) + \mathbf{10}\ (1.008) + \mathbf{4}(14.01) + 2(16) = 194.20\text{amu}$$

Practice exercise3.5

Calculate the molecular masses of methanol?

- **methanol $\mathbf{C\,H_4O}$**

$$= \mathbf{1}(12.01) + \mathbf{4}\ (1.008) + 1(16) = 32.4\ \text{amu}$$

n = number of moles

m = mass (atom or molecule)

M = molar mass (atomic mass or molecular mass)

What is the relation between them?

$$n = \frac{m}{M} = \frac{g}{g/mol} = mol$$

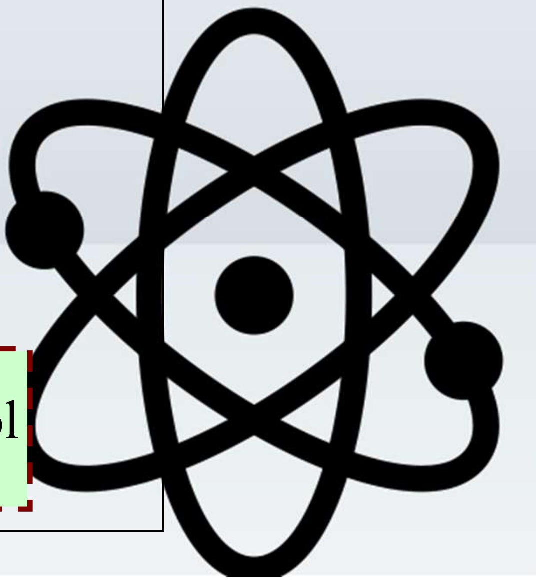
n = number of moles

N = number of atoms or molecules

N_A = Avogadro's number (atoms (or molecules)/mol)

What is the relation between them?

$$n = \frac{N}{N_A} = \frac{\text{atoms (or molecules)}}{\text{atoms (or molecules)/ mol}} = \text{mol}$$



EXAMPLE 3.2

How many moles of He atoms are in 6.46 g of He ?

$$n(He) = \frac{m}{M} = \frac{6.46g}{4.003g/mol} = 1.61mol$$

How many grams of Zn are in 0.356 mole of Zn?

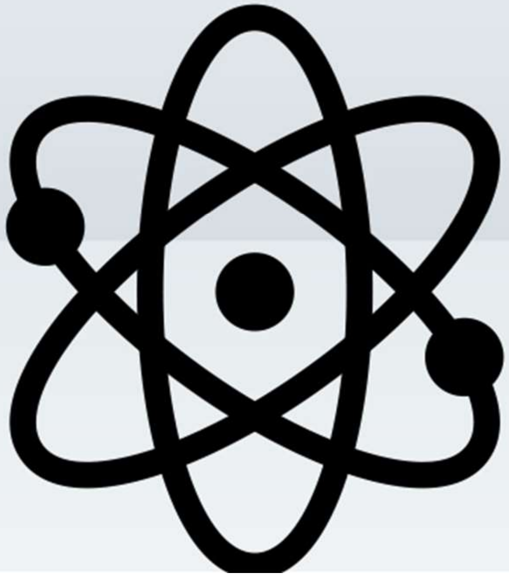
$$n(Zn) = \frac{m}{M} \Rightarrow m = nM$$

$$m = 0.356 \text{ mol } x 65.39 \text{ g/mol } = 23.3 \text{ g}$$

Example 3.6

- Methane is the principle component of natural gas . How many CH_4 are in 6.07 g of CH_4?

$$n(He) = \frac{m}{M} = \frac{6.06g}{16.04g/mol} = 0.378mol$$



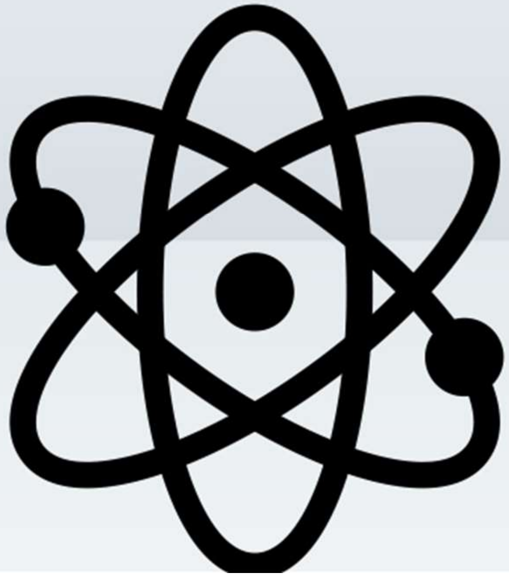
Example 3.4 p84:

How many S atoms are in 16.3 g of S?

Strategy:

1. How many moles in 16.3 g of S = X mol
2. 1 mole → 6.022×10^{23} S atoms

   X moles → ? atoms



Solution:

From the periodic Table: The atomic mass of S = 32.07 amu

The molar mass of S = 32.07 g/mol

Thus: 32.07 g → 1 mole of S

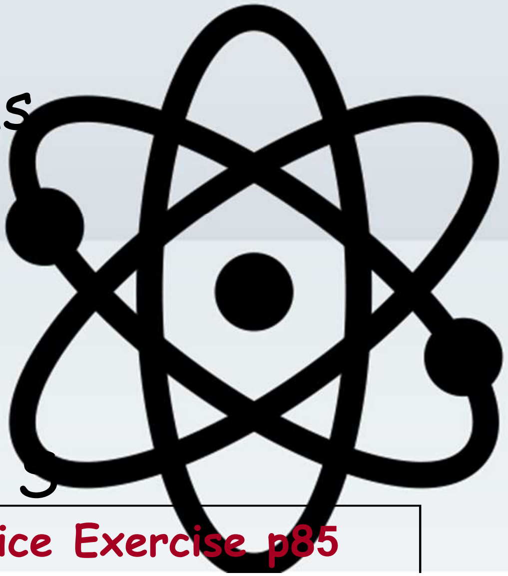
16.3 g of S → ? mole

$$n = \frac{1\,\text{mol x } 16.3\,\text{g}}{32.07\,\text{g}} = 0.508\,\text{mol}$$

We know: 1 mol of S → 6.022 x10^{23} S atoms

0.508 mole → ? S atoms

$$\text{number of S atoms} = \frac{6.022x10^{23}\text{ atoms x } 0.508\text{ mol}}{1\text{ mol}}$$
$$= 3.06x10^{23}\text{ S atoms}$$



There is 3.06 x10^{23} atoms of S in 16.3 g of S

H.W. Solve the Practice Exercise p85

حل مختصر

How many S atoms are in 16.3 g of S?

$$n(S) = \frac{m}{M} = \frac{16.3\ \text{g}}{32.07\ \text{g/mol}} = 0.508\ \text{mol}$$

$$n(S) = \frac{N}{N_A} \Rightarrow N = n x N_A$$

$$= 0.508\ \text{mol}\ x 6.022 x 10^{23}\ \text{atoms/mol}$$

$$= 3.06 x 10^{23}\ \text{atoms}$$

How many molecules of ethane (C_2H_6) are present in 0.334 g of C_2H_6?

(a) 2.01×10^{23}

(b) 6.69×10^{21}

(c) 4.96×10^{22}

(d) 8.89×10^{20}

$$\text{number of moles of } C_2H_6 = \frac{1 \text{ mole x } 0.334 \text{ g}}{30.068 \text{ g}} = 0.011 \text{ mol}$$

1 mole of C_2H_6 → 6.022×10^{23} molecules of C_2H_6

0.011 mole of C_2H_6 → ? molecules of C_2H_6

$$\text{number of molecules of } C_2H_6 = \frac{0.011 \text{ mol x } 6.022\text{x}10^{23} \text{ molecules}}{1 \text{ mole}}$$

$$= 6.624x10^{21} \text{ molecules}$$

Example 3.7 p87:

How many hydrogen atoms are present in 25.6 g of urea $[(NH_2)_2CO]$. The molar mass of urea is 60.06 g/mol.

$$n[(NH_2)_2CO] = \frac{m}{M} = \frac{25.6\,g}{60.06\,g/mol} = 0.426\,\text{mol}$$

$$n[(NH_2)_2CO] = \frac{N}{N_A}$$

$$\Rightarrow N = nxN_A = 0.426\,\text{mol}\,x6.022x10^{23}\ \text{molecules/mol}$$

$$N = 2.567x10^{23}\ \text{molecules}$$

$$1\,\text{molecule}\,[(\text{NH}_2)_2CO] \rightarrow 4\ \text{H atoms}$$

$$2.567\text{x}10^{23}\ [(\text{NH}_2)_2CO]\,\text{molecules} \rightarrow ?\text{H atoms}$$

$$\text{number of H atoms} = \frac{4\ \text{atomx}2.567\text{x}10^{23}\ \text{molecule}}{1\,\text{molecule}} = 1.03x10^{24}\ \text{atoms}$$

H.W. What is the mass, in grams, of one copper atom?

(a) 1.055×10^{-22} g

(b) 63.55 g

(c) 1 amu

(d) 1.66×10^{-24} g

(e) 9.476×10^{21} g

Atomic mass of Cu = 63.55 amu

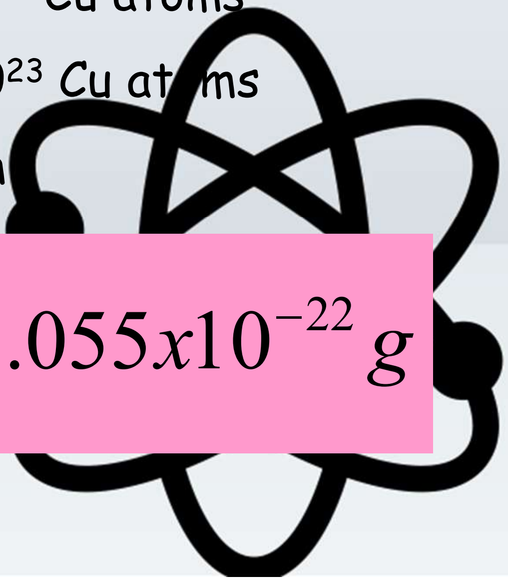
Molar mass of Cu = 63.55 g/mol

63.55 g of Cu → 1 mol of Cu

1 mol of Cu → 6.022×10^{23} Cu atoms

63.55g of Cu → 6.022×10^{23} Cu atoms

?g of Cu → 1 Cu atom



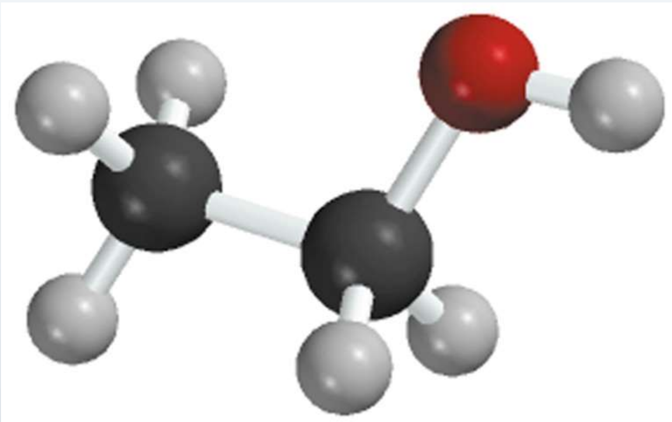
$$\text{grams of Cu} = \frac{1\ \text{atom x } 63.55\text{g}}{6.022\text{x}10^{23}\ \text{atom}} = 1.055x10^{-22}\ g$$

3.5 Percent composition of compounds

Percent composition of an element in compound =

$$\frac{n \times \text{molar mass of element}}{\text{molar mass of compound}} \times 100\%$$

n is the number of moles of the element in 1 mole of the compound



C_2H_6O

$$\%C = \frac{2 \times (12.01\text{ g})}{46.07\text{ g}} \times 100\% = 52.14\%$$

%H = $\frac{6 \times (1.008\text{ g})}{46.07\text{ g}}$ x 100% = 1[illegible].1[illegible]%

%O = $\frac{1 \times (16.00\text{ g})}{46.07\text{ g}}$ x [illegible]00% [illegible] [illegible]4.73%

Check the answer!

2.14% + 13.13% + 34.73% = 100.0%

Example 3.8

- Calculate the percent composition by mass of H , P, and O in H_3PO_4 acid ?

- Molar mass of H_3PO_4
- $= 3(1.008) + 1\,(30.97) + 4(16)$

$= 97.99$ amu

$$\%H = \frac{3(1.008)}{97.99} \times 100\% = 3.0864\%$$

$$\%P = \frac{1(30.97)}{97.99} \times 100\% = 31.61\ \%$$

$$\%O = \frac{4(16)}{97.99} \times 100\% = 65.31\%$$

PRACTIES EXERICISE 3.8

- Calculate the percent composition by mass of H , P, and O in H_2SO_4 acid ?

- Molar mass of H_2SO_4
- $= 2(1.008) + 1\,(32.7) + 4(16)$

$= 98.72$ amu

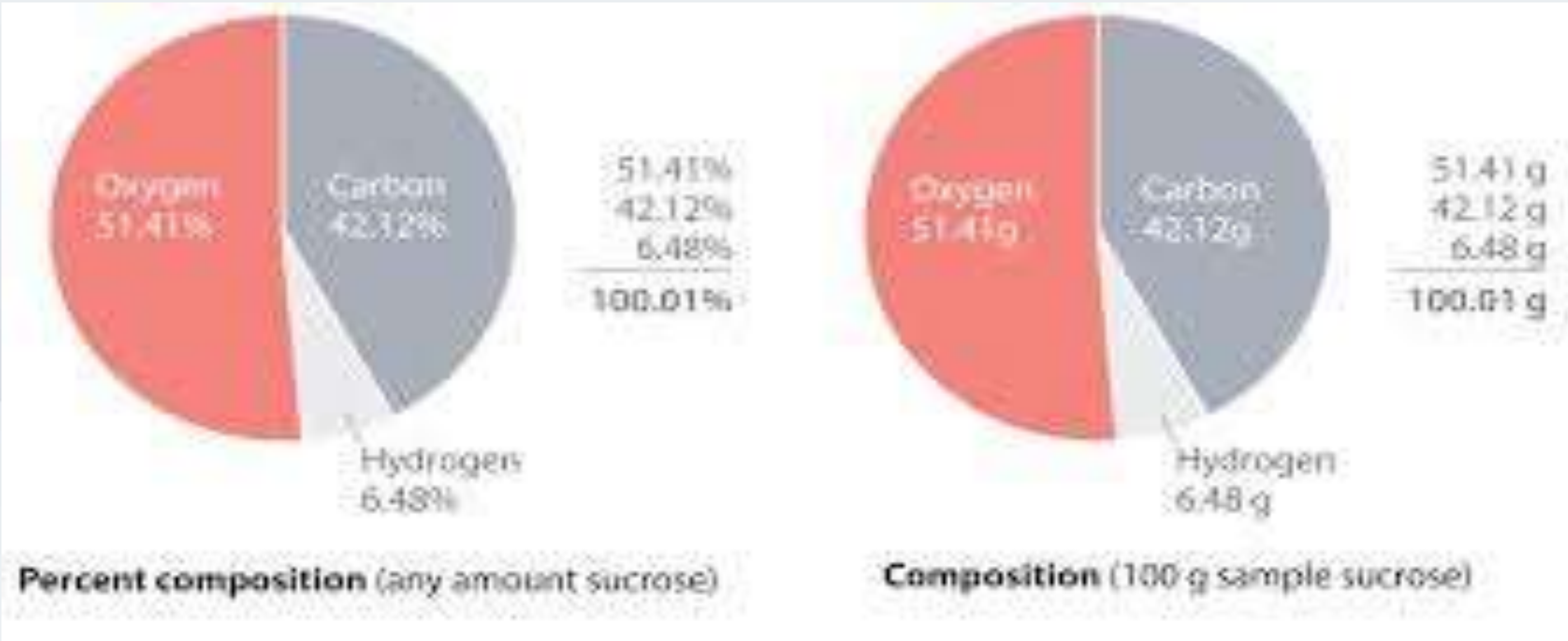
$$\%H = \frac{2(1.008)}{98.72} \times 100\% = 2.026\ \%$$

$$\%S = \frac{1(32.07)}{98.72} \times 100\% = 32.486\ \%$$

$$\%O = \frac{4(16)}{98.72} \times 100\% = 64.83\%$$

$$\%\ \text{Mass} = \frac{\text{mass of 1 element}}{\text{molar mass of compound}} \times 100$$


K2CrO4
32.9% O
40.3% K
26.8% Cr
K
Cr
O
K2Cr2O7
38.1% O
26.5% K
35.4% Cr
Potassium chromate
40.3% K 26.8% Cr 32.9% O
Potassium dichromate
26.5% K 35.4% Cr 38.1% O


Oxygen
51.41%
Carbon
42.12%
Hydrogen
6.48%
51.41%
42.12%
6.48%
100.01%
Percent composition (any amount sucrose)
Oxygen
51.41g
Carbon
42.12g
Hydrogen
6.48 g
51.41 g
42.12 g
6.48 g
100.01 g
Composition (100 g sample sucrose)

H.W. Calculate the percent of nitrogen in $Ca(NO_3)_2$:

a) 12.01%.
b) 17.10%.
c) 18%
d) 16%.

H.W. All of the substances listed below are fertilizers that contribute nitrogen to the soil. **Which of these is the richest source of nitrogen on a mass percentage basis?**

(a) Urea, $(NH_2)_2CO$
(b) Ammonium nitrate, NH_4NO_3
(c) Guanidine, $HNC(NH_2)_2$
(d) Ammonia, NH_3

(a) %N = 46.6%
(b) %N = 58%
(c) %N = 71.1%
(d) %N = 82.2%

ايا من هذه المواد هواغنى مصدر للنيتروجين على اساس احتوائه على اكبر نسبه وزنيه من النيتروجين؟

Percent Composition and Empirical Formulas

Q: Determine the empirical formula of a compound that has the following percent composition by mass: K 24.75, Mn 34.77, O 40.51 percent.

	K	Mn	O
% →100g	24.75g	34.77g	40.51g
n=m/M	24.75/39.10 =0.633mol	34.77/54.94 =0.6329mol	40.51/16.00 = 2.532mol
÷ on smallest no. of mole	0.633/0.632 =1	0.6329/0.632 = 1	2.532/0.632 =4
The empirical formula is	K_1	Mn_1	O_4
	$KMnO_4$		

خطوات الحل

1. ننشأ جدول نضع فيه العناصر المذكورة في السؤال
2. نعتبر أن النسبة المئوية معبر عنها بالجرام فلو كان عندنا 100 جرام من المركب فهذه ال 100 جرام موزعة على العناصر حسب نسبتها.
3. نوجد عدد المولات n لكل عنصر باستخدام القانون n=m/M .
4. نقسم عدد المولات على أصغر مول من العناصر.
5. الأرقام التي نحصل عليها تمثل empirical formula بشرط أن تكون أعداد صحيحة كما في المثال السابق.
6. في حالة ظهور أعداد عشرية نقوم بضرب الأرقام التي في الأسفل الموجودة في الصيغة بأعداد بدأ من 2، 3....... حتى نحصل على أعداد صحيحة .

Example 3.9 p90:

Ascorbic acid composed of 40.92% C, 4.58% H, and 54.50% O by mass. Determine its empirical formula.

	C	H	O
% →100g	40.92g	4.58g	54.50g
n=m/M	40.92/12.01 = 3.407mol	4.58/1.008 =4.54.mol	54.50/16.00 = **3.406 mol**
÷ on smallest no. of mole	3.407/3.406 = 1	0.4.54/3.406 = 1.33	3.406/3.406 = 1
Convert into integer x 3	3	3.99 = 4	3
The empirical formula is	C_3	H_4	O_3
	$C_3H_4O_3$		

خطوات الحل

1. ننشأ جدول نضع فيه العناصر المذكورة في السؤال
2. نعتبر أن النسبة المئوية معبر عنها بالجرام فلو كان عندنا 100 جرام من المركب فهذه ال 100 جرام موزعة على العناصر حسب نسبتها.
3. نوجد عدد المولات n لكل عنصر باستخدام القانون n=m/M .
4. نقسم عدد المولات على أصغر مول من العناصر.
5. الأرقام التي نحصل عليها تمثل empirical formula بشرط أن تكون أعداد صحيحة
6. في حالة ظهور أعداد عشرية كما في المثال السابق نقوم بضرب الأرقام التي في الأسفل الموجودة في الصيغة بأعداد بدأ من 2، 3....... حتى نحصل على أعداد صحيحة .

Determination of the Molecular Formula from the Percent Composition by Mass

Example 3.11 p93:

A sample compound contains 1.52g of N and 3.47g of O. The molar mass of this compound is between 92g . Determine the molecular formula.

Solution:

1.
$$n_N = \frac{1.52}{14.01} = 0.108 \text{ mol of N}$$
$$n_O = \frac{3.47}{16.00} = 0.217 \text{ mol of O}$$

2.
$$N : \frac{0.108}{0.108} = 1 \qquad O = \frac{0.217}{0.108} = 2$$

3. Thus the empirical formula is: NO_2

Present Composition by Mass

↓

Empirical Formula

↓

Molecular Formula

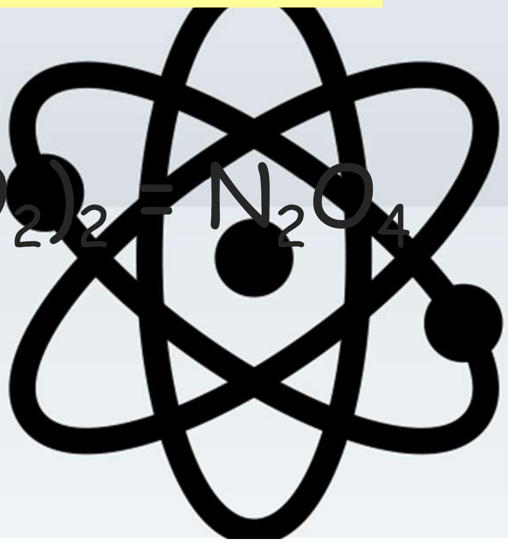
4. The molar mass of the empirical formula NO_2 = 14.01 + (2x16.00) = 46.01g

5. The ratio between the empirical formula and the molecular formula:

$$\text{Ratio} = \frac{\text{molar mass of compound}}{\text{empirical molar mass}}$$

$$\text{Ratio} = \frac{90}{46.01} \overset{1.956}{\approx} 2$$

6. The molecular formula is $(NO_2)_2 = N_2O_4$



PRACTIES EXERICISE 3.10

- **A sample of a compound containing born (B) and hydrogen (H) contains 6.444g of B and 1.803 g of (H). The molar mass of the compound is about 30g. What is its molecular formula?**

$$n_B = 6.444 \text{ g B} \times \frac{1 \text{ mol B}}{10.81 \text{ g B}} = 0.5961 \text{ mol B}$$

$$\text{B}: \frac{0.5961}{\mathbf{0.5961}} = 1.0$$

$$n_H = 1.803\text{g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 1.7888 \text{ mol H}$$

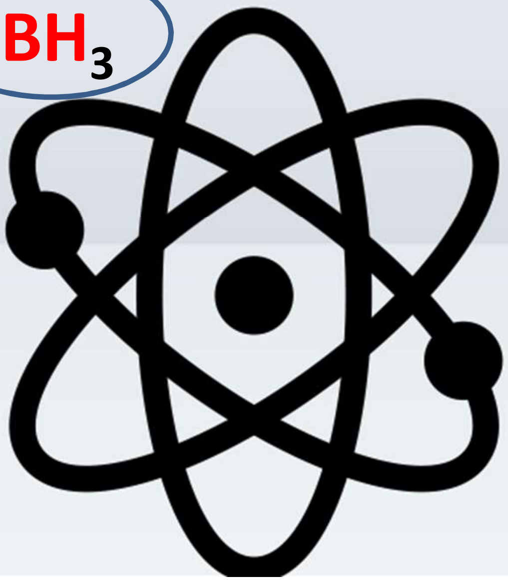
$$\text{H}: \frac{\mathbf{1.7888}}{\mathbf{0.5961}} = 3$$

BH_3

Molar mass of empirical formula = 10.81 + 3 x 1.008 = 13.834g

The ratio between molar mass and the molar mass of empirical formula
= molar mass / empirical formula = 30 g / 13.834 g ≈ 2

B_2H_6



3.7Chemical reactions and chemical equations

- A process in which one or more substances is changed into one or more new substances is a **chemical reaction**
- A ***chemical equation*** uses chemical symbols to show what happens during a chemical reaction

3 ways of representing the reaction of H_2 with O_2 to form H_2O

Two hydrogen molecules + One oxygen molecule ⟶ Two water molecules

$$2H_2 + O_2 \longrightarrow 2H_2O$$

reactants ⟶ products

Dr. Laila Al-Harbi

How to “Read” Chemical Equations

$2\,Mg + O_2$ $2\,MgO$

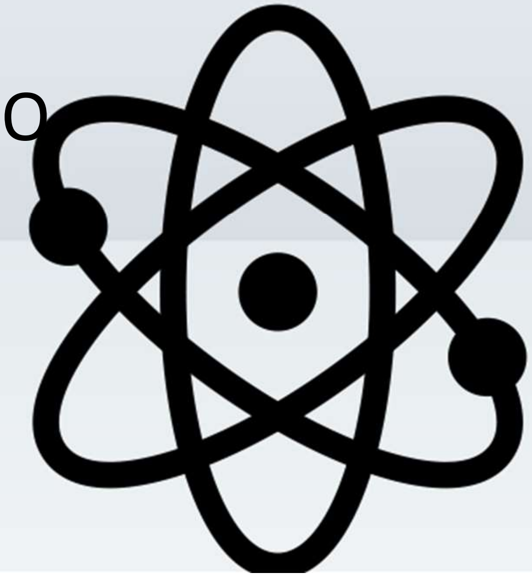
2 atoms Mg + 1 molecule O_2 makes 2 formula units MgO

2 moles Mg + 1 mole O_2 makes 2 moles MgO

48.6 grams Mg + 32.0 grams O_2 makes 80.6 g MgO

IS NOT

2 grams Mg + 1 gram O_2 makes 2 g MgO

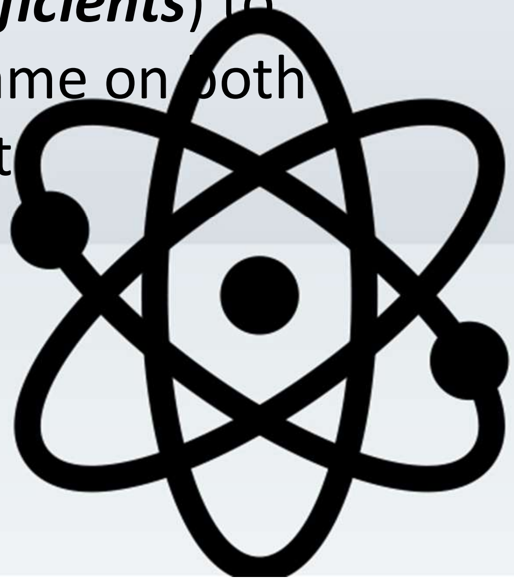


Balancing Chemical Equations

- 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**

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

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 subscript



$2C_2H_6$ NOT C_4H_{12}

طريقة وزن المعادلة الكيمائية

- نكتب الصيغه الصحيحه لكل متفاعل (على الطرف الايسر) ولكل ناتج (على الطرف الايمن)
- وزن المعادله الكيميائيه يكون بتغير الارقام التي بجانب الصيغه وليست التي تحتها بحيث يكون للعنصر نفس العدد على طرفي المعادله.
- توزن اولا العناصر الاقل ظهورا, ثم توزن العناصر الاكثر ظهورا
- الخطوه الاخيره هي التأكد من ان لديك نفس العدد من الذرات لكل عنصر على طرفي المعادله

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

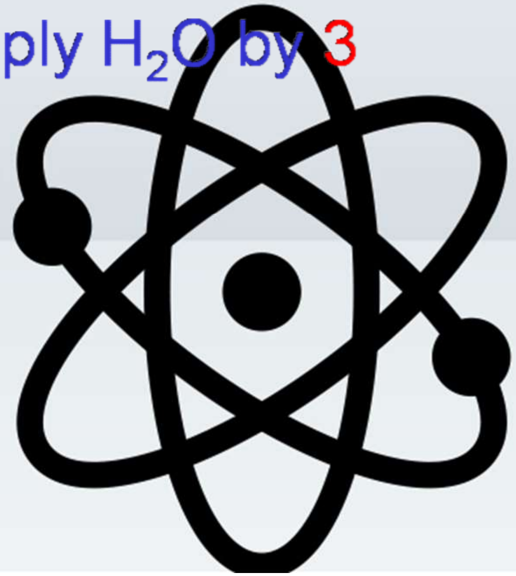
$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$ start with C or H but not O

1 carbon on right (↑ CO_2) multiply CO_2 by 2

$C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$

6 hydrogen on left (↑ C_2H_6) 2 hydrogen on right (↑ H_2O) multiply H_2O by 3

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



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

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

2 oxygen on left

4 oxygen (2x2) + 3 oxygen (3x1) = 7 oxygen on right

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

remove fraction
multiply both sides by 2

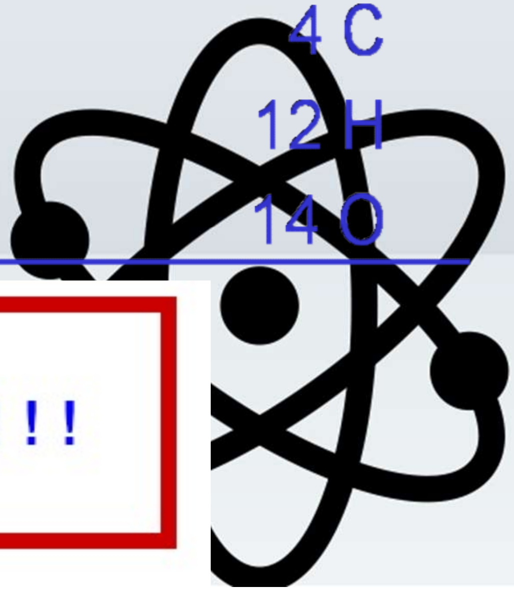
$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

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

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

4 C (2 x 2)	4 C
12 H (2 x 6)	12 H (6 x 2)
14 O (7 x 2)	14 O (4 x 2 + 6)

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



Never Work with Unbalanced Equations ! ! ! !

Example 3.12

$$Al + O_2 \rightarrow Al_2O_3$$

$$2Al + O_2 \rightarrow Al_2O_3$$

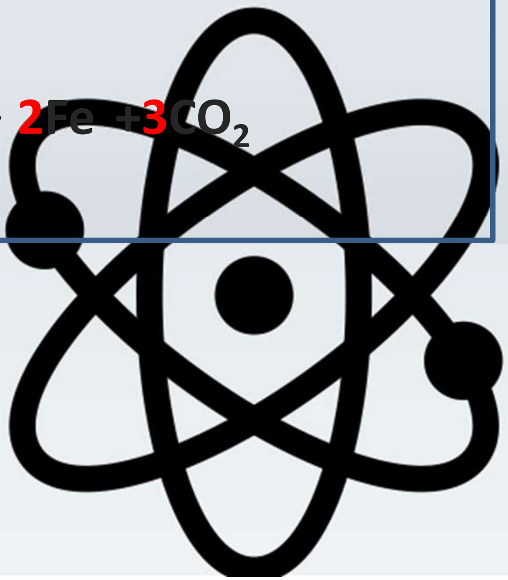
$$2Al + 3/2O_2 \rightarrow Al_2O_3$$

$$2(2Al + 3/2O_2 \rightarrow Al_2O_3)$$

$$4Al + 3O_2 \rightarrow 2Al_2O_3$$

PRACTIES EXERICISE 3.12

- $Fe_2O_3 + CO \rightarrow Fe + CO_2$
- $Fe_2O_3 + CO \rightarrow 2Fe + CO_2$
- $Fe_2O_3 + 1/3CO \rightarrow 2Fe + 1/3CO_2$
- $3(Fe_2O_3 + 1/3CO \rightarrow 2Fe + 1/3CO_2)$
- **$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$**



H.W. What is the coefficient of H_2O when the equation is balanced:

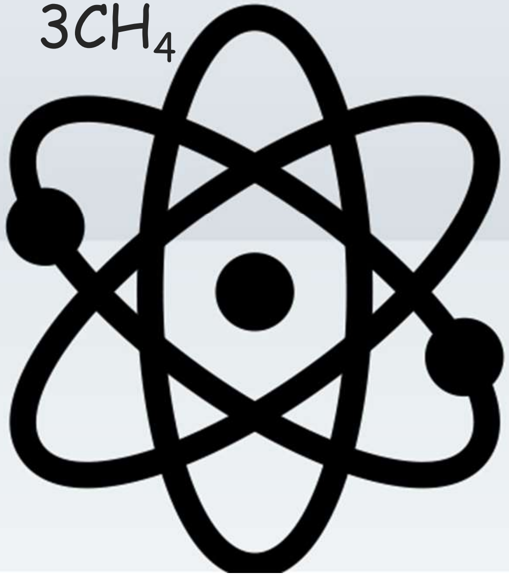
$$\_\ Al_4C_3 + \_\ H_2O \rightarrow \_\ Al(OH)_3 + 3CH_4$$

a. 13
b. 4
c. 6
d. 12

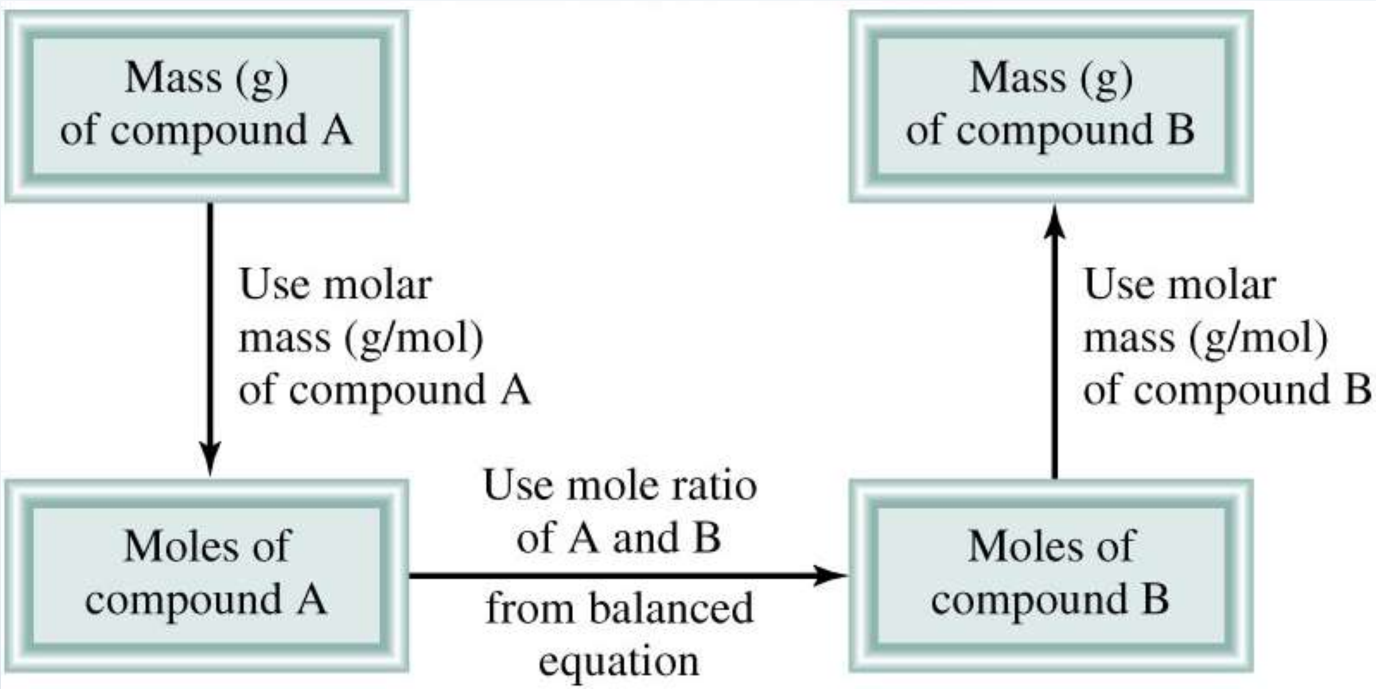
H.W. What are the coefficients of Al_4C_3, H_2O and $Al(OH)_3$, respectively, when the equation is balanced:

$$\_\ Al_4C_3 + \_\ H_2O \rightarrow \_\ Al(OH)_3 + 3CH_4$$

a. 4,1,5
b. 1,12,4
c. 1,24, 4
d. 4,12,1



Amounts of Reactants and Products



1. Write balanced chemical equation
2. Convert quantities of known substances into moles
3. Use coefficients in balanced equation to calculate the number of moles of the sought quantity
4. Convert moles of sought quantity into desired units

- لمعرفة المواد المتفاعلة أو الناتجة بمعلومية أحد المتفاعلات او الناتجة نقوم بالاتي:
- لابد أن تكون المعادلة موزونه
- حددي المادة المعطاه givenثم الماده المطلوبةRequired و اعملي علاقة بينهم و تجاهلي الباقي تماما
- العلاقة في المعادله الموزونه علاقة مولات
- فلوكانت الماده المعطاه بالجرامات نحولها الى مولات و اذاكانت بالمولات لا نحتاج الى هذه الخطوة
- اذا كانت الماده المطلوبة بالمولات نكتفى بهذا الحد
- اذا كانت الماده المطلوبة بالجرامات نحول المولات الى جرامات.

If **2 mol of** $C_6H_{12}O_6$ is burned , what is the number of moles **of** CO_2 produced?

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

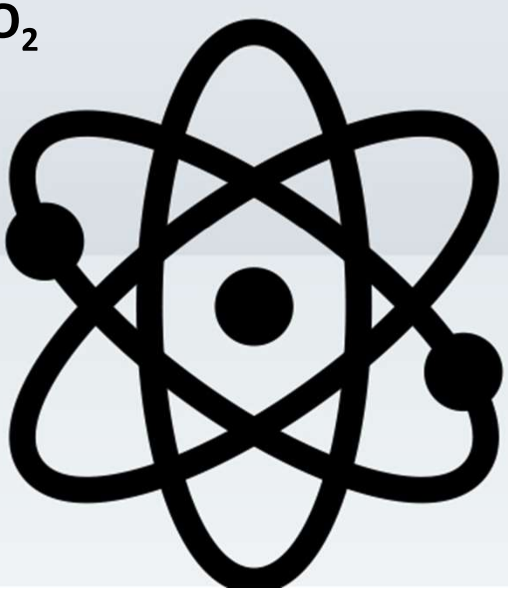
Given

Required

From the equation mole of $C_6H_{12}O_6$ → produce 6 mol CO_2

From the equation 2 mol $C_6H_{12}O_6$ → x mol CO_2

the number of moles of CO_2 produced = 2 × 6/ 1=12 mol



If 2 mol of $C_6H_{12}O_6$ is burned , what is the mass of CO_2 produced?

- If 2 mol of $C_6H_{12}O_6$ is burned , how many grams of CO_2 produced?

