AMS 105
APPLIED CHEMISTRY
COURSE MATERIAL
PRÉCIS

DEPARTMENT OF APPLIED MILITARY SCIENCE

ROYAL MILITARY COLLEGE OF CANADA

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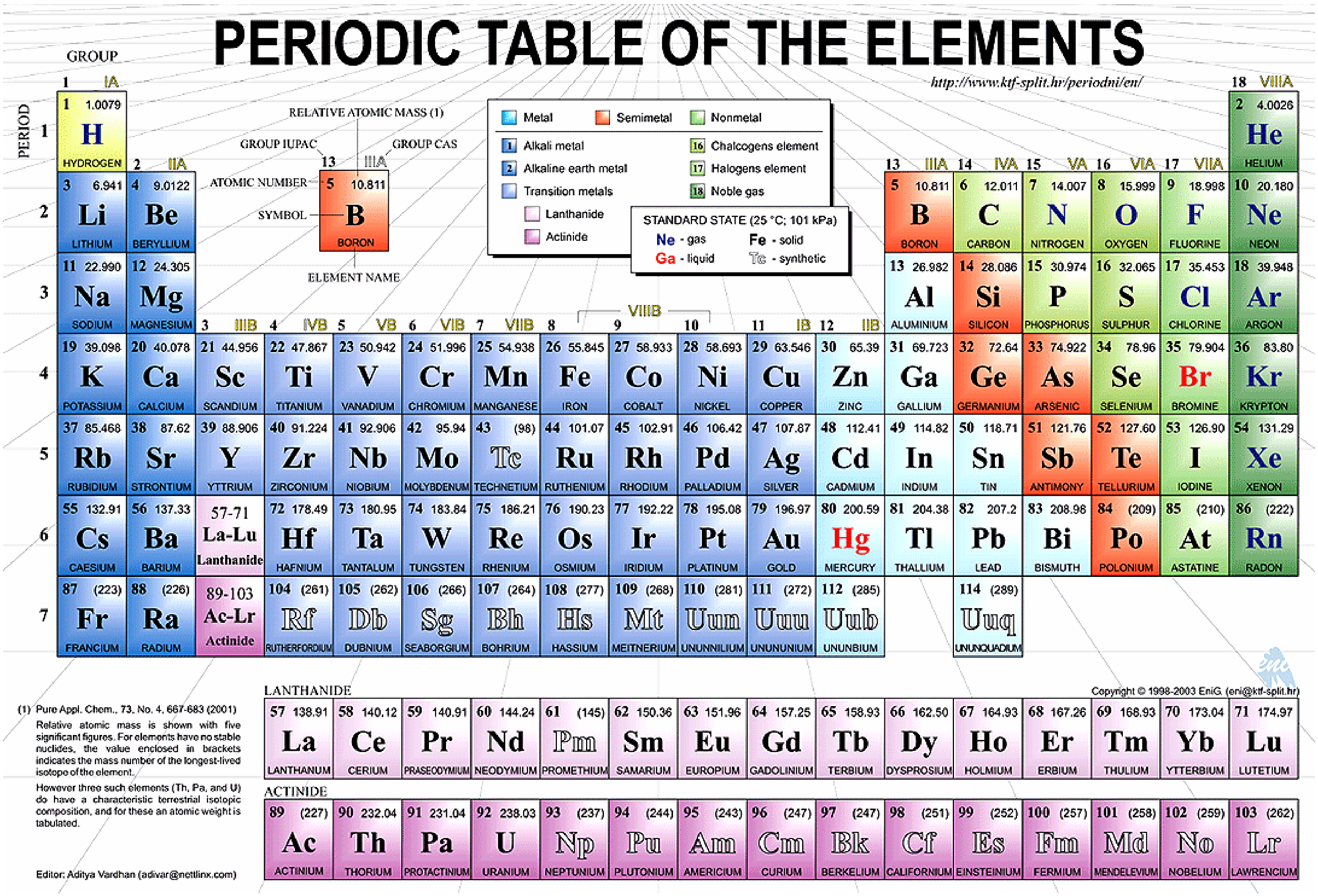
PERIODIC TABLE OF THE ELEMENTS
http://www.ktf-split.hr/periodni/en/
GROUP
PERIOD
1 IA
2 IIA
3 IIIB
4 IVB
5 VB
6 VIB
7 VIIB
8
9
10
VIIIB
11 IB
12 IIB
13 IIIA
14 IVA
15 VA
16 VIA
17 VIIA
18 VIIIA
RELATIVE ATOMIC MASS (1)
GROUP IUPAC
GROUP CAS
ATOMIC NUMBER
SYMBOL
ELEMENT NAME
13 IIIA
5 10.811 B BORON
Metal
Semimetal
Nonmetal
1 Alkali metal
2 Alkaline earth metal
Transition metals
Lanthanide
Actinide
16 Chalcogens element
17 Halogens element
18 Noble gas
STANDARD STATE (25 °C; 101 kPa)
Ne - gas
Fe - solid
Ga - liquid
Tc - synthetic
1 1.0079 H HYDROGEN
2 4.0026 He HELIUM
3 6.941 Li LITHIUM
4 9.0122 Be BERYLLIUM
5 10.811 B BORON
6 12.011 C CARBON
7 14.007 N NITROGEN
8 15.999 O OXYGEN
9 18.998 F FLUORINE
10 20.180 Ne NEON
11 22.990 Na SODIUM
12 24.305 Mg MAGNESIUM
13 26.982 Al ALUMINIUM
14 28.086 Si SILICON
15 30.974 P PHOSPHORUS
16 32.065 S SULPHUR
17 35.453 Cl CHLORINE
18 39.948 Ar ARGON
19 39.098 K POTASSIUM
20 40.078 Ca CALCIUM
21 44.956 Sc SCANDIUM
22 47.867 Ti TITANIUM
23 50.942 V VANADIUM
24 51.996 Cr CHROMIUM
25 54.938 Mn MANGANESE
26 55.845 Fe IRON
27 58.933 Co COBALT
28 58.693 Ni NICKEL
29 63.546 Cu COPPER
30 65.39 Zn ZINC
31 69.723 Ga GALLIUM
32 72.64 Ge GERMANIUM
33 74.922 As ARSENIC
34 78.96 Se SELENIUM
35 79.904 Br BROMINE
36 83.80 Kr KRYPTON
37 85.468 Rb RUBIDIUM
38 87.62 Sr STRONTIUM
39 88.906 Y YTTRIUM
40 91.224 Zr ZIRCONIUM
41 92.906 Nb NIOBIUM
42 95.94 Mo MOLYBDENUM
43 (98) Tc TECHNETIUM
44 101.07 Ru RUTHENIUM
45 102.91 Rh RHODIUM
46 106.42 Pd PALLADIUM
47 107.87 Ag SILVER
48 112.41 Cd CADMIUM
49 114.82 In INDIUM
50 118.71 Sn TIN
51 121.76 Sb ANTIMONY
52 127.60 Te TELLURIUM
53 126.90 I IODINE
54 131.29 Xe XENON
55 132.91 Cs CAESIUM
56 137.33 Ba BARIUM
57-71 La-Lu Lanthanide
72 178.49 Hf HAFNIUM
73 180.95 Ta TANTALUM
74 183.84 W TUNGSTEN
75 186.21 Re RHENIUM
76 190.23 Os OSMIUM
77 192.22 Ir IRIDIUM
78 195.08 Pt PLATINUM
79 196.97 Au GOLD
80 200.59 Hg MERCURY
81 204.38 Tl THALLIUM
82 207.2 Pb LEAD
83 208.98 Bi BISMUTH
84 (209) Po POLONIUM
85 (210) At ASTATINE
86 (222) Rn RADON
87 (223) Fr FRANCIUM
88 (226) Ra RADIUM
89-103 Ac-Lr Actinide
104 (261) Rf RUTHERFORDIUM
105 (262) Db DUBNIUM
106 (266) Sg SEABORGIUM
107 (264) Bh BOHRIUM
108 (277) Hs HASSIUM
109 (268) Mt MEITNERIUM
110 (281) Uun UNUNNILIUM
111 (272) Uuu UNUNUNIUM
112 (285) Uub UNUNBIUM
114 (289) Uuq UNUNQUADIUM
LANTHANIDE
57 138.91 La LANTHANUM
58 140.12 Ce CERIUM
59 140.91 Pr PRASEODYMIUM
60 144.24 Nd NEODYMIUM
61 (145) Pm PROMETHIUM
62 150.36 Sm SAMARIUM
63 151.96 Eu EUROPIUM
64 157.25 Gd GADOLINIUM
65 158.93 Tb TERBIUM
66 162.50 Dy DYSPROSIUM
67 164.93 Ho HOLMIUM
68 167.26 Er ERBIUM
69 168.93 Tm THULIUM
70 173.04 Yb YTTERBIUM
71 174.97 Lu LUTETIUM
ACTINIDE
89 (227) Ac ACTINIUM
90 232.04 Th THORIUM
91 231.04 Pa PROTACTINIUM
92 238.03 U URANIUM
93 (237) Np NEPTUNIUM
94 (244) Pu PLUTONIUM
95 (243) Am AMERICIUM
96 (247) Cm CURIUM
97 (247) Bk BERKELIUM
98 (251) Cf CALIFORNIUM
99 (252) Es EINSTEINIUM
100 (257) Fm FERMIUM
101 (258) Md MENDELEVIUM
102 (259) No NOBELIUM
103 (262) Lr LAWRENCIUM
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(1) Pure Appl. Chem., 73, No. 4, 667-683 (2001)
Relative atomic mass is shown with five significant figures. For elements have no stable nuclides, the value enclosed in brackets indicates the mass number of the longest-lived isotope of the element.
However three such elements (Th, Pa, and U) do have a characteristic terrestrial isotopic composition, and for these an atomic weight is tabulated.
Editor: Aditya Vardhan (adivar@nettlinx.com)

x

TABLEAU PÉRIODIQUE DES ÉLÉMENTS

PÉRIODE \ GROUPE	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 (VIIIB)	9 (VIIIB)	10 (VIIIB)	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 1.0079 H HYDROGÈNE																	2 4.0026 He HÉLIUM
2	3 6.941 Li LITHIUM	4 9.0122 Be BÉRYLLIUM											5 10.811 B BORE	6 12.011 C CARBONE	7 14.007 N AZOTE	8 15.999 O OXYGÈNE	9 18.998 F FLUOR	10 20.180 Ne NÉON
3	11 22.990 Na SODIUM	12 24.305 Mg MAGNÉSIUM											13 26.982 Al ALUMINIUM	14 28.086 Si SILICIUM	15 30.974 P PHOSPHORE	16 32.065 S SOUFRE	17 35.453 Cl CHLORE	18 39.948 Ar ARGON
4	19 39.098 K POTASSIUM	20 40.078 Ca CALCIUM	21 44.956 Sc SCANDIUM	22 47.867 Ti TITANE	23 50.942 V VANADIUM	24 51.996 Cr CHROME	25 54.938 Mn MANGANÈSE	26 55.845 Fe FER	27 58.933 Co COBALT	28 58.693 Ni NICKEL	29 63.546 Cu CUIVRE	30 65.38 Zn ZINC	31 69.723 Ga GALLIUM	32 72.64 Ge GERMANIUM	33 74.922 As ARSENIC	34 78.96 Se SÉLÉNIUM	35 79.904 Br BROME	36 83.798 Kr KRYPTON
5	37 85.468 Rb RUBIDIUM	38 87.62 Sr STRONTIUM	39 88.906 Y YTTRIUM	40 91.224 Zr ZIRCONIUM	41 92.906 Nb NIOBIUM	42 95.96 Mo MOLYBDÈNE	43 (98) Tc TECHNÉTIUM	44 101.07 Ru RUTHÉNIUM	45 102.91 Rh RHODIUM	46 106.42 Pd PALLADIUM	47 107.87 Ag ARGENT	48 112.41 Cd CADMIUM	49 114.82 In INDIUM	50 118.71 Sn ETAIN	51 121.76 Sb ANTIMOINE	52 127.60 Te TELLURE	53 126.90 I IODE	54 131.29 Xe XÉNON
6	55 132.91 Cs CÉSIUM	56 137.33 Ba BARYUM	57-71 La-Lu Lanthanides	72 178.49 Hf HAFNIUM	73 180.95 Ta TANTALE	74 183.84 W TUNGSTÈNE	75 186.21 Re RHÉNIUM	76 190.23 Os OSMIUM	77 192.22 Ir IRIDIUM	78 195.08 Pt PLATINE	79 196.97 Au OR	80 200.59 Hg MERCURE	81 204.38 Tl THALLIUM	82 207.2 Pb PLOMB	83 208.98 Bi BISMUTH	84 (209) Po POLONIUM	85 (210) At ASTATE	86 (222) Rn RADON
7	87 (223) Fr FRANCIUM	88 (226) Ra RADIUM	89-103 Ac-Lr Actinides	104 (267) Rf RUTHERFORDIUM	105 (268) Db DUBNIUM	106 (271) Sg SEABORGIUM	107 (272) Bh BOHRIUM	108 (277) Hs HASSIUM	109 (276) Mt MEITNERIUM	110 (281) Ds DARMSTADTIUM	111 (280) Rg ROENTGENIUM	112 (285) Cn COPERNICIUM	113 (. . .) Uut UNUNTRIUM	114 (287) Fl FLEROVIUM	115 (. . .) Uup UNUNPENTIUM	116 (291) Lv LIVERMORIUM	117 (. . .) Uus UNUNSEPTIUM	118 (. . .) Uuo UNUNOCTIUM