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

Given — Required

From the equation mole of $C_6H_{12}O_6$ → produce 6 mol CO_2

From the equation 2 mol $C_6H_{12}O_6$ → x mol CO_2

the number of moles of CO_2 produced = 2 × 6/ 1=12 mol

the mass of CO_2 produced = n × molecular mass of CO_2

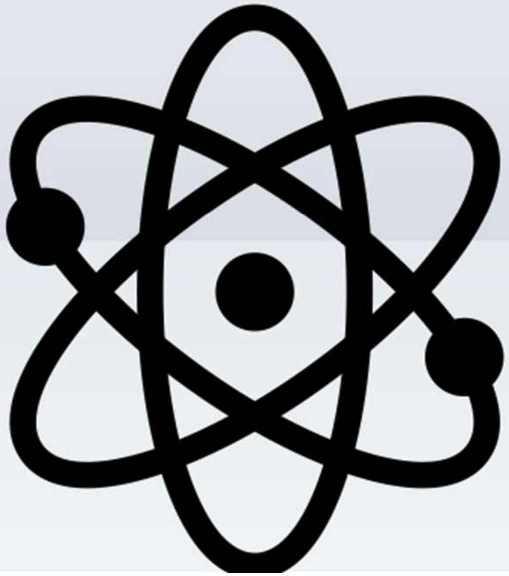
the mass of CO_2 produced = 12 ×44.01 = 528.12 g

Example 3.13

- **A general over all equation for this very complex process represents the degradation of glucose ($C_6H_{12}O_6$) to CO_2 and water. If 856 g of $C_6H_{12}O_6$ is consumed by person over a certain period, what is the mass of CO_2 produced?**

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6\ H_2O + 6CO_2$$

- **n = m/M = 856/180.2 = 4.75 mol**
- **From the equation mole of $C_6H_{12}O_6$ → produce 6CO_2**
- **From the equation 4.75 mol $C_6H_{12}O_6$ → x CO_2**
- **From the equation = 4.75 × 6 / 1= 28.5**
- **the mass of CO_2 produced = n × M**
- **the mass of CO_2 produced = 28.5 × 44.01 = 1254.35 g**



PRACTIES EXERICISE 3.13

Methanol burns in air according to the equation

If 209 g of methanol are used up in the combustion , what mass of water is produced?

$$2CH_3OH + 3O_2 \longrightarrow 2CO_2 + 4H_2O$$

Given ($2CH_3OH$) — Required ($4H_2O$)

n = m/M = 209/32= 6.53 mol

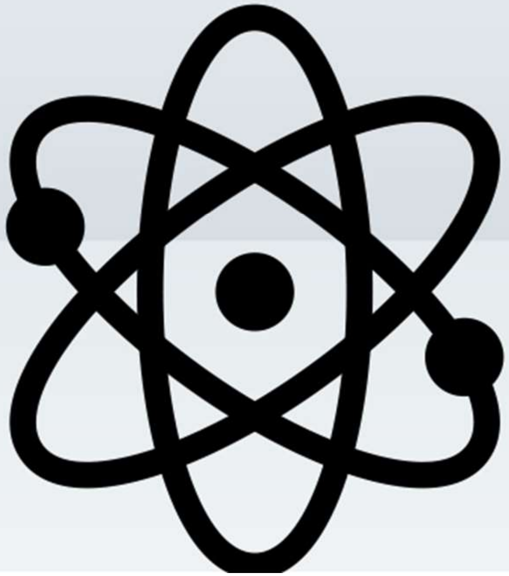
From the equation 2 moles of CH_3OH **→ produce 4 mol** H_2O

From the equation 6.53 mol CH_3OH **→ x mol** H_2O

the number of moles of H_2O**produced = 6.53 × 4/ 2=13.06 mol**

the mass of H_2O produced = n × **molecular mass of** CO_2

the mass of H_2O produced = 13.06 **×18 = 235g**



Example 3.14

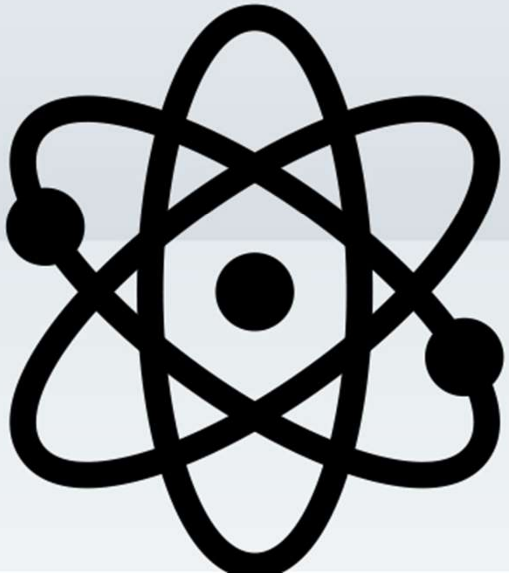
- **All alkali metals react with water to produce hydrogen gas and the corresponding alkali metal hydroxide. A typical reaction is that between lithium and water**

$$Li\ (s) + 2\ H_2O\ (l) \rightarrow 2\ Li\ OH\ (aq) + H_2\ (g)$$

How many grams of Li are needed to produce 9.89g of H_2 ?

- **From the equation 2 mole of Li → produce mole of H_2**
- **From the equation 2× 6.941 g Li → 2.016g H_2**
- **From the equation x g Li → 9.89 g CO_2**
- **the mass of CO_2 produced = 2× 6.941× 9.89 g / 2.016g**

= 68.1g Li



3.9 Limiting Reagent الكاشف المحدد

- **Limiting Reagent:** is the reactant used up first in a reaction and thus determine the amount of product
- **Excess Reagent** الكاشف الفائض: is the reactant present in quantities greater than necessary to react with the quantity of the limiting reagent (the one that is left at the end of the reaction).
- → Limiting reagent is in a reaction of more than one reactant!

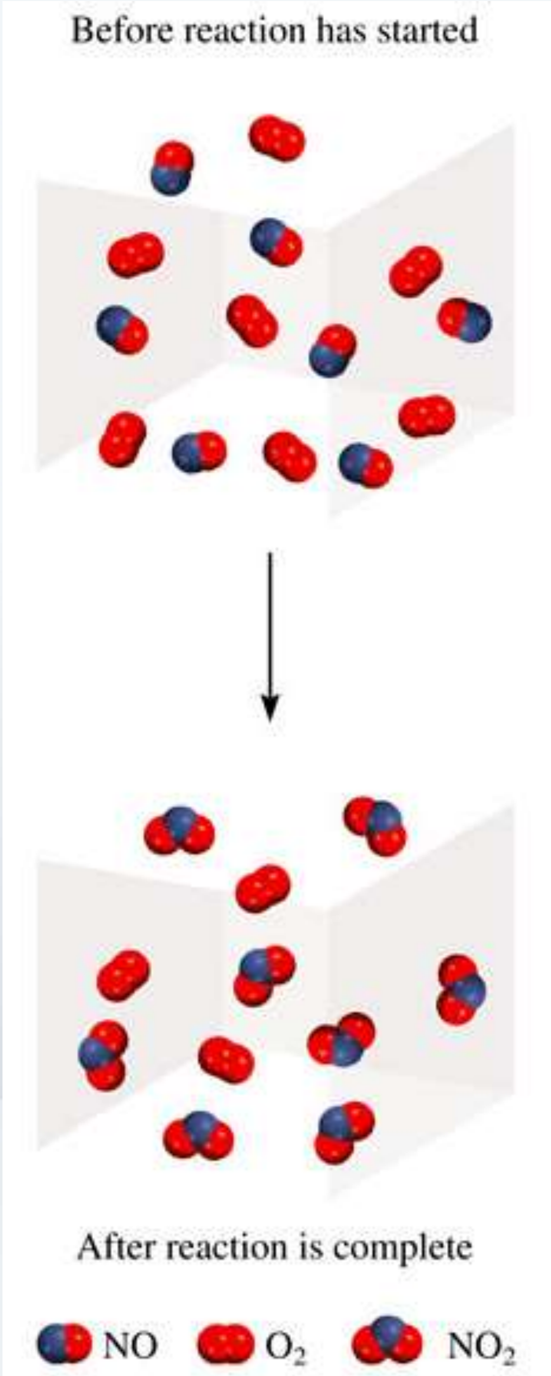
Limiting Reagent:

Reactant used up first in the reaction.

$$2NO + O_2 \longrightarrow 2NO_2$$

NO is the limiting reagent

O_2 is the excess reagent



Limiting Reactant

- الكاشف المحدد هو الكاشف الذي يحدد كمية الماده الناتجه
- لا يشترط أن يكون الكاشف المحدد هو نفسه كل مره
- دائما الكاشف المحدد موجود بعدد مولات أقل و الماده الاخرى موجوده بزياده
- مسألة الكاشف المحدد تختلف عن المسائل السابقة أنه يعطيك كلا المتفاعلين و يطلب الناتج
- لكل نحدد الكاشف المحدد نقوم بالخطوات التالية
- 1- نحول جرامات المواد المتفاعلات الى مولات
- 2- نقسم الجرامات الناتجة على معامل الماده في المعادلة الموزونه
- 3- الماده أقل عدد مولات هي الكاشف المحدد
- 4- نوجد الماده الناتجه حسب ما تعلمنا في الجزء السابق

Example:

- When 22.0 g NaCl and 21.0 g H_2SO_4 are mixed and react according to the equation below, which is the limiting reagent?

$$2NaCl + 1H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$$

- n NaCl = 22/58.5 = 0.376/2=0.188 mol

- n H_2SO_4 = 21/98 = 0.214/1 = 0.241 mol

- n NaCl (0.188 mol)<n H_2SO_4 (0.241 mol)

- So NaCl is the limiting reagent

Example

Consider the combustion of carbon monoxide (CO) in oxygen gas: $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

Starting with 3.60 moles of CO and 4 moles of O_2, calculate the number of moles of CO_2 produced ?

- n CO = 3.6/2=1.88 mol
- n O_2 = 4/1 = 4 mol
- n CO (1.88 mol)<n O_2 (4 mol)
- So CO is the limiting reagent

From chemical eq. 2 mole CO = 2 mol CO_2

3.6 mol = x

number of moles of CO_2 produced = 3.6 x 2/2 = 3.6 mol

Example

- 10.0g of aluminum reacts with 35.0 grams of chlorine gas to produce aluminum chloride. Which reactant is limiting, which is in excess, and how much product is produced?

$$2\ Al + 3\ Cl_2 \rightarrow 2\ AlCl_3$$

- n Al = 10/27=0.37 /2mol= 0.185 mol
- n Cl_2 = 35/ 71= 0.493 /3 = 0.164 mol
- n Cl_2 (0.185 mol)<n Al (0.164 mol)
- So Cl_2 is the limiting reagent

From chemical eq. 3 mole Cl_2 = 2 mol $AlCl_3$

0.493 mol = x

number of moles of $AlCl_3$ produced = 0.493 x 2/3 = 0.329 mol

mass of **$AlCl_3$ = 0.329 x 133.5 = 43.877 g**

Science Cl_2 is the LR so Al is the excess the amount remain

From chemical eq. 3 mole Cl_2 = 2 mol Al

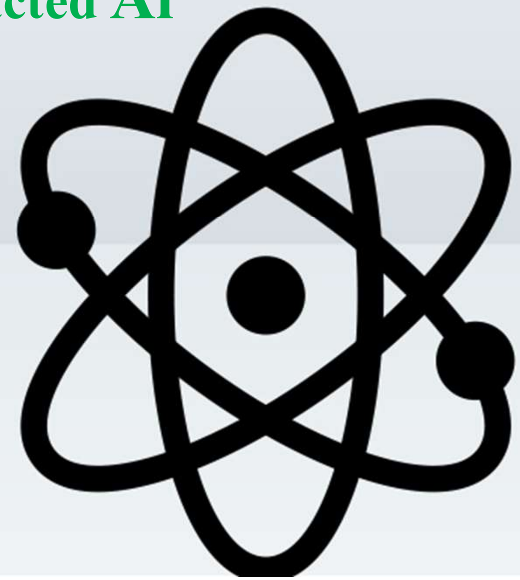
0.493 mol = x

number of moles of Al react = 0.493 x 2/3 = 0.329 mol

mass of Al = 0.329 x 27 = 8.8 83 g

The excess mass of Al = total mass Al – reacted Al

= 10 -8.883 = 1.117 g



- **Urea $(NH_2)_2CO$ is prepared by reacting ammonia with carbon dioxide**

$$2NH_3\ (g) + CO_2\ (g) \rightarrow (NH_2)_2CO\ (aq) + H2O\ (\iota)$$

- **In on process 637.2 g of NH_3 are treated with 1142 g of CO_2 a) which of the two limiting reagents? b) calculate the mass of $(NH_2)_2CO$ formed ? C) how much excess reagent (in gram) is left at the end of the reaction**
- n **NH_3** = 637.2 /17=37.482 /2mol= 18.74 mol
- n **CO_2 = 1142/ 44= 25.95 /1 = 25.95 mol**
- n **NH_3** (18.74 mol)<n **CO_2 (25.95 mol)**
- **So NH_3** is the limiting reagent

From chemical eq. 2 mole NH_3 = 1 mol $(NH_2)_2CO$

37.482 mol = x

number of moles of $AlCl_3$ produced = 37.82 x 1/2 = 18.94 mol

mass of **$AlCl_3$ =18.94 x 60.06 = 1125.6 g**

- C) how much excess reagent (in gram) is left at the end of the reaction

Science NH_3 is the LR so CO_2 is the excess the amount remain

From chemical eq. 2 mole NH_3 = 1 mol CO_2

37.482 mol = x

number of moles of CO_2 react = 0.493 x 1/2 = 18.74 mol

mass of CO_2 = 18.74 x 44 = 824.56 g

The excess mass of CO_2 = total mass CO_2 – reacted CO_2

= 1142 -823.4 = 318.6 g

PRACTIES EXERICISE 3.15

In one process, 124 g of Al are reacted with 601 g of Fe_2O_3

$$2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2Fe$$

Calculate the mass of Al_2O_3 formed.

- n Al = 124/27=4.59 /2mol= 2.296 mol
- n Fe_2O_3 = **601/ 159= 3.78 /1 = 3.78 mol**
- n Al $<$n Fe_2O_3 >>>> **So** Al is the limiting reagent

From chemical eq. 2 mole Al **= 2 mol Al_2O_3**

4.59 mol = x

number of moles of $AlCl_3$ produced = 4.59 x 1/2 = 2.296 mol

mass of $AlCl_3$ **= 2.296 x 102 = 234** g

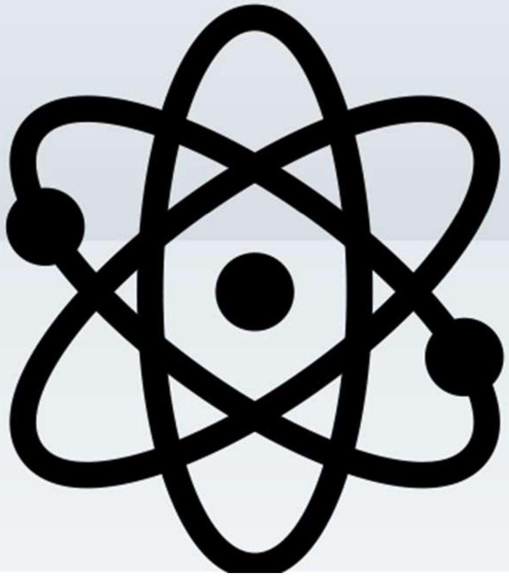
- n Al = 124/27=4.59 /2mol= 2.296 mol
- n Fe_2O_3 = **601/ 159= 3.78 /1 = 3.78 mol**
- n Al $<$n Fe_2O_3 >>>> **So** Al is the limiting reagent

From chemical eq. 2 mole Al **= 2 mol** Al_2O_3

4.59 mol = x

number of moles of $AlCl_3$ produced = 4.59 x 1/2 = 2.296 mol

mass of **$AlCl_3$ = 2.296 x 102 = 234 g**



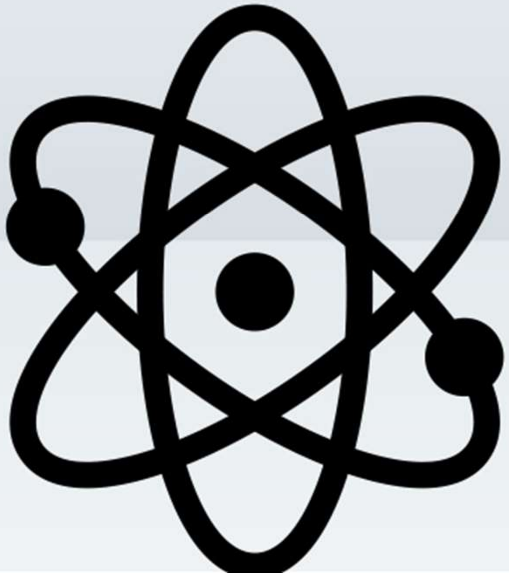
3.10 Reaction Yield

Theoretical Yield is the amount of product that would result if all the limiting reagent reacted.

Actual Yield is the amount of product actually obtained from a reaction.

$$\% \text{ Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

Actual Yield is always lees .



- When 22.0 g NaCl mixed with excess H_2SO_4 and 8.95 g HCl is formed .what is the %yield of HCl?

$$2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$$

n NaCl = 22/58.5 = 0.376/2=0.188 mol

From chemical eq. 2 mole NaCl = 2 mol HCl

0.376 mol = x

number of moles of HCl produced = 0.376 x 2/2 =0.376 mol

mass of HCl produced = n x MM(HCl) = 0.376 x 36.5 =13.724 g

- **%yield of HCl = practical/theoriotical x 100**

- **%yield of HCl = 8.95 /13.724 x 100 = 65.21%**

- When 22.0 g NaCl and 21.0 g H_2SO_4 are mixed and react according to the equation below 8.95 g HCl is formed .what is the %yield of HCl?

$$2NaCl + 1H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$$

- n NaCl = 22/58.5 = 0.376/2=0.188 mol
- n **H_2SO_4 = 21/98 = 0.214/1 = 0.241 mol**
- n NaCl (0.188 mol)<n **H_2SO_4 (0.241 mol)**
- **So** NaCl is the limiting reagent

From chemical eq. 2 mole NaCl = 2 mol HCl

0.376 mol = x

number of moles of HCl produced = 0.376 x 2/2 =0.376 mol

mass of HCl produced = n x MM(HCl) = 0.376 x 36.5 =13.724 g

- **%yield of HCl = practical/theoriotical x 100**
- **%yield of HCl = 8.95 /13.724 x 100 = 65.21%**

- **In one process, 3.54×10^7 g of $TiCl_4$ are reacted with 1.13×10^7 g of Mg a) Calculate the theoretical yield of the Ti? b) calculate the percent yield if 7.91×10^6 g of Ti are obtained ?**

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

n $TiCl_4$ = 3.54×10^7 g /189.68 = $1.87\text{x}10^5$ /1= $1.87\text{x}10^5$ mol

- n Mg = 1.13×10^7 g / **24.3** = $1.87\text{x}10^5$ /**2** = **2.32** x10^5**mol**

- n $TiCl_4$ <n Mg ,**So $TiCl_4$** is the limiting reagent

From chemical eq. 1 mole $TiCl_4$ = 1 mol Ti

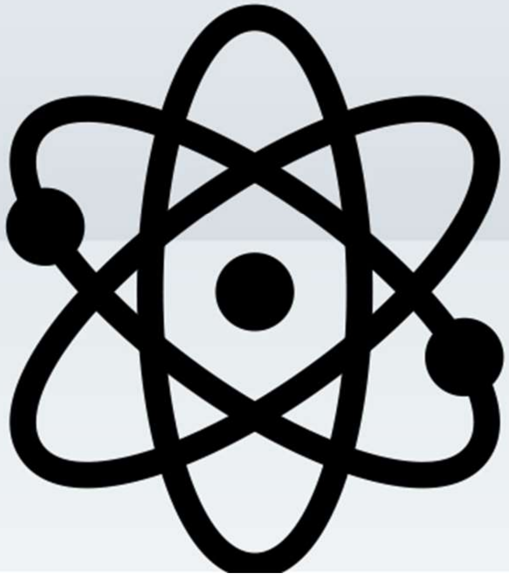
$1.87\text{x}10^5$ **mol = x**

number of moles of Ti produced = $1.87\text{x}10^5$ **x 1/1 =** $1.87\text{x}10^5$ **mol**

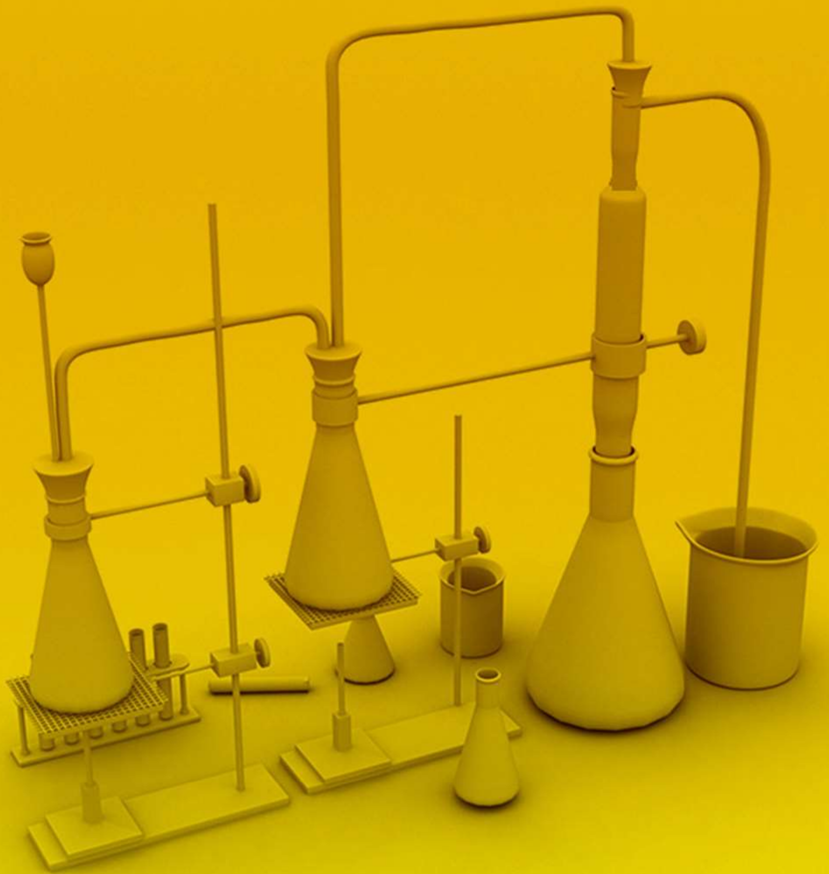
mass of Ti produced = n x MM(HCl) = $1.87\text{x}10^5$ **x 47.88 =** $8.93\text{x}10^9$ **g**

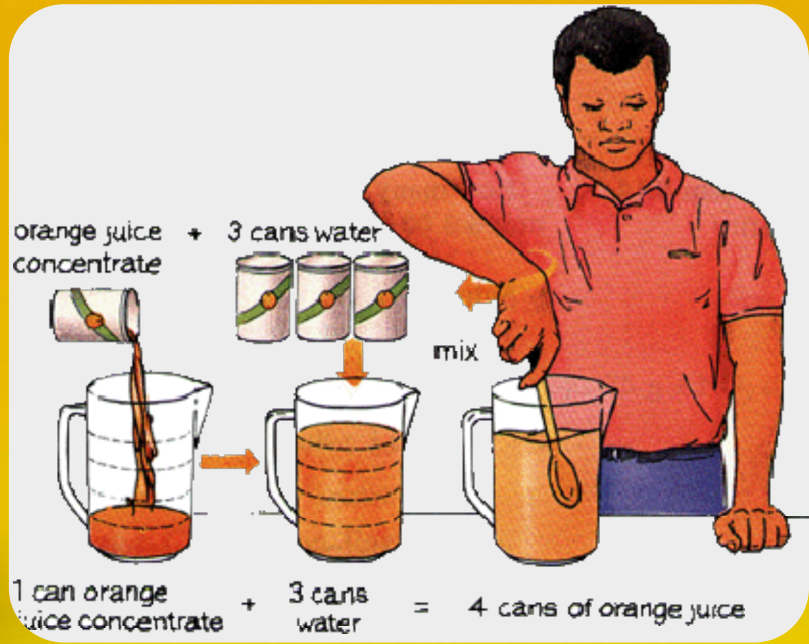
- **%yield of Ti = practical/theoriotical x 100**
- **%yield of HCl = 7.91×10^6** / $8.93\text{x}10^9$ **x 100 = 88.52%**

أشكر الزميلة د.نهى وزان لاني أستعنت بكثير من محاضراتها



Chapter 4
reaction in aqueous solution





4.5 Concentration of solutions

- dilution of solutions

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4.5 Concentration of solutions

✓ The concentration of a solution is the amount of solute present in a given amount of solvent , it can be expressed in terms of its molarity (molar concentration)

$$\textbf{Molarity } (\mathbf{\mathit{M}}) = \frac{\textbf{moles of solute}}{\textbf{volume of solution in liters}}$$

- Have mol and vol → molarity
- Have molarity and vol → mol of solute
- Have molarity and mol of solute → volume
- <u>AND:</u> mol of solute → grams of solute

Questions in Molarity

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

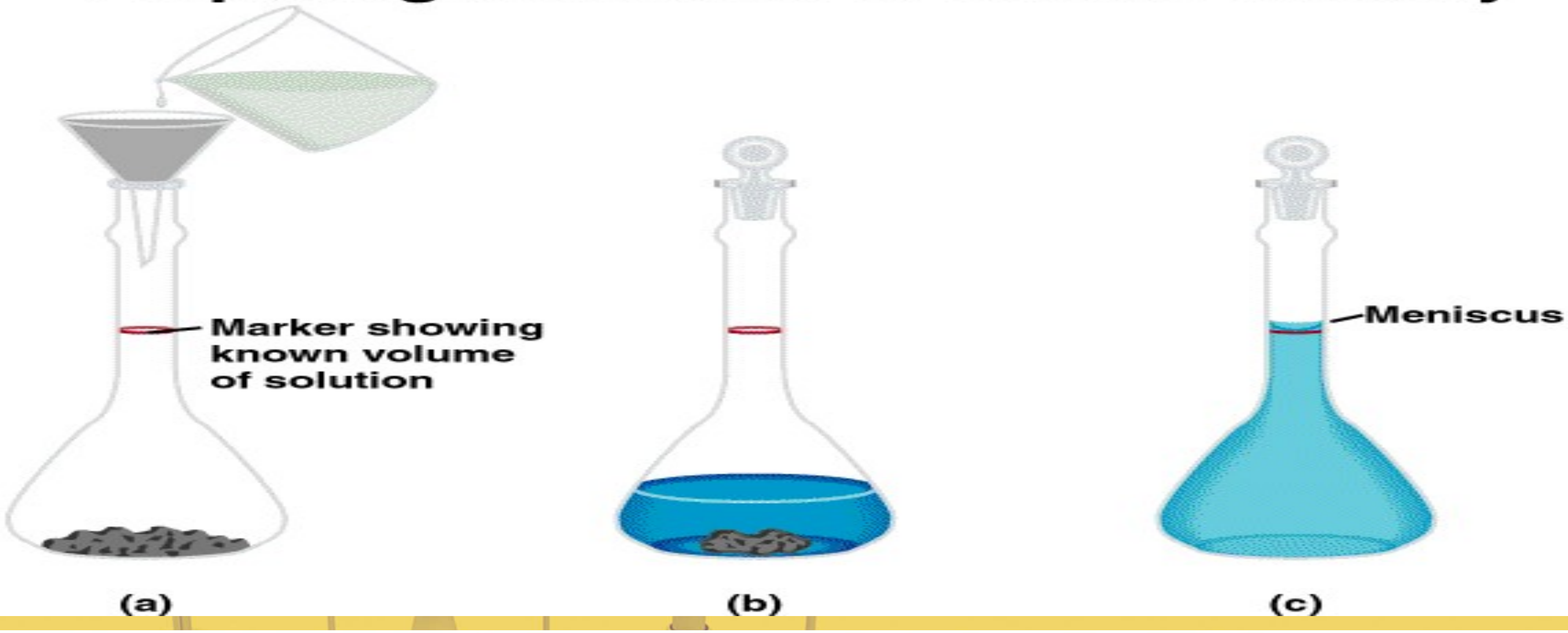
Given: n & V → M

Given: n & M → V

Given: V & M → n

$$n = \frac{m}{MM} \Rightarrow m = nMM$$

Preparing a Solution of Known Molarity



(a) (b) (c) (d)

Concentration of Solutions

What is the molar concentration of 0.730 mol glucose $C_6H_{12}O_6$ in 500 ml solution

$$C_6H_{12}O_6 = \frac{0.730 \text{ mol } C_6H_{12}O_6}{500 \text{ mL}} \times \frac{1000 \text{ mL soln}}{1 \text{L soln}}$$

$$= \mathbf{1.46\ M}\ C_6H_{12}O_6$$

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Example 4.6

DCM, Worked Examples

How many grams of potassium dichromate ($K_2Cr_2O_7$) are required to prepare a 250-mL solution whose concentration is 2.16 *M*?

Strategy How many moles of $K_2Cr_2O_7$ does a 1-L (or 1000 mL) 2.16 *M* $K_2Cr_2O_7$ solution contain? A 250-mL solution? How would you convert moles to grams?

Solution The first step is to determine the number of moles of $K_2Cr_2O_7$ in 250 mL or 0.250 L of a 2.16 *M* solution. Rearranging Equation (4.1) gives

$$\text{moles of solute} = \text{molarity} \times \text{L soln}$$

Thus,

$$\text{moles of } K_2Cr_2O_7 = \frac{2.16 \text{ mol } K_2Cr_2O_7}{1 \cancel{\text{L soln}}} \times 0.250 \cancel{\text{L soln}}$$
$$= 0.540 \text{ mol } K_2Cr_2O_7$$

The molar mass of $K_2Cr_2O_7$ is 294.2 g, so we write

$$\text{grams of } K_2Cr_2O_7 \text{ needed} = 0.540 \cancel{\text{mol } K_2Cr_2O_7} \times \frac{294.2 \text{ g } K_2Cr_2O_7}{1 \cancel{\text{mol } K_2Cr_2O_7}}$$
$$= 159 \text{ g } K_2Cr_2O_7$$

Check As a ball-park estimate, the mass should be given by [molarity (mol/L) × volume (L) × molar mass (g/mol)] or [2 mol/L × 0.25 L × 300 g/mol] = 150 g. So the answer is reasonable.

Practice Exercise What is the molarity of an 85.0-mL ethanol (C_2H_5OH) solution containing 1.77 g of ethanol?

Example 4.7

3.81g

DCM, Worked Examples

In a biochemical essay, a chemist needs to add 0.381 g of glucose to a reaction mixture. Calculate the volume in milliliters of a 2.53 *M* glucose solution she should use for the addition.

Strategy We must first determine the number of moles contained in 3.81 g of glucose and then use Equation (4.2) to calculate the volume.

Solution From the molar mass of glucose, we write

$$3.81 \text{ g } C_6H_{12}O_6 \times \frac{1 \text{ mol } C_6H_{12}O_6}{180.2 \text{ g } C_6H_{12}O_6} = 2.114 \times 10^{-2} \text{ mol } C_6H_{12}O_6$$

Next, we calculate the volume of the solution that contains 2.114×10^{-2} mole of the solute. Rearranging Equation (4.2) gives

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

$$= \frac{2.114 \times 10^{-2} \text{ mol } C_6H_{12}O_6}{2.53 \text{ mol } C_6H_{12}O_6/\text{L soln}} \times \frac{1000 \text{ mL soln}}{1 \text{ L soln}}$$

$$= 8.36 \text{ mL soln}$$

Check One liter of the solution contains 2.53 moles of $C_6H_{12}O_6$. Therefore, the number of moles in 8.36 mL or 8.36×10^{-3} L is $(2.53 \text{ mol} \times 8.36 \times 10^{-3})$ or 2.12×10^{-2} mol. The small difference is due to the different ways of rounding off.

Practice Exercise What volume (in milliliters) of a 0.315 *M* NaOH solution contains 6.22 g of NaOH?

Practice exercise 4.6

- What is the molarity of an 85 ml ethanol C_2H_5OH solution containing 1.77g of ethanol?

- Molar mass C_2H_5OH

  = 46.068g

n = 1.77g/ 46.068= 0.038 mol

- M=n/v= 0.038 mol/ 85 ml

- M= 0.452 M

Practice exercise 4.7

- What is the volume (in ml) of 0.315M NaOH solution contains 6.22g of NaOH?

- Molar mass NaOH= 40g

n = 6.22g /40g= 0.1555 mol

- v=n/M= 0.1555 mol / 0.315M

- v= 0.494l = 494 ml

Dilution is the procedure for preparing a less concentrated solution from a more concentrated solution.

Calculation based on that the number of moles of solute is constant we add only solvent

Moles of solute before dilution (i)	=	Moles of solute after dilution (f)

$$M_iV_i = M_fV_f$$

Practice

- How many mL of 5.0 M $K_2Cr_2O_7$ solution must be diluted to prepare 250 mL of 0.10 M solution?

V_i = ? **M_i = 5.0M** **V_f = 250 mL** **M_f = 0.10M**

$M_i = M_fV_f/V_i$ V_i = 250 ml ×0.1M/5ml **= 5 ml**

- If 10.0 mL of a 10.0 M stock solution of NaOH is diluted to 250 mL, what is the concentration of the resulting solution?

M_f= ? **V_i= 10.0 mL** **M_i = 10.0M** **V_f = 250 mL**

$M_i = M_fV_f/V_i$

M_i = 10ml ×10M/250ml **= 0.4 ml**

How would you prepare 60.0 mL of 0.200 *M* HNO_3 from a stock solution of 4.00 *M* HNO_3?

$$M_iV_i = M_fV_f$$

$M_i = 4.00$ $M_f = 0.200$ $V_f = 60$ ml $V_i = ?$ ml

$$V_i = \frac{M_fV_f}{M_i} = \frac{0.200 \times 60}{4.00} = 3 \text{ mL}$$

Example 4.8

DCM, Worked Examples

Describe how you would prepare 5.00×10^2 mL of a 1.75 M H_2SO_4 solution, starting with an 8.61 M stock solution of H_2SO_4.

Strategy Because the concentration of the final solution is less than that of the original one, this is a dilution process. Keep in mind that in dilution, the concentration of the solution decreases but the number of moles of the solute remains the same.

Solution We prepare for the calculation by tabulating our data:

$$M_i = 8.61\,M \qquad M_f = 1.75\,M$$
$$V_i = ? \qquad V_f = 5.00 \times 10^2\text{ mL}$$

Substituting in Equation (4.3),

$$(8.61\,M)(V_i) = (1.75\,M)(5.00 \times 10^2\text{ mL})$$
$$V_i = \frac{(1.75\,M)(5.00 \times 10^2\text{ mL})}{8.61\,M}$$
$$= 102\text{ mL}$$

Thus, we must dilute 102 mL of the 8.61 M H_2SO_4 solution with sufficient water to give a final volume of 5.00×10^2 mL in a 500-mL volumetric flask to obtain the desired concentration.

Check The initial volume is less than the final volume, so the answer is reasonable.

Practice Exercise How would you prepare 2.00×10^2 mL of a 0.866 M NaOH solution, starting with a 5.07 M stock solution?

Practice exercise 4.9

How would you prepare 200 mL of 0.866 *M* NaOH from a stock solution of 5.07 *M* NaOH?

$$M_iV_i = M_fV_f$$

M_i = 5.07 M_f = 0.866 V_f = 200 ml V_i = ? ml

$$V_i = \frac{M_fV_f}{M_i} = \frac{0.866 \times 200}{5.07} = 34.2 \text{ mL}$$

Problem 4.74 (page 163)

$M_1 = 0.568$ M $\quad V_1 = 46.2$ mL $= 46.2 \times 10^{-3}$ L

$\therefore$ moles for the first $Ca(NO_3)_2$ solution $(n_1) = M_1 \times V_1$

$= 0.568 \text{ M} \times 46.2 \times 10^{-3} \text{ L} = 0.026$ mol

$M_2 = 1.396$ M $\quad V_2 = 80.5$ mL $= 80.5 \times 10^{-3}$ L

$\therefore$ moles for the second $Ca(NO_3)_2$ solution $(n_2) = M_2 \times V_2$

$= 1.396 \text{ M} \times 80.5 \times 10^{-3} \text{ L} = 0.112$ mol

Total moles of $Ca(NO_3)_2$ in the final solution $= n_1 + n_2$

$= 0.026 + 0.112 = 0.138$ mol

Total volume of the final solution $= V_1 + V_2$

$= (46.2 \times 10^{-3} \text{ L}) + (80.5 \times 10^{-3} \text{ L}) = 126.7 \times 10^{-3}$ L

The concentration of the final solution $M_f = n/V$

$= 0.138 \text{ mol} / 126.7 \times 10^{-3} = 1.09$ M

Calculation based on that we have two solutions with different number of moles mixed together , then we will calculate the molarity of the new solution

$$\text{Molarity } (M) = \frac{\text{moles of solute}}{\text{volume of solution in liters}}$$

Molarity (M)

moles of solute = mass / molar mass

volume of solution in liters

Chapter 5 Gases

Chapter 5 Gases

- **5.1 Substances that exist s gases**
- **5.2 Pressure of the gas**
- **5.3 The gas laws**
- **5.4 Ideal gas equation**
- **5.5 Gas stoichiometry**
- **5.6 Dalton's Law of Partial Pressures**

5.1 substances that exist s gases

Elements that exist as gases at 25^{0}C and 1 atmosphere

1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg							

TABLE 5.1 Some Substances Found as Gases at 1 atm and 25°C

Elements	Compounds
H_2 (molecular hydrogen)	HF (hydrogen fluoride)
N_2 (molecular nitrogen)	HCl (hydrogen chloride)
O_2 (molecular oxygen)	HBr (hydrogen bromide)
O_3 (ozone)	HI (hydrogen iodide)
F_2 (molecular fluorine)	CO (carbon monoxide)
Cl_2 (molecular chlorine)	CO_2 (carbon dioxide)
He (helium)	NH_3 (ammonia)
Ne (neon)	NO (nitric oxide)
Ar (argon)	NO_2 (nitrogen dioxide)
Kr (krypton)	N_2O (nitrous oxide)
Xe (xenon)	SO_2 (sulfur dioxide)
Rn (radon)	H_2S (hydrogen sulfide)
	HCN (hydrogen cyanide)*

*The boiling point of HCN is 26°C, but it is close enough to qualify as a gas at ordinary atmospheri

Physical Characteristics of Gases

- Gases assume the volume and shape of their containers.
- Gases are the most compressible state of matter.
- Gases will mix evenly and completely when confined to the same container.
- Gases have much lower densities than liquids and solids.

5.2 Pressure of Gases and its Units


Blaise Pascal

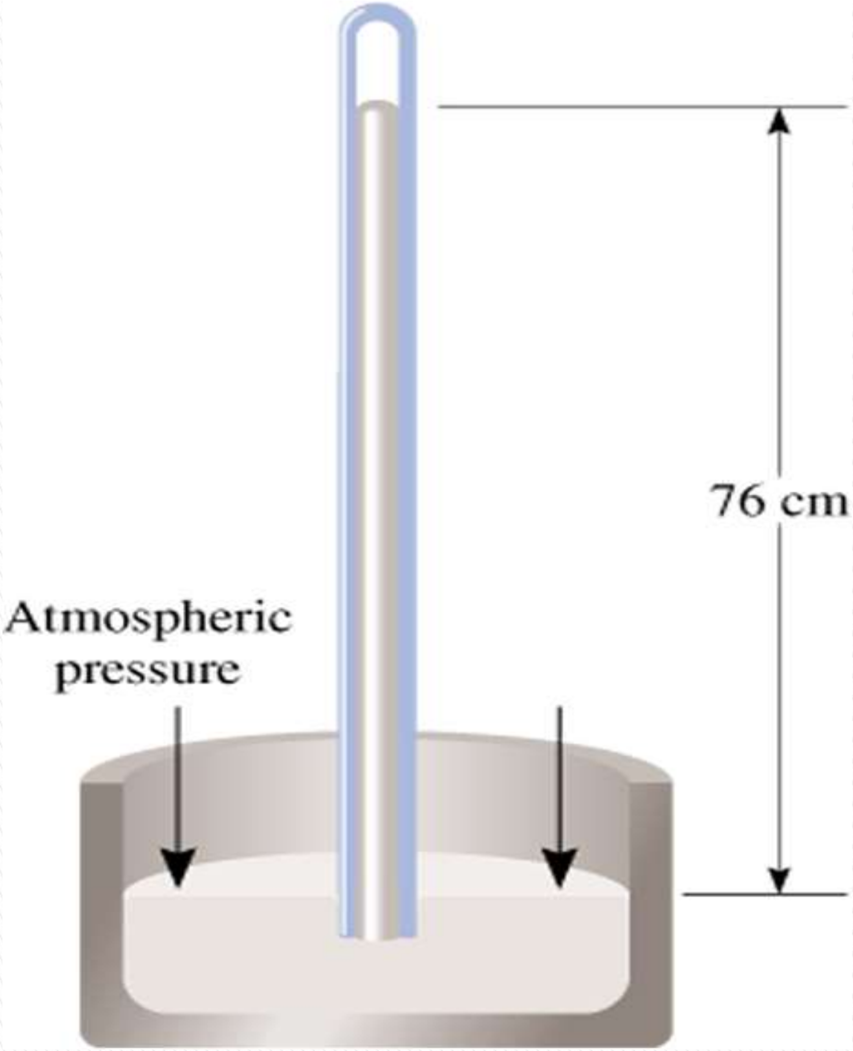
- Pressure is defined as the force applied per unit are

$$\textbf{Pressure} = \frac{\textbf{Force}}{\textbf{Area}} = \textbf{N/m}^2$$

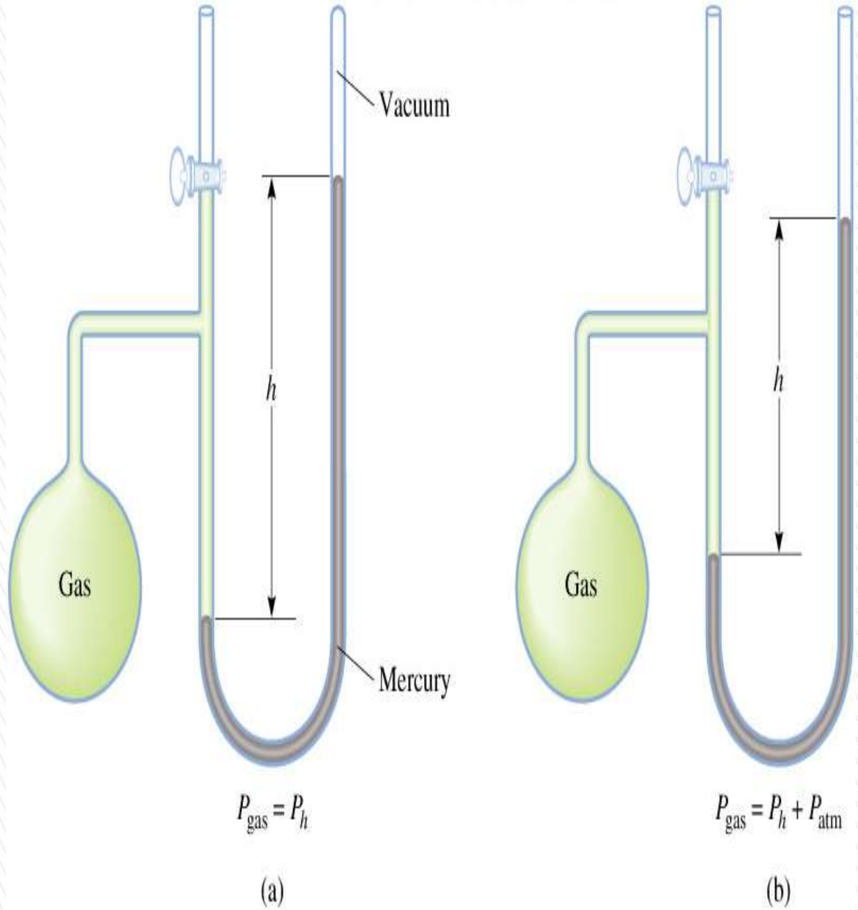
- The SI unit of pressure is Pascal (Pa) define as one Newton per square meter **(1Pa = N/m^2)**

- **Standard atmospheric pressure,** the pressure that supports a column of mercury exactly 760 mm high at 0 ^oC at sea level.

- Measured using a Barometer! – A device tha weigh the atmosphere above us!



Barometer



a Barometer! – A device that can weigh the atmosphere above us!

A simple manometer, a device for measuring the pressure of a gas in a container

Common Units of Pressure

Unit	Atmospheric Pressure	Scientific Field Used
Pascal (Pa); kilopascal (kPa)	1.01325×10^5 Pa 101.325 kPa	SI unit; physics, chemistry
Atmosphere (atm)	1 atm	Chemistry
Millimeters of mercury (mmHg)	760 mmHg	Chemistry, medicine biology
Torr	760 torr	Chemistry
Pounds per square inch (psl or lb/in^2)	14.7 lb/in^2	Engineering
Bar	1.01325 bar	Meteorology, chemistry

Common Units of Pressure

Remember the conversions for pressure:

760 mm Hg = 760 torr

1 atm = 760 mm Hg

760 mm Hg = 101.325 Pa

Convert 2.3 atm to torr:

Example ....

$$2.0 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 1520 \text{ torr}$$

http://www.unit-conversion.info/pressure.html

http://www.onlineconversion.com/pressure.htm

Example 5.1

DCM, Worked Examples

The pressure outside a jet plane flying at high altitude falls considerably below standard atmospheric pressure. Therefore, the air inside the cabin must be pressurized to protect the passengers. What is the pressure in atmospheres in the cabin if the barometer reading is 688 mmHg?

Strategy Because 1 atm = 760 mmHg, the following conversion factor is needed to obtain the pressure in atmospheres

$$\frac{1\text{ atm}}{760\text{ mmHg}}$$

Solution The pressure in the cabin is given by

$$\text{pressure} = 688\ \cancel{\text{mmHg}} \times \frac{1\text{ atm}}{760\ \cancel{\text{mmHg}}}$$

$$= 0.905\text{ atm}$$

Practice Exercise Convert 749 mmHg to atmospheres.

Example 5.2

DCM, Worked Examples

The atmospheric pressure in San Francisco on a certain day was 732 mmHg. What was the pressure in kPa?

Strategy Here we are asked to convert mmHg to kPa. Because

$$1 \text{ atm} = 1.01325 \times 10^5 \text{ Pa} = 760 \text{ mmHg}$$

the conversion factor we need is

$$\frac{1.01325 \times 10^5 \text{ Pa}}{760 \text{ mmHg}}$$

Solution The pressure in kPa is

$$\text{pressure} = 732 \text{ mmHg} \times \frac{1.01325 \times 10^5 \text{ Pa}}{760 \text{ mmHg}}$$
$$= 9.76 \times 10^4 \text{ Pa}$$
$$= 97.6 \text{ kPa}$$

Practice Exercise Convert 295 mmHg to kilopascals.

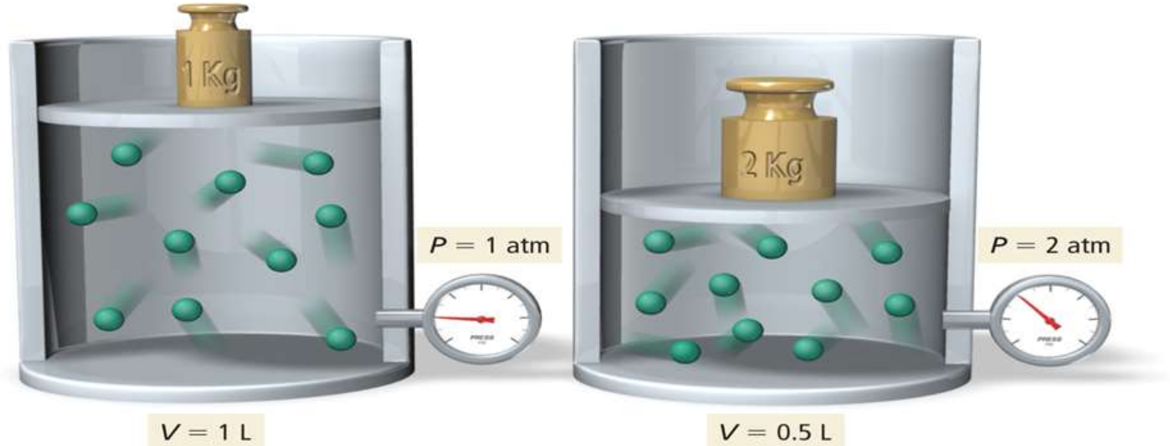
5.3 The gas laws

- Boyle's Law , V – P relationship

- Charles Law , V – T– relationship

- Avogadro's Law ,V and Amount

Boyle’s Law : P – V relationship



- Pressure is inversely proportional to Volume
- $P = \frac{k}{V}$ or $V = \frac{k}{P}$ or
- \
- $P_1V_1 = k$ $P_2V_2 = k'$

$$k = k'$$

Then : $P_1V_1 = P_2V_2$

Example A cylinder with a movable piston has a volume of 7.25 L at 4.52 atm. What is the volume at 1.21 atm?

Given:	V_1 = 7.25 L, P_1 = 4.52 atm, P_2 = 1.21 atm
Find:	V_2, L
Concept Plan:	V_1, P_1, P_2 ⇒ V_2
Relationships:	$P_1 \bullet V_1 = P_2 \bullet V_2$ $V_2 = \frac{P_1 \bullet V_1}{P_2}$
Solution:	$V_2 = \frac{P_1 \bullet V_1}{P_2}$ $= \frac{(4.52\ \text{atm}) \bullet (7.25\ \text{L})}{(1.21\ \text{atm})} = 27.1\ \text{L}$
Check:	since P and V are inversely proportional, when the pressure decreases ~4x, the volume should increase ~4x, and it does

A balloon is put in a bell jar and the pressure is reduced from 782 torr to 0.500 atm. If the volume of the balloon is now 2780 mL, what was it originally?

Given:	V_2 = 2780 mL, P_1 = 762 torr, P_2 = 0.500 atm
Find:	V_1, mL
Concept Plan:	V_1, P_1, P_2 → V_2
Relationships:	$P_1 \bullet V_1 = P_2 \bullet V_2$, 1 atm = 760 torr (exactly) $V_1 = \frac{P_2 \bullet V_2}{P_1}$
Solution:	$782 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.03 \text{ atm}$ $V_1 = \frac{P_2 \bullet V_2}{P_1} = \frac{(0.500 \text{ atm}) \bullet (2780 \text{ L})}{(1.03 \text{ atm})} = 1350 \text{ mL}$
Check:	since P and V are inversely proportional, when the pressure decreases ~2x, the volume should increase ~2x, and it does

Charles' Law

- volume is directly proportional to temperature
  - constant P and amount of gas

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

- as T increases, V also increases
- Kelvin T = Celsius T + 273

- V = constant x T
  - if T measured in Kelvin

A gas has a volume of 2.57 L at 0.00°C. What was the temperature at 2.80 L?

Given:	V_1 =2.57 L, V_2 = 2.80 L, t_2 = 0.00°C
Find:	t_1, K and °C
Concept Plan:	V_1, V_2, T_2 → T_1
Relationships:	$T_1 = T_2 \bullet \frac{V_1}{V_2}$ T(K) = t(°C) + 273.15, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Solution:	$T_2 = 0.00 + 273.15$ $T_2 = 273.15\,K$ $T_1 = \frac{T_2 \bullet V_1}{V_2}$ $= \frac{(273.15\,K)\bullet(2.57\,\cancel{L})}{(2.80\,\cancel{L})} = 29\underline{7}.6\,K$ $t_1 = T_1 - 273.15$ $t_1 = 29\underline{7}.6 - 273.15$ $t_1 = 24\,°C$
Check:	since T and V are directly proportional, when the volume decreases, the temperature should decrease, and it does

EXAMPLE

- Gas occupy 6L at 37^0C what will be its volume when its temperature decreased to the half?

$$V1=6L\ ,\ T1=37^0C$$

$$V2=???\ ,\ T2=\tfrac{1}{2}\ T1$$

$$V1T2=V2T1$$

$$V1\tfrac{1}{2}T1=V2T1$$

$$V2=\tfrac{1}{2}V1$$

$$V2=\tfrac{1}{2}(6)=3L$$

Avogadro's Law

- volume directly proportional to the number of gas molecules
  - V = constant x n
  - constant P and T
  - more gas molecules = larger volume
- count number of gas molecules by moles
- equal volumes of gases contain equal numbers of molecules
  - *the gas doesn't matter*

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

A 0.225 mol sample of He has a volume of 4.65 L. How many moles must be added to give 6.48 L?

Given:	V_1 = 4.65 L, V_2 = 6.48 L, n_1 = 0.225 mol
Find:	n_2, and added moles
Concept Plan:	V_1, V_2, n_1 ⇒ n_2 ; $n_1 \bullet \frac{V_2}{V_1} = n_2$
Relationships:	mol added = $n_2 - n_1$, $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

Solution:

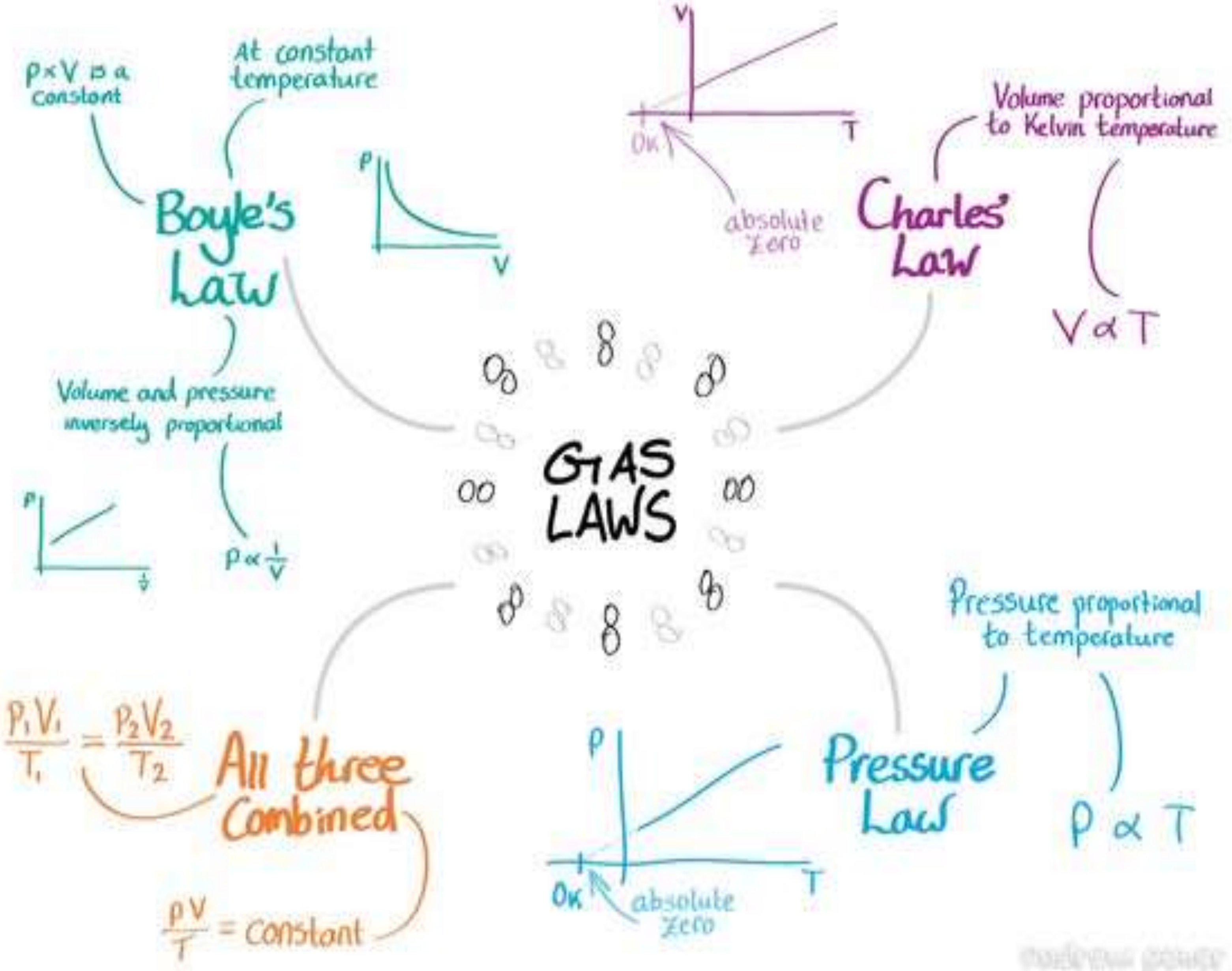
$$n_2 = \frac{n_1 \bullet V_2}{V_1}$$

$$= \frac{(0.225\text{ mol}) \bullet (6.48\text{ L})}{(4.65\text{ L})} = 0.314\text{ mol}$$

$$\text{moles added} = 0.314 - 0.225$$

$$\text{moles added} = 0.089\text{ mol}$$

Check: since n and V are directly proportional, when the volume increases, the moles should increase, and it does


GAS LAWS
Boyle's Law
P×V is a constant
At constant temperature
Volume and pressure inversely proportional
$P \propto \frac{1}{V}$
P
V
Charles' Law
Volume proportional to Kelvin temperature
$V \propto T$
absolute zero
0K
T
Pressure Law
Pressure proportional to temperature
$P \propto T$
absolute zero
0K
All three Combined
$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
$\frac{PV}{T}$ = constant

Combined Gas Law

- When we introduced Boyle's, Charles's, and Gay-Lussac's laws, we assumed that one of the variables remained constant.
- Experimentally, all three (temperature, pressure, and volume) usually change.
- By combining all three laws, we obtain the *combined gas law*:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Example 5.5

DCM, Worked Examples

An inflated helium balloon with a volume of 0.55 L at sea level (1.0 atm) is allowed to rise to a height of 6.5 km, where the pressure is about 0.40 atm. Assuming that the temperature remains constant, what is the final volume of the balloon?

Strategy The amount of gas inside the balloon and its temperature remain constant, but both the pressure and the volume change. What gas law do you need?

Solution We start with Equation (5.9)

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

Because $n_1 = n_2$ and $T_1 = T_2$,

$$P_1V_1 = P_2V_2$$

which is Boyle's law [see Equation (5.2)]. The given information is tabulated:

Initial Conditions	Final Conditions
$P_1 = 1.0$ atm	$P_2 = 0.40$ atm
$V_1 = 0.55$ L	$V_2 = ?$

Therefore,

$$\begin{aligned} V_2 &= V_1 \times \frac{P_1}{P_2} \\ &= 0.55\text{ L} \times \frac{1.0\text{ atm}}{0.40\text{ atm}} \\ &= 1.4\text{ L} \end{aligned}$$

Check When pressure applied on the balloon is reduced (at constant temperature), the helium gas expands and the balloon's volume increases. The final volume is greater than the initial volume, so the answer is reasonable.

Practice Exercise A sample of chlorine gas occupies a volume of 946 mL at a pressure of 726 mmHg. Calculate the pressure of the gas (in mmHg) if the volume is reduced at constant temperature to 154 mL.

5.4 Ideal Gas Law

- By combing the gas laws we can write a general equation
- **R** is called the **gas constant**
- the value of **R** depends on the units of P and V
  - we will use 0.08206 $\frac{atm \bullet L}{mol \bullet K}$ and convert P to atm and V to L
  - the other gas laws are found in the ideal gas law if two variables are kept constant
  - allows us to find one of the variables if we know the other 3

$$PV = nRT$$

Example 5.3

DCM, Worked Examples

Sulfur hexafluoride (SF_6) is a colorless, odorless, very unreactive gas. Calculate the pressure (in atm) exerted by 1.82 moles of the gas in a steel vessel of volume 5.43 L at 69.5°C.

Strategy The problem gives the amount of the gas and its volume and temperature. Is the gas undergoing a change in any of its properties? What equation should we use to solve for the pressure? What temperature unit should we use?

Solution Because no changes in gas properties occur, we can use the ideal gas equation to calculate the pressure. Rearranging Equation (5.8), we write

$$P = \frac{nRT}{V}$$

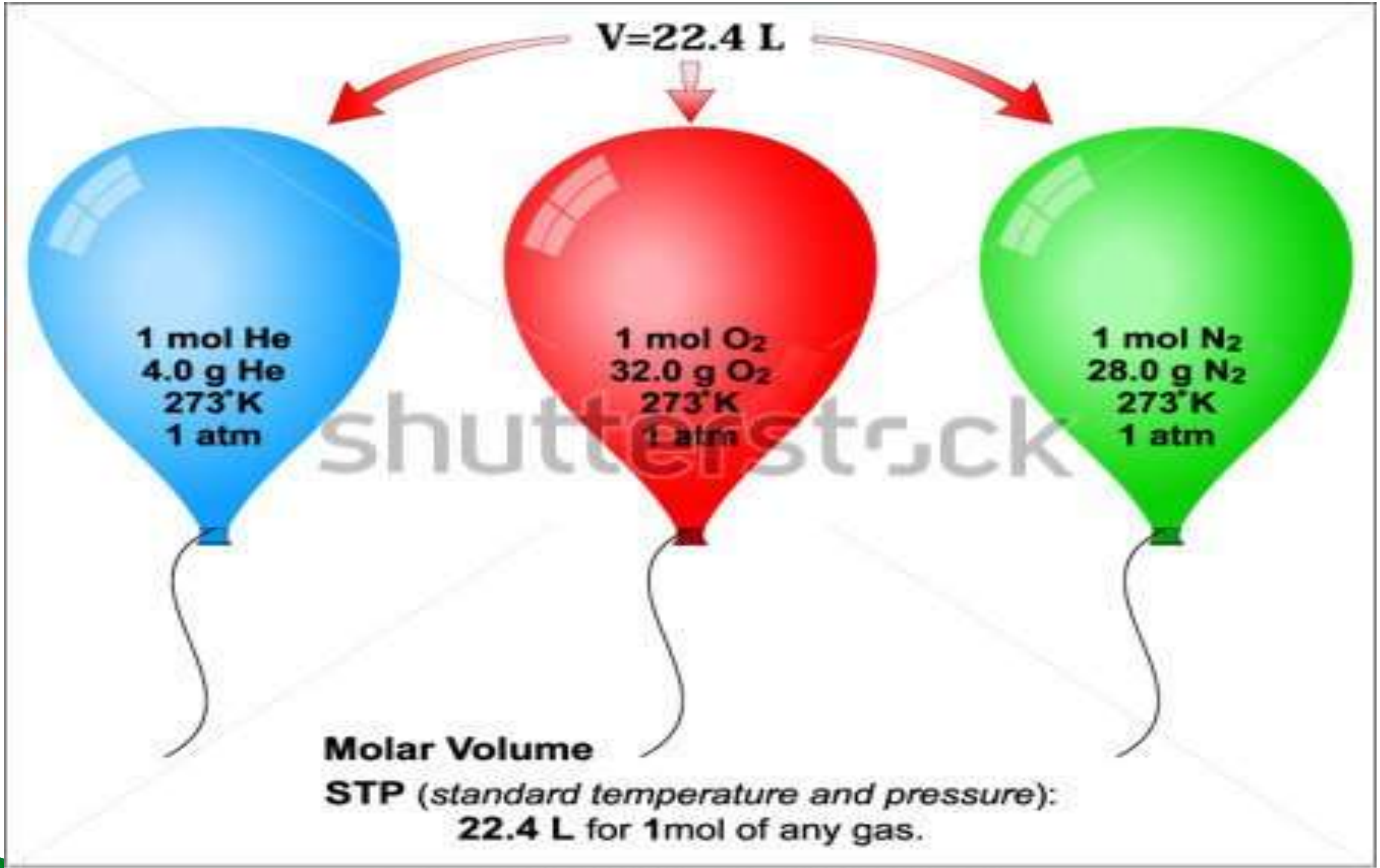
$$= \frac{(1.82 \text{ mol})(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(69.5 + 273) \text{ K}}{5.43 \text{ L}}$$

$$= 9.42 \text{ atm}$$

Practice Exercise Calculate the volume (in liters) occupied by 2.12 moles of nitric oxide (NO) at 6.54 atm and 76°C.

Standard Conditions

- since the volume of a gas varies with pressure and temperature, chemists have agreed on a set of conditions to report our measurements so that comparison is easy – we call these **standard conditions**
  - STP
- standard pressure = 1 atm
- standard temperature = 273 K = 0°C
- One mole of a gas occupy **22.41** L at STP


V=22.4 L
1 mol He
4.0 g He
273°K
1 atm
1 mol O2
32.0 g O2
273°K
1 atm
1 mol N2
28.0 g N2
273°K
1 atm
shutterstock
Molar Volume
STP (standard temperature and pressure):
22.4 L for 1mol of any gas.

Example 5-4

- Calculate the (volume in liters occupied by 7.40g of NH_3 at STP
- Solution 2
- $n_{NH3} = 7.4 / 17 = 0.435$ mol

- V = nRT/P
- V = 0.435 (0.0821) 273/1 = 9.74 L

- Solution 2
- 1 mole occupy 22.4 L at STP
- 0.435 mole x >>>>>>>>>> V = 0.435 X 22.4=9.74 L

EXAMPLE

- What is the volume of 2g O_2 gas at STP

$$PV = nRT$$

$$V = nRT/P$$

$$V = 2 \times 0.0821 \times 273/32 \times 1$$

$$V = 1.4 \text{ L}$$

- What is the volume of 2g O_2 gas at 4 atm and 35^0C

$$PV = nRT$$

$$V = nRT/P$$

$$T = 35 + 273 = 308 \text{ K}$$

$$V = 2 \times 0.0821 \times 308/ 32 \times 4$$

$$V = 0.395 \text{ L}$$

Gas Density

$$\text{mass} \times \frac{1\ \text{mol}}{\text{molar mass}} = \text{moles} \quad \therefore \quad \text{moles} = \frac{\text{mass}}{\text{molar mass}}$$

$$\text{density} = \frac{\text{mass in grams}}{\text{volume in liters}}$$

$$P \times V = n \times R \times T$$

$$P \times V = \frac{\text{mass}}{\text{molar mass}} \times R \times T$$

$$\frac{\text{mass}}{V} = \text{density} = \frac{P \times (\text{molar mass})}{R \times T}$$

- density is directly proportional to molar mass

Molar Mass of a Gas

- From density calculations
- $\mathcal{M} = dRT/P$

- From number of moles calculations
- $n = mass / \mathcal{M}$
- $PV = nRT$
- $PV = (mass / \mathcal{M})\,RT$

Example 5.8

DCM, Worked Examples

Calculate the density of carbon dioxide (CO_2) in grams per liter (g/L) at 0.990 atm and 55°C.

Strategy We need Equation (5.11) to calculate gas density. Is sufficient information provided in the problem? What temperature unit should be used?

Solution To use Equation (5.11), we convert temperature to kelvins ($T = 273 + 55 = 328$ K) and use 44.01 g for the molar mass of CO_2:

$$d = \frac{P\mathcal{M}}{RT}$$
$$= \frac{(0.990 \text{ atm})(44.01 \text{ g/mol})}{(0.0821 \text{ L}\cdot\text{atm/K}\cdot\text{mol})(328 \text{ K})} = \boxed{1.62 \text{ g/L}}$$

Alternatively, we can solve for the density by writing

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Assuming that we have 1 mole of CO_2, the mass is 44.01 g. The volume of the gas can be obtained from the ideal gas equation

$$V = \frac{nRT}{P}$$
$$= \frac{(1 \text{ mol})(0.0821 \text{ L}\cdot\text{atm/K}\cdot\text{mol})(328 \text{ K})}{0.990 \text{ atm}}$$
$$= 27.2 \text{ L}$$

Therefore, the density of CO_2 is given by

$$d = \frac{44.01 \text{ g}}{27.2 \text{ L}} = \boxed{1.62 \text{ g/L}}$$

Comment In units of grams per milliliter, the gas density is 1.62×10^{-3} g/mL, which is a very small number. In comparison, the density of water is 1.0 g/mL and that of gold is 19.3 g/cm^3.

Practice Exercise What is the density (in g/L) of uranium hexafluoride (UF_6) at 779 mmHg and 62°C?

اقرأ باسم ربك
جامعة الملك عبدالعزيز
King Abdulaziz University

Example 5-9

- A chemist synthesized a greensh-yellow gaseous compound of chlorine and oxygen and find that its denity is 7.7g/L at 36°C and 2.88 atm. Calculate the molar mass and determine its molecular formula.
- **Molar mass = dRT/ P**
- $\mathcal{m}$ = 7.7g/L ×0.0821×(36+273)/2.88 = 67.9 g/mol
- Mass of empirical formula (ClO)= 35.45+16= 51.45
- Ratio = **Molar mass** / Mass of empirical formula = 67.9/51.45= 1.3
- molecular formula. ClO_2

Example 5.10

DCM, Worked Examples

Chemical analysis of a gaseous compound showed that it contained 33.0 percent silicon (Si) and 67.0 percent fluorine (F) by mass. At 35°C, 0.210 L of the compound exerted a pressure of 1.70 atm. If the mass of 0.210 L of the compound was 2.38 g, calculate the molecular formula of the compound.

Strategy This problem can be divided into two parts. First, it asks for the empirical formula of the compound from the percent by mass of Si and F. Second, the information provided enables us to calculate the molar mass of the compound and hence determine its molecular formula. What is the relationship between empirical molar mass and molar mass calculated from the molecular formula?

Solution We follow the procedure in Example 3.9 (p. 88) to calculate the empirical formula by assuming that we have 100 g of the compound, so the percentages are converted to grams. The number of moles of Si and F are given by

$$n_{Si} = 33.0\ \cancel{g\ Si} \times \frac{1\ \text{mol Si}}{28.09\ \cancel{g\ Si}} = 1.17\ \text{mol Si}$$

$$n_{F} = 67.0\ \cancel{g\ F} \times \frac{1\ \text{mol F}}{19.00\ \cancel{g\ F}} = 3.53\ \text{mol F}$$

Therefore, the empirical formula is $Si_{1.17}F_{3.53}$, or, dividing by the smaller subscript (1.17), we obtain SiF_3.

To calculate the molar mass of the compound, we need first to calculate the number of moles contained in 2.38 g of the compound. From the ideal gas equation

$$n = \frac{PV}{RT}$$

$$= \frac{(1.70\ \text{atm})(0.210\ \text{L})}{(0.0821\ \text{L} \cdot \text{atm/K} \cdot \text{mol})(308\ \text{K})} = 0.0141\ \text{mol}$$

Because there are 2.38 g in 0.0141 mole of the compound, the mass in 1 mole, or the molar mass, is given by

$$\mathcal{M} = \frac{2.38\ \text{g}}{0.0141\ \text{mol}} = 169\ \text{g/mol}$$

The molar mass of the empirical formula SiF_3 is 85.09 g. Recall that the ratio (molar mass/empirical molar mass) is always an integer (169/85.09 ≈ 2). Therefore, the molecular formula of the compound must be $(SiF_3)_2$ or Si_2F_6.

Practice Exercise A gaseous compound is 78.14 percent boron and 21.86 percent hydrogen. At 27°C, 74.3 mL of the gas exerted a pressure of 1.12 atm. If the mass of the gas was 0.0934 g, what is its molecular formula?

5.5 Gas stoichiometry

- **Example 5.11**
- Calculate the volume of O_2(in L) requred for the complete combustion of 7.64 L of (C_2H_2) measured at the same T & P

$$2\ C_2H_{2\ (g)} + 5O_{2\ (g)} \rightarrow 4CO_{2\ (g)} + 2H_2O_{\ (l)}$$

- From Avogadro low v= Rn

- Volume of O_2 = 7.64 L × 5L O_2 /2L C_2H_2 = 19.1 L

Example 5.12

$$2NaN_3\,(S) \rightarrow 2\,Na\,(s) + 3N_2\,(g)$$

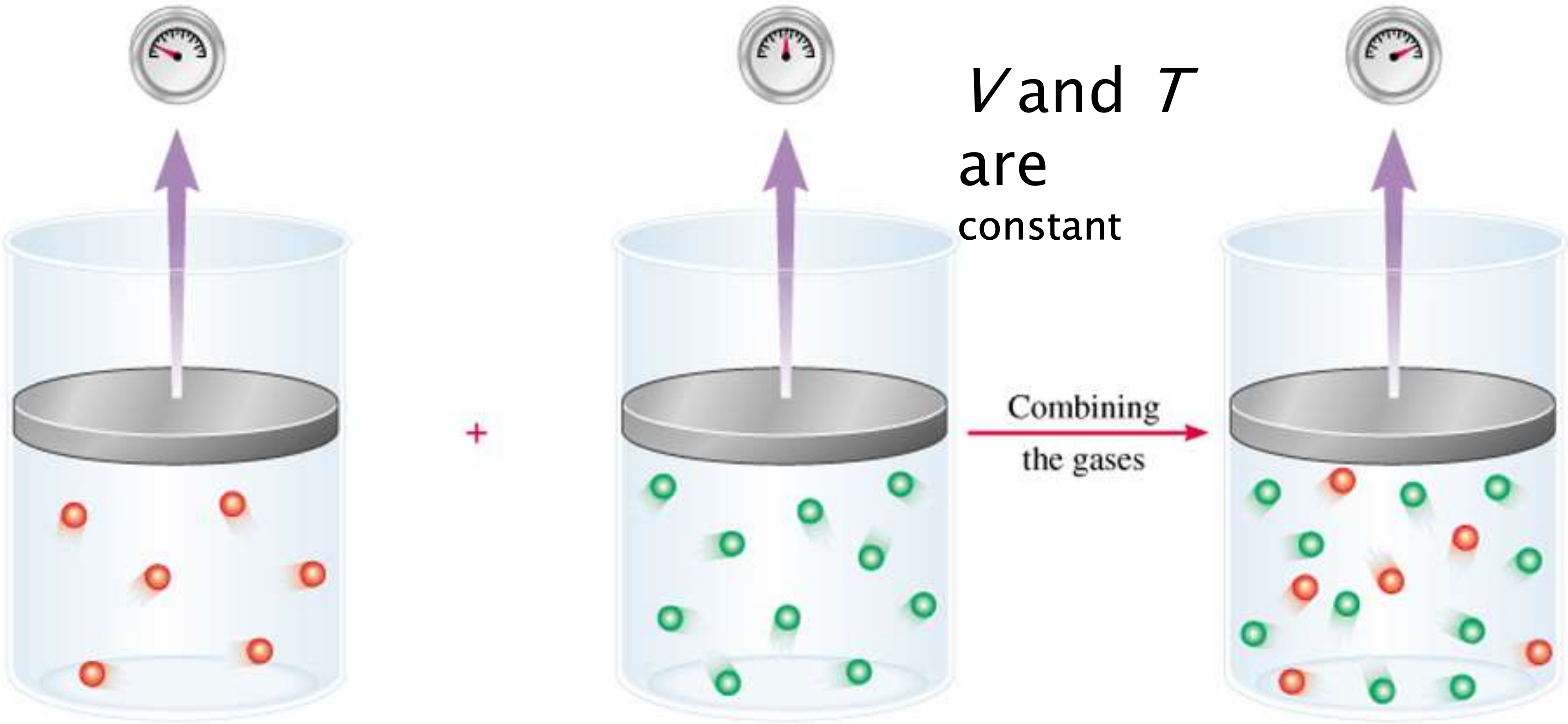
Calculate the volume of N_2 generate at 80°C and 823 mmHg by the decomposition of 60 g of NaN_3

- n of N_2 $= (60/65.02) \times 3/2 = 1.38$
- $PV=nRT \rightarrow V=nRT/P$
- $V= 1.38 \times 0.0821 \times (80+273)/(823/760)$

$$= 36.9\ L$$

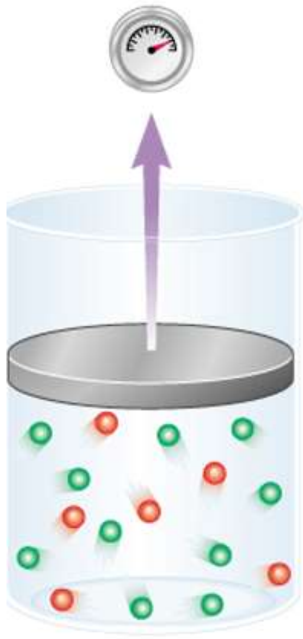
Dalton's Law of Partial Pressures

V and *T* are constant



P_A P_B $P_{total} = P_A + P_B$

Consider a case in which two gases, A and B, are in a container of



$$P_A = \frac{n_A RT}{V}$$

n_A is the number of moles of A

$$P_B = \frac{n_B RT}{V}$$

n_B is the number of moles of B

$$P_T = P_A + P_B \qquad X_A = \frac{n_A}{n_A + n_B} \qquad X_B = \frac{n_B}{n_A + n_B}$$

$$P_A = X_A P_T \qquad P_B = X_B P_T$$

$$\boxed{P_i = X_i P_T}$$

mole fraction $(X_i) = \dfrac{n_i}{n_T}$

Example 5.14

DCM, Worked Examples

A mixture of gases contains 4.46 moles of neon (Ne), 0.74 mole of argon (Ar), and 2.15 moles of xenon (Xe). Calculate the partial pressures of the gases if the total pressure is 2.00 atm at a certain temperature.

Strategy What is the relationship between the partial pressure of a gas and the total gas pressure? How do we calculate the mole fraction of a gas?

Solution According to Equation (5.14), the partial pressure of Ne (P_{Ne}) is equal to the product of its mole fraction (X_{Ne}) and the total pressure (P_T)

need to find

$$P_{Ne} = X_{Ne}P_T$$

want to calculate — given

Using Equation (5.13), we calculate the mole fraction of Ne as follows:

$$X_{Ne} = \frac{n_{Ne}}{n_{Ne} + n_{Ar} + n_{Xe}} = \frac{4.46 \text{ mol}}{4.46 \text{ mol} + 0.74 \text{ mol} + 2.15 \text{ mol}} = 0.607$$

Therefore

$$\begin{aligned} P_{Ne} &= X_{Ne}P_T \\ &= 0.607 \times 2.00 \text{ atm} \\ &= 1.21 \text{ atm} \end{aligned}$$

Similarly,

$$\begin{aligned} P_{Ar} &= X_{Ar}P_T \\ &= 0.10 \times 2.00 \text{ atm} \\ &= 0.20 \text{ atm} \end{aligned}$$

and

$$\begin{aligned} P_{Xe} &= X_{Xe}P_T \\ &= 0.293 \times 2.00 \text{ atm} \\ &= 0.586 \text{ atm} \end{aligned}$$

Check Make sure that the sum of the partial pressures is equal to the given total pressure; that is, (1.21 + 0.20 + 0.586) atm = 2.00 atm.

Practice Exercise A sample of natural gas contains 8.24 moles of methane (CH_4), 0.421 mole of ethane (C_2H_6), and 0.116 mole of propane (C_3H_8). If the total pressure of the gases is 1.37 atm, what are the partial pressures of the gases?

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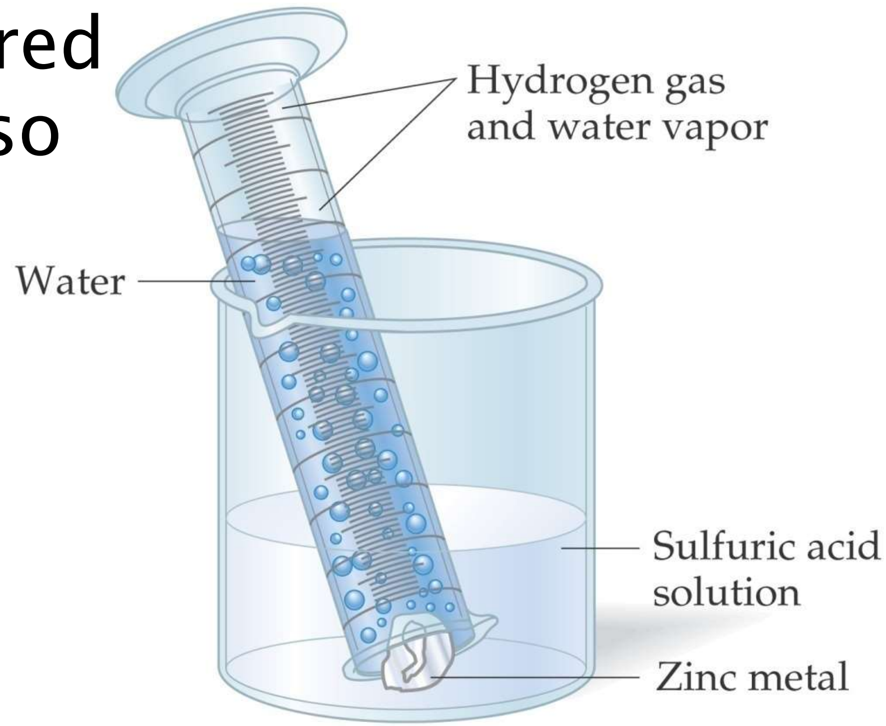
King Abdulaziz University

Collecting a Gas Over Water

- We can measure the volume of a gas by displacement.
- By collecting the gas in a graduated cylinder, we can measure the amount of gas produced.
- The gas collected is referred to as "wet" gas since it also contains water vapor.

$$P_T = P_{O2} + P_{H2O}$$



Example 5.15

DCM, Worked Examples

Oxygen gas generated by the decomposition of potassium chlorate is collected as shown in Figure 5.15. The volume of oxygen collected at 24°C and atmospheric pressure of 762 mmHg is 128 mL. Calculate the mass (in grams) of oxygen gas obtained. The pressure of the water vapor at 24°C is 22.4 mmHg.

Strategy To solve for the mass of O_2 generated, we must first calculate the partial pressure of O_2 in the mixture. What gas law do we need? How do we convert pressure of O_2 gas to mass of O_2 in grams?