LANTHANIDES

57 138.91 La LANTHANE	58 140.12 Ce CÉRIUM	59 140.91 Pr PRASÉODYME	60 144.24 Nd NÉODYME	61 (145) Pm PROMÉTHIUM	62 150.36 Sm SAMARIUM	63 151.96 Eu EUROPIUM	64 157.25 Gd GADOLINIUM	65 158.93 Tb TERBIUM	66 162.50 Dy DYSPROSIUM	67 164.93 Ho HOLMIUM	68 167.26 Er ERBIUM	69 168.93 Tm THULIUM	70 173.05 Yb YTTERBIUM	71 174.97 Lu LUTÉTIUM

ACTINIDES

89 (227) Ac ACTINIUM	90 232.04 Th THORIUM	91 231.04 Pa PROTACTINIUM	92 238.03 U URANIUM	93 (237) Np NEPTUNIUM	94 (244) Pu PLUTONIUM	95 (243) Am AMÉRICIUM	96 (247) Cm CURIUM	97 (247) Bk BERKÉLIUM	98 (251) Cf CALIFORNIUM	99 (252) Es EINSTEINIUM	100 (257) Fm FERMIUM	101 (258) Md MENDELÉVIUM	102 (259) No NOBÉLIUM	103 (262) Lr LAWRENCIUM

MASSE ATOMIQUE RELATIVE (1)
GROUPE IUPAC — 13
GROUPE CAS — IIIA
NOMBRE ATOMIQUE — 5
10.811
SYMBOLE — B
BORE
NOM DE L'ÉLÉMENT

Métaux · Métalloïdes · Non-métaux

- Métaux alcalins
- Métaux alcalino-terreux
- Métaux de transition
- Lanthanides
- Actinides
- Chalcogènes
- Halogènes
- Gaz nobles

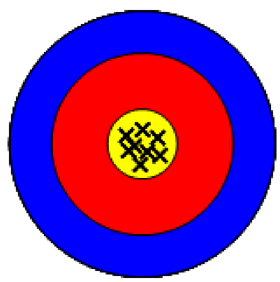
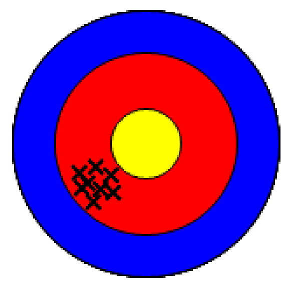
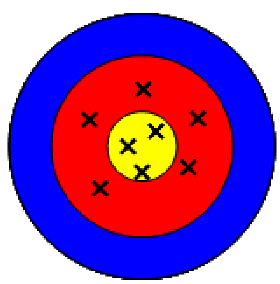
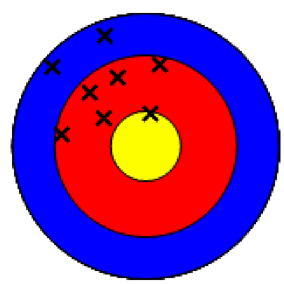
ETAT PHYSIQUE (25 °C; 101 kPa)
Ne - gaz · Fe - solide
Hg - liquide · Tc - synthétique

(1) Pure Appl. Chem., **81**, No. 11, 2131-2156 (2009)
La masse atomique relative est donnée avec cinq chiffres significatifs. Pour les éléments qui n'ont pas de nucléides stables, la valeur entre parenthèses indique le nombre de masse de l'isotope de l'élément ayant la durée de vie la plus grande. Toutefois, pour les trois éléments (Th, Pa et U) qui ont une composition isotopique terrestre connue, une masse atomique est indiquée.

SECTION 0 – NUMBERS IN SCIENCE

0.1 Types of Numbers

- A numerical quantity is presented in one of two ways:
  - Exact number (ex. counting bullets in an ammo box).
  - A measurement (ex. measuring the diameter of a round). There are two important aspects of measured numerical quantities
    - Accuracy: How close a measured result is to the true value.
    - Precision: The consistency or reproducibility of a measurement using identical experimental procedures.

	Accurate	Inaccurate (systematic error)
Precise		
Imprecise (reproducibility error)		

0.2 Errors in Measurements

0.2.1 Corrections

- Every set of experimental data must be corrected by:
  1. Comparison to a control data set (data obtained from a sample or experiment of known outcome). This helps moderate accuracy by establishing a correction factor for your data.
  2. Repeat testing using identical experimental procedures. The data collected during repeat testing will allow you to obtain a statistical average. This will moderate your experimental precision.

0.2.2 Experimental Error

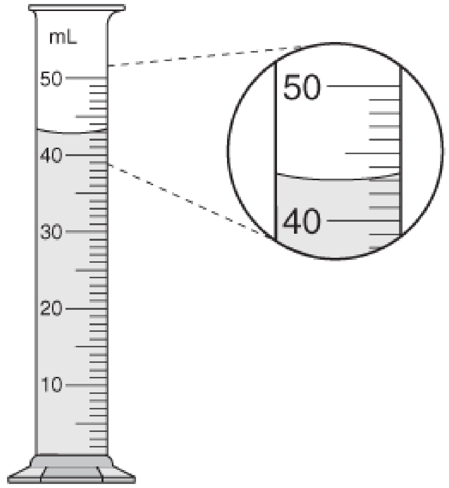
- No measurement is absolutely exact so measurements are considered to be an approximation to the true value.
- The confidence issues in experiments, tests or measurements are termed experimental errors.
- There are two types of experimental errors:
    - Systematic errors: Constant errors
        - They have the same algebraic sign as the measurement. They make measurement of a physical quantity either too large or too small.
        - Systematic errors affects accuracy so a control data set helps stem this problem.
        - ex. Barrel wear in a rifle and a site that is not zeroed.
    - Random errors: Unavoidable errors that result from chance variations.
        - They are generally small and have an equal probability of being positive or negative.
        - Random errors are associated with small irregularities in the precision of measuring equipment and/or human interpretation irregularities.
        - Repeat testing (maximizing the number of data sets) minimizes this problem.

0.2.3 Uncertainty

- Uncertainty Rule = Uncertainty lies in the last digit written in a measured number.
- Uncertainty can be expressed in many ways. It is usually shown with a ± sign (ex. $16.3 \pm 0.1\ ^{o}C$).
- It is generally accepted that the error attributed to a measurement is ± the smallest reported digit.

0.2.3.1 Example

- In the graduated cylinder shown on the next page, I estimate the volume to be 43.1 mL.
- But in truth, I can only say for certain that the volume is very close to 43 mL. Someone else might judge it to be 43.0 mL or 43.2 mL.
- So I present the value as 43.1 ± 0.1 mL, meaning that it is somewhere between 43.0 and 43.2 mL.



Self Quiz 0.2.3

1. What is the uncertainty in the actual diameter of the following rounds?
    a. 5.56 mm
    b. 155 mm
    c. 0.50 cal

Self Quiz 0.2.3 – Answers

1. What is the uncertainty in the actual diameter of the following rounds?
    a. 5.56 mm = ***0.01 mm***
    b. 155 mm = ***1 mm***
    c. 0.50 cal = ***0.01 cal***

0.3 Significant Figures

0.3.1 Definition

- The significant figures of a number are the digits that carry meaning contributing to its precision (ie. significant figures are the digits in a value that can be trusted).
- Significant figures are one way of keeping track of how much error there is in a measurement.

0.3.2 Example

- Referring to the graduated cylinder in section 0.2.3.1, my estimation of the volume, 43.1 mL, has three significant figures.

- If I wrote the volume as 43.1000 mL, there would be six significant figures, which would suggest that the value is more accurate than it really is.
- If I wrote the volume as 40 mL, there would only be one significant figure, which would suggest that the value is less accurate than it really is.

0.3.3 Rules for Identifying Significant Figures

1. All non-zero digits are significant (ex. 351 = 3 sig figs).
2. Zeros appearing anywhere between two non-zero digits are significant (ex. 7004 = 4 sig figs).
3. Leading zeros are not significant (ex. 004 = 1 sig fig).
4. Trailing zeros in a number containing a decimal point are significant (ex. 4.00 = 3 sig figs).
5. The significance of trailing zeros in a number without a decimal point is ambiguous; it is avoided by using scientific notation. If you are presented with trailing zeros, assume they are not significant (ex. 400 = 1 sig fig).
6. A number with all zero digits has no significant digits because the uncertainty is larger than the actual measurement (ex. 0.000 = no significance).

0.3.4 Round Decimal Numbers

1. If the first digit to be dropped is less than 5, the last retained digit is unchanged (ex. 37.4 becomes 37).
2. If the first digit to be dropped is greater than 5, the last retained digit is increased by 1 (ex. 37.7 becomes 38).
3. If the first digit to be dropped is 5:
    a. If the 5 is followed by zeroes, round the number so that it becomes even:
        i. If the last retained digit is even, it is left unchanged (ex. 38.50 becomes 38).
        ii. If the last retained digit is odd, it is increased by 1 (ex. 37.50 becomes 38).
    b. If the 5 is followed by non-zero digits, the last retained digit is increased by 1 (ex. 37.52 becomes 38).

0.3.5 Operations with Significant Figures

- Addition & Subtraction: The result should match the smallest number of decimal places in the least precise number (ex. 23.43 + 2.3 = 25.7).
    - Note: This is a measurement of decimal places, not significant figures.

- Multiplication & Division: The number of significant figures in the result should match the number of significant figures in the measurement with the fewest significant digits (ex. 23.43 x 2.3 = 54).

0.3.6 Pure Numbers and Defined Ratios

- A natural number, such as three, can be written as 3.0, 3.00, 3.000, … with any number of arbitrary trailing zeroes. It always means the same thing.
- Multiplication and division of a measurement by a natural number preserves the number of significant figures in the original measurement.
  - ex. If you measure that a nickel weighs 3.95 g and you are asked how much three nickels weigh: 3.95 g x 3 = 11.9 g
- This rule applies to defined (or known) ratios as well.
  - ex. If a standard wedding band weighs 9.9 grams and you are asked to calculate the combined weight of the wedding bands of five couples: 9.9 grams/person x 5 couples x 2 persons/couple = 99 grams (2 people/couple is a defined ratio)

0.4 Scientific Notation

0.4.1 Definition

- Scientists have developed a shorter method to express very large and very small numbers in a standard form called scientific notation.[1]
- The general form for scientific notation is: $A \times 10^{n}$
  - A = measured value
  - n = an exponent (power) of 10 that indicates the magnitude of the measurement
- ex. $3051 = 3.051 \times 10^{3}$

0.4.2 Rules for Scientific Notation

1. The A value should always contain a decimal point. This indicates the level of precision.
2. The 10^{n} portion of the equation indicates the magnitude of the measurement.
   - n = positive for large numbers (greater than 1)
   - n = negative for a small numbers (less than 1)

[1] For an example of the magnitude of numbers that can be represented by scientific notation, see section 6.1.2. Note that this example only shows you a large number represented with scientific notation but the notation can represent very small numbers as well.

0.4.3 Operations with Scientific Notation

- Addition & Subtraction: Before adding or subtracting two numbers expressed in scientific notation, their *n* value must be equal.
    - ex. $(3.5 \times 10^4) + (6.2 \times 10^3)$
      $= (3.5 \times 10^4) + (0.62 \times 10^4)$
      $= (3.5 + 0.62) \times 10^4$
      $= 4.12 \times 10^4$
- Multiplication: To multiply two powers with the same base, add their exponents (ex. $10^5 \times 10^3 = 10^{5+3} = 10^8$).
- So to multiply two numbers expressed in scientific notation, multiply their A values then add their *n* values.
    - ex. $(3.5 \times 10^4) \times (6.2 \times 10^3)$
      $= (3.5 \times 6.2) \times 10^{4+3}$
      $= 22 \times 10^7$
      $= 2.2 \times 10^8$
- Division: To divide two powers with the same base, subtract their exponents (ex. $10^5 \div 10^3 = 10^{5-3} = 10^2$).
- So to divide two numbers expressed in scientific notation, divide their A values then subtract their *n* values.

0.5 International System of Units

0.5.1 Definition

- The International System of Units (SI) is used by scientists worldwide.
- It is devised around the number ten.
- It is also the modern form of the metric system.
- The SI units are:
    - Length = metre (m)
    - Mass = kilogram (kg)
    - Time = second (s)
    - Electric current = ampere (A)
    - Temperature = Kelvin (K)
    - Luminous intensity = candela (cd)
    - Amount of substance = mole (mol)[2]

0.5.2 Orders of Magnitude

- Prefixes are added to SI units to indicate the magnitude of a measurement.