Solution From Dalton's law of partial pressures we know that

$$P_T = P_{O_2} + P_{H_2O}$$

Therefore

$$\begin{aligned} P_{O_2} &= P_T - P_{H_2O} \\ &= 762 \text{ mmHg} - 22.4 \text{ mmHg} \\ &= 740 \text{ mmHg} \end{aligned}$$

From the ideal gas equation we write

$$PV = nRT = \frac{m}{\mathcal{M}}RT$$

where m and $\mathcal{M}$ are the mass of O_2 collected and the molar mass of O_2, respectively. Rearranging the equation we obtain

$$\begin{aligned} m = \frac{PV\mathcal{M}}{RT} &= \frac{(740/760)\text{atm}(0.128 \text{ L})(32.00 \text{ g/mol})}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(273 + 24) \text{ K}} \\ &= 0.164 \text{ g} \end{aligned}$$

Check The density of the oxygen gas is (0.164 g/0.128 L), or 1.28 g/L, which is a reasonable value for gases under atmospheric conditions (see Example 5.8).

Practice Exercise Hydrogen gas generated when calcium metal reacts with water is collected as shown in Figure 5.15. The volume of gas collected at 30°C and pressure of 988 mmHg is 641 mL. What is the mass (in grams) of the hydrogen gas obtained? The pressure of water vapor at 30°C is 31.82 mmHg.

Chapter 7
Quantum Theory and the Electronic Structure of Atoms

- 7.1 From Classical Physics to Quantum Theory
- 7.3 Bohr's Theory of the Hydrogen Atom
- 7.6 Quantum Numbers
- 7.7 Atomic Orbital's
- 7.8 Electron Configurations
- 7.9 The Building-Up Principle

Home work

p312: 7.3, 7.8, 7.16, 7.18

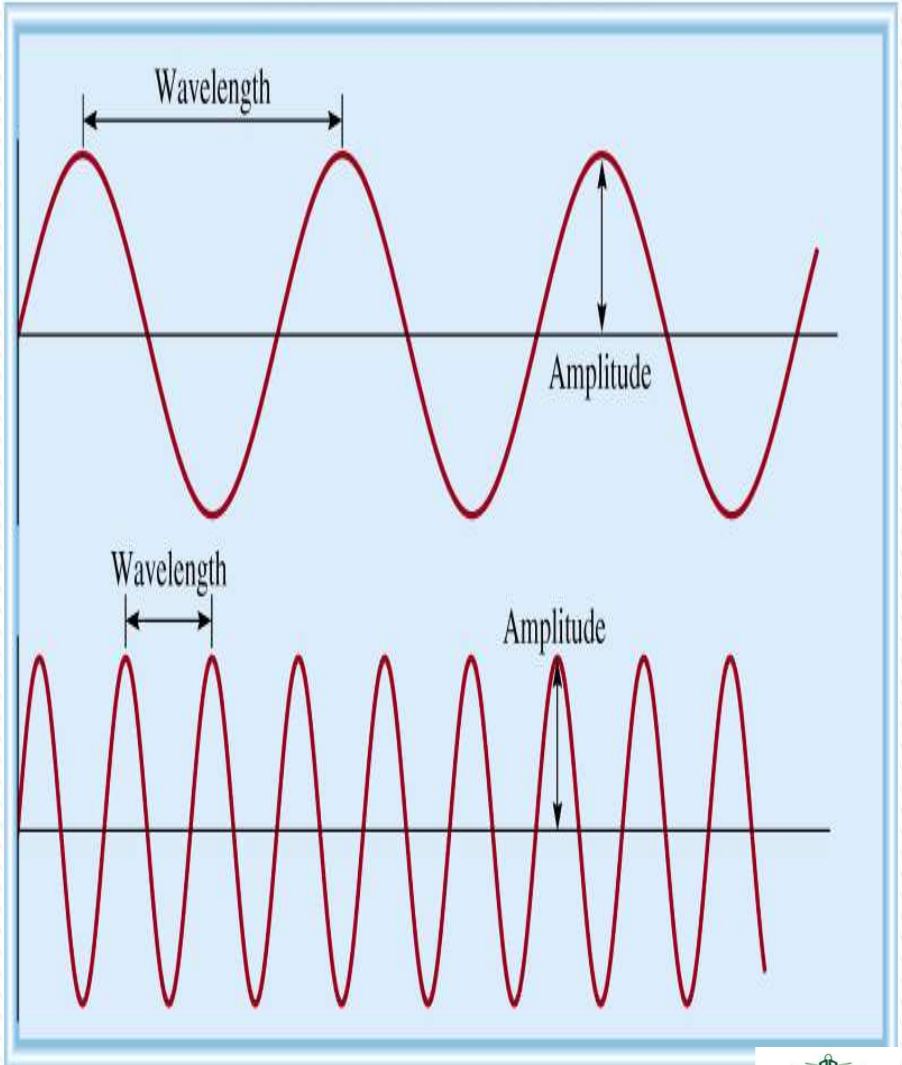
p313: 7.32, 7.34, 7.120

p314: 7.56, 7.58, 7.62, 7.66, 7.70

p315: 7.76, 7.78, 7.79, 7.84, 7.88, 7.90, 7.124

7.1 From Classical Physics to Quantum Theory

- Properties of Waves
- **Wavelength (λ)** is the distance between identical points on successive waves.
- **Amplitude** is the vertical distance from the midline of a wave to the peak or trough
- **Frequency (ν)** is the number of waves that pass through a particular point in 1 second (Hz = 1 cycle/s).



The speed (u) of the wave = λ x ν

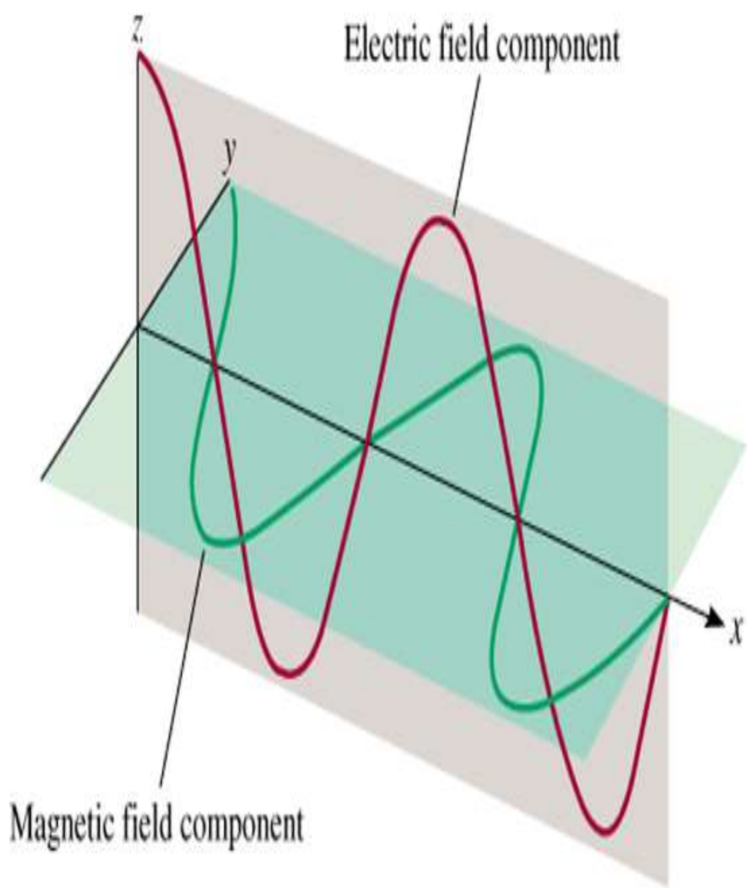
Electromagnetic radiation

- **Electromagnetic radiation** is the emission and transmission of energy in the form of electromagnetic waves.

- All electromagnetic radiation travels at the same velocity: the speed of light (*c*),

All electromagnetic radiation

$\lambda \times \nu = c$



Speed of light (*c*) in vacuum = 3.00×10^8 m/s

- **The wave length of the green light from a traffic signal is centered at 522nm . What is the frequency of this radiation?**

$$c = \lambda \times \nu$$

$$\nu = c/\lambda$$

$$= 3 \times 10^{8} / 522 \times 10^{-9}$$

$$= 5.57 \times 10^{14} \text{ Hz}$$

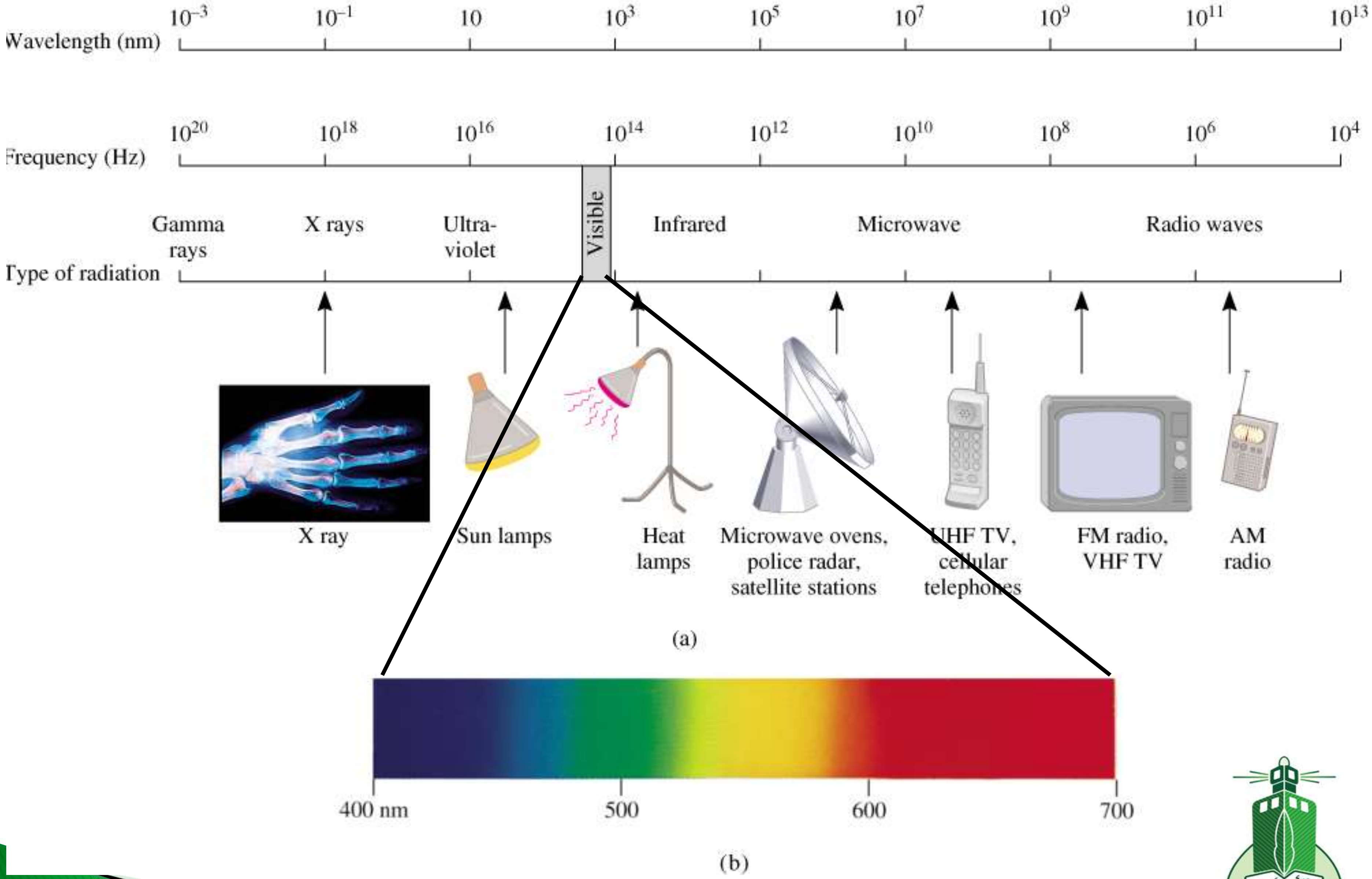
- **What is the λ in meter (m) of an electromagnetic waves whose frequncy is 3.64 x10^{7}Hz?**

$$c = \lambda \times \nu$$

$$\lambda = c/\nu$$

$$= 3 \times 10^{8} / 3.64 \times 10^{7}$$

$$= 7.25 \text{ m}$$


Wavelength (nm)
10^−3 10^−1 10 10^3 10^5 10^7 10^9 10^11 10^13
Frequency (Hz)
10^20 10^18 10^16 10^14 10^12 10^10 10^8 10^6 10^4
Type of radiation
Gamma rays
X rays
Ultra-violet
Visible
Infrared
Microwave
Radio waves
X ray
Sun lamps
Heat lamps
Microwave ovens, police radar, satellite stations
UHF TV, cellular telephones
FM radio, VHF TV
AM radio
(a)
400 nm
500
600
700
(b)

Planck's Quantum Theory

- Energy (light) is emitted or absorbed in discrete units (quantum).

$E = h \times \nu$

$E = h \times c/\lambda$

Planck's constant (h)
$h = 6.63 \times 10^{-34}$ J.s

Example 7.2

DCM, Worked Examples

Calculate the energy (in joules) of (a) a photon with a wavelength of 5.00×10^4 nm (infrared region) and (b) a photon with a wavelength of 5.00×10^{-2} nm (X ray region).

Strategy In both (a) and (b) we are given the wavelength of a photon and asked to calculate its energy. We need to use Equation (7.3) to calculate the energy. Planck's constant is given in the text and also on the back inside cover.

Solution (a) From Equation (7.3),

$$E = h\frac{c}{\lambda}$$

$$= \frac{(6.63 \times 10^{-34}\ \text{J}\cdot\text{s})(3.00 \times 10^8\ \text{m/s})}{(5.00 \times 10^4\ \text{nm})\dfrac{1 \times 10^{-9}\ \text{m}}{1\ \text{nm}}}$$

$$= 3.98 \times 10^{-21}\ \text{J}$$

This is the energy of a single photon with a 5.00×10^4 nm wavelength.

(b) Following the same procedure as in (a), we can show that the energy of the photon that has a wavelength of 5.00×10^{-2} nm is 3.98×10^{-15} J.

Check Because the energy of a photon increases with decreasing wavelength, we see that an "X-ray" photon is 1×10^6, or a million times, more energetic than an "infrared" photon.

Practice Exercise The energy of a photon is 5.87×10^{-20} J. What is its wavelength (in nanometers)?

Example

When copper is bombarded with high-energy electrons, X rays are emitted. Calculate the energy (in joules) associated with the photons if the wavelength of the X rays is 0.154 nm.

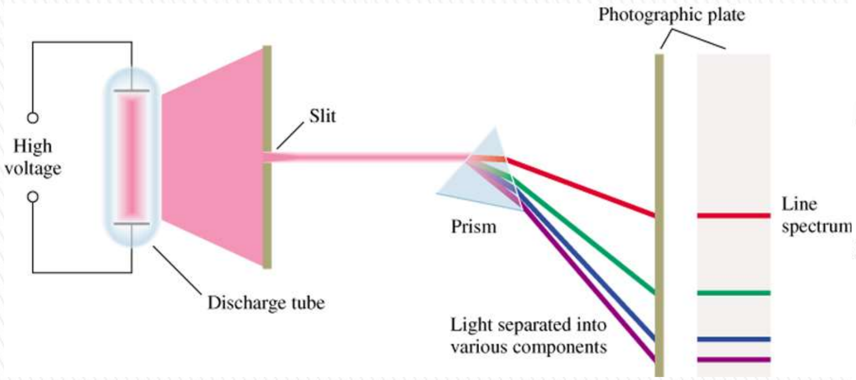
$$E = h \times \nu$$

$$E = h \times c / \lambda$$

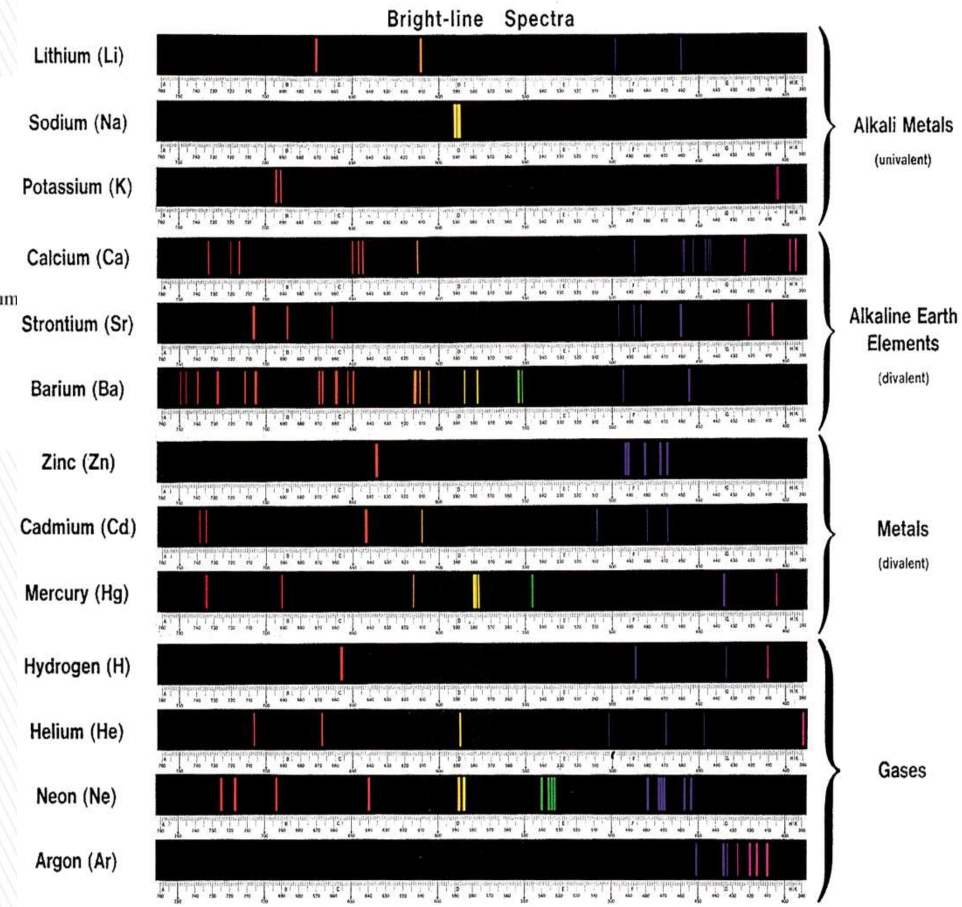
$$E = 6.63 \times 10^{-34} (J \cdot \cancel{s}) \times 3.00 \times 10^{8} (\cancel{m}/\cancel{s}) / 0.154 \times 10^{-9} (\cancel{m})$$

$$E = 1.29 \times 10^{-15}\ J$$

7.3Bohr's Theory of the Hydrogen Atom



Line Emission Spectrum of
Hydrogen Atoms



Emission Spectra

Emission Spectra of the Hydrogen atom

Dr Laila Al-Harbi

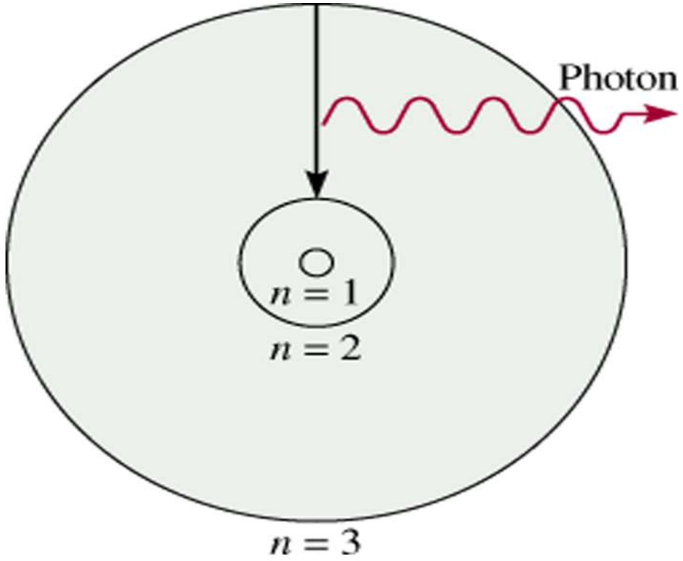
Bohr's Model of the Atom (1913)

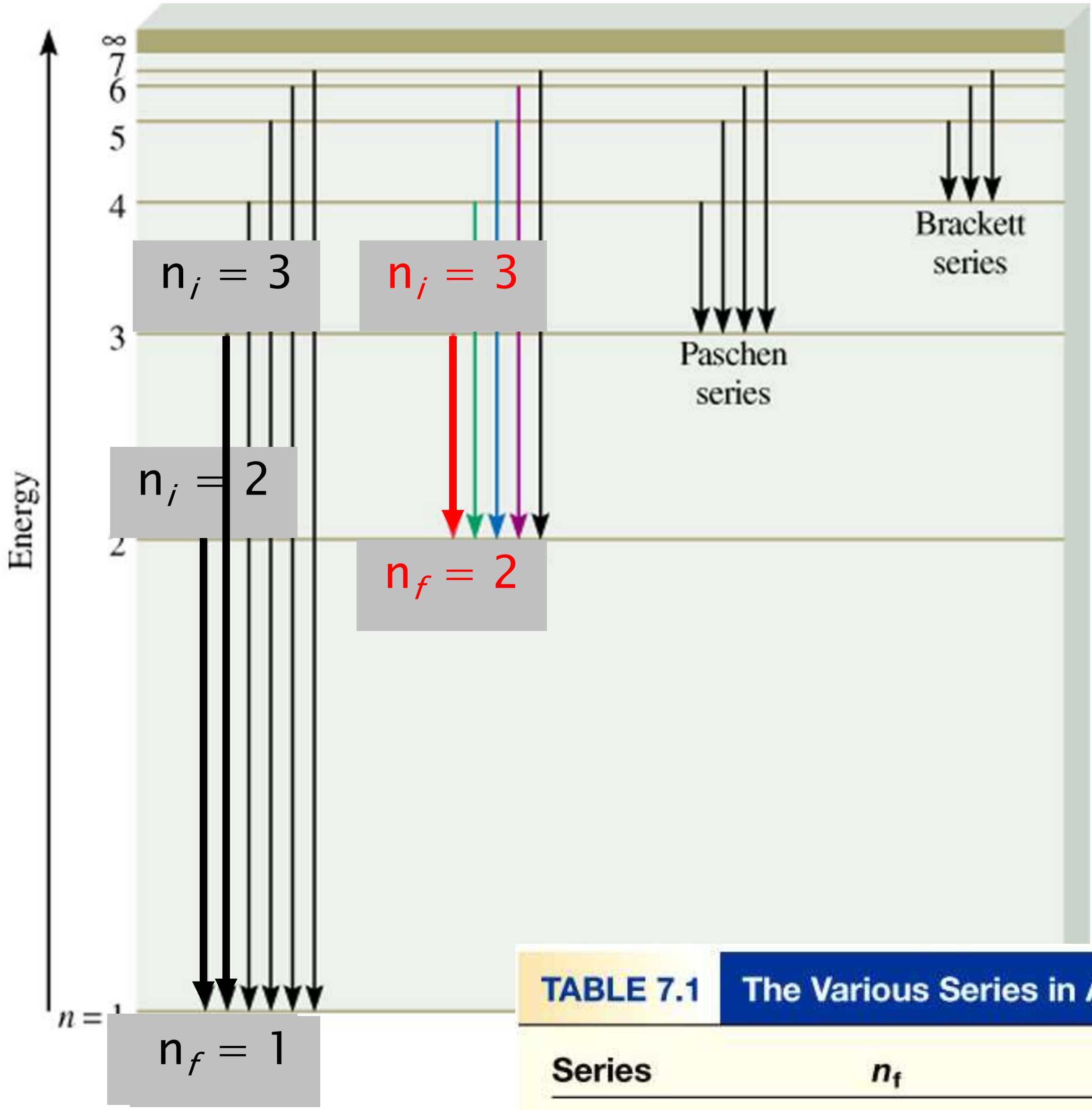
1. e^- can only have specific (quantized) energy values
2. light is emitted as e^- moves from one energy level to a lower energy level

$$E_n = -R_H\left(\frac{1}{n^2}\right)$$

n (principal quantum number) = 1,2,3,…

R_H (Rydberg constant) = 2.18×10^{-18}J





$$E_{photon} = \Delta E = E_f - E$$

$$E_f = -R_H\left(\frac{1}{n_f^2} \right)$$

$$E_i = -R_H\left(\frac{1}{n_i^2} \right)$$

$$\Delta E = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

TABLE 7.1 The Various Series in Atomic Hydrogen Emission Spectrum

Series	n_f	n_i	Spectrum Region
Lyman	1	2, 3, 4, . . .	Ultraviolet
Balmer	2	3, 4, 5, . . .	Visible and ultraviolet
Paschen	3	4, 5, 6, . . .	Infrared
Brackett	4	5, 6, 7	Infrared

$$\Delta E = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

- When photon is emitted
- $n_i > n_f$
- ΔE is negative
- Energy is lost to the surrounding

- When photon is absorbed
- $n_i < n_f$
- ΔE is positive
- Energy is gained from the surrounding

Example 7.4

DCM, Worked Examples

What is the wavelength of a photon (in nanometers) emitted during a transition from the $n_i = 5$ state to the $n_f = 2$ state in the hydrogen atom?

Strategy We are given the initial and final states in the emission process. We can calculate the energy of the emitted photon using Equation (7.6). Then from Equations (7.2) and (7.1) we can solve for the wavelength of the photon. The value of Rydberg's constant is given in the text.

Solution From Equation (7.6) we write

$$\begin{aligned}\Delta E &= R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)\\ &= 2.18 \times 10^{-18}\,\text{J}\left(\frac{1}{5^2} - \frac{1}{2^2}\right)\\ &= -4.58 \times 10^{-19}\,\text{J}\end{aligned}$$

The negative sign indicates that this is energy associated with an emission process. To calculate the wavelength, we will omit the minus sign for ΔE because the wavelength of the photon must be positive. Because $\Delta E = h\nu$ or $\nu = \Delta E/h$, we can calculate the wavelength of the photon by writing

$$\begin{aligned}\lambda &= \frac{c}{\nu}\\ &= \frac{ch}{\Delta E}\\ &= \frac{(3.00 \times 10^8\,\text{m/s})(6.63 \times 10^{-34}\,\text{J}\cdot\text{s})}{4.58 \times 10^{-19}\,\text{J}}\\ &= 4.34 \times 10^{-7}\,\text{m}\\ &= 4.34 \times 10^{-7}\,\text{m} \times \left(\frac{1\,\text{nm}}{1 \times 10^{-9}\,\text{m}}\right) = 434\,\text{nm}\end{aligned}$$

Check The wavelength is in the visible region of the electromagnetic region (see Figure 7.4). This is consistent with the fact that because $n_f = 2$, this transition gives rise to a spectral line in the Balmer series (see Figure 7.6).

Practice Exercise What is the wavelength (in nanometers) of a photon emitted during a transition from $n_i = 6$ to $n_f = 4$ state in the H atom?

7.6 Quantum Numbers

- Electrons in multi-electron atoms can be classified into a series of:

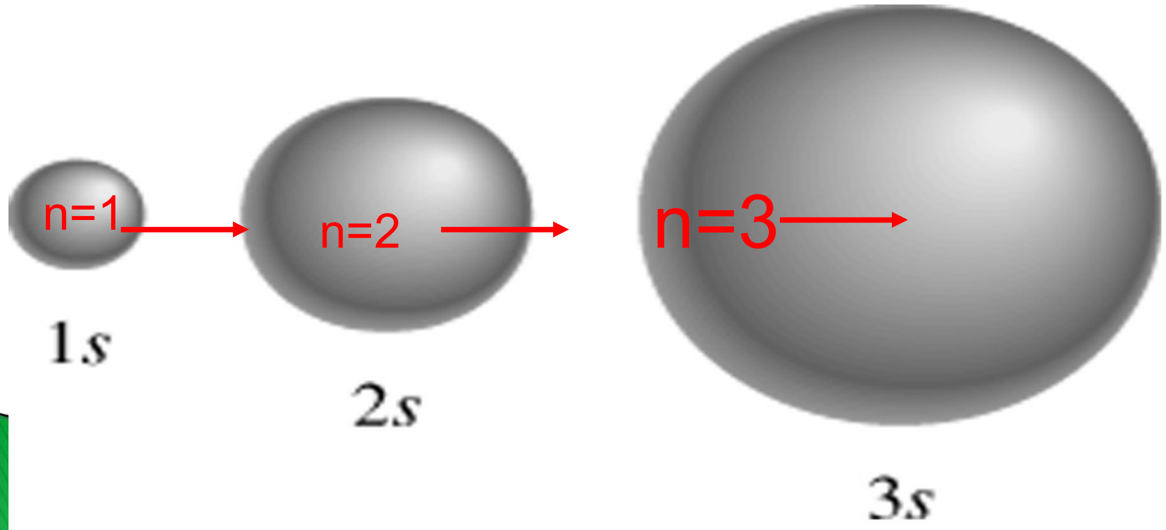
  shells → subshells → orbitals

- Three quantum numbers are required to describe the distribution of electrons in hydrogen and other atoms.

- A fourth quantum the spin quantum number – describe the behavior of a specific electron and completes the description of electron in the atom

Principal Quantum Number, *n*

- The principal quantum number, *n*, describes the energy level on which the orbital resides and to distance from nucleus (size)
- The maximum number of electrons in principle quantum number $n = 2n^2$
- The values of n are integers ≥ 0.
- possible values of ***n* = 1, 2, 3, 4, ...**



Angular momentum Quantum Number, ℓ

- related to shape of various **subshells within a given shell**

- Allowed values of ℓ are integers ranging from 0 to $n - 1$.

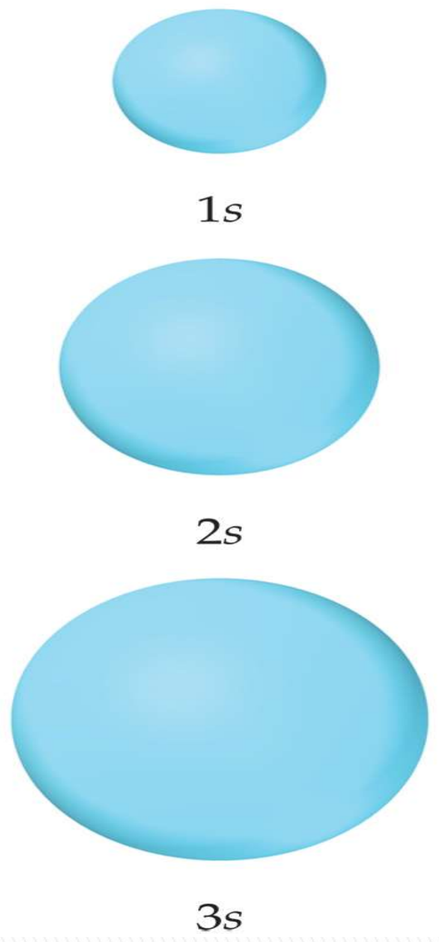
- We use letter designations to communicate the different values of ℓ and, therefore, the shapes and types of orbitals.

Value of ℓ	**0**	**1**	**2**	**3**
Type of orbital	**s**	**p**	**d**	**f**

values of *n*	values of ℓ	orbitals
1	0	1*s*
2	0, 1	2*s*, 2*p*
3	0, 1, 2	3*s*, 3*p*, 3*d*

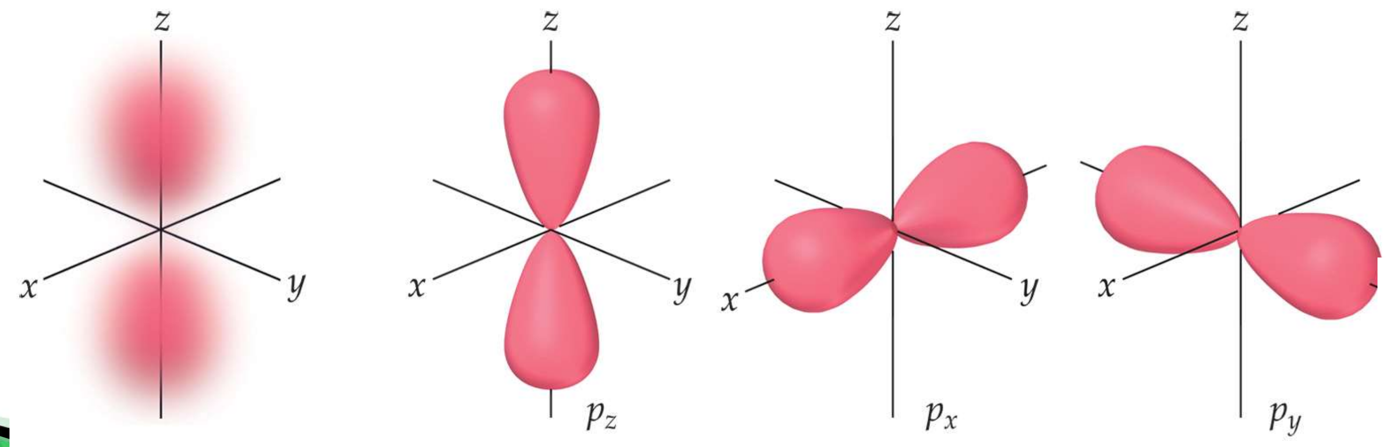
7.7 Atomic orbitals

- *s* Orbitals
- Value of $\ell = 0$.
- Spherical in shape.
- Radius of sphere increases with increasing value of *n*.



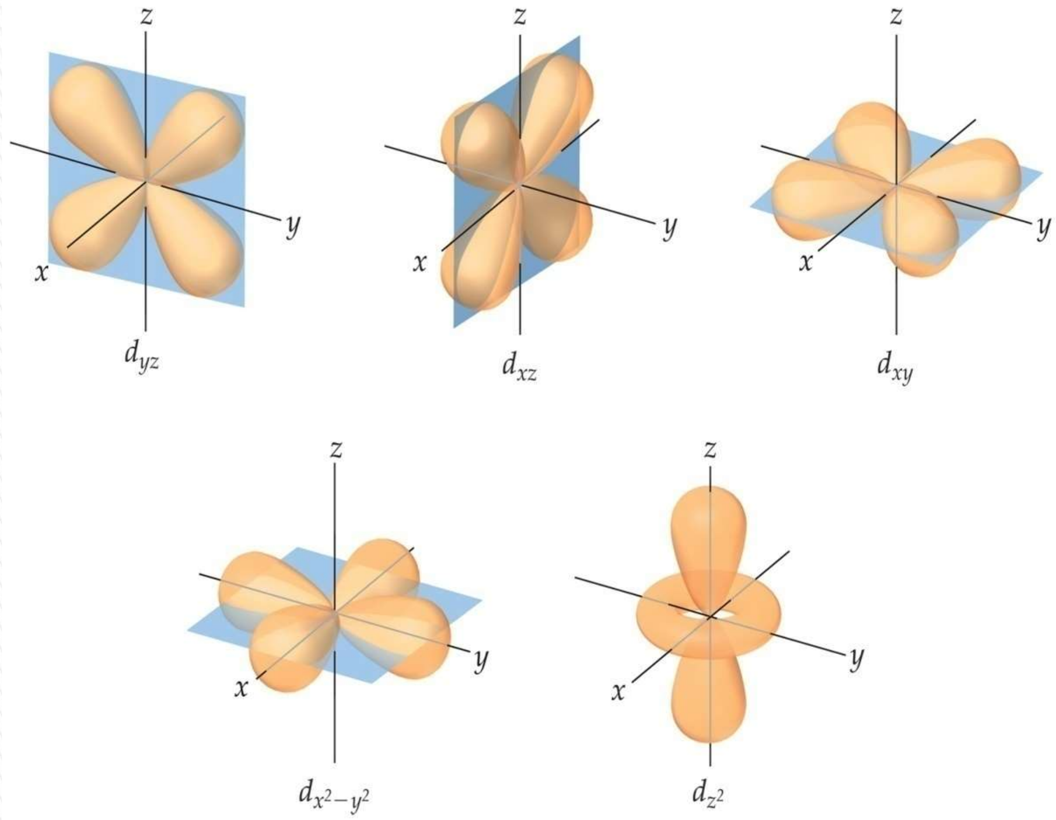
p Orbitals

- Value of $\ell = 1$.
- Have two lobes with a node between them.
- $m_\ell = 2\ell + 1 = 2\times1+1 = 3$
- Value of $m\ell$ = 1,0,-1 (P_x , P_z , P_y)



d Orbitals

- Value of ℓ is 2.
- $m_\ell = 2\ell + 1 = 2\times2+1 = 5$
- Value of $m\ell$ = 2,1,0,-1,-2
- (d_{xy} , d_{zy} , d_{xz} ,d_{z^2} $d_{x^2-y^2}$)
- Complex stracture



Magnetic Quantum Number, m_ℓ

- Describes the three-dimensional **orientation** of the orbital.
- Values are integers ranging from - ℓ to ℓ :

$$-\ell \le m_l \le \ell.$$

- The number of orbitals in each subshell ℓ *equal* $=2\ell+1$

$$m_\ell = 2\ell + 1$$

- Therefore, on any given energy level, there can be up to 1 *s* orbital, 3 *p* orbitals, 5 *d* orbitals, 7 *f* orbitals, etc

Value of ℓ	**0 (s)**	**1(p)**	**2(d)**	**3(f)**
Value of mℓ	0	-1, 0 ,+1	-1, 0 ,+1	-1, 0 ,+1
№ of orbitals	2(0)+1=1 2e	2(1)+1=3 6e	2(2)+1=5 10e	2(3)+1=7 14e

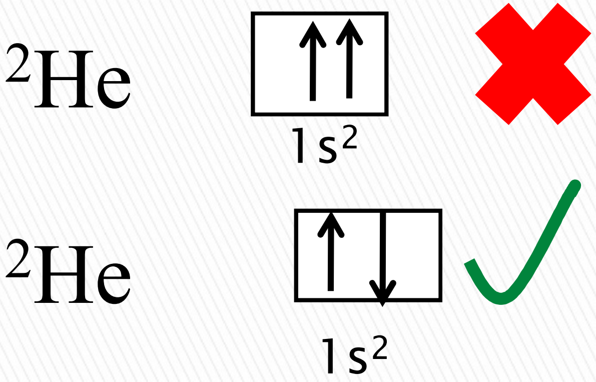
The Pauli exclusion principle

- The “spin” of an electron describes its magnetic field, which affects its energy.

- **The spin quantum number has only 2 allowed values:**
  m_s = **+1/2 and −1/2.**

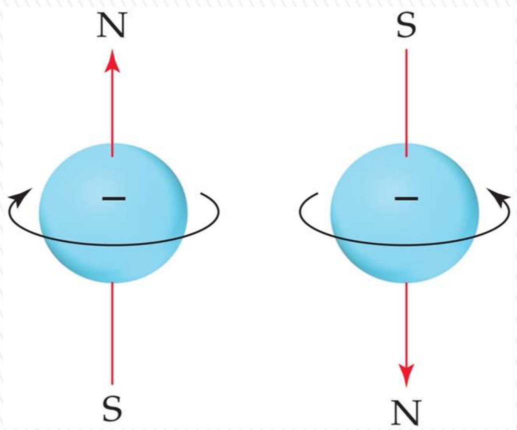
- **The Pauli exclusion principle**
- **“No two electrons in the same atom can have identicalvalues for all four of their quantum numbers”.**

Orbital diagram



n =1 , ℓ=0 , mℓ=0 , ms = +½

n =1 , ℓ=0 , mℓ=0 , ms = -½



Dr Laila Al-Harbi

Quantum Numbers

Principal Quantum Number	Angular momentum Quantum Number	Magnetic Quantum Number	Spin Quantum Number
n	**ℓ**	**mℓ**	ms
the energy level (size)	shape of various **subshells**	orientation of the orbital	The "spin" of an electron
integers ≥ 0 n = 1, 2, 3, 4, .	from 0 to n − 1	−ℓ ≤ ml ≤ ℓ	+1/2 and −1/2

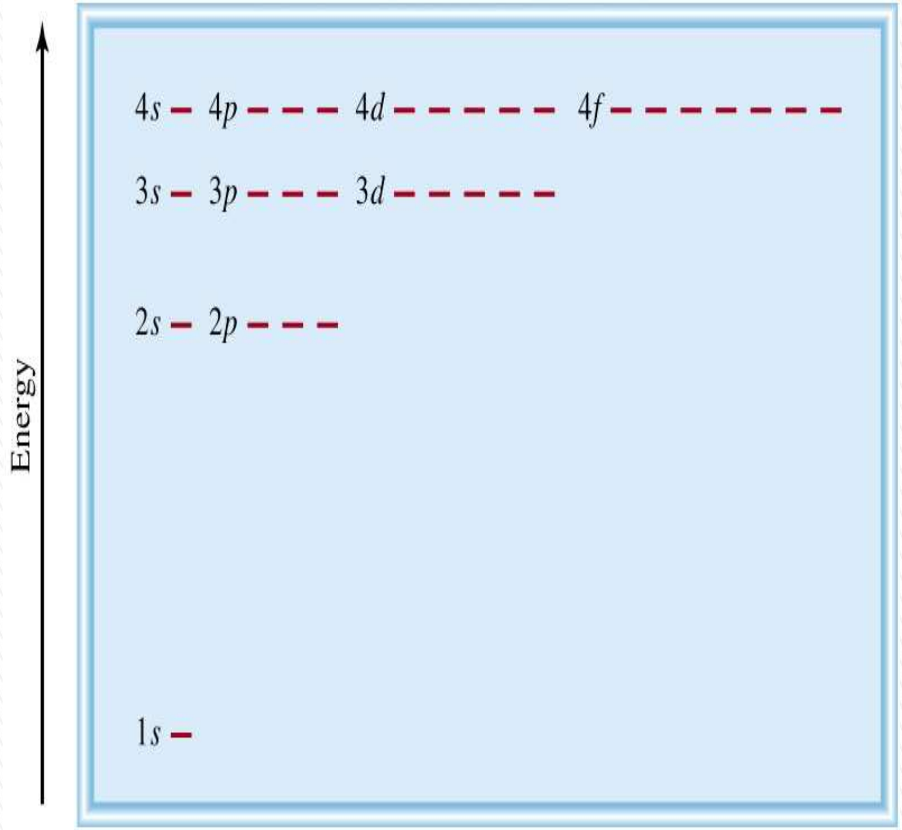
№ of electrons $n = 2n^2$

Value of ℓ	**0**	**1**	**2**	**3**
Type of orbital	**s**	**p**	**d**	**f**

№ of orbitals in each subshell ℓ equal =2 ℓ +1

Energies of Orbitals

- For a one-electron hydrogen atom, orbitals on the same energy level have the same energy.
- That is, they are degenerate.



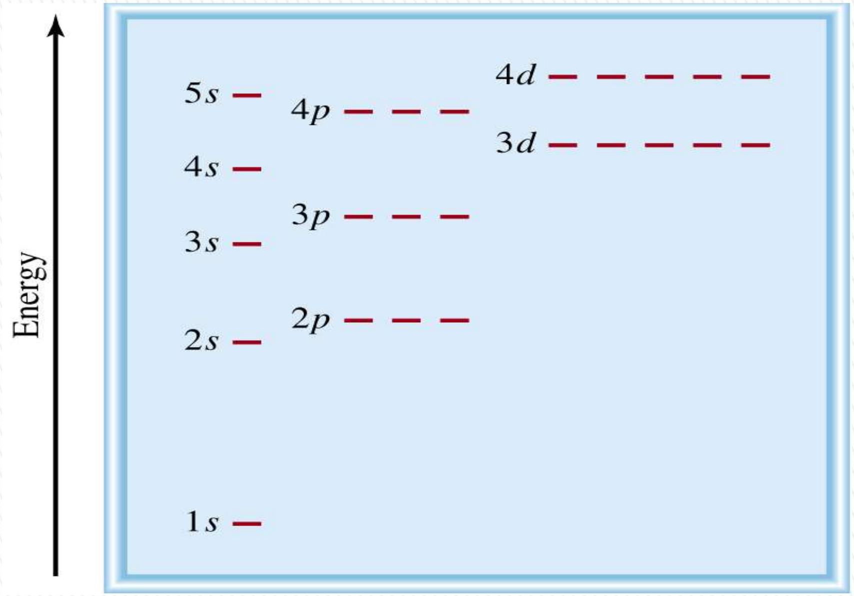
- The energies of H orbitals increase as follows
- $1s<2s<3s=3p=3d<4s=4p4d=4f$

The Shielding Effect (many electron atoms)

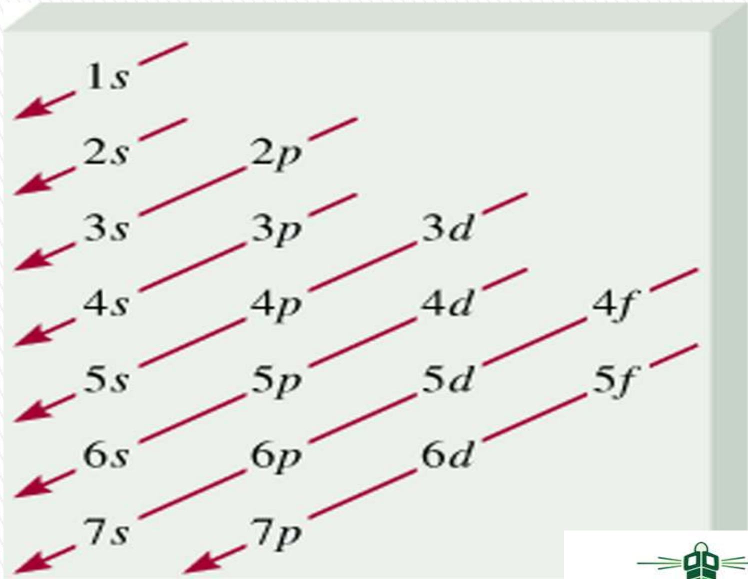
- Electrons in the smaller orbitals (lower energy) are closerto nucleus (e.g., 1s) than electrons in larger orbitals (e.g., 2p, 3s)
- Thus they are "shielded" from the attractive forces of the nucleus.
- This causes slight increase in energy of the more distant electrons.
- thus 4s orbital is lower in energy than the 3d orbital .

- As the number of electrons increases, though, so does the repulsion between them.

- Therefore, in many-electron atoms, orbitals on the same energy level are no longer degenerate.



- Order of orbitals (filling) in multi-electron atom
- 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s

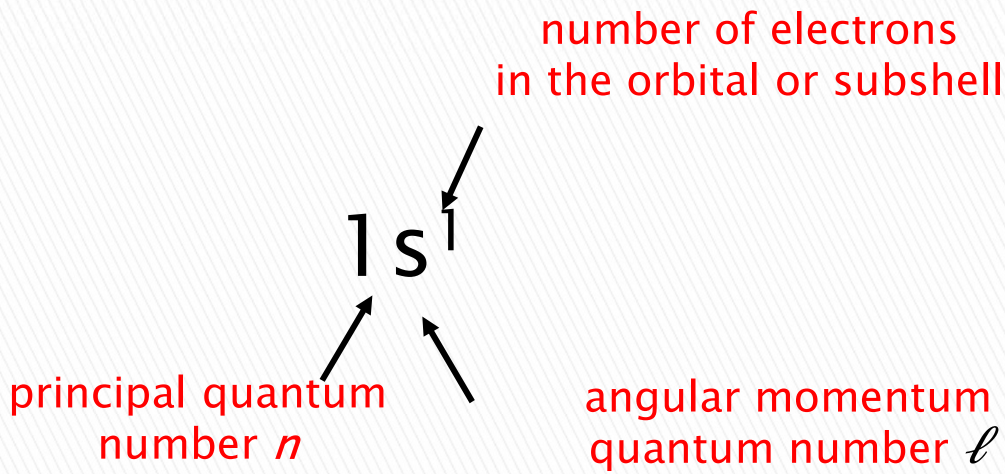


7.8 Electron configuration

- is how the electrons are distributed among the various atomic orbitals in an atom.

- The four quantum numbers n, ℓ, m_ℓ and m_s enable us to label completely an electron in any orbital in any atom.

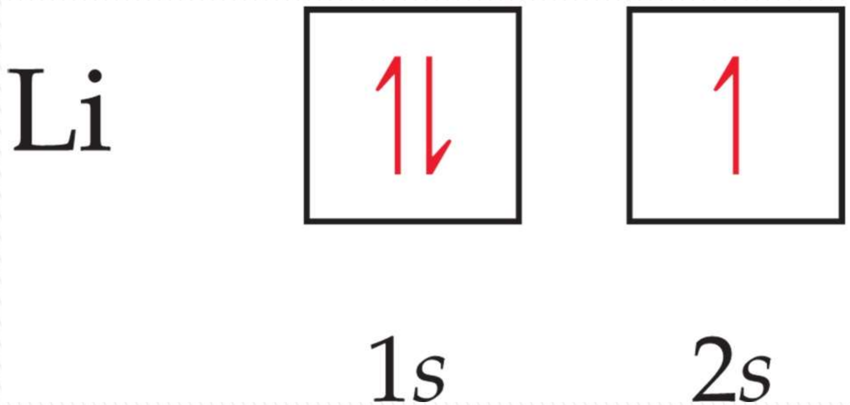
- The value of m_s has no effect on the energy ,size, shape , or orintation of an orbital, but it determines , how electron are arranged in an orbital.



Orbital diagram

Orbital Diagrams

- Each box represents one orbital.
- Half-arrows represent the electrons.
- The direction of the arrow represents the spin of the electron.



Example 7.6

DCM, Worked Examples

List the values of n, ℓ, and m_ℓ for orbitals in the 4*d* subshell.

$n = 4$, $\ell = 2$, $m_\ell = -2, -1, 0, 1, 2$

Practice Exercise Give the values of the quantum numbers associated with the orbitals in the 3*p* subshell.

$n = 3$, $\ell = 1$, $m_\ell = -1, 0, 1$

Example 7.7

DCM, Worked Examples

What is the total number of orbitals associated with the principal quantum number $n = 3$?

n = 3, ℓ = 0 , mℓ =0

n = 3, ℓ = 1, mℓ = -1, 0, 1

n = 3, ℓ = 2 , mℓ =-2, -1, 0, 1,2

9

Total number of orbitals = $n^2 = (3)^2 = 9$

Total number elecrons in orbitals = $2n^2 = 2(3)^2 = 18$ electrons

Practice Exercise What is the total number of orbitals associated with the principal quantum number $n = 4$?

n = 4, ℓ = 0 , mℓ =0

n = 4, ℓ = 1, mℓ = -1, 0, 1

n = 4, ℓ = 2 , mℓ =-2, -1, 0, 1,2

n = 4, ℓ = 3 , mℓ =03,-2, -1, 0, 1,2,3

16

Total number of orbitals = $n^2 = (4)^2 = 16$

Total number elecrons in orbitals = $2n^2 = 2(4)^2 = 32$ electrons

Example 7.8

DCM, Worked Examples

Write the four quantum numbers for an electron in a $3p$ orbital.

Strategy What do the "3" and "p" designate in $3p$? How many orbitals (values of m_ℓ) are there in a $3p$ subshell? What are the possible values of electron spin quantum number?

Solution To start with, we know that the principal quantum number n is 3 and the angular momentum quantum number ℓ must be 1 (because we are dealing with a p orbital).

For $\ell = 1$, there are three values of m_ℓ given by -1, 0, and 1. Because the electron spin quantum number m_s can be either $+\frac{1}{2}$ or $-\frac{1}{2}$, we conclude that there are six possible ways to designate the electron using the (n, ℓ, m_ℓ, m_s) notation:

$$(3, 1, -1, +\tfrac{1}{2}) \qquad (3, 1, -1, -\tfrac{1}{2})$$
$$(3, 1, 0, +\tfrac{1}{2}) \qquad (3, 1, 0, -\tfrac{1}{2})$$
$$(3, 1, 1, +\tfrac{1}{2}) \qquad (3, 1, 1, -\tfrac{1}{2})$$

Check In these six designations we see that the values of n and ℓ are constant, but the values of m_ℓ and m_s can vary.

Practice Exercise Write the four quantum numbers for an electron in a $5p$ orbital.

Example 7.9

DCM, Worked Examples

What is the maximum number of electrons that can be present in the principal level for which $n = 3$?

Solution When $n = 3$, $\ell = 0$, 1, and 2. The number of orbitals for each value of ℓ is given by

Value of ℓ	Number of Orbitals $(2\ell + 1)$
0	1
1	3
2	5

The total number of orbitals is nine. Because each orbital can accommodate two electrons, the maximum number of electrons that can reside in the orbitals is 2×9, or 18.

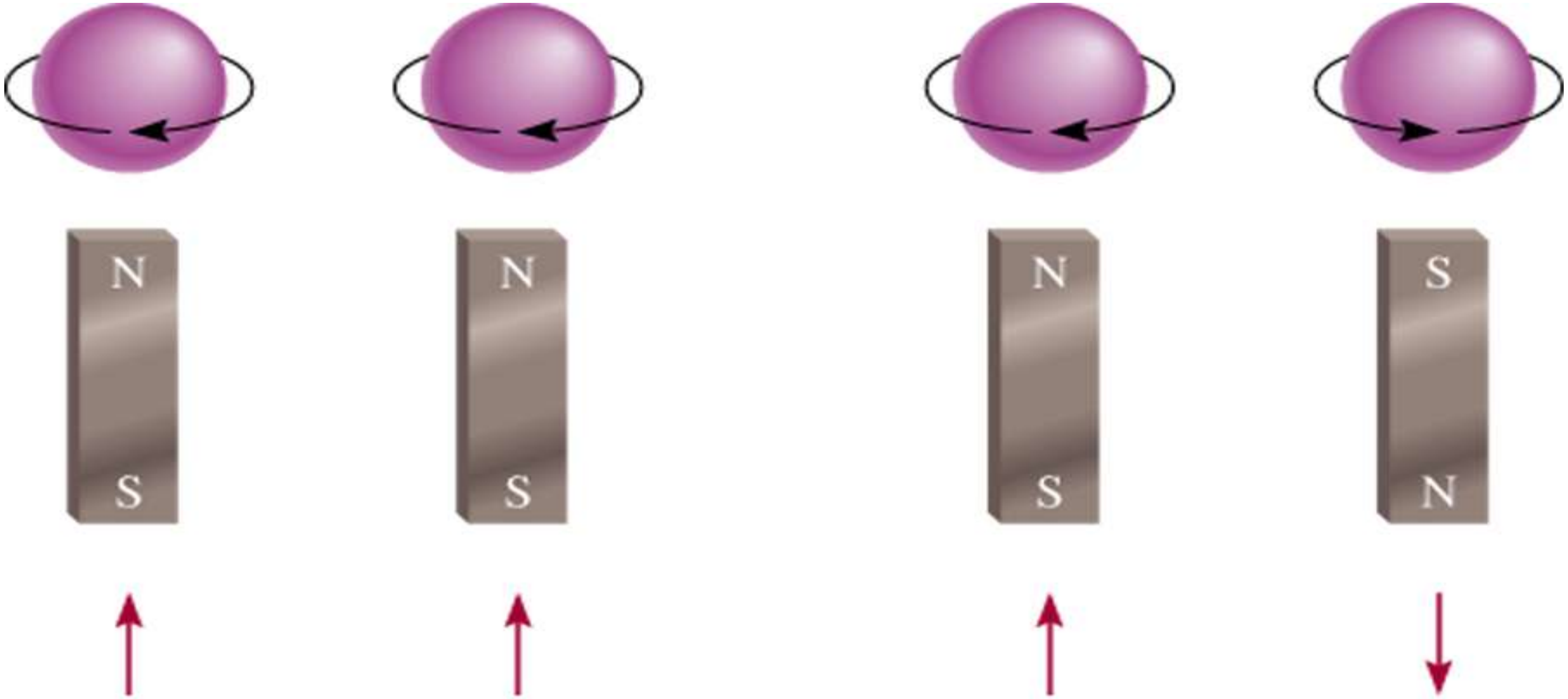
Total number of orbitals = $n^2 = (3)^2 = 9$

Practice Exercise Calculate the total number of electrons that can be present in the principal level for which $n = 4$.

Value of ℓ	Number of Orbitals $(2\ell + 1)$
0	1
1	3
2	5
3	**7**

Total number of orbitals = $n^2 = (4)^2 = 16$

Diamagnetism and paramagnetism



Paramagnetic

unpaired electrons

↑__ ↑__ ___

2p

attracted by a magnetic

Diamagnetic

all electrons paired

↑↓ ↑↓ ↑↓

2p

repelled by magnet

1-Which of the following is paramagnetic

- Mg
- Ar
- He
- N

2- Which of the following is dimagnatic

- Mg
- Na
- N

3- How many unpaired electrons in N atom

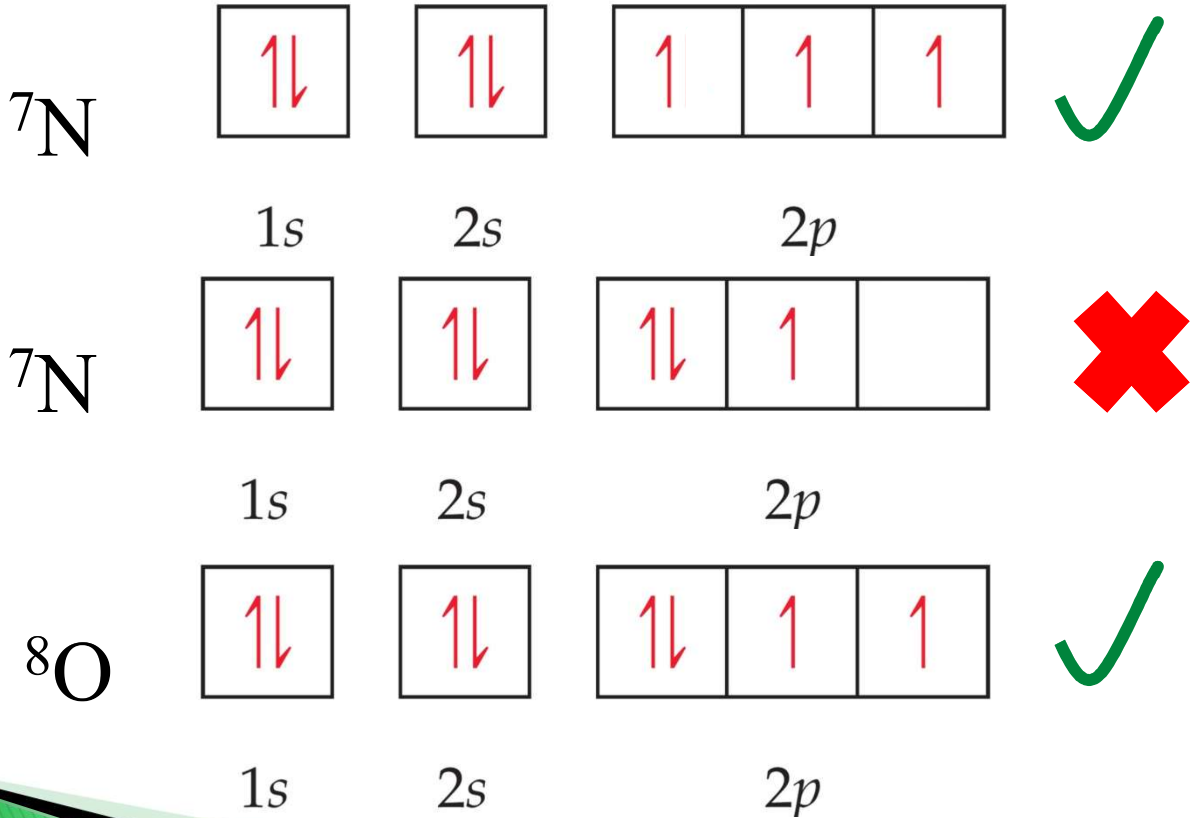
a) 2
b) 5
c) 4
d) 3

4- How many unpaired electrons in Mg

a) 2
b) 0
c) 4
d) 3

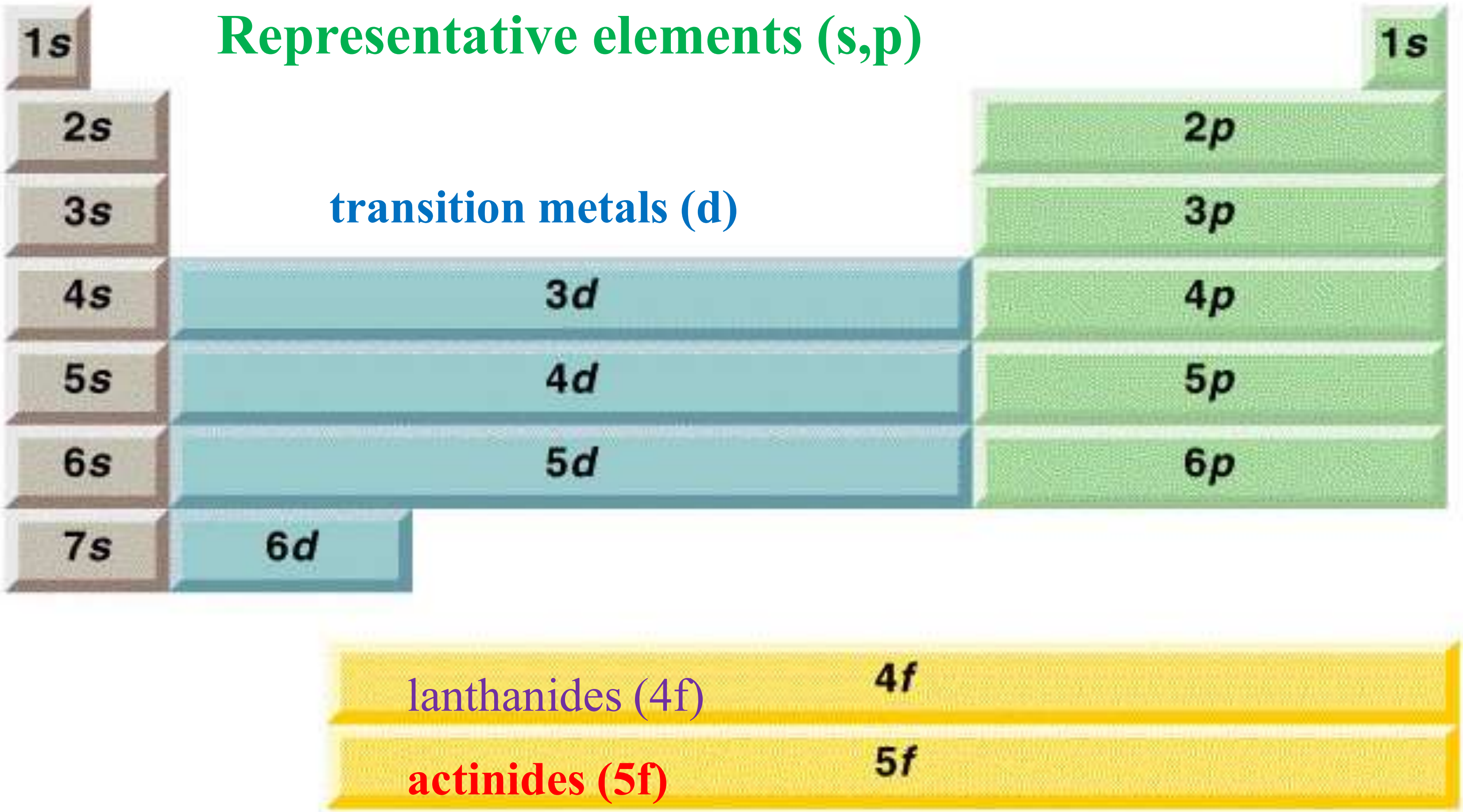
Hund's Rule

- The most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins



7.9 The Building-Up Principle

- The Aufbau principle dictates that as protons are added one to the nucleus to build up the elements, electrons are similarly added to the atomic orbitals.



Ions derived from representative

Ions derived from main group elements lose or gain electrons to have noble gas electron configuration $ns^2\ np^6$.

isoelectronic – are atoms or ions have the same number of electrons

Elements in same group have same valence shell configurations

e.g., group V:

^{7}N	$2s^2\ 2p^3$
^{15}P	$3s^2\ 3p^3$
^{33}As	$4s^2\ 4p^3$
^{51}Sb	$5s^2\ 5p^3$
^{83}Bi	$6s^2\ 6p^3$

ELECTRON CONFIGURATIONS - THE GROUND STATE

$_1H$	$1s^1$	1 A	General Configurations of 1A
$_3Li$	$1s^2\ 2s^1$		ns^1
$_{11}Na$	$1s^22s^22p^63s^1$		
$_{19}K$	$1s^22s^22p^63s^23p^64s^1$		Paramagnetic , 1 unpaired electrons
$_4Be$	$1s^22s^2$	2 A	General Configurations of 2A
$_{12}Mg$	$1s^22s^22p^63s^2$		ns^2
$_{20}Ca$	$1s^22s^22p^63s^23p^64s^2$		Diamagnetic , 0 unpaired electrons
$_5B$	$1s^22s^22p^1$	3 A	General Configurations of 3A
$_{13}Al$	$1s^22s^22p^63s^23p^1$		ns^2np^1
			Paramagnetic , 1 unpaired electrons
$_6C$	$1s^22s^22p^2$	4 A	ns^2np^2
$_{14}Si$	$1s^22s^22p^63s^23p^2$		Paramagnetic , 2 unpaired electrons

ELECTRON CONFIGURATIONS - THE GROUND STATE

$_7N$	$1s^2 2s^2 2p^3$	5 A	ns^2np^3
$_{15}P$	$1s^2 2s^2 2p^6 3s^2 3p^3$		Paramagnetic , 3 unpaired electrons
$_8O$	$1s^2 2s^2 2p^4$	6 A	ns^2np^4
$_{16}S$	$1s^2 2s^2 2p^6 3s^2 3p^4$		Paramagnetic , 2 unpaired electrons
^9F	$1s^2 2s^2 2p^5$	7 A	ns^2np^5
$_{17}Cl$	$1s^2 2s^2 2p^6 3s^2 3p^5$		Paramagnetic , 1 unpaired electrons
$_2He$	$1s^2$	8 A	ns^2np^6
$_{10}Ne$	$1s^2 2s^2 2p^6$		
$_{18}Ar$	$1s^2 2s^2 2p^6 3s^2 3p^6$		Diamagnetic , 0 unpaired electrons

ELECTRON CONFIGURATIONS - ions

ions

ns^2

ns^2np^6

Ion	Configuration	Group
$_{1}H^{-1}$	$1s^2$ = [He]	
$_{3}Li^{+1}$	$1s^2$ = [He]	1 A
$_{11}Na^{+1}$	$1s^22s^22p^6$ = [Ne]	Lose 1e
$_{19}K^{+1}$	$1s^22s^22p^63s^23p^6$ = [Ar]	+1

Paramagnetic , 1 unpaired electrons

Ion	Configuration	Group
$_{4}Be^{+2}$	$1s^2$ = [He]	2 A
$_{12}Mg^{+2}$	$1s^22s^22p^6$ = [Ne]	Lose 2e +2

General Configurations of 2A

ns^2np^6

Diamagnetic , 0 unpaired electrons

Ion	Configuration	Group
$_{5}B^{+3}$	$1s^2$ = [He]	3 A
$_{13}Al^{+3}$	$1s^22s^22p^6$ = [Ne]	Lose 3e +3

General Configurations of 3A

ns^2np^6

Diamagnetic , 0 unpaired electrons

Atom	Configuration	Group
$_{6}C$	$1s^22s^22p^2$	4 A
$_{14}Si$	$1s^22s^22p^63s^23p^2$	±4

ns^2np^6

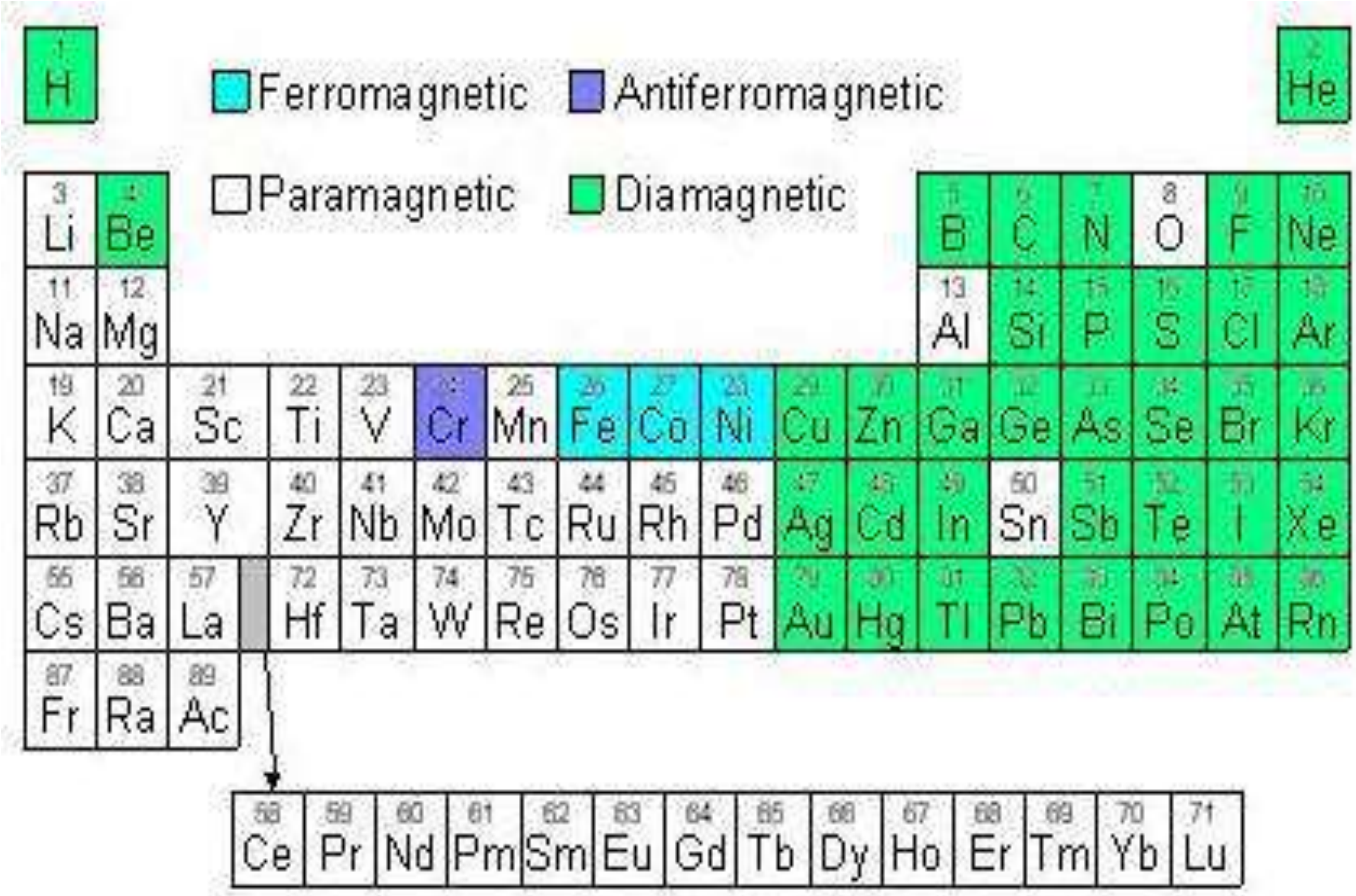
Diamagnetic , 0 unpaired electrons

ELECTRON CONFIGURATIONS - THE GROUND STATE

$_{7}N^{-3}$	$1s^22s^22p^6$ = [Ne]	**5 A** gain 3e -3	ns^2np^6 Diamagnetic , 0 unpaired electrons
$_{15}P^{-3}$	$1s^22s^22p^63s^23p^6$ = [Ar]		
$_{8}O^{-2}$	$1s^22s^22p^6$ = [Ne]	**6 A** gain 2e -2	ns^2np^6 Diamagnetic , 0 unpaired electrons
$_{16}S^{-2}$	$1s^22s^22p^63s^23p^6$ = [Ar]		
$^{9}F^{-}$	$1s^22s^22p^6$ = [Ne]	**7 A** gain 1e -1	ns^2np^6 Diamagnetic , 0 unpaired electrons
$_{17}Cl^{-}$	$1s^22s^22p^63s^23p^6$ = [Ar]		

Na^+, Al^{3+}, F^-, O^{2-}, and N^{3-} are all ***isoelectronic*** with Ne

$^{10}Na^+$, $^{10}Al^{3+}$, $^{10}F^-$, $^{10}O^{2-}$, and $^{10}N^{3-}$ *isoelectronic* with Ne


Ferromagnetic
Antiferromagnetic
Paramagnetic
Diamagnetic
H
He
Li
Be
B
C
N
O
F
Ne
Na
Mg
Al
Si
P
S
Cl
Ar
K
Ca
Sc
Ti
V
Cr
Mn
Fe
Co
Ni
Cu
Zn
Ga
Ge
As
Se
Br
Kr
Rb
Sr
Y
Zr
Nb
Mo
Tc
Ru
Rh
Pd
Ag
Cd
In
Sn
Sb
Te
I
Xe
Cs
Ba
La
Hf
Ta
W
Re
Os
Ir
Pt
Au
Hg
Tl
Pb
Bi
Po
At
Rn
Fr
Ra
Ac
Ce
Pr
Nd
Pm
Sm
Eu
Gd
Tb
Dy
Ho
Er
Tm
Yb
Lu

Electron Configurations of Cations and Anions Of Representative Elements

Na $[Ne]3s^1$ — Na^+ $[Ne]$

Ca $[Ar]4s^2$ — Ca^{2+} $[Ar]$

Al $[Ne]3s^23p^1$ — Al^{3+} $[Ne]$

Atoms lose electrons so that cation has a noble-gas outer electron configuration.

Atoms gain electrons so that anion has a noble-gas outer electron configuration.

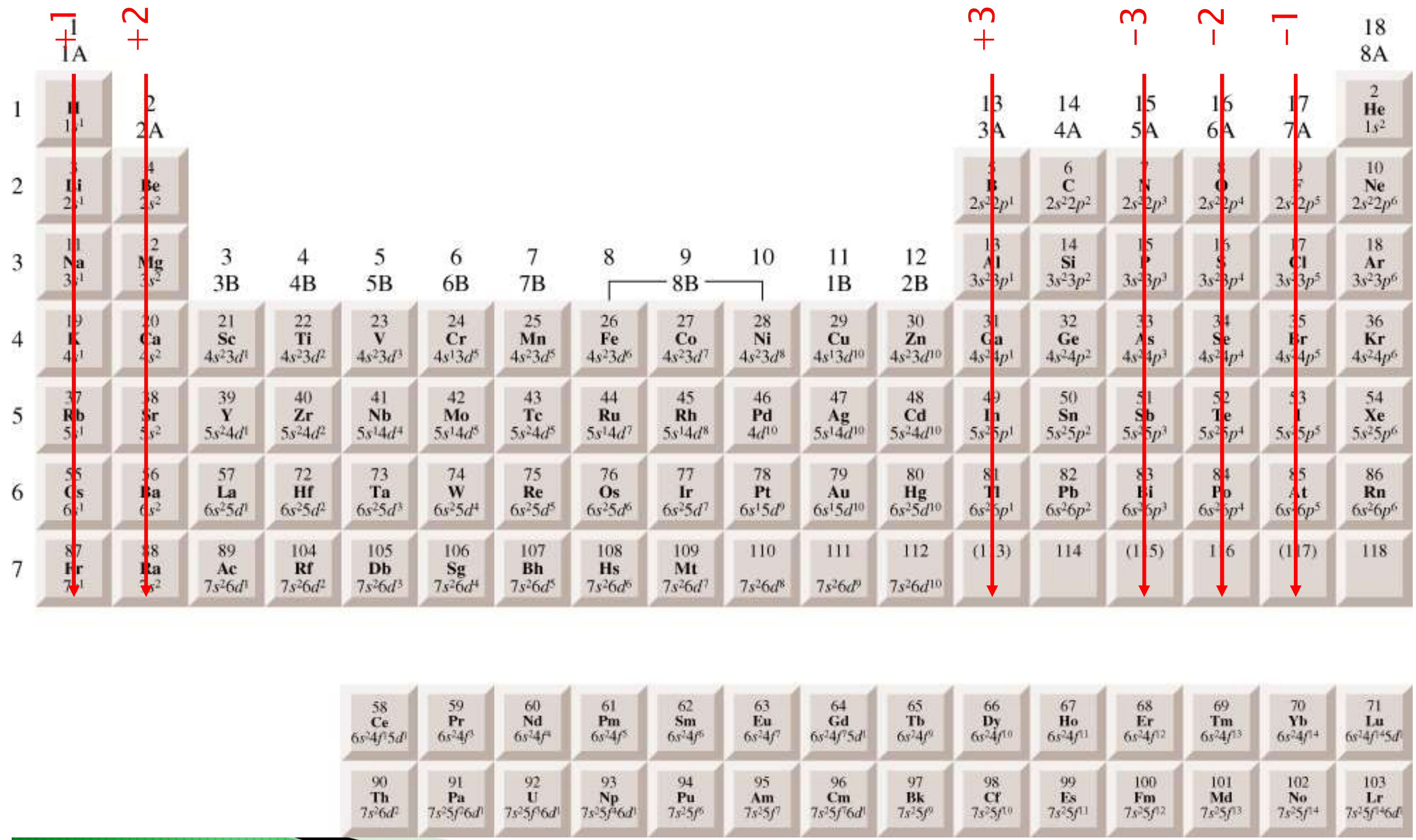
H $1s^1$ — H^- $1s^2$ or $[He]$

F $1s^22s^22p^5$ — F^- $1s^22s^22p^6$ or $[Ne]$

O $1s^22s^22p^4$ — O^{2-} $1s^22s^22p^6$ or $[Ne]$

N $1s^22s^22p^3$ — N^{3-} $1s^22s^22p^6$ or $[Ne]$

Cations and Anions Of Representative Elements



Group	General Configurations	unpaired	p/d	ions		
1A	ns^1	1	Paramagnetic	+1	Diamagnetic , 0 unpaired electrons	ns^2np^6
2A	ns^2	0	Diamagnetic	+2		
3A	ns^2np^1	1	Paramagnetic	+3		
4A	ns^2np^2	2	Paramagnetic	±4		
5A	ns^2np^3	3	Paramagnetic	−3		
6A	ns^2np^4	2	Paramagnetic	−2		
7A	ns^2np^5	1	Paramagnetic	−1		

- The general formula of an element in group IA is

A. S^2

B. S^1

C. S^2p^1

D. s^1p^1

- Which of the following species is isoelectronic with Cl^- (17+1 =18)

(a) F^- (9+1 =10)

(b) O^{2-} (8+2 =10)

(c) K^+ (19−1=18)

(d) Na^+ (11−1 =10)

- S^2p^6 is the general formula of an element in group

A. 1A

B. 2A

C. 6A

D. 8A

Which of the following are have the same number of electrons (isoelectronic)?