---

[2] Yes, the symbol for "mole" is "mol". Moles will be described in detail in section 6.1.

    - ex. It is inconvenient to give the distance from Kingston to Orlando in metres so the prefix kilo is used to make kilometres (1 km = 1000 m); Orlando is 2100 km (or 2,061,000 metres) from Toronto.
- Prefixes from 10^{-18} to 10^{18} are shown below.

Power	Prefix	Abbreviation
10^{-18}	atto-	a
10^{-15}	femto-	f
10^{-12}	pico-	p
10^{-9}	nano-	n
10^{-6}	micro-	μ
10^{-3}	milli-	m
10^{-2}	cent-	c
10^{-1}	deci-	d
10^{1}	deka	da
10^{3}	kilo-	k
10^{6}	mega-	M
10^{9}	giga-	G
10^{12}	tera-	T
10^{15}	peta	P
10^{18}	exa-	E

0.6 Units

0.6.1 Conversion Factors

- All values in a question must have the same units before operations can be performed (ex. you cannot add 30 minutes + 2 hours).
- If values have different units, use conversion factors to put all values in the same units (ex. 60 min/hour).
- You can recognize conversion factors because they have two units of the same quantity (ex. time/time).
- Always check that the appropriate units cancel to arrive at a solution with the correct dimensions when performing calculations.
- Example:

  30 min + 2 hours
  = 30 min + (2 ~~hours~~ x 60 min/~~hour~~)
  = 30 min + 120 min
  = 150 min

Self Quiz 0.6.1

1. How many seconds are there in 2.0 years?
2. A convoy traveling at 45.0 km/h has to make a trip of 100.0 km. How many minutes will the trip take?
3. You are the leader of a small recce patrol that is being launched by rowboat from a ship 12 NM from shore. It is 03h00 and sunrise is at 07h00. You must reach shore 1 hour before sunrise. Based on the sea state the patrol will average 4.0 miles/h. Will your patrol make it?
4. Mach number is a term used to describe the velocity of an aircraft or missile. It is a dimensionless number giving the ratio of the velocity of the object compared to the speed of sound in air (under a standard conditions). The speed of sound is 761.2 mph. An officer states that an anti-tank missile is faster than any tank round fired from a Leopard II C1 tank. The missile's velocity is mach 3.00. Standard tank rounds have velocities of: 2000 m/s (APFSDS), 1425 m/s (HEAT), 960 m/s (MPHE). Is he right?

You will need the following conversions for this quiz:

1 nautical mile = 1 NM = 1.151 miles = 1.852 km
1 mile = 1.609 km

Self Quiz 0.6.1 – Answers

1. How many seconds are there in 2.0 years?

   2.0 yrs x 365.25 days/yr x 24 hr/day x 60 min/hr x 60 sec/hr = 6.3 x 10^7 seconds

2. A convoy traveling at 45.0 km/h has to make a trip of 100.0 km. How many minutes will the trip take?

   Time = distance ÷ speed
   = 100.0 km ÷ 45.0 km/h
   = 2.22 hrs

3. You are the leader of a small recce patrol that is being launched by rowboat from a ship 12 NM from shore. It is 03h00 and sunrise is at 07h00. You must reach shore 1 hour before sunrise. Based on the sea state the patrol will average 4.0 miles/h. Will your patrol make it?

   Distance = 12 NM x (1.151 miles/NM) = 13.812 miles[3]

[3] Keep all the digits in this number, even though there are too many significant figures, because it is an intermediate value in the question.

Time = distance ÷ speed
= 13.812 miles ÷ 4.0 miles/h
= 3.4 hrs[4]

Your patrol must complete the journey in 3 hours to arrive on time so no, you will not make it.

4. Mach number is a term used to describe the velocity of an aircraft or missile. It is a dimensionless number giving the ratio of the velocity of the object compared to the speed of sound in air (under a standard conditions). The speed of sound is 761.2 mph. An officer states that an anti-tank missile is faster than any tank round fired from a Leopard II C1 tank. The missile's velocity is mach 3.00. Standard tank rounds have velocities of: 2000 m/s (APFSDS), 1425 m/s (HEAT), 960 m/s (MPHE). Is he right

   Missile velocity = 3.00 x 761.2 mph = 2283.6 miles/hr

   Convert hours to seconds: 2283.6 miles/hr ÷ 60 min/hr ÷ 60 sec/min = 0.63433 miles/sec

   Convert miles to metres: 0.63433 miles/sec x 1.609 km/mile x 1000 m/km = 1020.6 m/s

   Therefore, the APFSDS and HEAT tank rounds are faster than the anti-tank missile but the missile is faster than the MPHE rounds

0.6.2 Units of Temperature

0.6.2.1 Kelvin

- The SI unit for temperature is Kelvin.
- Most countries use SI units for everything except temperature.
- Zero Kelvin (0 K) is called absolute zero. There is a complete absence of energy; it is not possible for a colder temperature to exist.
- So it is impossible to have a negative temperature in Kelvin.
- The common unit for temperature is Celsius.
- The conversion between Celsius and Kelvin is: °C + 273.15 = K
  K – 273.15 = °C
- Examples:
  - Water freezes at 0 °C + 273.15 = 273 K
  - Water boils at 373.15 K – 273.15 = 100 °C

[4] Use correct significant figures because this is the final answer.

0.6.2.2 Fahrenheit

- The USA unit for temperature is Fahrenheit. They are the only country in the world that uses this temperature scale.
- The conversion between Fahrenheit and Celsius is awkward:
  °C = [°F – (32 °F)] ÷ (1.8 °F/°C)
  °F = [°C x (1.8 °F/°C)] + (32 °F)
- Examples:
  - Water freezes at [0 °C x (1.8 °F/°C)] + (32 °F) = 32 °F
  - Water boils at [212 °F – (32 °F)] ÷ (1.8 °F/°C) = 100 °C

0.7 Manipulating Equations

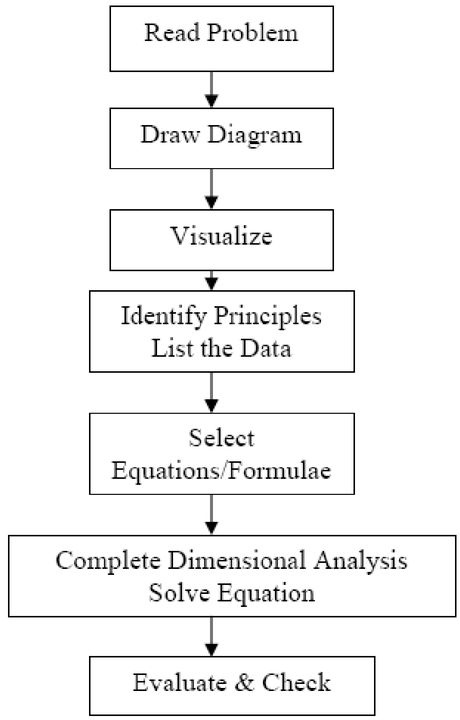
- Mathematical equations are usually written in a format that shows how to solve for one variable.
  - Example: $a = b + c$ shows how to solve for a
- These equations can be manipulated to solve for other variables.
  - Example: If you know a and c, you can manipulate the above equation to solve for b: $b = a – c$
- In order to manipulate an equation, you must do the same operation(s) on both sides of the equals sign. Perform whatever operation(s) are required to isolate the variable of interest.

0.7.1 Examples

- Solve $a = b + c$ for c
  - Subtract b from both sides: $a – b = b + c – b$
  - Simplify: $a – b = c$
- Solve $d \times e = f$ for d
  - Divide both sides by e: $\frac{d \times e}{e} = \frac{f}{e}$
  - Simplify: $d = \frac{f}{e}$
- Solve $g = \frac{h}{i}$ for i
  - First, take i out of denominator so multiply both sides by i: $g \times i = \frac{h}{i} \times i$
  - Simplify: $g \times i = h$
  - Second, isolate i by dividing both sides by g: $\frac{g \times i}{g} = \frac{h}{g}$
  - Simplify: $i = \frac{h}{g}$

0.8 Problem Solving Strategies

- There are many ways to solve problems. The approach below is designed for math problems relating to science and technology.

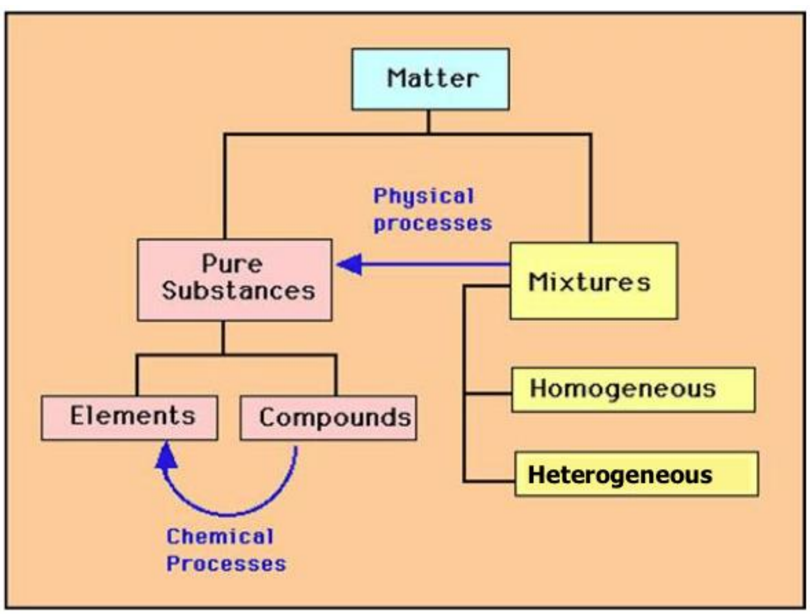


SECTION I – MATTER AND ATOMS

CHAPTER 1 – MATTER AND TRANSFORMATIONS

1.1 Definitions

- Chemistry: The study of matter (its composition and characteristics) and the transformations that it undergoes.
- Matter: Anything that occupies space and has mass (ex. people, books, oxygen gas in air). Excludes energy (ex. heat, light, sound, electricity). Matter is divided into two categories:
    - Pure substances: Contain a single type of matter (ex. solid table salt, water). Pure substances can be further classified:
        - Elemental substances: Made entirely of one element. (ex. Pure (24 carat) gold is composed entirely of gold (Au) atoms).
        - Compounds: Pure substances made up of more than one element (ex. Water contains hydrogen and oxygen).
    - Mixtures: Contain two or more pure substances. Mixtures can be further classified:
        - Homogeneous mixture (also called solution): A mixture with a uniform composition (ie. the separate components of the mixture are not visible) (ex. salt water).
        - Heterogeneous mixture: A mixture in which the components are physically distinct (ex. salad dressing).
- Atom: The smallest component of an element that has the chemical properties of the element. Atoms are composed of sub-atomic particles (described in chapter 3).



Self Quiz 1.1

1. Classify the following examples of matter as pure substances, heterogeneous mixtures or homogeneous mixtures.
    a. piece of tree bark
    b. iron nail
    c. rusty iron nail
    d. well-stirred mixture of food dye in water
    e. bees wax and candle wax mixed together by hand
    f. bees wax and candle wax melted together, stirred well, then allowed to solidify
2. Which of the following are elemental substances and which are compounds?
    a. fluorine (F_2)
    b. bromine chloride ($BrCl_3$)
    c. phosphorus (P_4)
    d. acetylene (C_2H_2)
    e. hydrogen chloride (HCl)
    f. argon (Ar)
    g. aluminum oxide (Al_2O_3)
    h. aluminum (Al)

Self Quiz 1.1 – Answers

1. Classify the following examples of matter as pure substances, heterogeneous mixtures or homogeneous mixtures.
    a. piece of tree bark = ***heterogeneous mixture***
    b. iron nail = ***pure substance***
    c. rusty iron nail = ***heterogeneous mixture***
    d. well-stirred mixture of food dye in water = ***homogeneous mixture***
    e. bees wax and candle wax mixed together by hand = ***heterogeneous mixture***
    f. bees wax and candle wax melted together, stirred well, then allowed to solidify = ***homogeneous mixture***
2. Which of the following are elemental substances and which are compounds?
    a. fluorine (F_2) = ***elemental substance***
    b. bromine chloride ($BrCl_3$) = ***compound***
    c. phosphorus (P_4) = ***elemental substance***
    d. acetylene (C_2H_2) = ***compound***
    e. hydrogen chloride (HCl) = ***compound***
    f. argon (Ar) = ***elemental substance***
    g. aluminum oxide (Al_2O_3) = ***compound***
    h. aluminum (Al) = ***elemental substance***

1.2 Elements

- Matter is made up of elements.
- There are 90 natural elements and up to 28 synthetic (manmade) elements.
    - Some of the synthetic elements have not been fully verified and so there is disagreement as to whether or not they truly exist.
- Every element has a name (ex. gold) and a symbol (ex. Au). The symbols are often derived from the element's Latin name.
- Elements are organized in the Periodic Table of the Elements.
    - The Periodic Table is regulated by the IUPAC (International Union of Pure and Applied Chemistry).
    - The Periodic Table contains a large amount of information. It is discussed in more detail later in chapter 4.

1.3 Representing Elements

1.3.1 Chemical Formulae

- The components of a compound are presented in a chemical formula, which has the form A_nB_m.
- A and B are element symbols (ex. Au, K).
- *n* and *m* show the number of atoms of each element present in one group of the compound.
    - Subscripts (*n, m*) are not included if the value is one (by convention).
- ex. Water = H_2O (one group of water contains two hydrogen atoms and one oxygen atom)
- ex. Glucose = $C_6H_{12}O_6$ (one group of glucose contains six carbon atoms, 12 hydrogen atoms and six oxygen atoms).

Self Quiz 1.3.1

1. List the full names of the elements present in formaldehyde (CH_2O). How many of each element is present in one group of formaldehyde?
2. Do the same for methane (CH_4).

Self Quiz 1.3.1 – Answers

1. List the full names of the elements present in formaldehyde (CH_2O). How many of each element is present in one group of formaldehyde?
    - ***Carbon (1 atom)***
    - ***Hydrogen (2 atoms)***
    - ***Oxygen (1 atom)***
2. Do the same for methane (CH_4).
    - ***Carbon (1 atom)***
    - ***Hydrogen (4 atoms)***

1.3.2 Chemical Reactions

- A chemical reaction is presented in the following form: $A + B \rightarrow C + D$
    - A and B are reactants; C and D are products
- The arrow signifies that a reaction is occurring. Note: You cannot replace the arrow with an equal sign.
- Only reactants are present at the start of the reaction; only products are present at the end of the reaction (assuming a complete reaction).
- ex. $Na + Cl_2 \rightarrow NaCl$
    - Na and Cl_2 are reactants; NaCl is a product

1.3.3 State

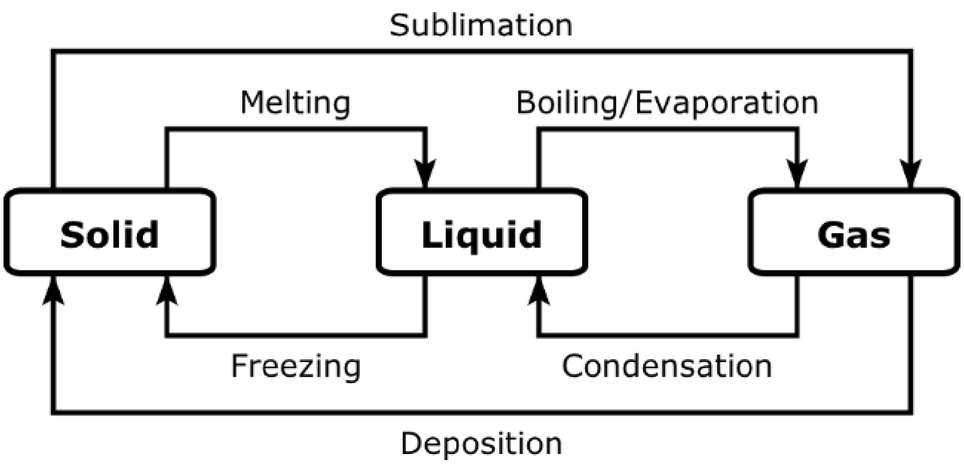
- The state (solid, liquid, gas or aqueous) of a compound is denoted with a bracketed subscript after the chemical formula.
    - Solid = s
    - Liquid = ℓ
    - Gas = g
    - Aqueous = aq (‘aqueous’ means ‘dissolved in water’)[5]
- Example[6] – CO_2:
    - gaseous $CO_2 \rightarrow CO_{2(g)}$
    - solid CO_2 (dry ice) $\rightarrow CO_{2(s)}$
    - aqueous CO_2 (such as occurs in pop) $\rightarrow CO_{2(aq)}$

[5] Aqueous solutions are discussed in detail in chapter 13.
[6] Liquid CO_2 is not included in this example because CO_2 undergoes deposition at atmospheric pressure. Liquid CO_2 can only be produced a pressures above 5 atm.

1.4 Transformations of Matter

1.4.1 Physical Transformations

- In a physical transformation, the chemical composition of the substance is unchanged but its physical state (solid, liquid or gas) is altered.
- Physical transformations are reversible.
- The physical transformations are:
    - Condensation = gas to liquid
    - Evaporation/boiling = liquid to gas
    - Freezing/solidification = liquid to solid
    - Melting = solid to liquid
    - Sublimation = solid to gas
    - Deposition = gas to solid



1.4.2 Chemical Transformations

- In a chemical transformation, a new substance with different chemical properties is formed.
    - ex. $Na + Cl_2 \rightarrow NaCl$
      Na and Cl_2 are both harmful to humans but NaCl is edible table salt.
- Usually irreversible.
- Note: Chemical transformations turn one (or more) compounds into different compounds; the elements within the compounds do not change.
    - ex. You can change a chlorine atom to be bonded to sodium (NaCl) instead of another chlorine atom (Cl_2) but you cannot turn lead into gold.

Self Quiz 1.4.2

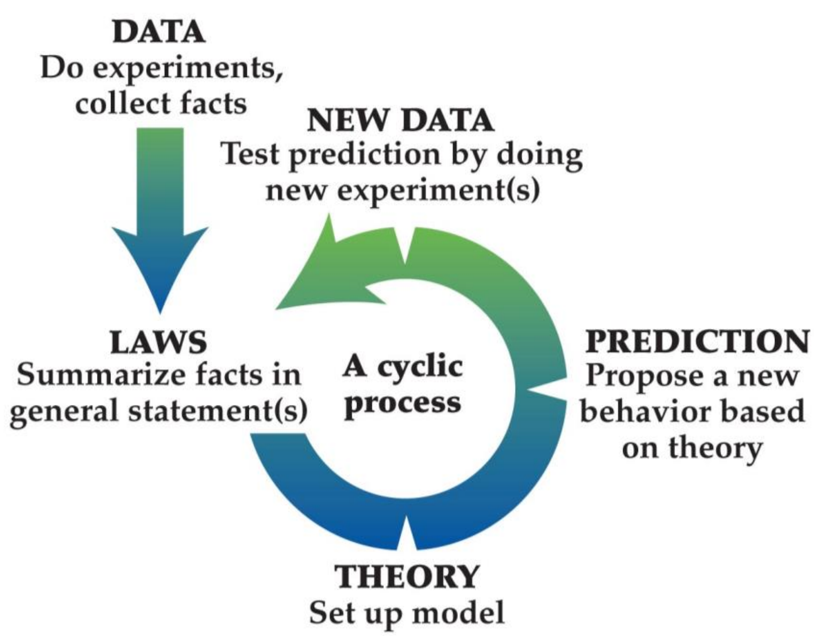
1. What is the name of the process by which matter changes directly from the solid state to the gas state?
2. At atmospheric pressure, the freezing point of ethanol is -117.3 °C and its boiling point is 78.5 °C. What is the melting point of ethanol?
3. The detonation and explosion of a stick of dynamite is a _________ change.

Self Quiz 1.4.2 – Answers

1. What is the name of the process by which matter changes directly from the solid state to the gas state? ***Sublimation***
2. At atmospheric pressure, the freezing point of ethanol is -117.3 °C and its boiling point is 78.5 °C. What is the melting point of ethanol? ***-117.3 °C***
3. The detonation and explosion of a stick of dynamite is a ***chemical*** change.

1.5 The Scientific Method

1. Experiment: Allows study, measure, comparison and collection of data on physical or chemical transformations.
2. Scientific law: Summarizes and predicts the outcome of an experiment.
3. Theory: An attempt to explain why a law exists.



1.5.1 Bias

- Bias is a strong preference or inclination of the researcher that inhibits impartial judgment.
- It occurs when investigators believe so strongly in a particular theory that they interpret data so that it will fit their predicted theory instead of developing a theory based on the data.
- Bias can be a large problem during the scientific method.

CHAPTER 2 – ATOMIC THEORY

2.1 Dalton's Atomic Theory

2.2.1 The Theory

- John Dalton was the first scientist to develop an atomic theory (a way of describing matter) in 1803.
- His theory was based on several main postulates.[7] Today, many flaws are known in these postulates but Dalton's theory is still praised as the first one to identify atoms as the smallest component of matter (that has the chemical properties of an element).
- The postulates are listed below with modern corrections shown in brackets.

2.2.2 Postulates of Dalton's Atomic Theory (with modern corrections)

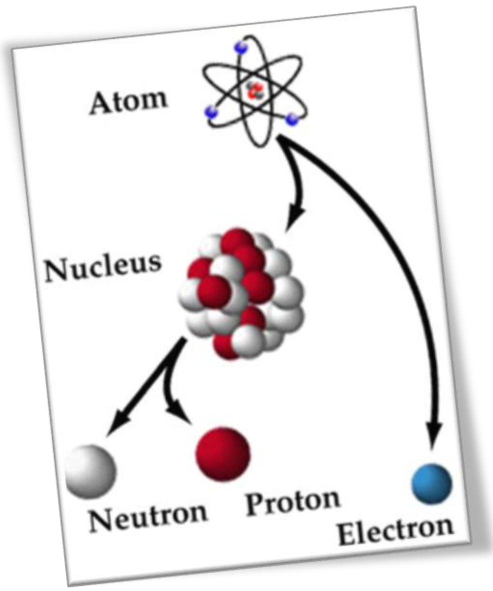
1. Matter is composed of microscopic particles called "atoms" (*true*).
2. Atoms are indivisible (*false = atoms are composed of subatomic particles*).
3. Atoms can neither be created nor destroyed (*false = subatomic particles cannot be created or destroyed; atoms can be manipulated in nuclear reactions (fission and fusion), chapter 11*).
4. Atoms of an element are identical in size, shape, mass and other properties (*false = does not account for isotopes, section 3.4*).
5. Atoms of different elements have different properties (*true*).
6. Atoms of different elements may combine with each other in a fixed, simple, whole number ratios to form compound atoms (*false = large compounds can have very large, complex ratios*).
7. All chemical reactions are due to the combination or separation of atoms (*true*).

2.2 Discovery of Subatomic Particles and the Nucleus

- The problems with Dalton's theory arose because he did not realize that atoms are not the smallest unit of matter.
- We now know that atoms are composed of sub-atomic particles.
- J. J. Thomson and James Chadwick identified the three sub-atomic particles that compose an atom between 1897 and 1932.

[7] The number of postulates varies depending on who is presenting Dalton's theory. The general information from all sources is usually the same.

- The subatomic particles are:
  - Neutron: no charge, relative mass of one (1 amu)[8]
  - Proton: +1 charge, relative mass of 0.9986 (1 amu)
  - Electron: –1 charge, relative mass of 0.0005 (0 amu)
- Ernest Rutherford then made the first step in understanding how subatomic particles interact in an atom by identifying the nucleus.
  - Nucleus: A small, dense positively-charged mass in the centre of the atom that contains the protons and neutrons.



2.3 Bohr's Atomic Theory

2.3.1 The Theory

- Niels Bohr developed a new model of the atom in 1913 to explain the location of electrons.
- Bohr said that electrons orbit the nucleus in fixed paths with specific energy levels called shells or orbits.[9]
- The energy of an electron must be correct for the orbit it is in.
  - ex. For electrons to be in the first orbit around hydrogen, the must have 1.0 eV of energy.
- If electrons move between orbits, they must absorb or emit the exact amount of energy equal to the change.
- Electrons can exist only in these orbits, not between them.

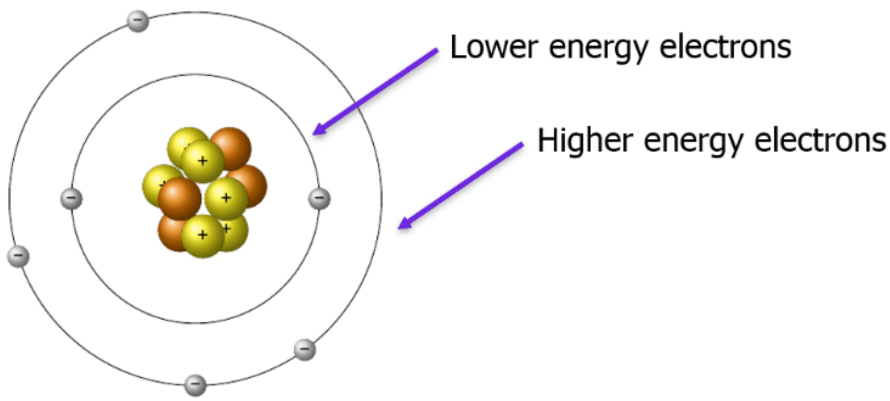
---

[8] Atomic mass units (amu) are described in section 3.1.
[9] The terms 'orbit' and 'shell' will be used interchangeably for the rest of this text.