$^{14}_{7}N^{3-}$ $^{19}_{9}F$ $^{20}_{10}Ne$ $^{24}_{12}Mg$ $^{27}_{13}Al^{3+}$

1. $^{19}_{9}F$ $^{20}_{10}Ne$
2. $^{20}_{10}Ne$ $^{24}_{12}Mg$
3. $^{19}_{9}F$ $^{20}_{10}Ne$ $^{27}_{13}Al^{3+}$
4. $^{14}_{7}N^{3-}$ $^{20}_{10}Ne$ $^{27}_{13}Al^{3+}$
5. None of the above

Chapter 8
Periodic Relationships Among the Elements

- 8.2 Periodic Classification of the Elements
- 8.3 Periodic Variation in Physical Properties
- Effective nuclear charge
- ❖ Atomic Radius
- ❖ Ionic Radius
- 8.4 Ionization Energy
- 8.5 Electron Affinity
- ❖ Electronegativity (ch.9 p. 377-378)

p357: 8.5, 8.8, 8.12, 8.20, 8.24, 8.26, 8.28, 8.30, 8.32

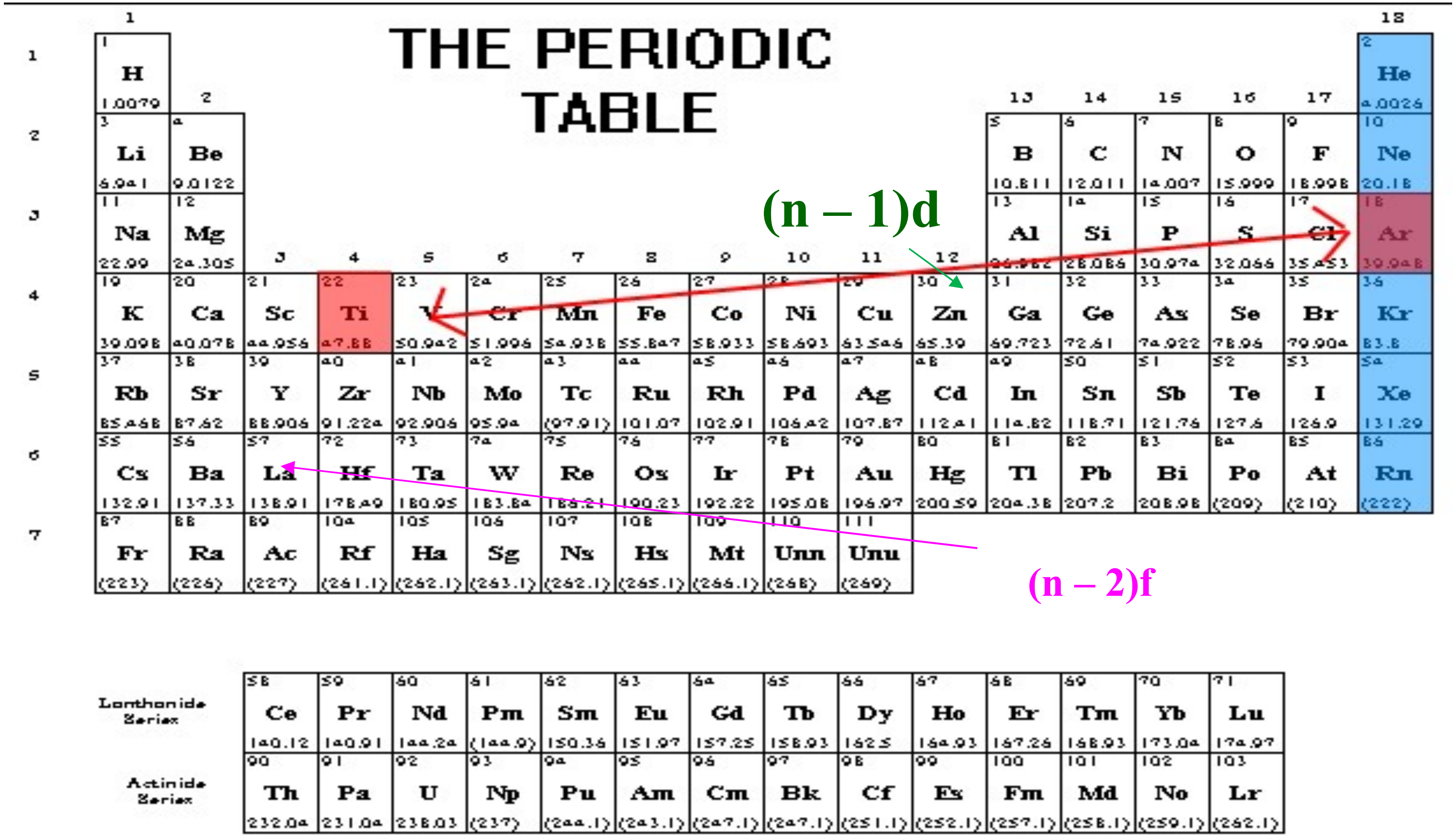
p358: 8.36, 8.38, 8.40, 8.44, 8.46

p358: 8.52, 8.54, 8.62, 8.64

Short-hand notation

- Instead of using complete electronic configuration Short-hand notation is useful
- show preceding inert gas configuration plus the additional electrons- [noble gas]$_{\text{pervious period}}$ additional electron(use general electronic configuration (A)), for d electrons (n – 1)d orbitals.
- Remember , starting from period 4 the (n – 1)d orbitals will appear , example

$_{31}Ga$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$	[Ar] $4s^2 3d^{10} 4p^1$
$_{32}Ge$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$	[Ar] $4s^2 3d^{10} 4p^2$
$_{33}As$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$	[Ar] $4s^2 3d^{10} 4p^3$
$_{34}Se$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$	[Ar] $4s^2 3d^{10} 4p^4$
$_{35}Br$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$	[Ar] $4s^2 3d^{10} 4p^5$
$_{36}Kr$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$	[Kr]



$_{22}Ti$ $1s^22s^22p^63s^23p^64s^23d^2$ $[Ar]\ 4s^23d^2$

$_{46}Pd$ $[Kr]\ 4d^{10}$

$_{62}Pb$ $[Xe]\ 4f^{14}5d^2\ 6s^26p^2$

$_{78}Pt$ $[Xe]\ 4f^{14}\ 5d^9\ 6s^1$

electronic configuration of transition metals

- There is a tendency toward half-filled and completely filled d subshells. This is a consequence of the closeness of the 3d and the 4s orbital energies.
- Some irregularities occur when there are enough electrons to half-fill *s* and *d* orbitals on a given row.

- For instance, the electron configuration for copper ^{29}Cu is

[Ar] $4s^1\ 3d^{10}$ rather than the expected [Ar] $4s^2\ 3d^9$.

- the electron configuration for ^{24}Cr is

[Ar] $4s^1\ 3d^5$ rather than the expected [Ar] $4s^2\ 3d^4$.

- Additional exceptions are **Mo $5s^14d^5$**; **Ag $5s^14d^{10}$**; **Au $6s^15d^{10}$**
- That is reasonable considering their position on the periodic chart.

$_{21}Sc$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$ [Ar] $4s^2 3d^1$, *para- 1 unpaired e-*

$_{22}Ti$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$ [Ar] $4s^2 3d^2$, *para- 2 unpaired e-*

$_{23}V$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$ [Ar] $4s^2 3d^3$ *para- 3 unpaired e-1*

$_{24}Cr$ $1s^2 2s^2 2p^6 3s^2 3p^6 \boxed{4s^1 3d^5}$ [Ar] $4s^1 3d^5$ NOT [Ar] $4s^2 3d^4$ *para- 6 unpaired e-1*

$_{25}Mn$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$ [Ar] $4s^2 3d^5$ *para- 5 unpaired e-1*

$_{26}Fe$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ [Ar] $4s^2 3d^6$ *para- ,4 unpaired e-1*

$_{27}Co$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$ [Ar] $4s^2 3d$ *para- ,3 unpaired e-1* ^{7}

$_{28}Ni$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$ [Ar] $4s^2 3d^8$ *para- ,2 unpaired e-1*

$_{29}Cu$ $1s^2 2s^2 2p^6 3s^2 3p^6 \boxed{4s^1 3d^{10}}$ [Ar] $4s^1 3d^{10}$ NOT [Ar] $4s^2 3d^9$

$_{30}Zn$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$ [Ar] $4s^2 3d^{10}$ *diamagnetic - ,0 unpaired e-1*

Electron Configurations of Cations of Transition Metals

When a cation is formed from an atom of a transition metal, electrons are always removed **first from the *ns*** orbital and then from **the $(n-1)d$ orbitals**.

Fe: [Ar]$4s^2 3d^6$, para- ,4 unpaired e-1

Fe^{2+}: [Ar]$4s^0 3d^6$ or [Ar]$3d^6$

para- ,4 unpaired e-1

Fe^{3+}: [Ar]$4s^0 3d^5$ or [Ar]$3d^5$

para- ,5 unpaired e-1

Mn: [Ar]$4s^2 3d^5$

para- ,5 unpaired e-1

Mn^{2+}: [Ar]$4s^0 3d^5$ or [Ar]$3d^5$

para- ,5 unpaired e-1

keep in mid that most transition metals can form more than one cation and frequently the cations are not isoeletronic with the preceding noble gases

1- How many unpaired electrons in Fe^{+2}

a) 2
b) 4
c) 4
d) 3

2- How many unpaired electrons in Fe^{+3}

a. 2
b. 5
c. 4
d. 3

3- How many unpaired electrons in Mn^{+2}

- 2
- 5
- 4
- 3

4- How many unpaired electrons in Mn

- 2
- 5
- 4
- 3

- What is the ground-state electron configuration of Mn?
- $3d^5$
- $4s^1\ 3d^5$
- $4s^2\ 3d^6$
- $4s^2\ 3d^5$

- What is the ground-state electron configuration of Mn^{+2}
- $3d^5$
- $4s^1\ 3d^5$
- $4s^2\ 3d^6$
- $4s^2\ 3d^5$

- What is the ground-state electron configuration of Fe^{+2}
- $3d^6$
- $4s^1\ 3d^5$
- $4s^2\ 3d^6$
- $4s^2\ 3d^5$

- What is the ground-state electron configuration of Fe^{+3}
- $3d^5$
- $4s^1\ 3d^5$
- $4s^2\ 3d^6$
- $4s^2\ 3d^5$

Gallium element is found in the periodic table in

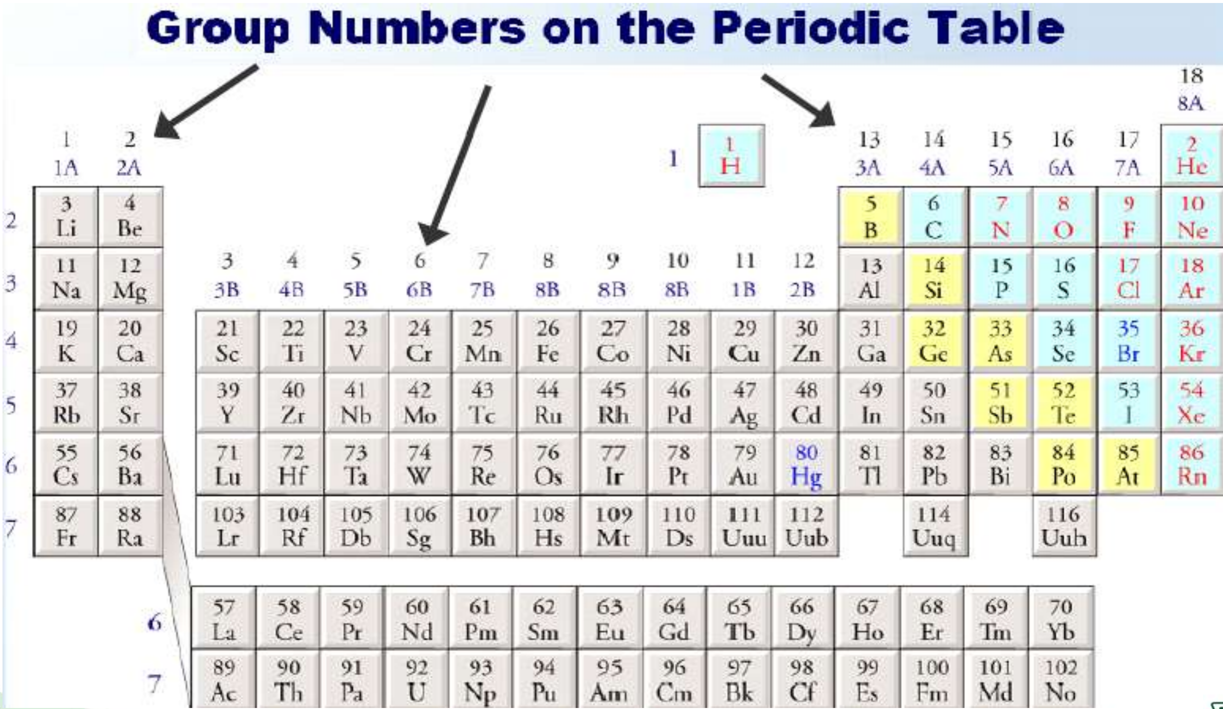
(a) period 3, group 1B

(b) period 3A, group 4

(c) period 4, group 1A

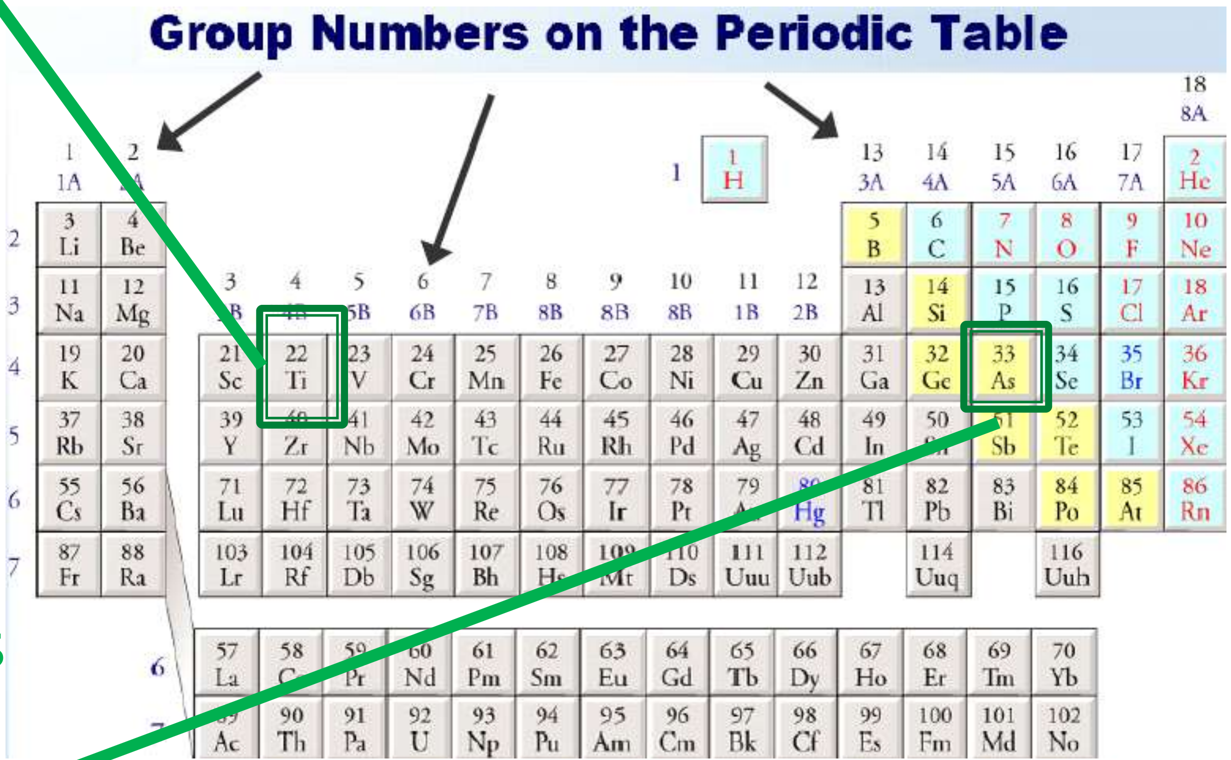
(d) period 4, group 3A

Answer (d)



Titanium (Ti) element is found in the periodic table in

(a) s-block
(b) P-block
(c) d-block
(d) f-block



The 15th element in the period 4 is

(a) s-block
(b) P-block
(c) d-block
(d) f-block

Example 8.1 p328

- An atom of a certain element has 15 electrons. Without consulting a periodic table, answer the following questions: **(P)**

(a) What is the ground-state electron configuration of this element?

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

(b) How should be element be classified?

Period 3, group 5A

The element is representative element.

(c) Is the element diamagnetic or paramagnetic

paramagnetic

8.3 Periodic Variation in Physical Properties

- Effective nuclear charge
- Atomic Radius
- Ionic Radius

- Effective nuclear charge
- lower effective charge on nucleus
- inner electrons shield outer electrons from nucleus
- shielding effect of electrons reduces the attraction between the nucleus and the electrons
- repulsive forces between electrons offset the attractive forces

Effective nuclear charge (Z_{eff}) is the "positive charge" felt by an electron.

$$Z_{eff} = Z - \sigma \qquad 0 < \sigma < Z \text{ (s = shielding constant)}$$

$$Z_{eff} \approx Z - \text{number of inner or core electrons}$$

	Z	Core	Z_{eff}	Radius
Na	11	10	1	186
Mg	12	10	2	160
Al	13	10	3	143
Si	14	10	4	132

Within a Period as Z_{eff} increases radius decreases

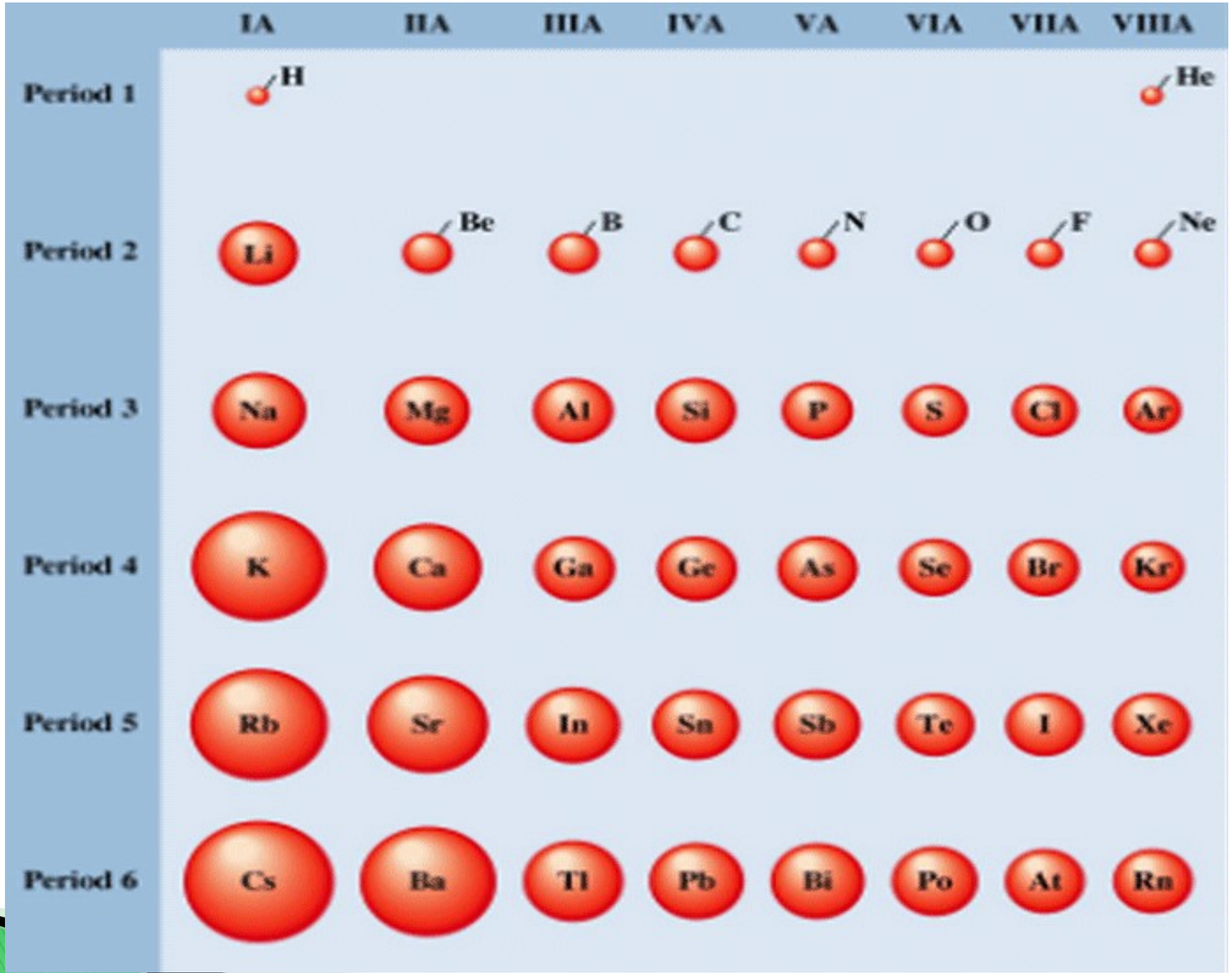
decreases

The atomic radius

- The atomic radius is ½ the distance between the 2 nuclei of the adjacent atoms.
- Atomic radius - a number of physical properties of elements are related to the size of an atom
- Atomic radius, in general, decreases as we move from left to right (→) in a row of the periodic table a Period.
- Atomic radius increases from top to bottom ↓ in a family or **group**.

Within a Period atomic radius decreases

Within a group atomic radius increase

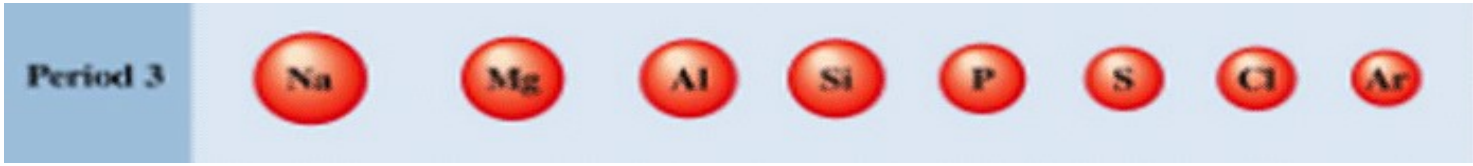


Example 8.2 p332

- Referring to a periodic table, arrange the following atoms in order of increasing atomic radius: P , Si , N

increasing … small to large

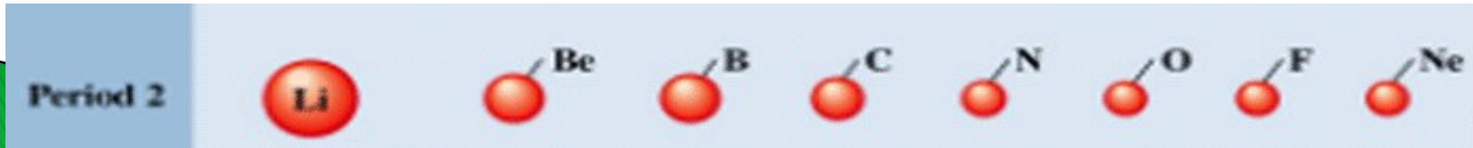
(small) N < P < Si (large)



- arrange the following atoms in order of decreasing radius: C , Li, Be

- decreasing … large to small

(large) Li > Be > C (small)

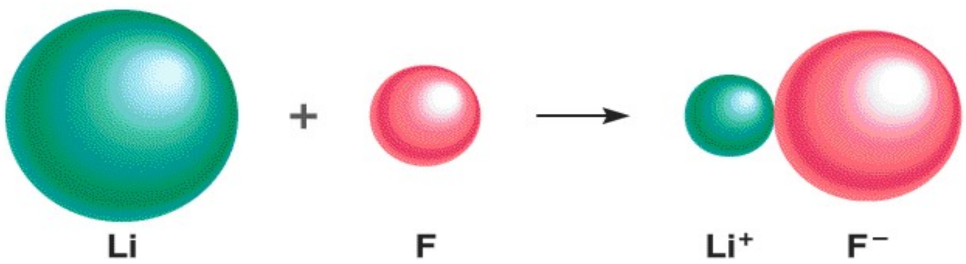


Ionic radius is the radius of anions and cations

- The ionic radius is the radius of anions and Cations

- Anions>>gain electrons >>> ionic radius increase because the nuclear charge remain the same but the repulsion resulting from the additional electrons enlarges the domain of the electron

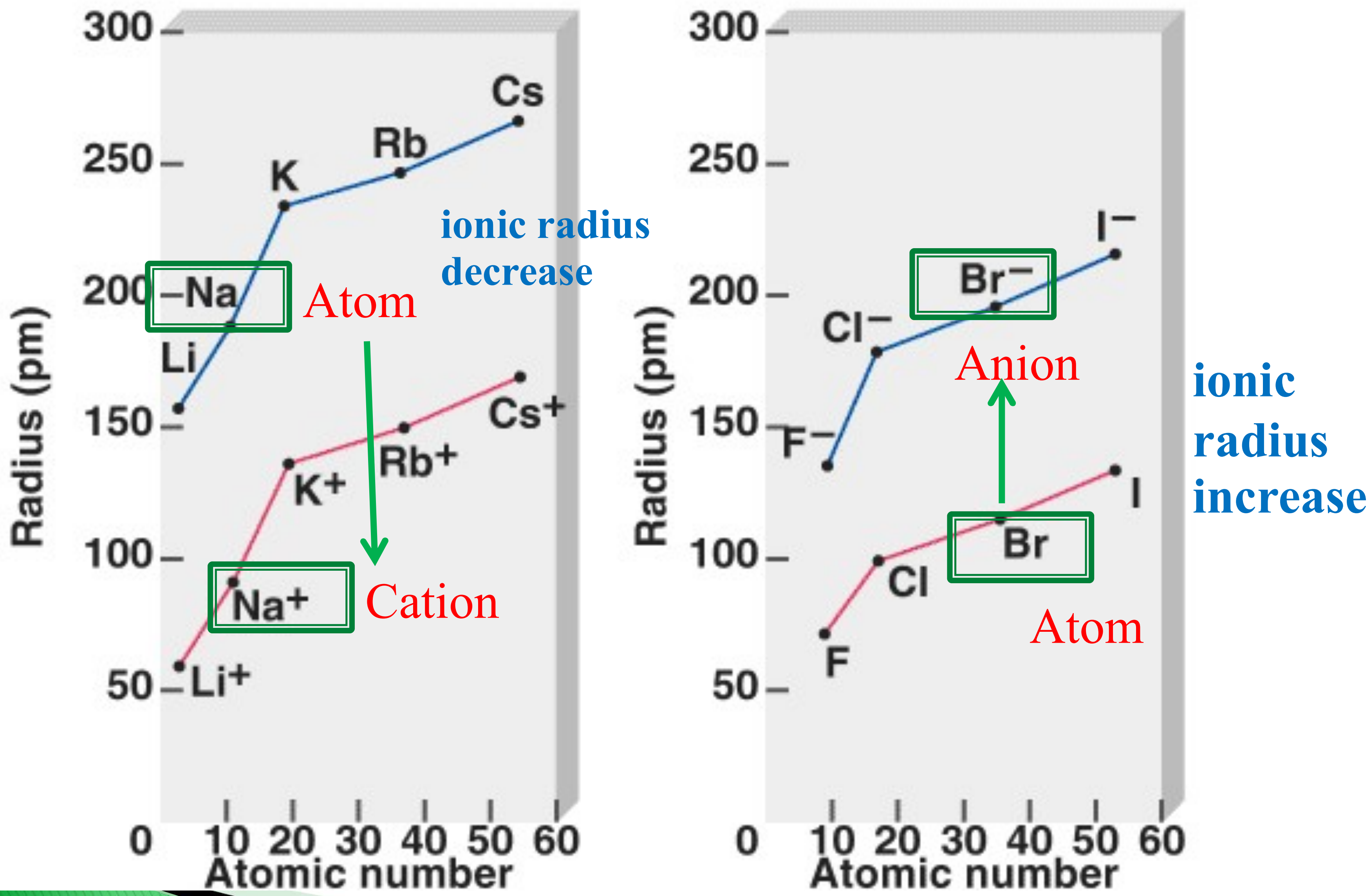
- Cations… lose electron …ionic radius decrease because removing one or more electron from an atom reduces electron-electron repulsion but the nuclear charge remains the same so the electron clouds shrinks , and the cation is smaller than atom



Cation is always **smaller** than atom from which it is formed.

Anion is always **larger** than atom from which it is formed.

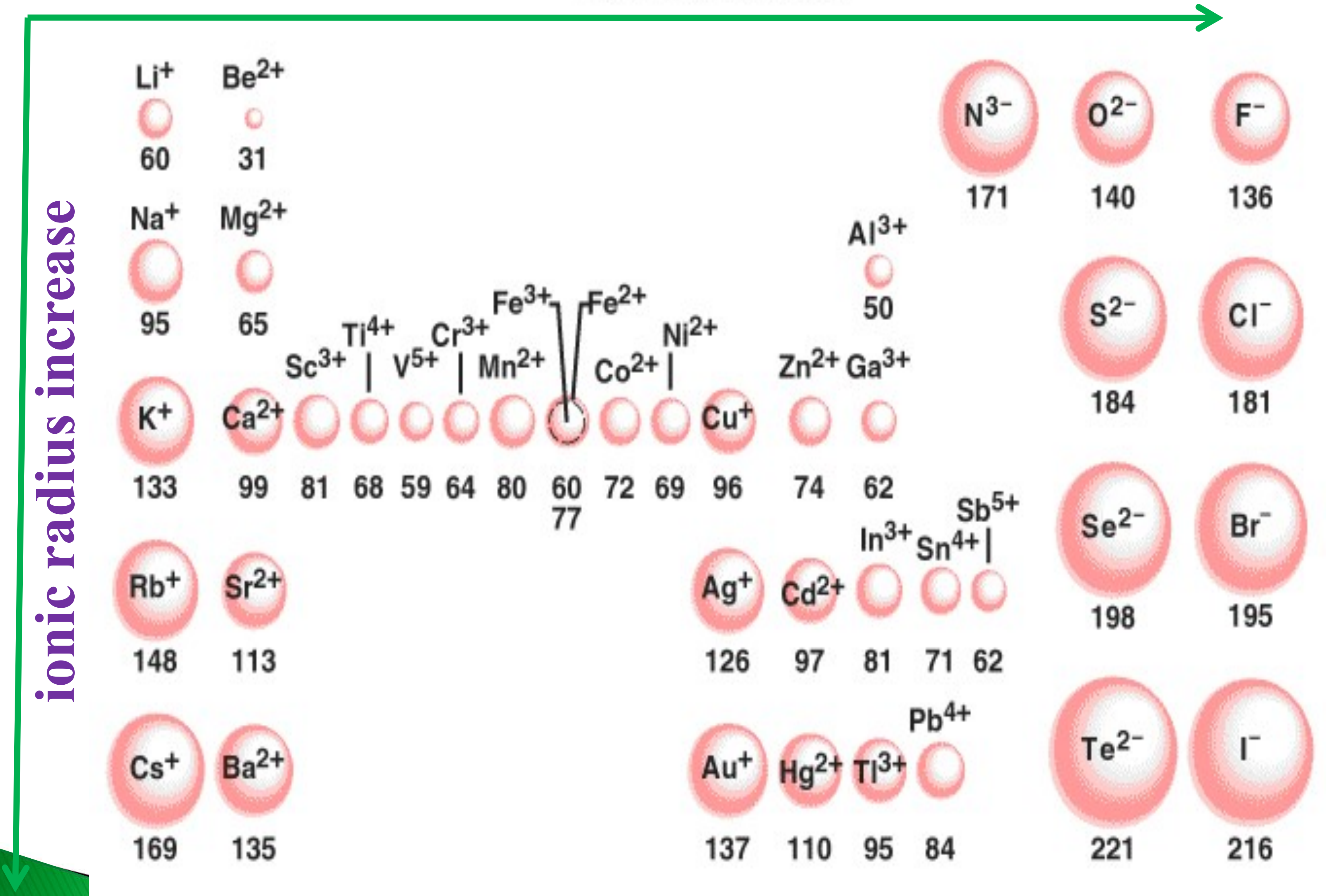
Comparison of Atomic Radii with Ionic Radii



- Isoelectronic ions
- Cations is smaller than anions ($^{10}Na^{+} < {}^{10}F^{-}$)
- The greater effective nuclear charge of $^{10}Na^{+}$ results in smaller radius.
- Isoelectronic cations
- $^{10}Al^{+3} < {}^{10}Mg^{+2} < {}^{10}Na^{+}$

- Isoelectronic anions
- $^{10}F^{-} < {}^{10}O^{-2} < {}^{10}N^{-3}$

Ionic Radii



ionic radius increase
Li+ 60
Be2+ 31
Na+ 95
Mg2+ 65
K+ 133
Ca2+ 99
Sc3+ 81
Ti4+ 68
V5+ 59
Cr3+ 64
Mn2+ 80
Fe3+ 60
Fe2+ 77
Co2+ 72
Ni2+ 69
Cu+ 96
Zn2+ 74
Ga3+ 62
Al3+ 50
Rb+ 148
Sr2+ 113
Ag+ 126
Cd2+ 97
In3+ 81
Sn4+ 71
Sb5+ 62
Cs+ 169
Ba2+ 135
Au+ 137
Hg2+ 110
Tl3+ 95
Pb4+ 84
N3− 171
O2− 140
F− 136
S2− 184
Cl− 181
Se2− 198
Br− 195
Te2− 221
I− 216

Example 8.3

- For each of the following pair ,indicate which is larger
- A) $^{10}F^{-}$,$^{10}N^{-3}$

  $^{10}N^{-3}$

- B) $^{10}Mg^{+2}$, $^{18}Ca^{+2}$

  $^{18}Ca^{+2}$

- C) Fe^{+2} , Fe^{+3}

  Fe^{+2}

- For each of the following pair ,indicate which is smaller
- A) $^{18}K^{+}$,$^{2}Li^{+}$

  $^{2}Li^{+}$

- B) $^{10}N^{-3}$, $^{18}P^{-3}$

  $^{10}N^{-3}$

- C) Au^{+} , Au^{+3}

  Au^{+3}

Ionization energy

- **Ionization energy** (IE) is the minimum energy (kJ/mol) required to remove an electron from a gaseous atom in its ground state.
- The higher ionization energy, the more difficult it is to remove the electrons.
- **The first ionization energy** is the amount of energy required to remove the 1st electron from an atom in the gaseous state.

$I_1 + X_{(g)} \longrightarrow X^{+}_{(g)} + e^{-}$ I_1 first ionization energy

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

$I_3 + X_{(g)} \longrightarrow X^{3+}_{(g)} + e^{-}$ I_3 third ionization energy

$$I_1 < I_2 < I_3$$

- When electron is removed from atom, repulsion among the remaining electrons decrease, because nuclear charge remains constant. More energy is needed to remove another electron from the positively charged ion.

- The IE for nonmetal is higher than metal , IE for metalloid fall between metals and nonmetals (highest value for 8A).

- The first IE increase from left to right (→) in period.
- The first IE decrease from top to bottom (↓) in group.

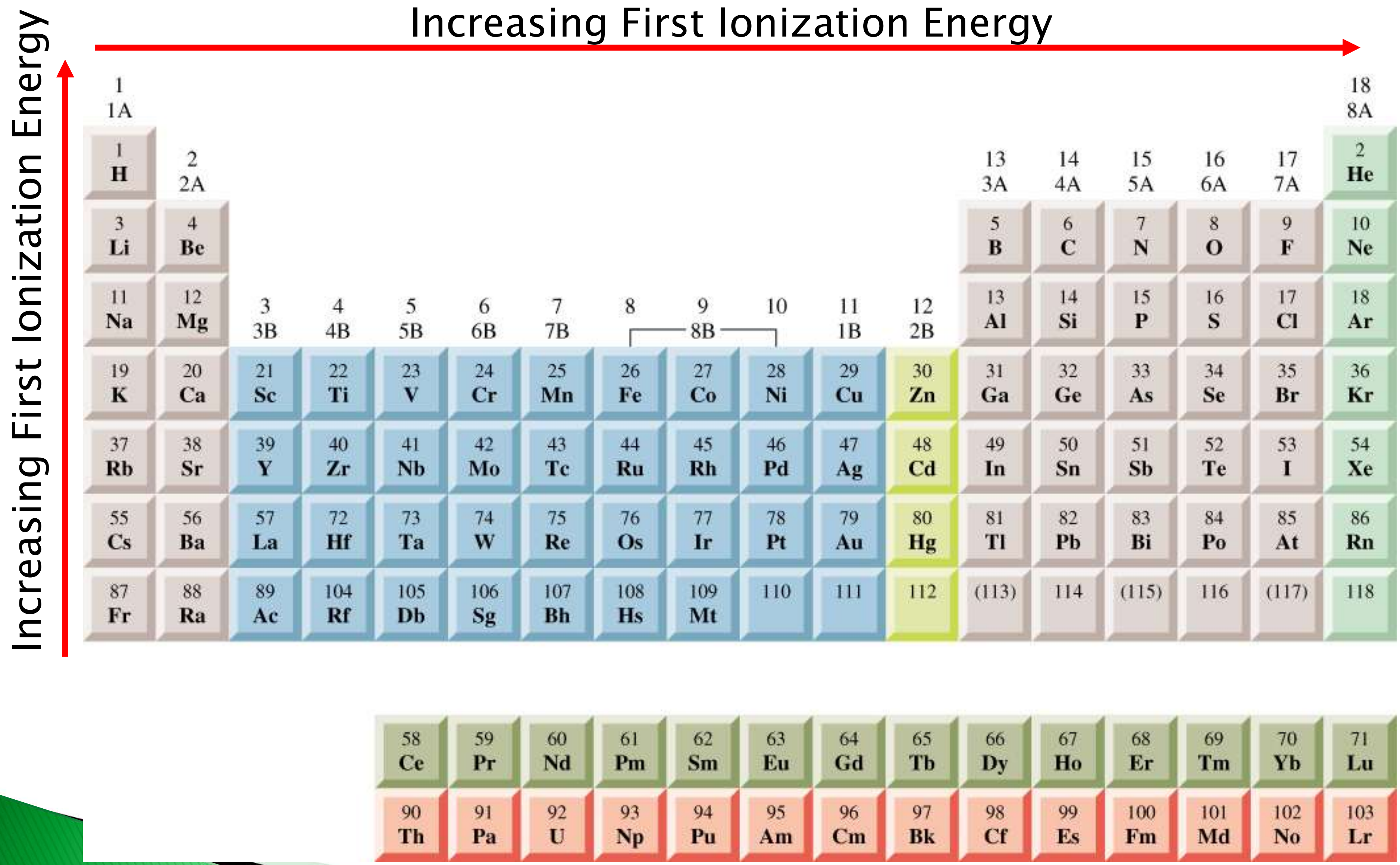
But there is some exceptions

A) Group 2A (ns^2) higher than 3A (ns^2 np^1) in the same period

B) Group 5A (ns^2 np^3) higher than 6A (ns^2 np^4) in the same period

(remember no exceptions in group).

General Trend in First Ionization Energies



Examples

which of the following has greatest ioniz1ation energy

a) **Na**
b) **K**
c) **Li**
d) **Rb**

- **which of the following has <u>greatest</u> ionization energy**

a) **C**
b) **N**
c) **F**
d) **Ne**

- ✓ عودي للجدول الدوري حددي مكان العناصر ثم حددي هي في نفس المجموعه أم الدوري
- ✓ تذكري الشذوذ فقط في IE&EA و في الدوره و ليس المجموعه
- ✓ في حال وجود عنصر من عناصر المجموعه 8 فأنه الأعلى في طاقة التأين و الأقل في الألفة الإلكترونية
- ✓ الشذوذ في في طاقة التأين في عناصر المجموعتين (5A&6A), (2A & 3A)
- ✓ لا يؤخذ الشذوذ في الأعتبار إلا في حال وجود عنصرين من مجموعتي الشذوذ

Examples

- which of the following has greatest ionization energy

a) C
b) B
c) Li
d) O

- which of the following has <u>greatest</u> ionization energy

a) C
b) N
c) O
d) Ne

- which of the following has greatest ionization energy

a) C
b) N
c) B
d) O

- Arrange the following elements in order of increasing IE(C,N,O,Ne)

a) C<N<O<Ne
b) C<N<O<Ne
c) Ne<N<O<C
d) C<O<N<Ne

Increasing (lowest to highest)
Decreasing (highest to lowest)

Electron affinity

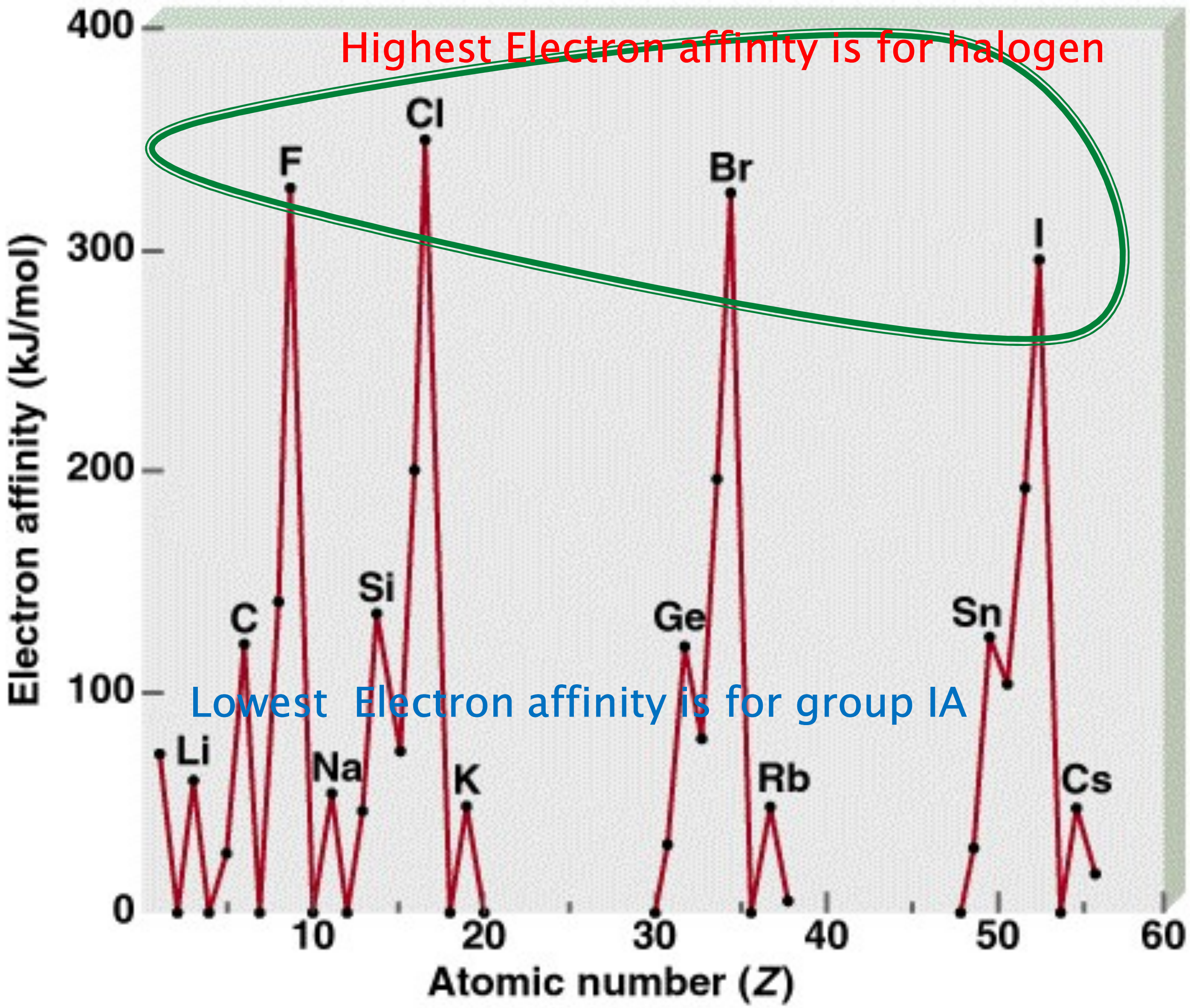
- Electron affinity is the negative of the energy change that occurs when an electron is accepted by an atom in the gaseous state to form an anion.

$$X_{(g)} + e^- \longrightarrow X^-_{(g)}$$

$F_{(g)} + e^- \longrightarrow X^-_{(g)}$ $\Delta H = -328$ kJ/mol EA = +328 kJ/mol

$O_{(g)} + e^- \longrightarrow O^-_{(g)}$ $\Delta H = -141$ kJ/mol EA = +141 kJ/mol

Electron Affinity Versus Atomic Number



- The EA for nonmetal is higher than metal , IA for metalloid fall between metals and nonmetals.