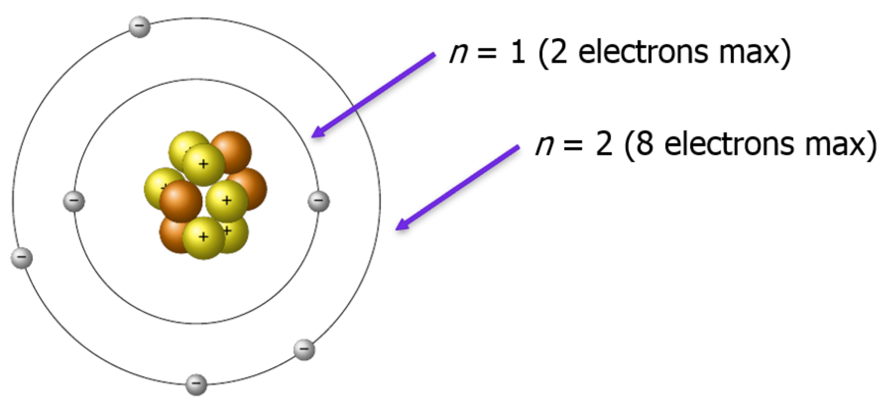
2.3.2 Electron Orbits

2.3.2.1 Definition

- The electron orbits that are closer to the nucleus (smaller orbit radius) host electrons with lower energy.
- So the farther the electron is from the nucleus, the more energy it is carrying.



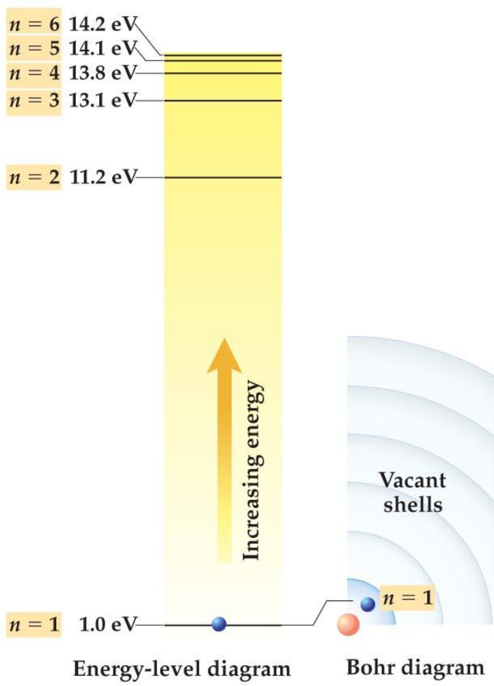
- Each shell is numbered with an *n* value. The shell closest to the nucleus is $n = 1$. The second shell is $n = 2$, etc.
- Each shell has a maximum number of electron it can hold, which is calculated by the equation: max # electrons = $2n^2$.
    - $n = 1$ can hold $2 \times 1^2 = 2 \times 1 = 2$ electrons max,
    - $n = 2$ can hold $2 \times 2^2 = 2 \times 4 = 8$ electrons max,
    - $n = 3$ can hold $2 \times 3^2 = 2 \times 9 = 18$ electrons max…
- The inner shells must be filled before electrons start filling higher shells.



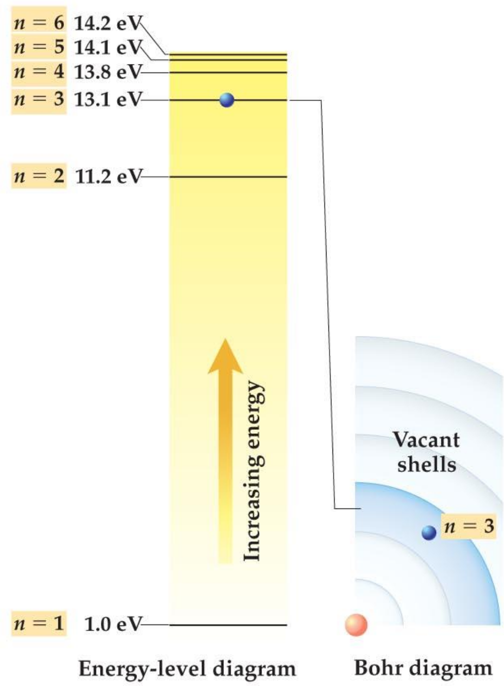
- The electron shells are very far apart so most of an atom is actually blank space.
- If each proton and neutron were 10 centimeters in diameter, the entire atom would have a diameter of 10 kilometers.

2.3.2.2 Energy Level Diagram

- It is difficult to draw the Bohr diagram of an atom showing the energy levels of the orbits.
- Atoms can instead be represented with energy level diagrams.
- This diagram represents the orbits as lines with increasing energy as they move away from the nucleus (the nucleus is assumed to be at the bottom of the plot).
- An example of the energy level diagram for ground state and excited state hydrogen is shown below.[10]



Hydrogen ground state



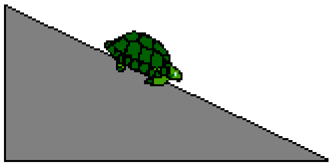
Hydrogen excited state

2.3.3 Quantized Energy

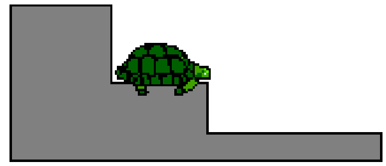
- Bohr's realization that electrons can only exist in specific orbits eventually lead to the development of quantum mechanics.[11]
- The energy of electrons is quantized, meaning that electrons can have only specific energy values.

[10] Ground and excited states are discussed in section 2.7.

[11] Quantum mechanics (QM) is the science of the very small; objects behave differently at the quantum level. QM is beyond the scope of this course but a brief understanding of it is require to understand the behaviour of matter and its interactions with energy on the scale of atoms and subatomic particles.



Non-quantum turtle



Quantum turtle

- This is like saying that you can climb on the rungs of a ladder but you cannot stand between them.

2.4 Orbitals

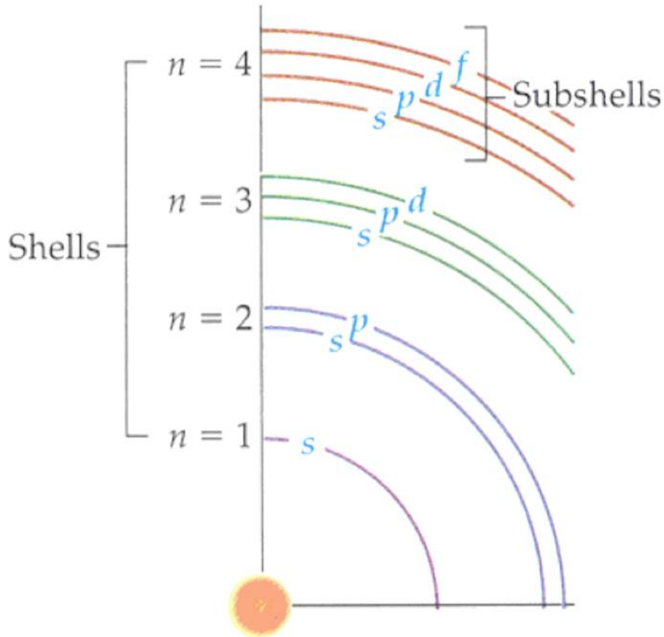
2.4.1 Definition

- Bohr's atomic theory was later modified by Arnold Sommerfeld, who discovered orbitals (subshells).[12]
- Each electron orbit is composed of orbitals, which are labelled with letters.
    - The first orbital type is *s*, the second is *p*, the third is *d*, the fourth is *f*.
    - After that, the letters continue alphabetically: *h, l, j, k…*[13]
- The number of subshell types corresponds to the shell number.
- The n = 1 shell has one subshell type (*s*), the n = 2 shell has two (*s, p*), the n = 3 shell has three (*s, p, d*), the n = 4 shell has four (*s, p, d, f*).

---

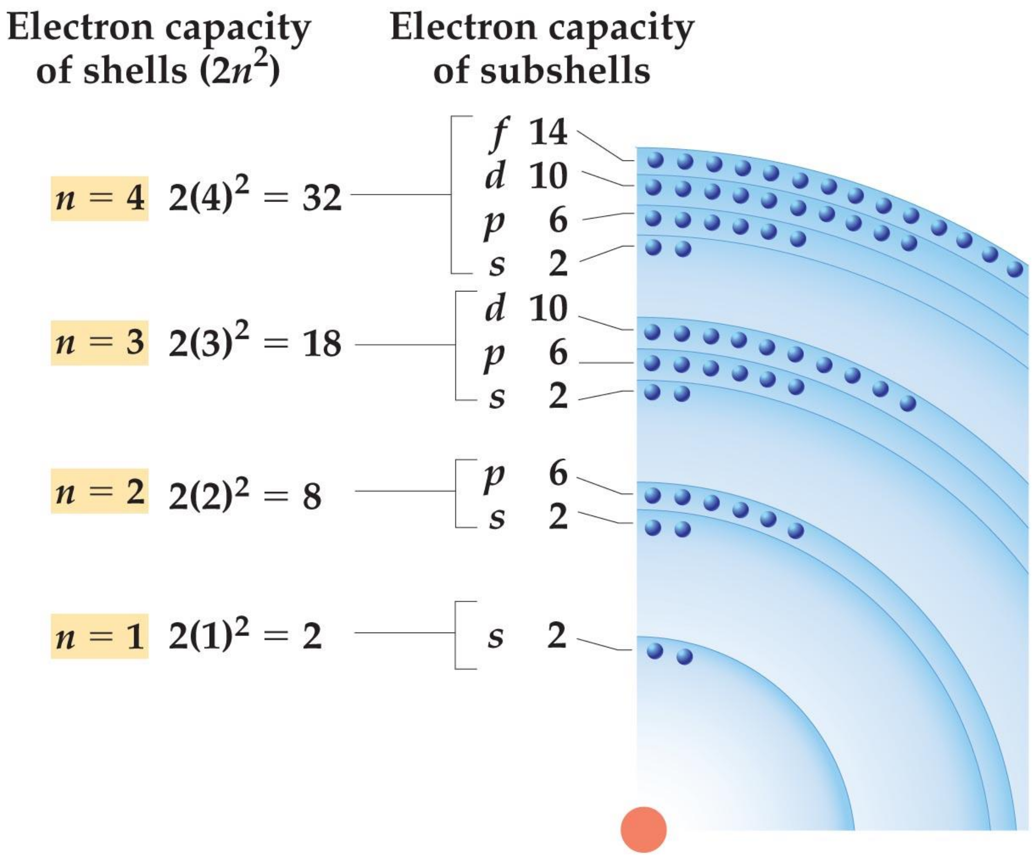
[12] Like orbit/shell, the terms 'orbital' and 'subshell' will be used interchangeably for the rest of this text.

[13] Subshells larger than *f* are not occupied in ground state atoms. Electrons can be excited to these shells but that is not discussed in this course. Ground and excited states are explained in section 2.7.

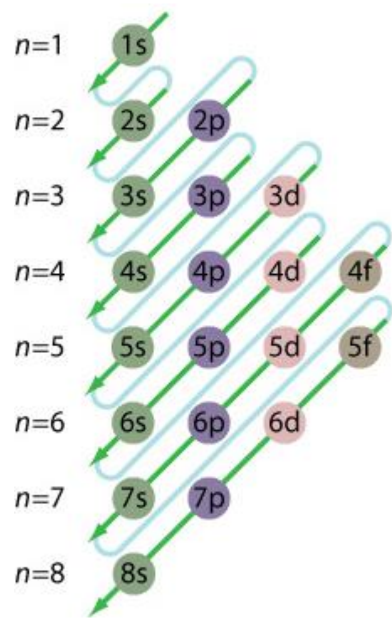


2.4.2 Orbital Capacity

- There are multiple orbitals of some types.
  - There is one *s* orbital, three *p* orbitals, five *d* orbitals, seven *f* orbitals.
- Each subshell can hold 2 electrons.
  - *s* orbital can hold 1 x 2 = 2 electrons
  - *p* orbitals can hold 3 x 2 = 6 electrons
  - *d* orbitals can hold 5 x 2 = 10 electrons
  - *f* orbitals can hold 7 x 2 = 14 electrons
- This fits with the $2n^2$ rule from section 2.3.2.
  - $n = 1$ shell can hold $2(1)^2 = 2$ electrons (it contains one *s* subshell, which holds 2 electrons)
  - $n = 2$ shell can hold $2(2)^2 = 8$ electrons (it contains one *s* subshell and three *p* subshells, which hold 2 electrons each: 2 x (1 + 3) = 8)
  - $n = 3$ shell can hold $2(3)^2 = 18$ electrons (it contains one *s* subshell, three *p* subshells and five *d* subshells, which hold 2 electrons each: 2 x (1 + 3 + 5) = 18)
  - $n = 4$ shell can hold $2(4)^2 = 32$ electrons (it contains one *s* subshell, three *p* subshells, five *d* subshells and seven *f* subshells, which hold 2 electrons each: 2 x (1 + 3 + 5 + 7) = 32)



2.4.3 Orbital Filling Order



- The orbitals don't always fill in order.
- As the orbit number becomes larger and the number of orbitals increases, the orbitals become so close together (in size and energy) that there is cross-over between orbitals from different orbits.
- Logically, orbitals should fill in the following order:
  1s → 2s → 2p → 3s → 3p → 3d → 4s → 4p → …
- But the true order is:
  1s → 2s → 2p → 3s → 3p → **4s → 3d** → 4p → …
- These crossovers are very important because they change the order that the orbitals are filled, which affects the properties of large elements.
- The Orbital Filling Tree (right) can be used to help determine the order in which the orbitals are filled.

2.5 Electronic Configuration

- Subshells are written by first listing the *n* number of the shell followed by the letter of the subshell.
    - ex. The *s* subshell in the *n* = 2 shell is 2*s*.
- To show the number of electrons in the subshell, add it as a superscript.
    - ex. If there is one electron in the *s* subshell in the *n* = 2 shell, it is written as $2s^1$.
- To show all the electrons in an atom, list the subshells side-by-side. This is called the electronic configuration of an element.
    - ex. The 11 electrons in sodium are written as: $1s^22s^22p^63s^1$

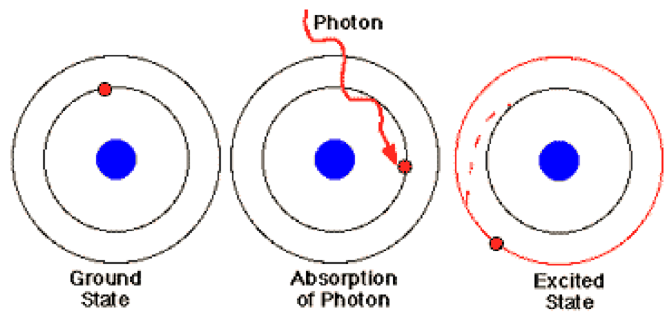
2.6 Schrödinger Atomic Model

- In 1925, Erwin Schrödinger proposed a quantum mechanical model of electron behaviour and the orbital concept.
- His model says that electrons are not solid particles behaving like the planets orbiting around the sun on defined orbit paths.
- Instead, electrons exist simultaneously as waves and particles within their orbital.
- Schrödinger also said different types of orbitals have different shapes.
- These shapes define zones of high probability of finding an electron.
- This is the quantum mechanical model of the atom.
- The general concept of Bohr's theory with quantized shells and subshells is considered to properly explain the chemical behaviour of the elements, but Schrödinger's quantum mechanical model better describes the actual electron behaviour.

2.7 Ground & Excited State Atoms

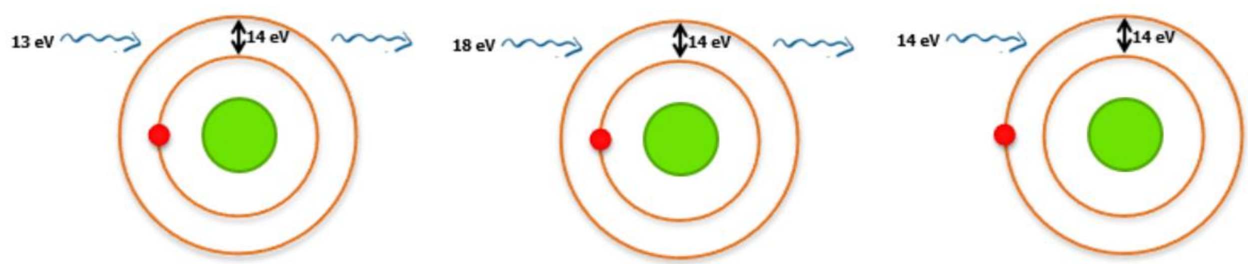
2.7.1 Definition

- Atoms exist in the ground state under normal conditions. This means that all of the electrons are in the lowest possible orbit.
- If an atom is exposed to energy (heat, electricity, light, etc.), the electrons absorb the energy and jump to a higher orbit (if the energy absorbed is equal to the change and there is room in the higher orbit).
    - This energy is usually referred to as a photon (a quantum packet of light or other electromagnetic radiation).
- If this happens, the atom referred to as being in an excited state.



2.7.2 Quantized Energy

- Because electron shells are quantized, an electron must absorb an amount of energy that is exactly equal to the gap between its current shell and a higher shell to be excited.
- If the energy available is more or less than this amount, the electron will not absorb it and the atom will not become excited.



2.7.3 Relaxation

- From an excited state, the electron will naturally relax back to its natural ground state.
- The excited electron can go directly back to its original shell (ex. n = 3 to n = 1) or go step-wise, jumping from shell to shell (ex. n = 3 to n = 2 to n = 1).
- When an electron relaxes back to a lower energy shell, it releases the corresponding energy photon.
- Since the shells have very specific (quantum) energy levels, the relaxation of an electron releases a very specific amount of energy (photon), corresponding to very specific wavelength of light.[14]
- All the colours of light released when excited states of an element relax collectively make up the light spectra of that element.

[14] The connection between energy and wavelength is discussed in section 2.8.

Self Quiz 2.7

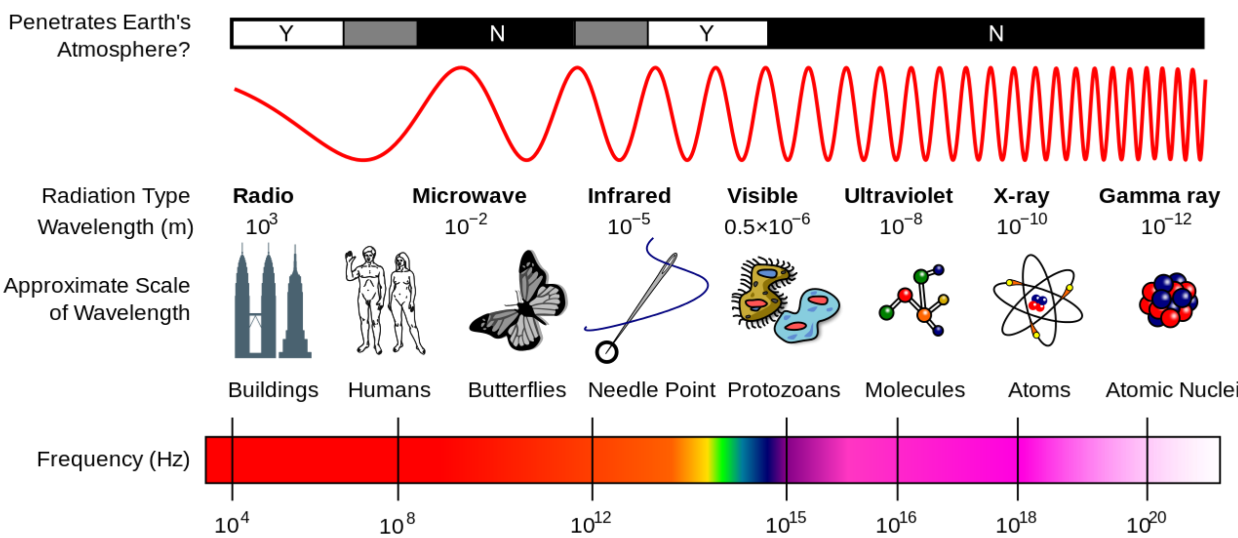
1. Use the energy level diagrams of hydrogen in section 2.3.2.2 to answer the following questions:
   a. How can you tell that the right diagram is of excited state hydrogen?
   b. If an electron is in the *n* = 1 shell, what is the electron's energy in eV?
   c. If an electron is in the *n* = 2 shell, what is the electron's energy in eV?
   d. How much energy would it take to transfer an electron from the *n* = 1 shell to *n* = 2 shell?
   e. Would it take the same, more or less energy to take an electron from the *n* = 2 shell to the *n* = 3 shell?
   f. If the atom is exposed to photons with an energy of 10.2 eV, will anything happen? Why or why not?

Self Quiz 2.7 – Answers

1. Use the energy level diagrams of hydrogen in section 2.3.2.2 to answer the following questions:
   a. How can you tell that the right diagram is of excited state hydrogen? ***The electron in this diagram is in the n = 3 orbit. Electrons always go to the orbit of lowest energy. Since both the n = 1 and n = 2 orbits are empty, the electron would not be in n = 3 unless it had been excited.***
   b. If an electron is in the *n* = 1 shell, what is the electron's energy in eV? ***1.0 eV***
   c. If an electron is in the *n* = 2 shell, what is the electron's energy in eV? ***11.2 eV***
   d. How much energy would it take to transfer an electron from the *n* = 1 shell to the *n* = 2 shell? ***10.2 eV***
   e. Would it take the same, more or less energy to take an electron from the *n* = 2 shell to the *n* = 3 shell? ***Less energy is required to move an electron from n = 2 to n = 3 than n = 1 to n = 2 (1.9 eV vs. 10.2 eV).***
   f. If the atom is exposed to photons with an energy of 10.2 eV, will anything happen? Why or why not? ***The electron in the n = 1 orbit (ground state) will absorb the energy and be excited to the n = 2 orbit. This is possible because the photon has an energy exactly equal to the gap between the n = 1 and n = 2 orbits.***

2.8 Electromagnetic Spectrum

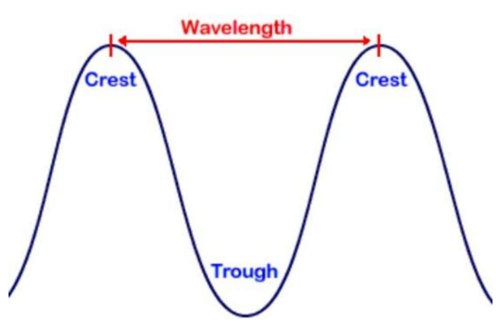
- The electromagnetic spectrum is the the range of wavelengths over which electromagnetic radiation extends.[15]
- Electromagnetic radiation is a type of radiation in which electric and magnetic fields vary simultaneously.
- This includes visible light, radio waves, x-rays, UV radiation and all other radiation familiar to humans.



2.8.1 Properties

2.8.1.1 Wavelength

- A wavelength is the distance between two crests (or two troughs) in a wave.
- Wavelength is inversely proportional to energy (so long waves have low energy and short waves have high energy).



[15] You only need to have a brief understanding of the electromagnetic spectrum for this course.

2.8.1.2 Energy

- The energy content of a light wave is calculated from its wavelength:

$$E = \frac{hc}{\lambda}$$

  - h = Planck's constant = 6.626×10^{-34} J·s
  - c = Speed of light = 3.0×10^{8} m/s
  - λ = wavelength
  - E = energy

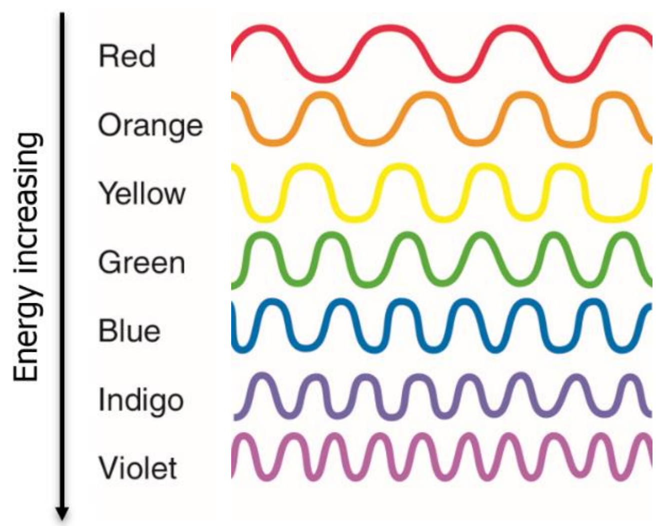
2.8.2 Light

2.8.2.1 Definition

- Light is an electromagnetic wave (a wave that contains energy).
- It is composed of photons (quantized packets of energy) travelling at the speed of light.
  - speed of light, c = 3.0×10^{8} m/s
- Different types of light are characterized by their unique wavelengths, λ.
- The wavelength defines the colour of the light and the amount of energy it carries.

2.8.2.2 Visible Spectrum

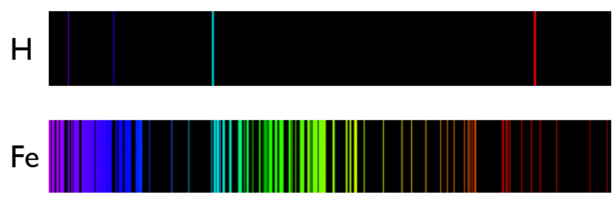
- Our eyes detect wavelengths of 400 – 750 nm as coloured light.
  - Red light (750 nm) has the lowest energy.
  - Violet light (400 nm) has the highest energy.