- The EA increase from left to right (→) in period.
- The first EA decrease from top to bottom (↓) in group.

- <u>But there is some exceptions</u>

A) Group 2A (ns^2) **lower than** 1A (ns^1) in the same period

B) Group 5A (ns^2 np^3) **lower than** 4A (ns^2 np^2) in the same period

(remember no exceptions in group).

Examples

which of the following has greatest EA

a) Na
b) He
c) Li
d) Rb

- which of the following has lowest EA

a) C
b) N
c) F
d) Ne

- عودي للجدول الدوري حددي مكان العناصر ثم حددي هي في نفس المجموعه أم الدوري
- تذكري الشذوذ فقط في IE&EA و في الدوره و ليس المجموعه
- في حال وجود عنصر من عناصر المجموعه 8 فأنه الأعلى في طاقة التأين و الأقل في الألفة الإلكترونية
- الشذوذ في الألفة الألكترونية في عناصر المجموعتين (5A&4A), (2A & 1A)
- لا يؤخذ الشذوذ في الأعتبار إلا في حال وجود عنصرين من مجموعتي الشذوذ

Examples

- **which of the following has greatest EA**

a) **C**
b) **B**
c) **Li**
d) **O**

- **which of the following has <u>greatest</u> EA**

a) **C**
b) **N**
c) **O**
d) **Ne**

- **which of the following has greatest EA**

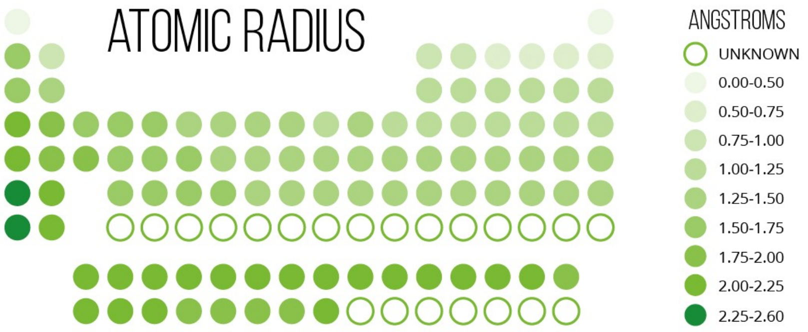
a) **C**
b) **N**
c) **B**
d) **Li**

- **Arrange the following elements in order of increasing EA (C,N,O,Ne)**

a) **C<N<O<Ne**
b) **C<N<O<Ne**
c) **Ne<N<C<O**
d) **C<O<N<Ne**

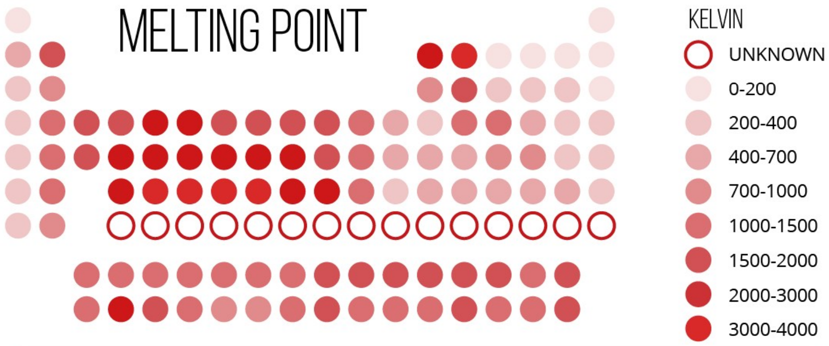
Increasing (lowest to highest)
Decreasing (highest to lowest)

PERIODICITY: TRENDS IN THE PERIODIC TABLE



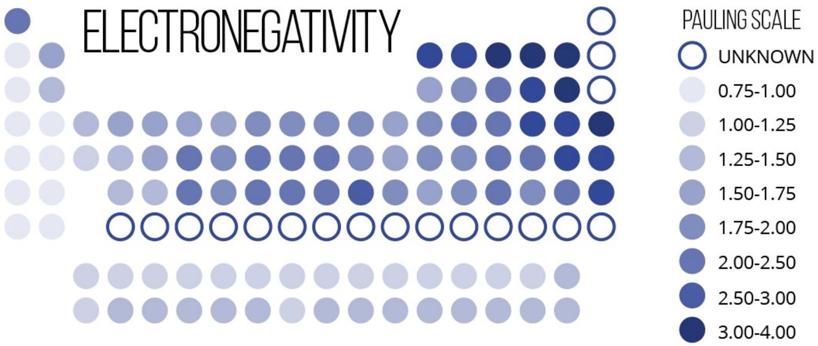
Atomic radius decreases across a period as nuclear charge increases but shielding effects remain approximately constant, resulting in electrons being drawn closer to the nucleus.

Atomic radius increases down a group as valence electrons become increasingly distant from the nucleus, and shielding also increases. This leads to a increase in atomic radius despite the increasing nuclear charge down a group.



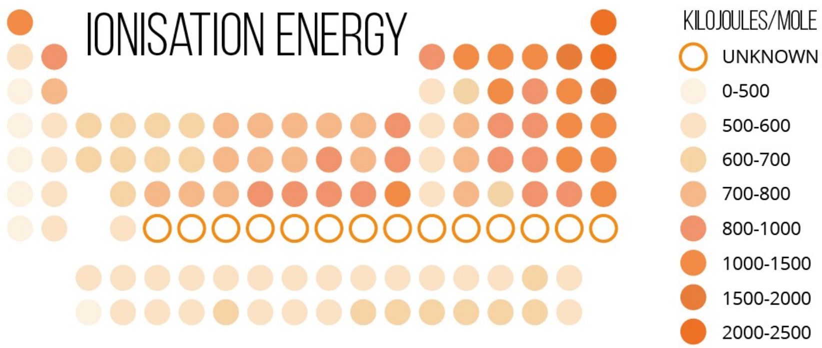
Metallic bonded and macromolecular substances tend to have high melting points. For both, this is due to the fact that the bonds require a lot of energy to break.

The majority of non-metals have a simple molecular structure. Simple molecular substances have low melting points as only weak intermolecular forces must be overcome in order to melt them. Strength of these is determined by the size of the molecule.



Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. Generally, electronegativity increases moving towards the top right of the Periodic Table.

This increase in electronegativity across a period is due to the increased nuclear charge and approximately constant shielding effects resulting in a greater force of attraction to the nucleus of the atom felt by the bonding electrons.



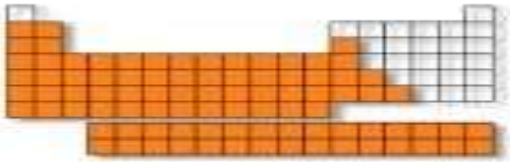
The first ionisation energy generally increases from left to right across a period, as the electron is drawn closer to the nucleus by the increased nuclear charge and becomes harder to remove.

Electrons in p orbitals are slightly easier to remove than those in s orbitals of the same energy level. Paired electrons in the same orbital can lead to repulsion, again making an electron easier to remove. Both of these factors can lead to lower than expected first ionisation energies.

NAME________________________________

TEST: The Periodic Table, Properties, and Positions

<u>DIRECTIONS:</u> The Periodic Table is broken into three main *GROUPS*, each with specific properties. Correctly name these groups and identify their properties using the provided choices page.



▸**Group:** ____________________

▸**Properties:** _____

▸**Group:** ____________________

▸**Properties:** _____

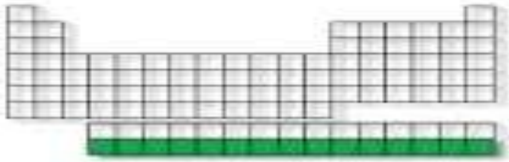
▸**Group:** ____________________

▸**Properties:** _____

<u>DIRECTIONS:</u> The Periodic Table is classified into nine *FAMILIES*, each with specific properties. Correctly name these families and identify their properties using the provided choices page.

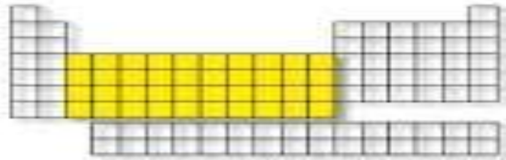
▸**Family:** ____________________

▸**Properties:** _____



▸**Family:** ____________________

▸**Properties:** _____

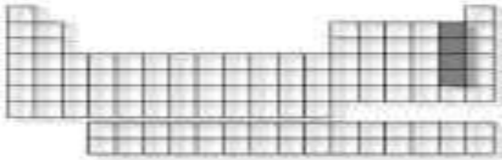


▸**Family:** ____________________

▸**Properties:** _____

▸**Family:** ____________________

▸**Properties:** _____



▸**Family:** ____________________

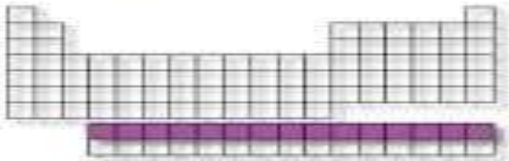
▸**Properties:** _____

▸**Family:** ____________________

▸**Properties:** _____

▸**Family:** ____________________

▸**Properties:** _____



▸**Family:** ____________________

▸**Properties:** _____

▸**Family:** ____________________

▸**Properties:** _____

Chapter 9

Chemical Bonding I: Basic Concepts

- 9.1 Lewis dot symbols
- 9.2 the ionic bond
- 9.4 the covalent bond
- 9.5 Electroegativity
- 9.6 Writing Lewis structures
- 9.7 formal charge and Lewis structures
- 9.8 the concept of resonance
- 9.9 the exception of octate rules

9.1 Lewis dot symbols

- When atoms interact to form chemical bond, only their outer region are in contact

- The Octet Rule: in forming chemical bonds, atoms usually gain, lose or share electrons until they have 8 in the outer shell to reach the same electronic configuration of the noble gasses (ns^2 np^6) (except hydrogen, helium and lithium).

- Lewis Dot Representation: In the representation of an atom, the valence electrons of an atom (**outer most shell electrons**) are represented by dots.

Lewis Dot Symbols

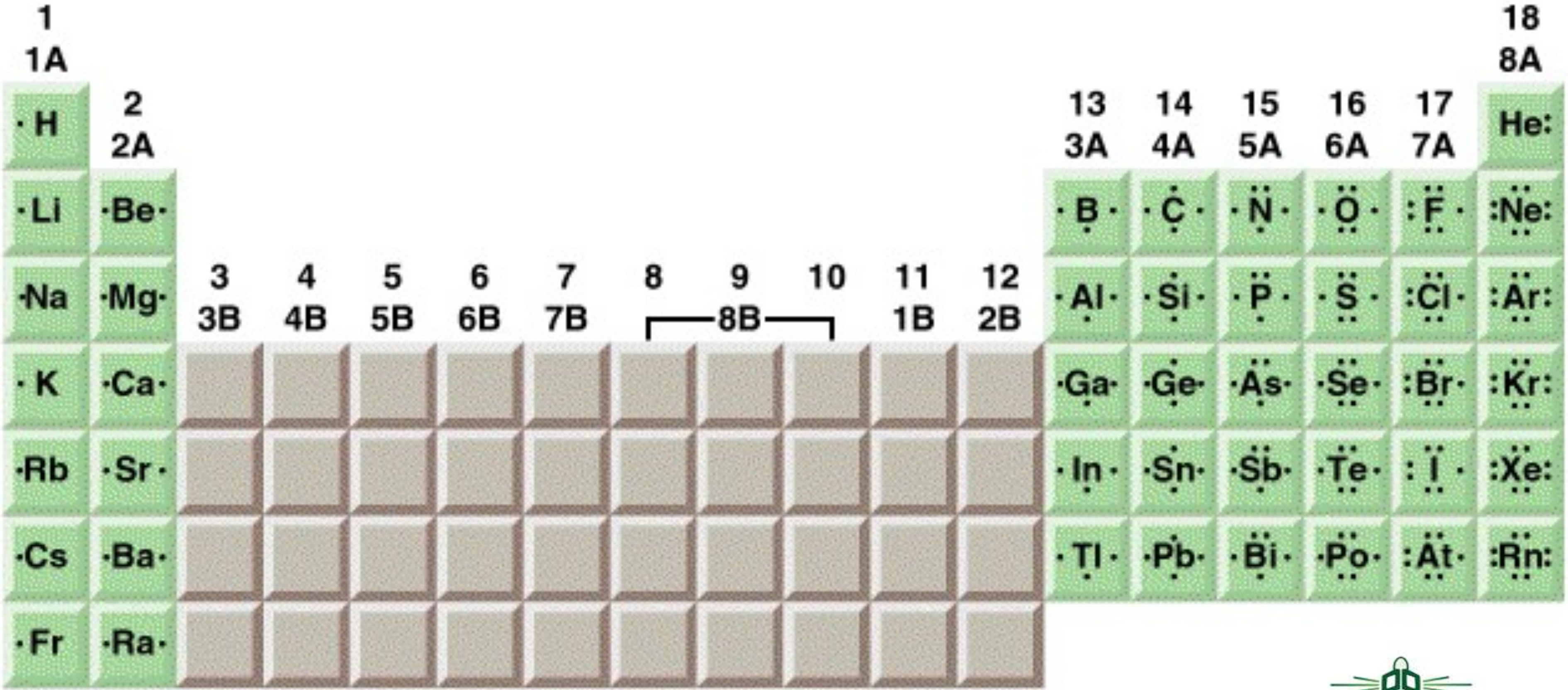
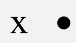
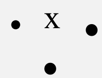
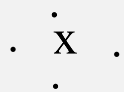
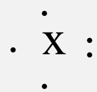
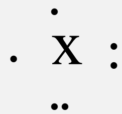
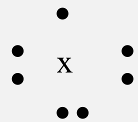
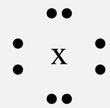


Table 9-1

Lewis dot symbols

Group	1A	2A	3A	4A	5A	6A	7A	8A
Lewis Dot								
Bonding electrons	1	2	3	4	3	2	1	0
nonbonding electrons (pair of nonbonding electrons)	0	0	0	0	2e 1pair	4e 2pairs	6e 3pairs	8 e 4pairs

examples

- What is Lewis dot structure of element in group 5

X.

.X.

. x •
   •

  .
. X :
  .

- What is Lewis dot structure of element Z=5

X.

.X.

. x •
   •

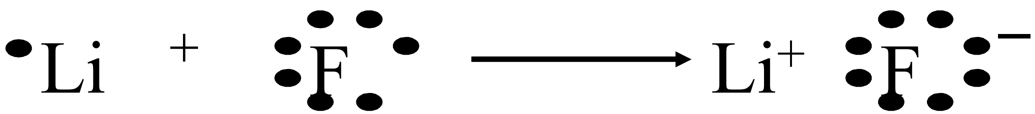
  .
. X :
  .

Types of Bonds

Types of Atoms	Type of Bond	Bond Characteristic
metals to nonmetals	Ionic	electrons transferred
nonmetals to nonmetals	Covalent	electrons shared

9.2 the ionic bond

- ionic bond is the electrostatic force that hold ions together in an ionic compound

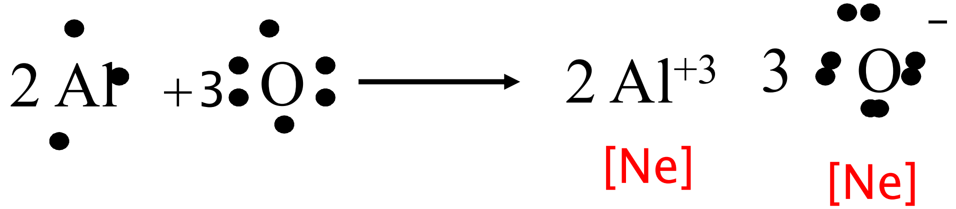


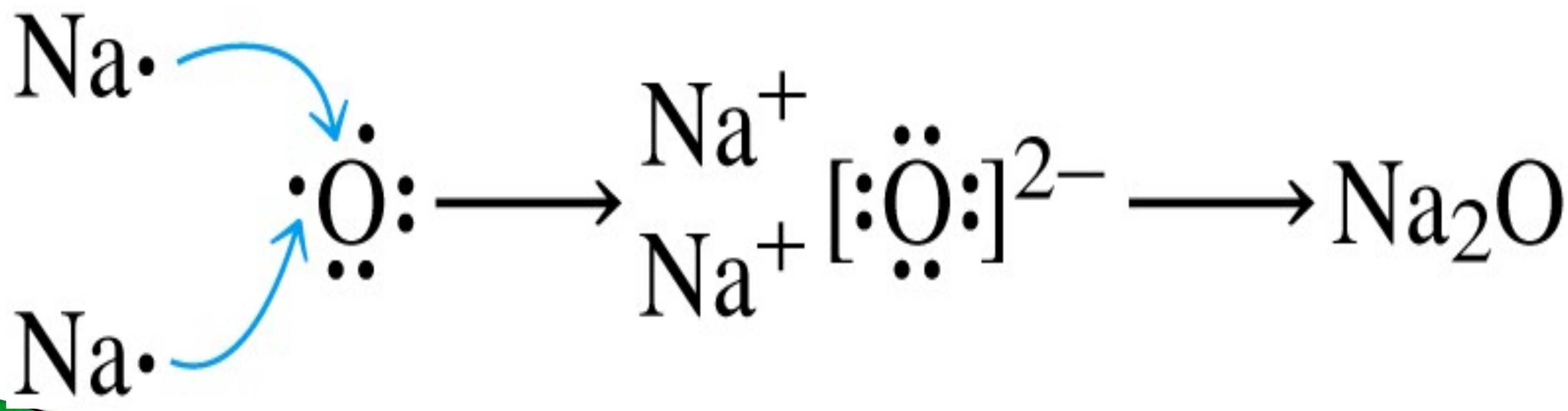
$1s^2 2s^1$ $1s^2 2s^2 2p^5$ [He] $1s^2 2s^2 2p^6$ [Ne]

- the resulting anions & cations attract each other in such a ratio that the charges cancel out.
- **Note: Do not show the charges in the final product.**
  Example: KI NOT K^+I^-

Example 9.1

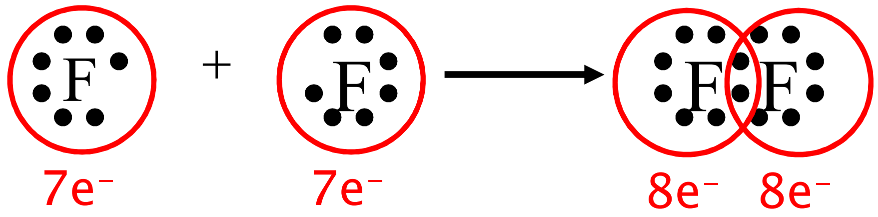
- Use Lewis dot symbol to show formation of Al_2O_3



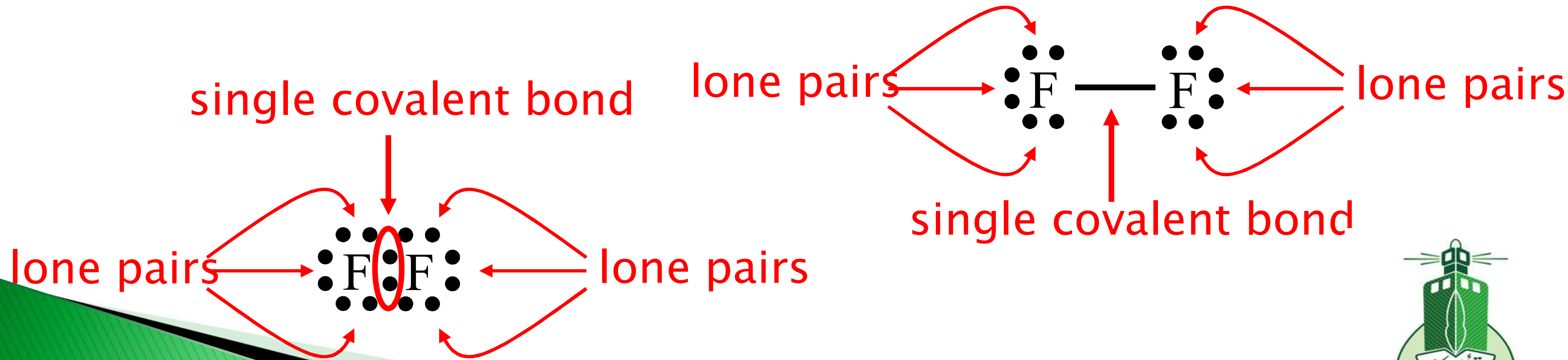


9.4 the covalent bond

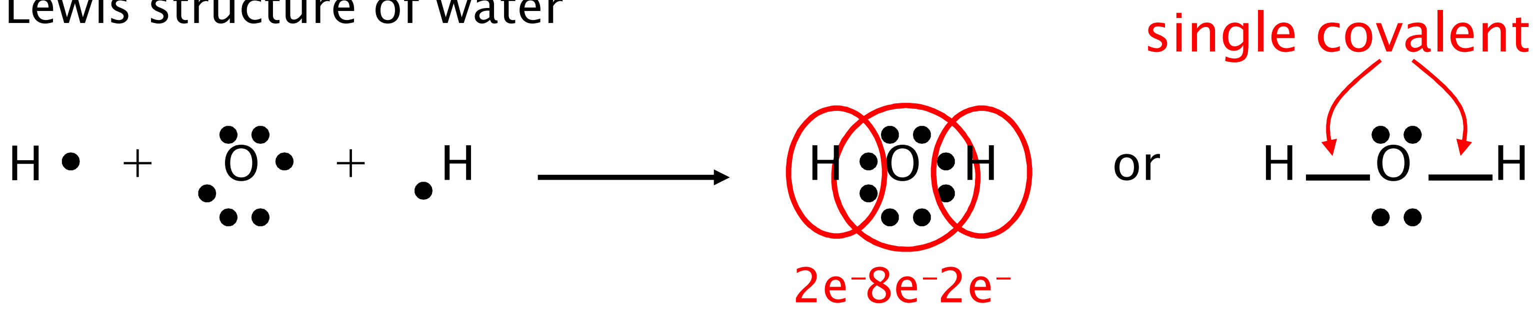
- A **covalent bond** is a chemical bond in which two or more electrons are shared by two atoms.



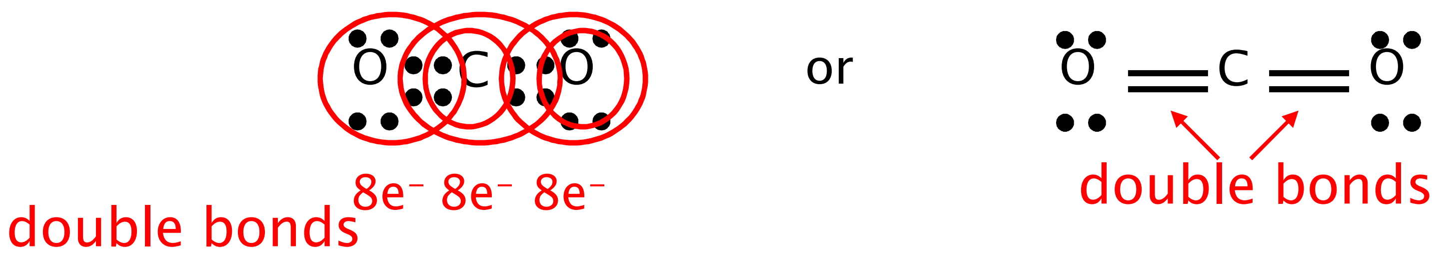
Lewis structure of F_2



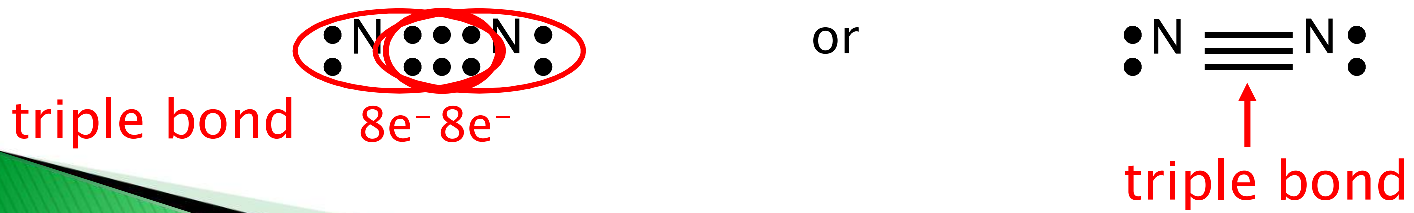
Lewis structure of water



Double bond – two atoms share two pairs of electrons



Triple bond – two atoms share three pairs of electrons



Electronegativity

- **Electronegativity** is the ability of an atom to attract toward itself the electrons in a chemical bond.
- High electronegativity →pick up electron easily

- Electronegativity increase from left to right in(→) period.
- Electronegativity decrease from top to bottom (↓)in group .
- Transition metals don’t follow these trend.

- Nonmetals have high electronegativity, metals have low electronegativity.

- high difference in electronegativity (2 or more), element tend to form ionic bond **(metal + nonmetal)** .(NaCl)

- small difference in electronegativity (less than 2), element tend to form **polar covalent** bond **(nonmetal + nonmetal)** .(HCl)

- Same electronegative of the same elements (equal 0) from pure covalent bond **(non polar covalent)** **(nonmetal + nonmetal)** (H_2) (F_2) (N_2)

Example 9.2

- Classify the following bonds as ionic, polar covalent, or covalent

- A) HCl =3-2.1=0.9

  Polar covalent

(nonmetal + nonmetal)

- b) KF =4-0.8=3.2

  Ionic

(metal + nonmetal)

- c) C-C =2.5-2.5=0

  Non polar covalent

- Classify the following bonds as ionic, polar covalent, or covalent

- A) CsCl =3-1=2

  Ionic

  (metal + nonmetal)

- b) H_2S =2.5-2.1=0.4

  Polar covalent

(nonmetal + nonmetal)

- c) N-N =3-3=0

  Non polar covalent

9.6 Writing Lewis structures

1. Write the skeletal structure of the compounds, using chemical symbol and placing bonded atoms next to one another.
2. (A) determine the total number of electrons in the valence shells of all of the atoms of the molecule, add electrons (if molecule have net –ve charge add electrons, if molecule have net +ve charge subtract electrons) …

∑ № of atoms (group №)

1. (B) Complete an octet for all atoms ***except*** hydrogen

∑ № of atoms (8) ***except*** hydrogen ∑ № of atoms (2)

4. Find the number of bonds by C = B-A/2
5. Find the number of lone pair of electron by D=B-C

Example 9.3

Write the Lewis structure of nitrogen trifluoride (NF_3).

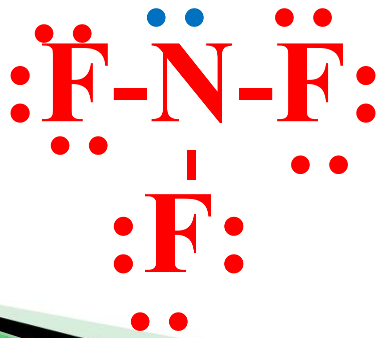
Step 1 – N is less electronegative than F, put N in center

Step 2 – A= 5X1 + 7X3 = 26 valance electrons

Step 3 – B = 8X1+8X3 =32 electrons

Step 4 - C = 32-26 =6/2=3 bonds

Step 5 - D= 26-6 =20 nonbonding electrons or 10 pair of electrons



Lewis structure of NF_3 consist of 3 single bond , 20 nonbonding electrons or 10 pair of electrons

Example

Write the Lewis structure of nitrogen trifluoride (NH_3).

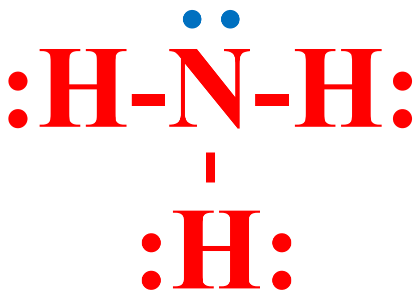
Step 1 – N is less electronegative than F, put N in center

Step 2 – A= 5X1 + 1X3 = 8 valance electrons

Step 3 – B = 8X1+**2**X3 =14 electrons

Step 4 - C = 14-8 =6/2=3 bonds

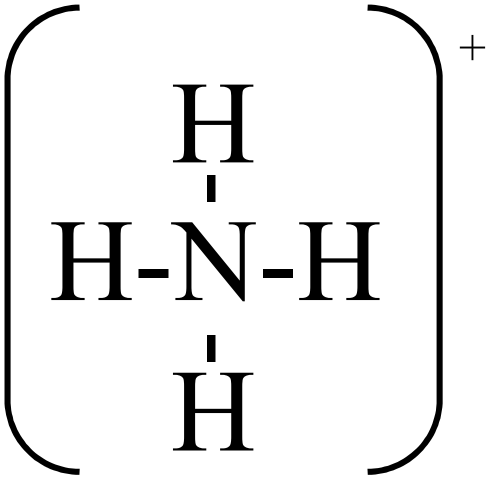
Step 5 - D= 8-6 =2 nonbonding electrons or 1 pair of electrons



Lewis structure of NH_3 consist of 3 single bond , 20nonbonding electrons or 1 pair of electrons

NH_4^+

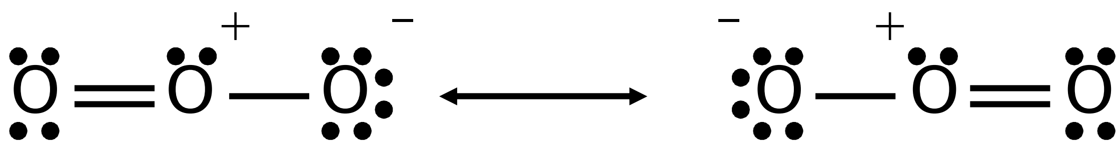
- Step 2 – A= 5X1 + 1X4 -1 = 8 valance electrons
- Step 3 – B = 8X1+2X4 =16 electrons
- Step 4 - C = 16-8 =8/2=4 bonds
- Step 5 - D= 8-4 =4 non bonding electrons , 2 pair of electrons



9.8 the concept of resonance

- A **resonance structure** is one of two or more Lewis structures for a single molecule that cannot be represented accurately by only one Lewis structure (**after** formal charge has been determined).

- When number of bonds is more than number of atoms around central atom , in this case the extra bond positions to give resonance structure



Ozone have 2resonance structure

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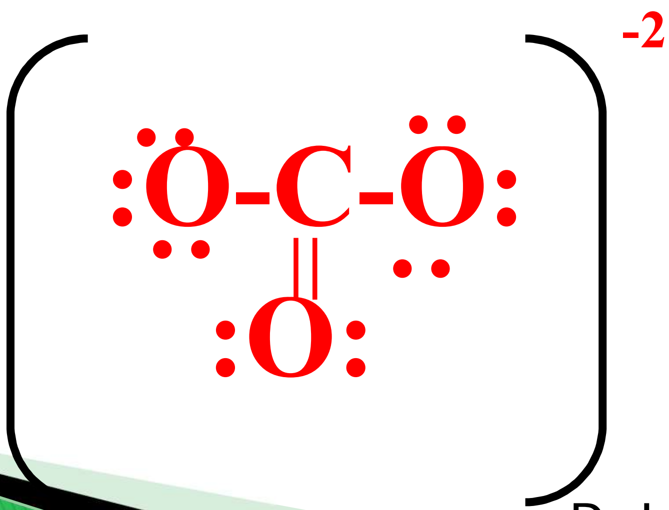
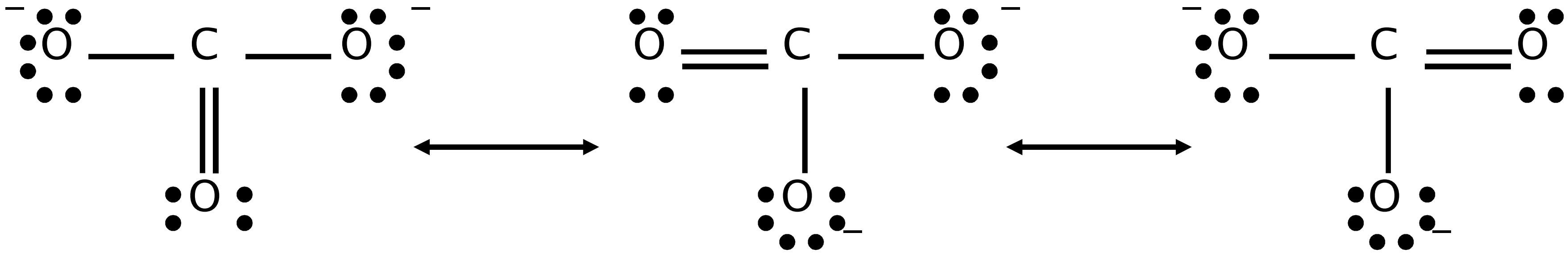
Example 9.5

- Write the Lewis structure of carbon dioxide $[CO_3]^{-2}$
- Step 1 – C is less electronegative than O, put C in center
- Step 2 – A= 4X1 + 6X3 +2 = 24 valance electrons
- Step 3 – B = 8X1+8X3 = 32 electrons
- Step 4 - C = 32-24 =8/2=4 bonds
- Step 5 - D= 24-8 =16 nonbonding electrons or 8 pair of electrons
- -



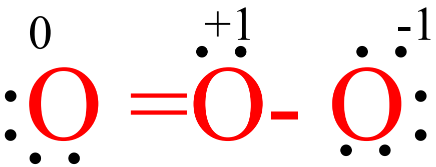


Carbonate ion have 3 resonance structure

9.7 formal charge and Lewis structures

- formal charge is the difference between the number of valence electrons in an isolated atom and the number of electrons assigned to that atom in a Lewis structure.

$$\text{formal charge on an atom in a Lewis structure} = \text{total number of valence electrons in the free atom} - \left(\text{total number of nonbonding electrons} + \text{Number of bond}\right)$$



6 6 6

(4+2) (2+3) (6+1)

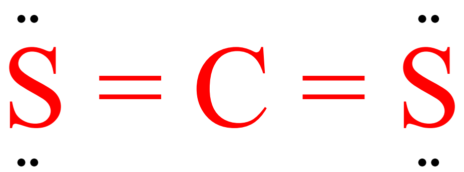
0 +1 -1

- For molecules , the sum of the charges should be zero
- For ion , the sum of the charges should be -ve for anions
- For ion , the sum of the charges should be +ve for cations
- formal charge and Lewis structures

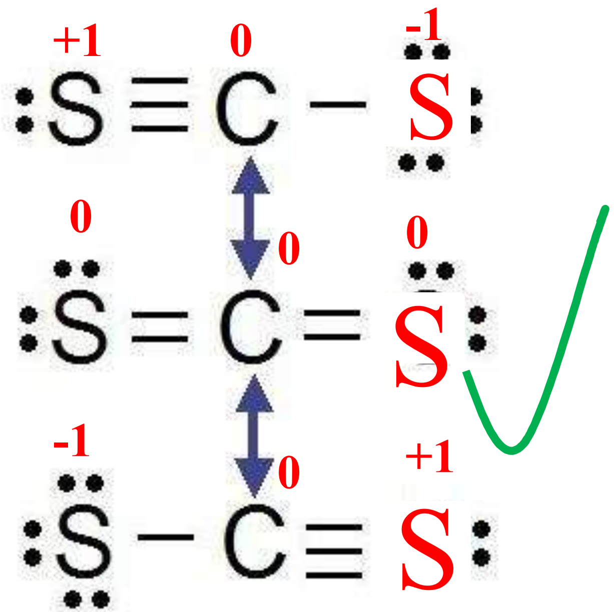
1. For neutral molecules, a Lewis structure in which there are no formal charges is preferable to one in which formal charges are present.
2. Lewis structures with large formal charges are less plausible than those with small formal charges.
3. Among Lewis structures having similar distributions of formal charges, the most plausible structure is the one in which negative formal charges are placed on the more electronegative atoms

Example 9.3

- Write the Lewis structure of carbon disulfide (CS_2).
- Step 1 – C is less electronegative than S, put C in center
- Step 2 – A= 4X1 + 6X2 = **16** valance electrons
- Step 3 – B = 8X1+8X2 =**24** electrons
- Step 4 - C = 24-16 =8/2=**4** bonds
- Step 5 - D= 16-8 =**8** nonbonding electrons or 4 pair of electrons



3 resonance structure

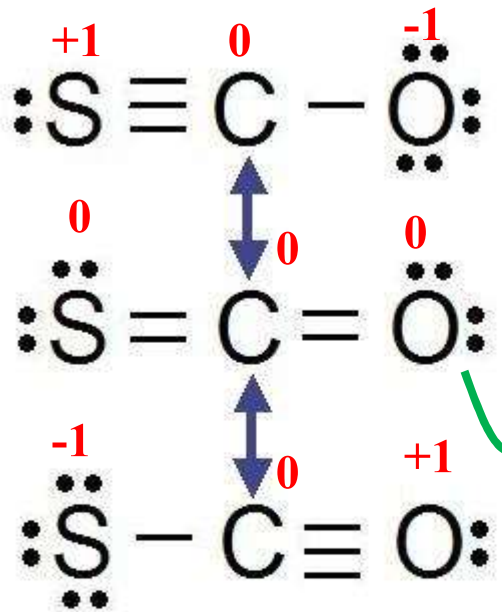


Example

- Write the Lewis structure of carbon disulfide (COS).
- Step 1 – C is less electronegative than S&O, put C in center
- Step 2 – A= 4X1 + 6X1+ 6X1 = **16** valance electrons
- Step 3 – B = 8X1+8X1 +8X1=**24** electrons
- Step 4 - C = 24-16 =8/2=**4** bonds
- Step 5 - D= 16-8 =**8** nonbonding electrons or 4 pair of electrons

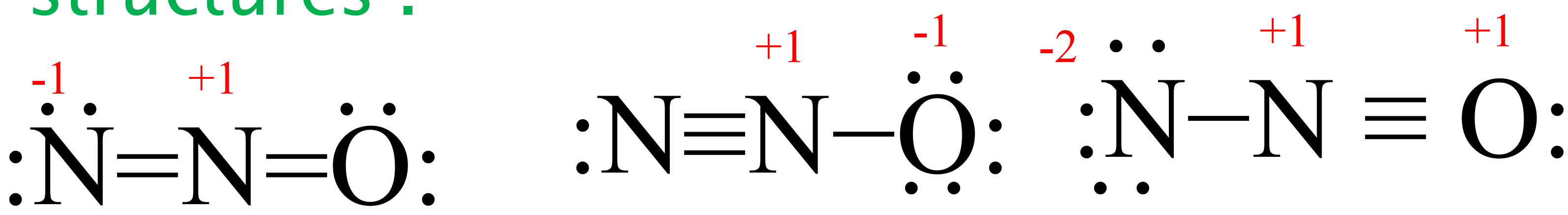
$$\ddot{\underset{..}{S}} = C = \ddot{\underset{..}{O}}$$

3 resonance structure



Example 9.8

- Draw three resonance structure for N2O (NNO),indicate formal charge rank the structures .