2.8.3 Emission & Adsorption Spectra

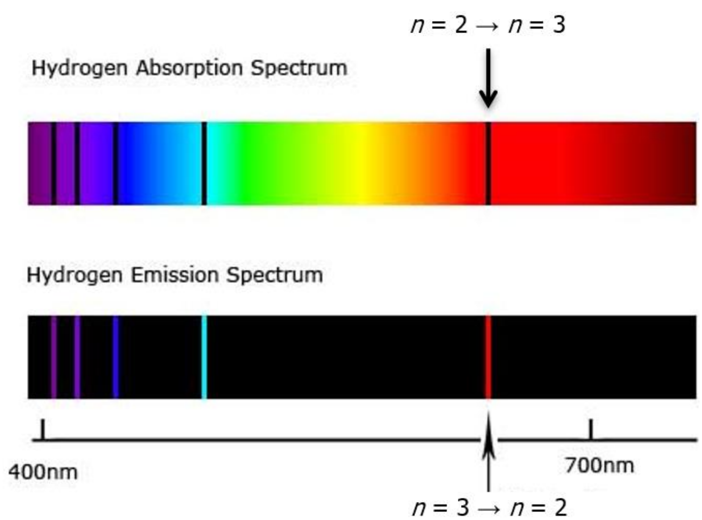
2.8.3.1 Emissions Spectra

- The emission spectrum of an element shows the wavelengths of light that the element emits when excited electrons return to the ground state.
- One element will have several lines on its emission spectrum because several excited to ground state transitions are possible.
  - ex. For hydrogen, the most noticeable transitions are:
    - $n = 5 \rightarrow n = 2$ (indigo)
    - $n = 4 \rightarrow n = 2$ (light blue)
    - $n = 3 \rightarrow n = 2$ (red)
- There are also many transitions that emit wavelengths that are not in the visible spectrum.
- Transition metals, lanthanide and actinides have more complex emission spectra than main group elements because they have more possible transitions due to their partially full *d* and *f* orbitals.



2.8.3.2 Adsorption Spectra

- The adsorption spectrum of an element shows all the wavelengths except those that are absorbed by an atom when its electrons are excited.
- The adsorption spectrum is the reverse of the emission spectrum.



CHAPTER 3 – PROPERTIES OF AN ATOM

3.1 Properties of an Atom

- Atomic number, Z = # of protons in an atom
    - The atomic number defines the identify of an atom.
        - ex. Carbon atoms always have six protons so any atom with six protons is carbon.
    - Atomic number is the only way to identify an element because the number of electrons can change (in ions, section 3.6) and the number of neutrons can change (in isotopes, section 3.4).
- Mass number, A = # of protons + neutrons (the mass of electrons is negligible compared to protons and neutrons so they are not counted)
- Atomic mass unit, amu = unit of mass of an atom (each protons weighs 1 amu, each neutron weighs 1 amu and each electron weighs 0 amu)
    - ex. 1 carbon atoms contains 6 protons, 6 electrons and 6 neutrons so it weighs 6 + 0 + 6 = 12 amu
- 1 amu = 1.66×10^{-24} grams

Self Quiz 3.1

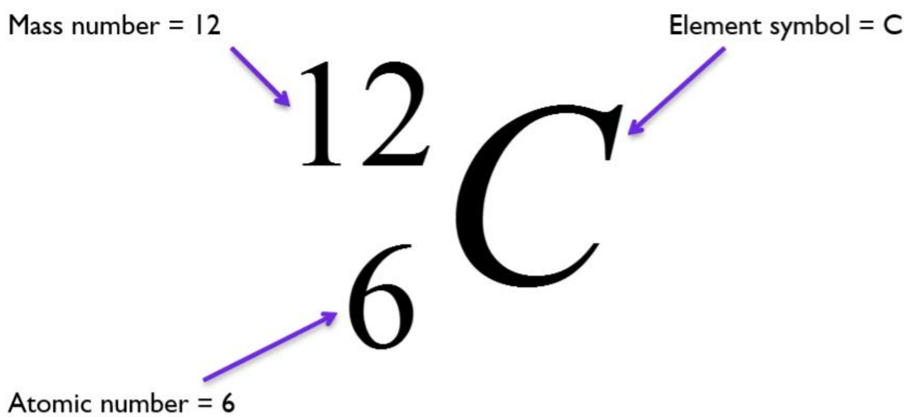
1. Suppose you are told that an atom has a mass number of 16 amu. Is that enough information to identify the element? If yes, what is the element?
2. Suppose you are told that an atom with a mass number of 27 amu contains 14 neutrons. Is that enough information to identify the element? If yes, what is the element?

Self Quiz 3.1 – Answers

1. Suppose you are told that an atom has a mass number of 16 amu. Is that enough information to identify the element? If yes, what is the element? ***No, because multiple elements can have the same mass. It is the atomic number (number of protons) that is unique to each element.***
2. Suppose you are told that an atom with a mass number of 27 amu contains 14 neutrons. Is that enough information to identify the element? If yes, what is the element? ***Yes, because the number of neutrons can be subtracted from the mass number to determine the atomic number (number of protons), which is unique for each element. The atomic number of this element is 27 – 14 = 13, which is aluminum.***

3.2 Representing Elements

- The conventional way to represent an element is to give the elemental symbol (X) with the mass number (A) as a superscript and the atomic number (Z) as a subscript: $^{A}_{Z}X$
- Variations:
  - Sometimes the locations of A and Z are reversed. Remember that the mass is always bigger than (or equal to for hydrogen-1) the atomic number.
  - Sometimes Z is not written. This is fine because writing Z is redundant (ex. all carbon atoms have an atomic number of 6 so you do not need to write the elemental symbol C and the atomic number 6).
  - In a sentence, the superscript/subscript format is uncommon. Instead, elements are usually identified using the full name then a hyphen followed by A (ex. hydrogen-1 is a hydrogen atom with a mass of 1 amu).



Self Quiz 3.2

1. Answer the following questions for this element: $^{131}_{54}Xe$
   a. What is the full name of the element?
   b. How many protons, electrons and neutrons does it contain?
   c. Represent this element using another format.

Self Quiz 3.2 – Answers

1. Answer the following questions for this element: $^{131}_{54}Xe$
   a. What is the full name of the element? ***Xenon***

b. How many protons, electrons and neutrons does it contain?

 # protons = atomic number = 54

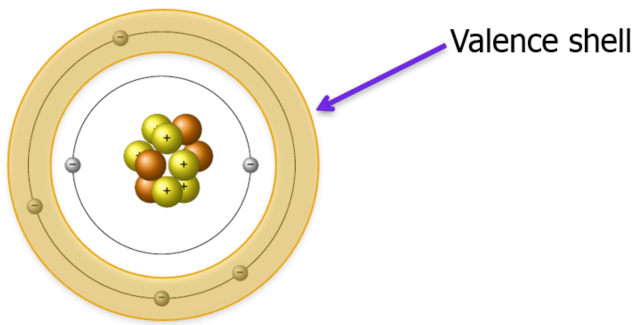
 # neutrons = mass number – atomic number = 131 – 54 = 77

 # electrons = atomic number – charge = 54 – 0 = 54

c. Represent this element using another format. ***Xenon-131 or ^{131}Xe***

3.3 Valence Shell

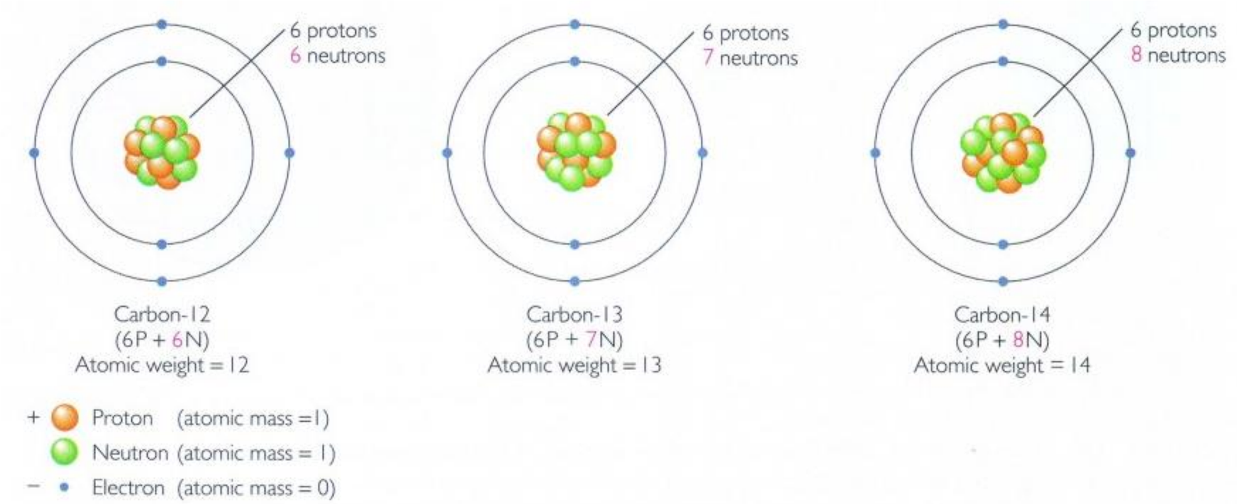
- The outermost shell of an atom that contains electrons in the ground state is called the valence shell.



- Recall from section 2.4 that electron orbits are sub-divided into orbitals that do not fill in sequential order. Valence electrons are still only electrons in orbit with the highest *n* number.
    - ex. When the filling order is considered, the electronic configuration of bromine is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
    - Even though 3*d* is filled after 4*s*, only the electrons in the *n* = 4 shell are valence electrons: $1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{4s^2} 3d^{10} \mathbf{4p^5}$
    - So bromine has seven valence electrons, which it should since it is in column 7 of the Periodic Table.

3.4 Isotopes

- The number of neutrons can vary between different atoms of the same element.
- So different atoms of the same element can have different masses.
- The number of neutrons in an atom is usually similar (but not necessarily equal) to the number of protons.
- Multiple atoms with the same number of protons but different number of neutrons are called isotopes.
    - ex. All carbon atoms have 6 protons. Most carbon atoms have 6 neutrons but some have 7 or 8 neutrons. So there are isotopes of carbon that weigh 12, 13 or 14 amu.



- Usually, only some of the isotopes of an element are stable and one is the most common.
  - ex. Carbon-12 is stable and common, carbon-13 is stable but uncommon, carbon-14 is unstable (radioactive).
- Some elements do not have any stable isotopes (ex. Uranium).

3.5 Average Mass

3.5.1 Definition

- So far, we've presented the atomic mass as a whole number.
- This is true when describing an individual atom (ex. a carbon atom has mass of 12 amu) but because of isotopes, different atoms of the same element can have different masses (ex. two carbon atoms have masses of 12 amu and 13 amu respectively).
- So when describing a group of atoms, the mass is an average of the isotope mass numbers relative to the natural abundance of each isotope.
- This is called the average (or weighted) mass.
- Examples:
  - The average mass of a carbon atom in a group of carbon atoms containing 50% carbon-12 and 50% carbon-13 is (12 amu x 0.5) + (13 amu x 0.5) = 12.5 amu.
  - The average mass of a carbon atom in a group of C atoms containing 75% carbon-12 and 25% carbon-13 is (12 amu x 0.75) + (13 amu x 0.25) = 12.25 amu.
- Remember that the mass of an individual atom must still be a whole number so if I pick up a single carbon atom, it will weigh either 12 amu or 13 amu but if I pick up a handful of carbon atoms, their average mass will be 12.25 amu (in the second example above).

3.5.2 Natural Abundance

3.5.2.1 Earth

- Every isotope has a natural abundance on Earth. Some isotopes are more common than others.
- The average mass of an element accounts for the natural abundance of the isotopes of that element on Earth.
- Example: Carbon
    - Natural abundance of carbon-12 is 98.93%
    - Natural abundance of carbon-13 is 1.07%
    - Natural abundance of carbon-14 is trace
    - Carbon, average mass = (12 amu x 0.9893) + (13 amu x 0.0107) + (14 amu x 0) = 12.01 amu

3.5.2.2 Space

- The natural abundance of isotopes varies throughout space.
- Most notably, the ratio to oxygen-16 to oxygen-17 to oxygen-18 is different between rocks found on Earth, on Mars and on asteroids (and other parts of the universe).
- This is how scientists determine if a rock found on Earth is from an asteroid or if it is a regular Earth rock. If it is an asteroid then the average mass of oxygen will be different than that found on Earth.