A			B			C		
5	5	6	5	5	6	5	5	6
6	4	6	5	4	7	7	4	5
-1	1	0	0	1	-1	-2	+1	+1

A B ✓ C

3 resonance structure

B< A < C

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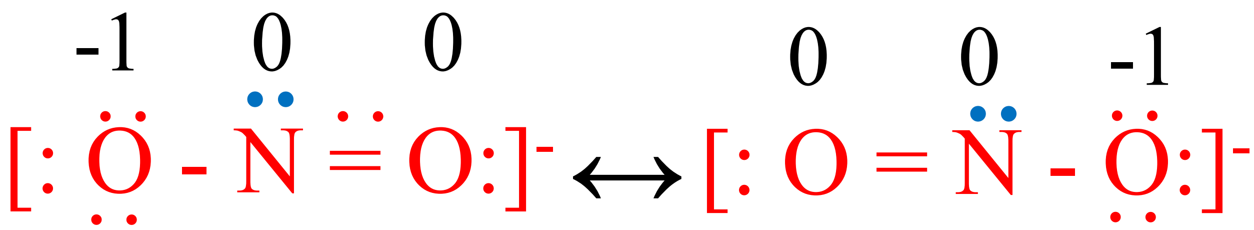
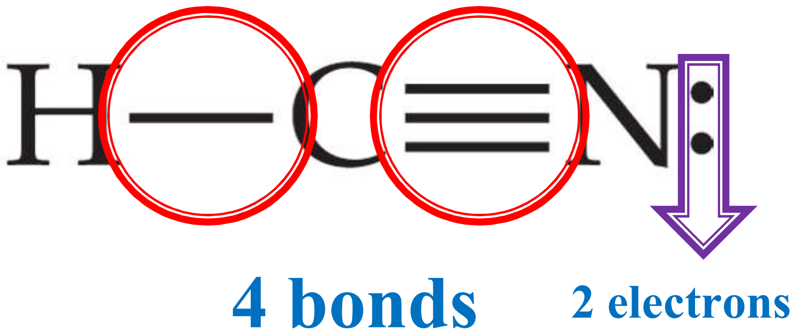
Example 9.5

- Write the Lewis structure of nitrogen dioxide ion $[NO_2]^{-1}$
- Step 1 – N is less electronegative than O, put N in center
- Step 2 – A= 5X1 + 6X2 +1 = 18 valance electrons
- Step 3 – B = 8X1+8X2 = 24 electrons
- Step 4 - C = 24-18 =6/2=3 bonds
- Step 5 - D= 18-6 =12 nonbonding electrons or 6 pair of electrons



2 resonance structure

Writing Lewis Structures



A = 1X1+4X1+5X1 =10 valance electrons

B = 1X2+8X1+8X1 =18electrons

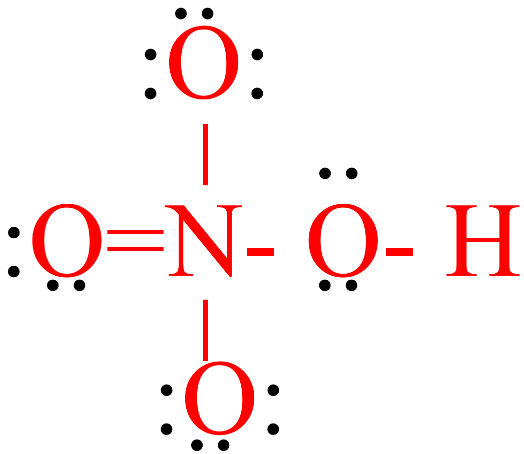
C = 18-10 =8/2=4 bonds

D= 10-8 =2 electrons

Lewis structure of HCN consist of 4 bond , 1 triple bond , 0 double bond , 2 nonbonding electrons or 1 pair of electrons

Example 9.4

- Write the Lewis structure for nitric acid (HNO_3) in which the three O atoms are bonded to the central N atom and ionizable H atom is bonded to one of the O atom.
- Step 1 –put N in center ,surrounded by 3O atoms , H bonded to one of the O
- Step 2 – Count valence electrons 5 + (3 x 6) +1 = 24 nonbonding electrons or 12 pair of electrons

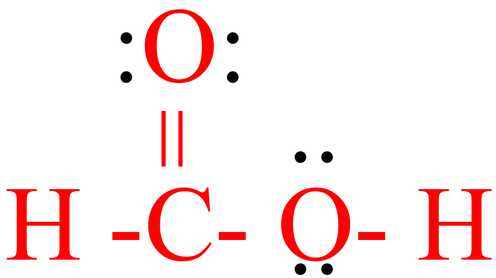


3 resonance structure

Example 9.4

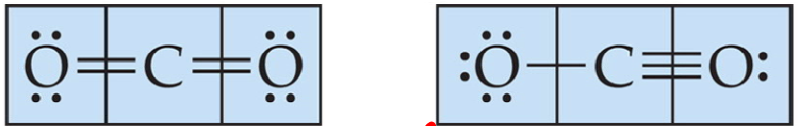
- Write the Lewis structure of formic acid (HCOOH).
- Step 1 –put C in center ,surrounded by 2O atoms , H Step 2 – A= 4X1 + 6X2 +2x1 = **18** valance electrons
- Step 3 – B = 8X1+8X2 +2 x2 =**28** electrons
- Step 4 - C = 28-18 =10/2=**5** bonds
- Step 5 - D= 18-10 =**8** nonbonding electrons or 4 pair of electrons

3 resonance structure



3 resonance structure

- Write the Lewis structure of carbon dioxide [CO_2]
- Step 1 – C is less electronegative than O, put C in center
- Step 2 – A= 4X1 + 6X2 = **16** valance electrons
- Step 3 – B = 8X1+8X2 = **24** electrons
- Step 4 - C = 24-18 =8/2=**4** bonds
- Step 5 - D= 16-8 =**8** nonbonding electrons or 4 pair of electrons

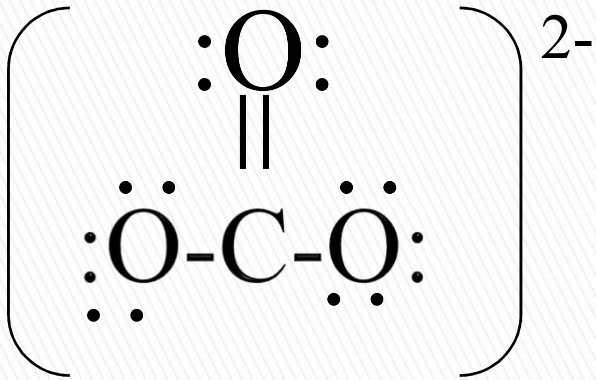


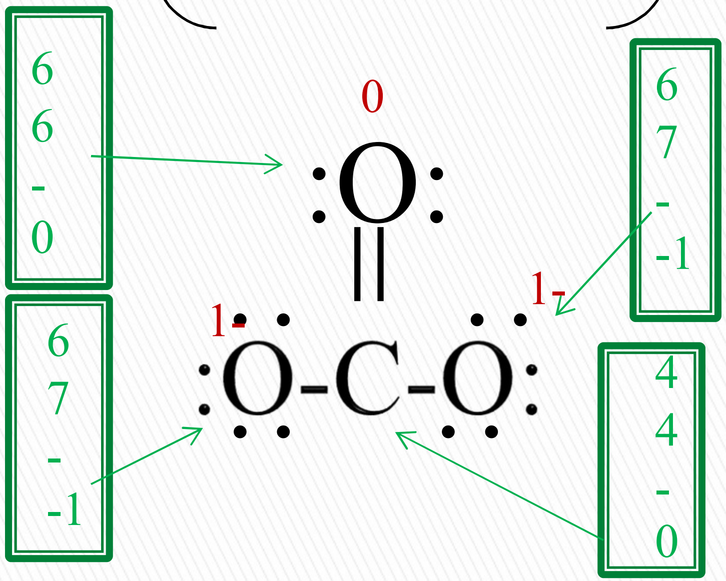
	O	C	O	O	C	O
Valence electrons:	6	4	6	6	4	6
−(Electrons assigned to atom):	6	4	6	7	4	5
Formal charge:	0	0	0	−1	0	+1

3 resonance structure

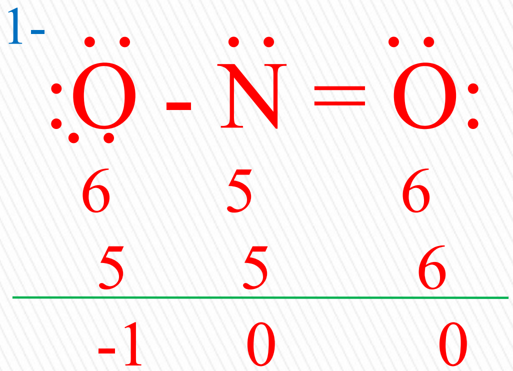
Example 9.6

- Write the formal charge for the carbonate ion?





- Write the formal charge for the NO_2^- ion?



9.9 the exception of octate rules

- There are three types of ions or molecules that do not follow the octet rule: (central atom)

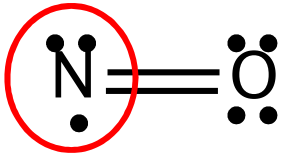
1. Ions or molecules with an odd number of electrons
2. Ions or molecules with less than an octet (the incomplete Octet)
3. Ions or molecules with more than eight valence electrons (an expanded octet)

Ions or molecules with an odd number of electrons

- Though relatively rare and usually quite unstable and reactive, there are ions and molecules with an odd number of electrons(radical).

NO

$$\begin{array}{r} N - 5e^- \\ O - 6e^- \\ \hline 11e^- \end{array}$$



$A = 5X1 + 6X1 = 11$

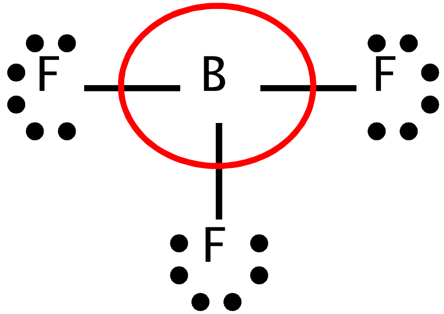
The incomplete Octet

- Covalent compounds containing Group 3 atoms may be satisfied with 6 valence electrons (Be , B , Al)

BF_3

B – $3e^-$
3F – $3 \times 7e^-$

$24e^-$



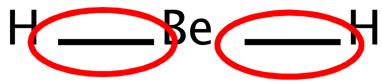
3 single bonds (3x2) = 6
9 lone pairs (9x2) = 18

Total = 24

BeH_2

Be – $2e^-$
2H – $2 \times 1e^-$

$4e^-$



An expanded octet

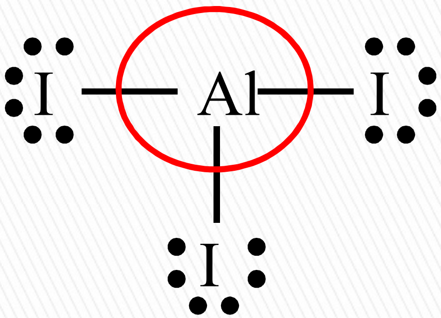
- Usually occurs in element in 3rd period and beyond
  - More than 4 bonds
  - Elements ≥ row 3 can use s, p & d orbitals and have > 8 VE
- **P: 8 OR 10**
- **S: 8, 10, OR 12**
- **Xe: 8, 10, OR 12**
- Examples

SF_6 PF_5 XeF_4

Example 9-9

- Write Lewis structure

AlI_3



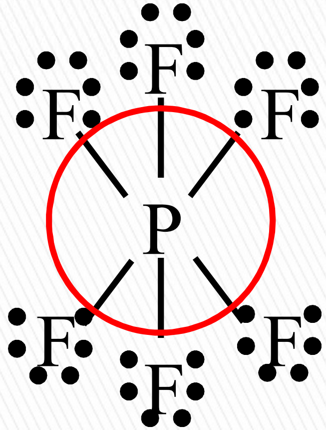
- Write Lewis structure

BeF_2

Example 9-10

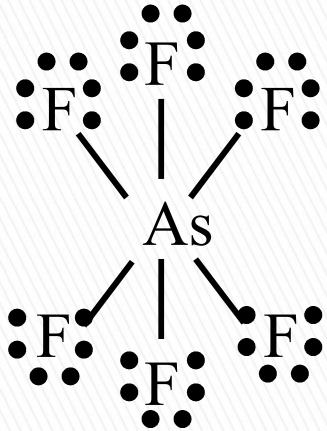
- Write Lewis structure

PF_5



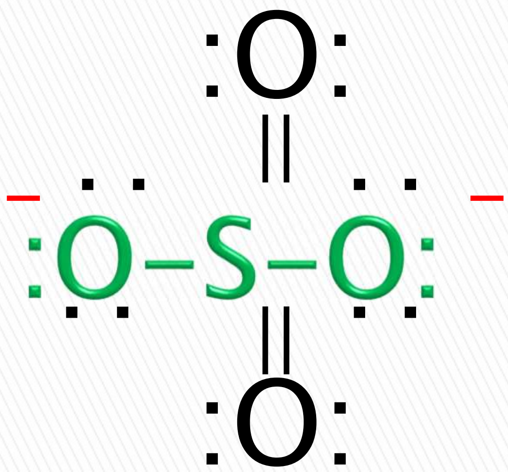
- Write Lewis structure

AsF_5

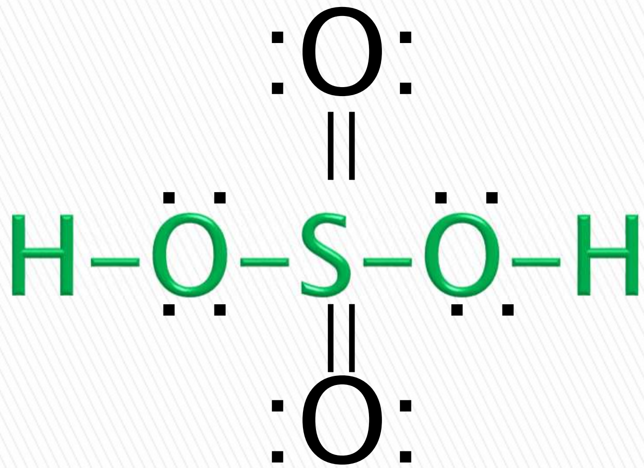


Example 9-11

- Write Lewis structure $[SO_4]^{-2}$



- Write Lewis structure H_2SO_4



Chapter 14
Chemical Equilibrium

- 14.1 the concept of equilibrium and the equilibrium constant
- 14.4 writing equilibrium constant expression
- 14.4 what does the equilibrium constant tell us
- 14.5 factors that effect chemical equilibrium

The Equilibrium Constant

Equilibrium is a state in which there are no observable changes as time goes by.

$$a\,A_{(g,l,s)} + b\,B_{(g,l,s)} \leftrightarrow c\,C_{(g,l,s)} + d\,D_{(g,l,s)}$$

conc. of products

$$K = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b}$$

conc. of reactants

equilibrium constant

- [A], [B], etc. are the equilibrium concentrations
- K ≈ [products] / [reactants]
- **K is a constant at a given temperature**
- **Solids drop out of the expression & water drops out when the solvent is water**
- **K has no unit**

- **K >> 1 ; favors products** >>> Lie to the right
- **K << 1 favors reactants** >>> Lie to the left
- K ≈ 1: roughly equal concentration of reactants and products

14.2 Writing Equilibrium Constant Expressions

$$N_2O_4\,(g) \rightleftharpoons 2NO_2\,(g)$$

$$K_c = \frac{[NO_2]^2}{[N_2O_4]} \qquad K_p = \frac{P^2_{NO_2}}{P_{N_2O_4}}$$

In most cases $K_c \neq K_p$

$$aA\,(g) + bB\,(g) \rightleftharpoons cC\,(g) + dD\,(g)$$

$$\mathbf{K_p = K_c(RT)^{\Delta n}}$$

$$\Delta n = (c + d)_{\text{products}} - (a + b)_{\text{reactants}}$$

$K_p = K_c$, when $\Delta n = 0$

Δn = moles of gaseous products – moles of gaseous reactants

- **Homogeneous Equilibria** (all species are in the same phase)

$$CH_3COOH\,(aq) + H_2O\,(l) \rightleftharpoons CH_3COO^-\,(aq) + H_3O^+\,(aq)$$

$$K'_c = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH][H_2O]}$$

$[H_2O]$ = constant

$$K_c = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = K'_c\,[H_2O]$$

General practice **not** to include units for the equilibrium constant.

B. Heterogeneous Equilibria applies to reactions in which reactants and products are in different phases.

$$CaCO_3\,(s) \rightleftharpoons CaO\,(s) + CO_2\,(g)$$

$$K'_c = \frac{[CaO][CO_2]}{[CaCO_3]}$$

$[CaCO_3]$ = constant
$[CaO]$ = constant

$$K_c = [CO_2] = K'_c \times \frac{[CaCO_3]}{[CaO]}$$

$$K_p = P_{CO_2}$$

The concentration of **solids** and **pure liquids** are not included in the expression for the equilibrium constant.

Example 14-1

- Write the equilibrium constant expression for the following reactions:

(a) $HF\ (aq) + H_2O\ (\ell) \rightleftarrows H_3O^+\ (aq) + F^-\ (aq)$ **Homogeneous**

(b) $2\ NO\ (g) + O_2\ (g) \rightleftarrows 2\ NO_2\ (g)$ **Homogeneous**

(c) $CH_3COOH\ (aq) + C_2H_5COH\ (aq) \rightleftarrows CH_3COOC_2H_5 + H_2O\ (\ell)$

$$K_c = \frac{[H_3O^+][F^-]}{[HF]}$$

$$K_c = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

$$K_c = \frac{[CH_3COOC_2H_5]}{[CH_3COOH]\ [C_2H_5COH]}$$

$$K_p = \frac{P^2_{NO2}}{P^2_{NO}\ P^2_{O2}}$$

Example 14-2

- The equilibrium concentrations

$$2\ NO\ (g) + O_2\ (g) \rightleftarrows 2\ NO_2\ (g)$$ **Homogeneous**

at 230⁰C are [NO] = 0.0542 *M* , $[O_2]$ = 0.127 *M*, and $[NO_2]$ = 15.5 *M*. Calculate the equilibrium constant K_c.

$$K_c = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

$$K_c = \frac{[15.5]^2}{[0.0542]^2[0.127]} = 6.44 \times 10^5$$

Example 14-3

- The equilibrium constant K_p for the reaction

$$2NO_2\,(g) \rightleftarrows 2NO\,(g) + O_2\,(g)$$ **Homogeneous**

- is 158 at 1000K. What is the equilibrium pressure of O_2 if the $P_{NO} = 0.400$ atm and $P_{NO} = 0.270$ atm?

$$K_p = \frac{P_{NO}^{2} P_{O2}}{P_{NO2}^{2}}$$

$$P_{O2} = K_p \frac{P_{NO2}^{2}}{P_{NO}^{2}}$$

$$P_{O2} = 158 \text{ x } (0.400)^2/(0.270)^2 = 347 \text{ atm}$$

Example 14-4

- Methanol is manufactured industrially by the reaction

$$CO\ (g) + 2\ H_2\ (g) \rightleftarrows CH_3OH(g)$$ **Homogeneous**

K_c = 10.5 at 220°C. What is the value of K_p at this temperature

$$K_p = K_c\ (RT)^{\Delta n}$$

$$\Delta n = 1\text{-}3 = \text{-}2$$

$$K_p = K_c\ (0.0821 \times 493)^{-2} = 6.41 \times 10^{-3}$$

Example 14-5

- Write the equilibrium constant expression for the following reactions:

(a) $(NH_4)_2Se\ (s) \rightleftarrows 2NH_3\ (g) + H_2Se\ (g)$ Heterogeneous

(b) $AgCl\ (s) \rightleftarrows Ag^+\ (aq) + Cl^-\ (aq)$ Heterogeneous

(c) $P_4\ (s) + 6Cl_2(g) \rightleftarrows 4PCl_3(\ell)$ Heterogeneous

$$K_c = [NH_3]^2\ [H_2Se]$$

$$K_p = P_{NH3}{}^2\ P_{H2Se}$$

$$K_c = [Ag^+\ ]^2\ [Cl^-]$$

$$K_c = \frac{1}{[Cl_2]^6}$$

Example 14-6

Consider the following equilibrium at 295 K:

$$NH_4HS\,(s) \rightleftarrows NH_3\,(g) + H_2S\,(g)$$ Heterogeneous

The partial pressure of each gas is 0.265 atm. Calculate K_p and K_c for the reaction?

$$K_p = P_{NH_3} P_{H_2S} = 0.265 \text{ x } 0.265 = 0.0702$$

$$K_p = K_c(RT)^{\Delta n}$$

$$K_c = K_p(RT)^{-\Delta n}$$

$$\Delta n = 2 - 0 = 2 \qquad T = 295 \text{ K}$$

$$K_c = 0.0702 \text{ x } (0.0821 \text{ x } 295)^{-2} = 1.20 \text{ x } 10^{-4}$$

Q: What is K_p in terms of K_c for the following reaction ?

$$2NO(g) + O_2(g) \leftrightarrow 2NO_2(g)$$

A. $K_p = K_cRT$
B. $K_p = K_c/RT$
C. $K_p = K_cR/T$
D. $K_p = K_c$
E. $K_p = K_c/(RT)^2$

Solution:

$$K_p = K_c(RT)^{\Delta n}$$

$$\Delta n = 2 - 3 = -1$$

$$K_p = K_c(RT)^{-1} = \frac{K_c}{RT}$$

14.4 what does the equilibrium constant tell us

- A) Predicting the direction of a reaction
- The ***reaction quotient (Q_c)*** is calculated by substituting the initial concentrations of the reactants and products into the equilibrium constant (K_c) expression.
- **IF**

- $Q_c > K_c$ system proceeds from right to left to reach equilibrium
- $Q_c = K_c$ the system is at equilibrium
- $Q_c < K_c$ system proceeds from left to right to reach equilibrium

- B) Calculating equilibrium concentration

1. Express the equilibrium concentrations of all species in terms of the initial concentrations and a single unknown x, which represents the change in concentration.
2. Write the equilibrium constant expression in terms of the equilibrium concentrations. Knowing the value of the equilibrium constant, solve for x.
3. Having solved for x, calculate the equilibrium concentrations of all species.

A) Predicting the direction of a reaction

Q has the same form as K, . . . but uses existing concentrations

n-Butane $\rightleftharpoons$ iso-Butane

0.25 0.35

Kc =2.5

$$Q = \frac{[iso]}{[n]} = \frac{0.35}{0.25} = 1.40$$

Since Q (1.4) < Kc (2.5) , the system at equilibrium

To reach equilibrium [iso-Butane] must increase and [n-Butane] must decrease.

Example 14.8

$$N_2\,(g) + 3H_2(g) \rightleftarrows 2NH_3\,(g)$$

Homogeneous

- 0.249mol,N_2 , 3.2×10^{-2} mol H_2 , 6.42×10^{-4} mol NH_3in A 3.50 L at 375°C , kc= 1.2, Decide whether the system is at equilibrium .if it is not predict which way the net reaction will proceed.

$$Q\,c= \frac{[NH_3]^2}{[N_2]\,[H_2]^3}$$

Since the molarity = number of moles /volume in L , **M=n/Vin I**

$[N_2]_\circ = 0.249/3.50\ L = 0.0711\ M$

$[H_2]_\circ = 3.2\times10^{-2} /3.50\ L = 9.17 \times10^{-3}\ M$

$[NH_3]_\circ = 6.42\times10^{-4} /3.50\ L = 1.83 \times10^{-4}\ M$

$$Q\,c= \frac{[NH_3]^2}{[N_2]\,[H_2]^3} = \frac{[1.83 \times10^{-4}]^2}{[0.0711][9.17 \times10^{-3}]^2} = 0.611$$

Since Q (0.611) < Kc (1.2) system , the system at equilibrium

To reach equilibrium $[NH_3]$ must increase and $[N_2]$, $[H_2]$ must decrease.

The net reaction will proceed from left to right untial equilibrium is reached .

B) Calculating equilibrium concentration

cis-stilbene ⇌ trans-stilbene Kc =24

Step 1 Define equilibrium condition in terms of initial condition and a change variable

	[cis-stibene]	[trans-stibene]
Initial	0.85	0
At equilibrium	0.85-x	x

Step 2 Put equilibrium Conic. into K_c .

$$\frac{[trans-stibene]}{[cis-stibene]} = \frac{x}{0.850 - x}$$

Step 3. Solve for x. 24(0.85-x) = x = 20.4 – 24x=x

25x= 20.4 >>>>>> x = 0.816 M

At equilibrium

[cis-stilbene] = 0.85 – 0.816 = 0.034 M

[trans-stilbene] = x = 0.816 M

Example 14.9

$H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$ $K_c = 54.3$

M= 0.5/1 =0.5 mol

Step 1 Define equilibrium condition in terms of initial condition and a change variable

	$[H_2]$	$[I_2]$	$[HI]$
Initial	x	x	0
At equilibrium	0.5-x	0.5-x	2x

Step 2 Put equilibrium Conic. into K_c .

$$\frac{[HI]^2}{[H_2]\,[I_2]}$$

Step 3. Solve for x. $54.3 = (2x)^2/(0.5-x)^2$

Square root of both sides & solve gives: $7.369 = 2x/0.5-x$

x = 0.393

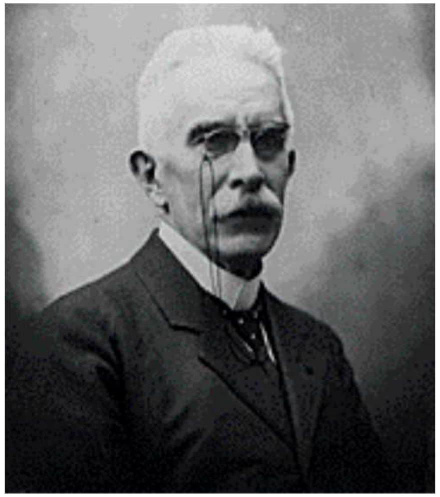
At equilibrium

$[H_2] = [I_2] = 0.5 - 0.393 = 0.107$ M

$[HI] = 2x = 0.786$ M

14.5 factors that effect chemical equilibrium

- Le Chatelier's principle , If an external stress is applied to a system at equilibrium, the system adjusts in such a way that the stress is partially offset as the system reaches a new equilibrium position.



factors that effect chemical equilibrium

1. Changes in Concentration
2. Changes in Volume and Pressure
3. Changes in Temperature
4. Adding a Catalyst

Changes in Concentration

$$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$$

↑ Add NH_3

Equilibrium shifts left to offset stress ← Add NH_3

Change	Shifts the Equilibrium
Increase concentration of product(s)	left
Decrease concentration of product(s)	right
Increase concentration of reactant(s)	right
Decrease concentration of reactant(s)	left

Example 14.11

- $N_2\,(g) + 3H_2(g) \rightleftarrows 2NH_3\,(g)$ >>>> $Kc=2.37\times10^{-3}$, $T=720^{\circ}C$

Equilibrium shifts left to offset stress ← Add NH_3 (↑) **3.65 M**

$[N_2] = 0.683$ M

$[H_2] = 8.800$M

$[NH_3] = 1.05$M the concentration increase to 3.65 M

Calculate Qc compare it value with Kc

$$\frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(3.65^2)}{0.683 \times (8.8)^3} = 2.86 \times 10^{-2}$$

because Qc (2.86×10^{-2}) > Kc (2.37×10^{-3}) , The net reaction direction from right to left until Qc = Kc

Changes in Volume and Pressure

$$A\,(g) + B\,(g) \rightleftharpoons C\,(g)$$

$\Delta n = n_{products} - n_{reactants}$

Δn = is the number of moles for substance in gaseous products &gaseous reactants

Note: Pressure and volume are inversely proportional.

Because the pressure of gases is related directly to the concentration by $P = n/V$, changing the pressure by increasing/decreasing the volume of a container will disturb an equilibrium system.

Change	**Shifts the Equilibrium**
Increase pressure (decrease volume)	Side with fewest moles of gas
Decrease pressure (Increase volume)	Side with most moles of gas

Example 14.12

- Predict the net reaction direction (increasing P & decreasing V)

(a) 2 PbS (*s*) + 3O$_2$ (g) ⇄ 2PbO (s) + 2HS$_2$ (g)

$$\Delta n = n_{products} < n_{reactants} = 2<3$$

When the volume of an equilibrium mixture of gases is reduced, a net change occurs in the direction that produces fewer moles of gas (left to right toward product).

(b) PCl$_5$ (g) ⇄ PCl$_3$(g) + Cl$_2$ (g)

$$\Delta n = n_{products} > n_{reactants} = 2>1 =1$$

When the volume of an equilibrium mixture of gases is reduced, a net change occurs in the direction that produces fewer moles of gas (right to left toward reactant).

(c) H$_2$ (*g*) + CO$_2$(*g*) ⇄ H$_2$O(g) + CO(*g*)

$$\Delta n = n_{products} = n_{reactants} = 2\text{-}2 = 0$$

The change in P,v has no effect on the equilibrium .

Example

- Which of the following increasing pressure will shift equilibrium to left

(a) 2 PbS (*s*) + $3O_2$ (g) ⇄ 2PbO (s) + $2HS_2$ (g)

(b) PCl_5 (g) ⇄ PCl_3(g) + Cl_2 (g)

(c) H_2 (*g*) + CO_2(*g*) ⇄ H_2O(g) + CO(*g*)

(d) No correct answer

- Which of the following increasing pressure will cause no change in equilibrium

(a) 2 PbS (*s*) + $3O_2$ (g) ⇄ 2PbO (s) + $2HS_2$ (g)

(b) PCl_5 (g) ⇄ PCl_3(g) + Cl_2 (g)

(c) H_2 (*g*) + CO_2(*g*) ⇄ H_2O(g) + CO(*g*)

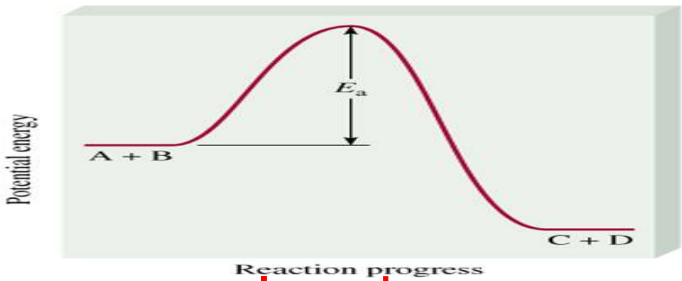
(d) No correct answer

Changes in Temperature

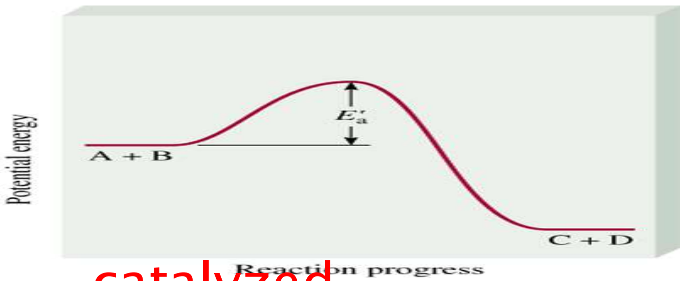
Change	Exothermic - ΔH	Endothermic +ΔH
Increase temperature	*K* decreases	*K* increases
Decrease temperature	*K* increases	*K* decreases

•Adding a Catalyst

- does not change *K*
- does not shift the position of an equilibrium system
- system will reach equilibrium sooner



uncatalyzed



catalyzed

Catalyst lowers E_a for both forward and reverse reactions

Catalyst does not change equilibrium constant or shift equilibrium.

Le Châtelier's Principle

Change	Shift Equilibrium	Change Equilibrium Constant
Concentration	yes	no
Pressure	yes	no
Volume	yes	no
Temperature	yes	yes
Catalyst	no	no

LE CHATELIER'S PRINCIPLE

STRESS	SHIFT	WHY?
increase concentration of a substance	*away* from substance	extra concentration needs to be used up
decrease concentration of a substance	*towards* substance	need to produce more of substance to make up for what was removed
increase pressure of system	towards *fewer* moles of gas	for gas: pressure increase = volume decrease
decrease pressure of system	towards *more* moles of gas	for gas: pressure decrease = volume increase
increase temperature of system	*away* from heat/ energy *exothermic* reaction is favored	extra heat/ energy must be used up
decrease temperature of system	*towards* heat/ energy *exothermic* reaction is favored	more heat/ energy needs to be produced to make up for the loss
add a catalyst	NO SHIFT	The rates of both the forward and reverse reactions are increased by the same amount.

Example 14.13

- Predict the net reaction direction a) if RXN heated at constant V, b) some N_2F_4 removed at constant T&V c) Decrease P? d) catalyst is added

$$N_2F_4\,(g) \rightleftarrows 2\,NF_2\,(g) \quad >> \Delta H = 38.5\ kJ/mol$$

- a) ΔH>0 >>endothermic reaction ,T increase, *K* increase, a net change occurs in the direction is from left to right toward product).

- b) Conc. of the reactant decrease the system shift right to left (some **NF_2 combines to produce N_2F_4)**

- c) **P** decrease the system shift left to right .

- d) if catalyst is added to reaction mixture ,the reaction will reach equilibrium faster but no change in the change equilibrium constant or shift equilibrium.

Q: Which of the following will result in an equilibrium shift to the right?

$$PCl_3(g) + Cl_2(g) \leftrightarrow PCl_5(g).........\quad \Delta H = -87.9\ \text{kJ/mol}$$

1) Increase temperature/increase volume
2) Increase temperature/decrease volume
3) Decrease temperature/increase volume
4) Decrease temperature/decrease volume
5) None of the above

Chapter 15
Acids and Bases

- 15.2 the Acids and Bases properties of water
- 15.3 PH- a measure of acidity

Dr Laila Al-Harbi

Acids and Bases

- **Acid: Substance that produces hydrogen ions in water solution.**

$$HCl\ (aq) \rightarrow H^{+}(aq) + Cl^{-}(aq)$$

- **Base: Substance that produces hydroxide ions in water solution.**

$$NaOH\ (aq) \rightarrow Na^{+}(aq) + OH^{-}(aq)$$

- **An acid neutralizes a base**

$$H^{+}(aq) + OH^{-}(aq) \rightarrow H2O(\ell)$$

15.2 the Acids and Bases properties of water

- water is unique solvent , it can act as acid or base.
- In pure water, a few molecules act as bases and a few act as acids.

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

acid(1) + base (1) ⇄ acid(1) + base (1)

- This is referred to as autoionization of water
- The equilibrium expression for this process is

$$K_c = [H_3O^+] [OH^-]$$

- This special equilibrium constant is referred to as the ion constant for water, K_w.

$$\text{At } 25°C,\ K_w = 1.0 \times 10^{-14}$$

In pure water,

$$K_w = [H_3O^+] [OH^-] = 1.0 \times 10^{-14}$$

Because in pure water $[H_3O^+] = [OH^-]$,

$$[H_3O^+] = (1.0 \times 10^{-14})^{1/2} = 1.0 \times 10^{-7}$$

In acidic solution

$$[H3O^+] > [OH^-]$$

In basic solution

$$[H3O^+] < [OH^-]$$

Example 15.2

- Calculate the $[H^+]$ ions in ammonia , $[OH^-]$ =0.0025 M
- $K_w = [H_3O^+]\ [OH^-] = 1.0 \times 10^{-14}$
- $[H_3O^+] = 1.0 \times 10^{-14} / [OH^-]$
- $[H_3O^+] = 1.0 \times 10^{-14} / 0.0025$
- $[H_3O^+] = 4.0 \times 10^{-12}$ M

- Calculate the $[OH^-]$ ions in a 1.3 M HCl.
- $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$
- $[OH^-] = 1.0 \times 10^{-14} / [H_3O^+]$
- $[OH^-] = 1.0 \times 10^{-14} / 1.3$
- $[OH^-] = 7.7 \times 10^{-15}$ M

15-3 pH - A Measure of Acidity

- pH is defined as the negative base-10 logarithm of the hydronium ion concentration.

$$pH = -\log [H_3O^+] \;....\; [H_3O^+] = 10^{-pH}$$

- In the same manner

$$pOH = -\log [OH^-] \;.....\; [OH^-] = 10^{-pOH}$$

- In the same manner

$$pK_w = -\log [14\times10^{-14}] = 14$$

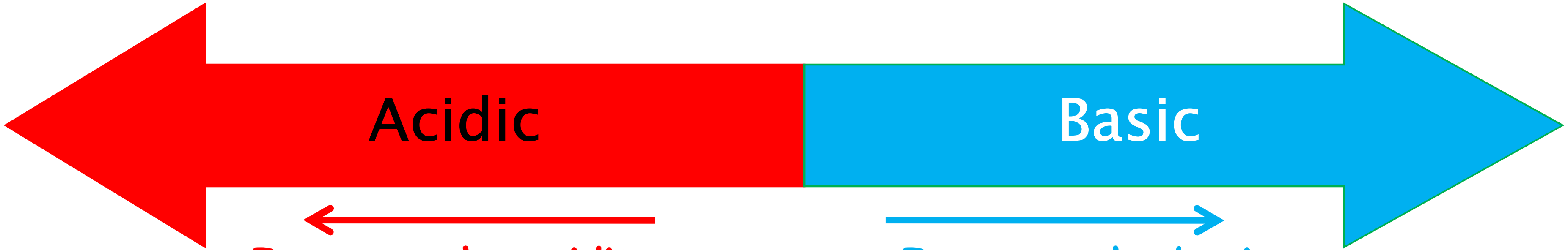
- <u>In pure water,</u>

$$\mathbf{pH + pOH = 14}$$

$$\mathbf{pH = pOH = 7}$$

pH Range

0 1 2 3 4 5 6 7 8 9 10 11 12 13 14



Increase the acidity — Neutral — Increase the basisty

Acidic	Neutral	Basic
$[H^+] > [OH^-]$	$[H^+] = [OH^-]$	$[OH^-] > [H^+]$
$pH < 7$	$pH = pOH = 7$	$pH > 7$

Example 15.3

- The $[H^+]=3.2 \times 10^{-4}$ M.
- The $[H^+]=1.0 \times 10^{-3}$ M.
  What is the pH in the two occasions.
- $pH = -\log [H_3O^+]$
- $pH = -\log 3.2 \times 10^{-4} = 3.49$
- $pH = -\log 1.0 \times 10^{-3} = 3.00$
- $[H_3O^+]$ increase ,pH decrease >>> more acidic

- The $[H^+]=0.76$ M, nitric acid solution ,What is the pH .
- $pH = -\log [H_3O^+]$
- $pH = -\log 0.76 = 0.12$

Example 15.4

- The pH = 4.82 , What is the $[H^+]$ of the rain water .
- $[H_3O^+] = 10^{-pH}$
- $[H_3O^+] = 10^{-4.82}$
- $[H_3O^+] = 1.5 \times 10^{-5}M$

- The pH = 3.33 , What is the $[H^+]$ of orange juice
- $[H_3O^+] = 10^{-pH}$
- $[H_3O^+] = 10^{-3.33}$
- $[H_3O^+] = 4.7 \times 10^{-4}M$

Example 15.5

- The $[OH^-]=2.9 \times 10^{-4}$ M. What is the pH of the NaOH solution
- $pOH = -\log [OH^-]$
- $pOH = -\log 2.9 \times 10^{-4} = 3.54$

$pH + pOH = 14$

$pH = 14 - pOH$

$= 14\text{-}3.54 = \underline{10.46}$

- The $[OH^-]=2.5 \times 10^{-7}$ M. What is the pH of solution the blood?
- $pOH = -\log [OH^-]$
- $pOH = -\log 2.5 \times 10^{-7}$ =

$pH + pOH = 14$

$pH = 14 - pOH$

$= 14\text{-}3.54 = \underline{7.4}$

Acidic solution — basic solution

$$[H_3O^+] \times [OH^-] = 1.0 \times 10^{-14}$$

$$pH + pOH = 14$$

Determining pH, pOH, $[OH^-]$, $[H_3O^+]$

Remember $[H^+] = [H_3O^+]$

Use this chart to determine unknowns
given one value (25°C)
$K_w = 1.01 \cdot 10^{-14}$