Self Quiz 3.4–3.5

1. There are three hydrogen isotopes: hydrogen-1, hydrogen-2 (deuterium), hydrogen-3 (tritium).
    a. The average mass of hydrogen is 1.008 amu. From this, what is the most common hydrogen isotope on Earth?
    b. The CANDU nuclear reactor (a Canadian invention) is praised worldwide because it uses natural uranium (instead of enriched uranium). The CANDU reactor uses heavy water as a neutron moderator. Heavy water (H_2O) is water containing deuterium instead of hydrogen. What is it called heavy water?

Self Quiz 3.4–3.5 – Answers

1. There are three hydrogen isotopes: hydrogen-1, hydrogen-2 (deuterium), hydrogen-3 (tritium).

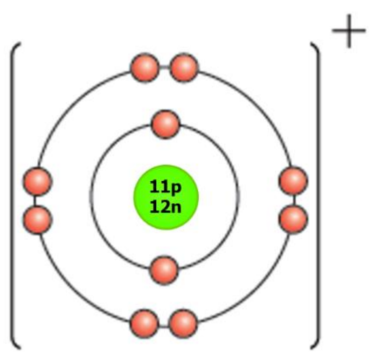
a. The average mass of hydrogen is 1.008 amu. From this, what is the most common hydrogen isotope on Earth? ***Hydrogen-1 is the most common isotope because the average mass of hydrogen is very close to 1 amu.***

b. The CANDU nuclear reactor (a Canadian invention) is praised worldwide because it uses natural uranium (instead of enriched uranium). The CANDU reactor uses heavy water as a neutron moderator. Heavy water (H_2O) is water containing deuterium instead of hydrogen-1. What is it called heavy water? ***Water containing deuterium has a mass of 2 amu while hydrogen-1 has a mass of 1 amu. Therefore, water containing deuterium is heavier than water containing hydrogen-1.***

3.6 Ions

3.6.1 Definition

- An atom is neutral (no charge) if the number or protons equals the number of electrons.
- An atom has a charge (is an ion) if it has a different amount of protons and electrons.
  - Anion = Negatively charged ion (more electrons than protons)
  - Cation = Positively charged ion (fewer electrons than protons)
- Ions are presented by writing the elemental symbol followed by the charge in a superscript.
  - ex. Sodium missing one electron is Na^+ (pictured below)
  - ex. Oxygen with two additional electrons is O^{2-}
- Remember: Atoms cannot gain or lose protons. Only electrons can move.



3.6.2 Calculating Charge

- The charge on an ion is calculated by: charge = # protons – # electrons
- The number of electrons in an ion is calculated by:
  # electrons = atomic number – charge

Self Quiz 3.6

1. Fill in the following table:

	$^{14}N^{3-}$	$^{24}Mg^{2+}$	$^{23}Na^{+}$	$^{56}Fe^{3+}$
Mass Number				
Atomic Number				
# Protons				
# Neutrons				
# Electrons				

2. What do N^{3-}, Mg^{2+} and Na^{+} have in common?

Self Quiz 3.6 – Answers

1. Fill in the following table:

	$^{14}N^{3-}$	$^{24}Mg^{2+}$	$^{23}Na^{+}$	$^{56}Fe^{3+}$
Mass Number	14	24	23	56
Atomic Number	7	12	11	26
# Protons	7	12	11	26
# Neutrons	14 – 7 = 7	24 – 12 = 12	23 – 11 = 12	56 – 26 = 30
# Electrons	7 – (-3) = 10	12 – 2 = 10	11 – 1 = 10	26 – 3 = 23

2. What do N^{3-}, Mg^{2+} and Na^{+} have in common? ***All three ions have 10 electrons total with 8 valence electrons. These ions now have a full valence octet (section 4.7), which means that they are stable (non-reactive).***

3.7 Oxidation State

- Oxidation state (also called oxidation number) represents the number of electrons that an atom can gain, lose, or share in order to produce a full valence shell during bonding with another element.
- It is an integer and can be positive, negative or zero.
- Oxidation states are important for understanding bonding types and knowing how many atoms of each element are involved in a bond.
- But oxidation state is not necessarily the same as formal charge.
- Transition metals can often have multiple stable oxidation states because they have partially filled *d* or *f* orbitals, which can be stable with different orbital fillings (half full orbitals are often stable).
- Example: Iron can have oxidation states of +2 or +3.
    - o The highest energy electrons in non-bonded iron are $4s^2 3d^6$. Iron can be stable with $4s^1 3d^5$ (half filled 4*s* and 3*d* orbitals; oxidation state of +2) or $4s^0 3d^5$ (empty 4*s*, half filled 3*d*; oxidation state of +3).
- The common oxidation states for the first row transitional metals are shown below.

Element Name and Symbol	Atomic Number	Common Oxidation States	Electron Configuration	
Scandium (Sc)	21	+3	Sc: [Ar] $4s^2 3d^1$	Sc: [Ar] $\underset{4s}{\uparrow\downarrow}$ $\underbrace{\uparrow\ \_\ \_\ \_\ \_}_{3d}$
Titanium (Ti)	22	+4	Ti: [Ar] $4s^2 3d^2$	Ti: [Ar] $\underset{4s}{\uparrow\downarrow}$ $\underbrace{\uparrow\ \uparrow\ \_\ \_\ \_}_{3d}$
Vanadium (V)	23	+2, +3, +4, +5	V: [Ar] $4s^2 3d^3$	V: [Ar] $\underset{4s}{\uparrow\downarrow}$ $\underbrace{\uparrow\ \uparrow\ \uparrow\ \_\ \_}_{3d}$
Chromium (Cr)	24	+2, +3, +6	Cr: [Ar] $4s^1 3d^5$	Cr: [Ar] $\underset{4s}{\uparrow}$ $\underbrace{\uparrow\ \uparrow\ \uparrow\ \uparrow\ \uparrow}_{3d}$
Manganese (Mn)	25	+2, +3, +4, +6, +7	Mn: [Ar] $4s^2 3d^5$	Mn: [Ar] $\underset{4s}{\uparrow\downarrow}$ $\underbrace{\uparrow\ \uparrow\ \uparrow\ \uparrow\ \uparrow}_{3d}$
Iron (Fe)	26	+2, +3	Fe: [Ar] $4s^2 3d^6$	Fe: [Ar] $\underset{4s}{\uparrow\downarrow}$ $\underbrace{\uparrow\downarrow\ \uparrow\ \uparrow\ \uparrow\ \uparrow}_{3d}$
Cobalt (Co)	27	+2, +3	Co: [Ar] $4s^2 3d^7$	Co: [Ar] $\underset{4s}{\uparrow\downarrow}$ $\underbrace{\uparrow\downarrow\ \uparrow\downarrow\ \uparrow\ \uparrow\ \uparrow}_{3d}$
Nickel (Ni)	28	+2	Ni: [Ar] $4s^2 3d^8$	Ni: [Ar] $\underset{4s}{\uparrow\downarrow}$ $\underbrace{\uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow\ \uparrow\ \uparrow}_{3d}$
Copper (Cu)	29	+1, +2	Cu: [Ar] $4s^1 3d^{10}$	Cu: [Ar] $\underset{4s}{\uparrow}$ $\underbrace{\uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow}_{3d}$
Zinc (Zn)	30	+2	Zn: [Ar] $4s^2 3d^{10}$	Zn: [Ar] $\underset{4s}{\uparrow\downarrow}$ $\underbrace{\uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow}_{3d}$

3.8 Radioactivity

3.8.1 Definition

- Unstable isotopes are called radioactive.
- In radioactive isotopes, the ratio of protons to neutrons is unstable so the nucleus breaks apart naturally over time.
    - ex. Carbon-14 decays to nitrogen-14 (plus an electron).
    - ex. Uranium-235 decays to thorium-231 (plus an alpha particle).[16]

3.8.2 Half-Life

- Some radioactive isotopes decay faster than others. How quickly a radioactive isotope decays is measured by its half-life.
- Half-life is the amount of time in which 50% of a sample of radioactive material decays.
    - ex. The half-life of carbon-14 is 5730 years. So if you begin with 100 atoms of carbon-14, after 5730 years you will have 50 atoms of carbon-14 (and 50 atoms of nitrogen-14).
    - ex. The half-life of uranium-238 is 4.5 billion years.
- The half-life of carbon-14 is quite small so it is very uncommon to find carbon-14 on Earth.
- But some isotopes have very large half-lives and so are common to find on Earth (ex. uranium-238).

[16] This is the natural decay of uranium-235, which takes hundreds of millions of years. You are used to seeing the decay of U-235 to krypton-92 and barium-141 (plus two neutrons). This decay process occurs in a nuclear reactor after U-235 has been bombarded with neutrons (U-235 becomes U-236, which quickly decomposes to Kr-92 and Ba-141). This does not occur in nature.

CHAPTER 4 – PERIODIC TABLE OF THE ELEMENTS

4.1 Developing the Periodic Table

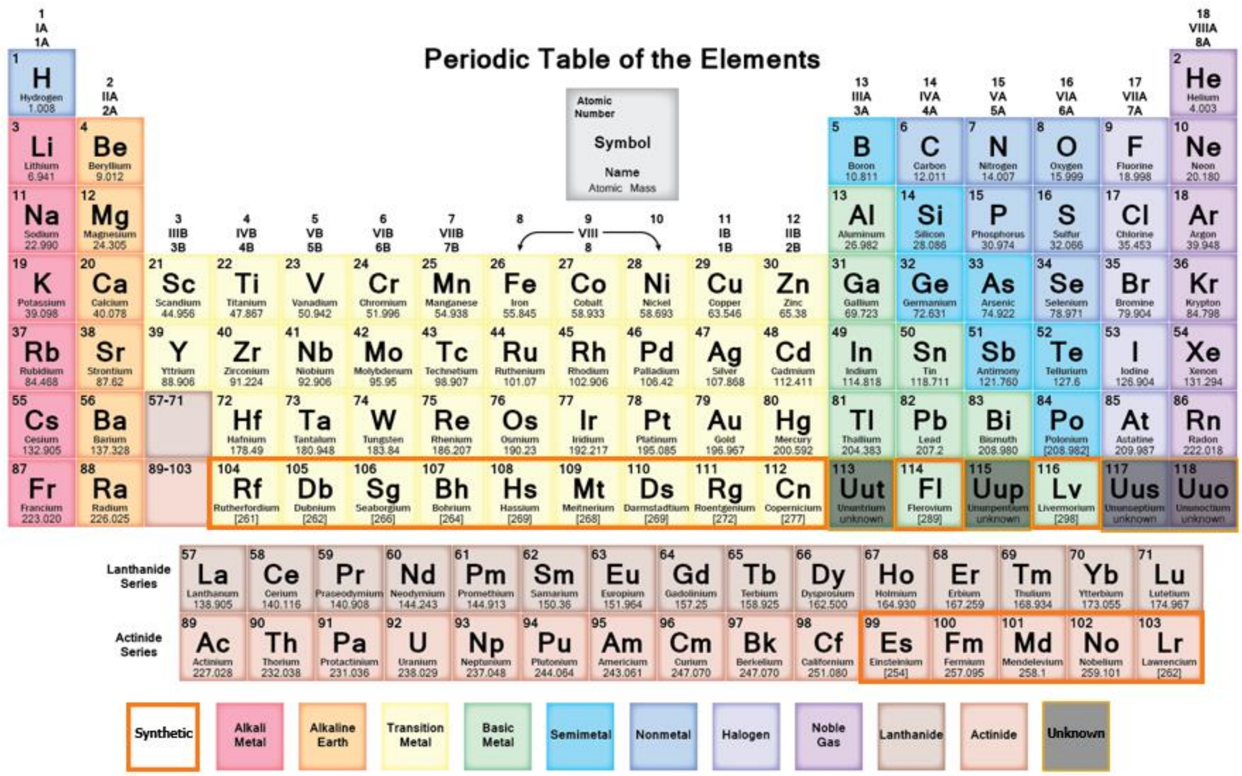
- Dmitri Mendeleev was the first person to organize the elements.
- In 1869, it identified patterns in the chemical and physical properties of elements, called periodicity.
- Mendeleev found that by organizing the elements by increasing mass with breaks every eight elements[17], all the elements in a column had recurring properties.
- Mendeleev's organization of the elements led to the Law of Mendeleev (also called the Periodic Law): "Properties of the elements recur in regular cycles (periodically) when the elements are arranged in order of increasing mass."
- It was later realized that the elements should be organized by increasing atomic number not mass but mass was a close approximation.
- Mendeleev's Periodic Table from 1898 is shown below (for interest).

0	I	II	III	IV	V	VI	VII	
He 4	Li 7	Be 9·1	B 11	C 12	N 14	O 16	F 19	—
Ne 20	Na 23	Mg 24·4	Al 27·1	Si 28·4	P 31·0	S 32·1	Cl 35·5	—
Ar 40	K 39·1	Ca 40	Sc 44·1	Ti 48·1	V 51·2	Cr 52.1	Mn 55·0	Fe 56, Ni 58·7, Co 59
Kr > 45	Cu 63·6	Zn 65·4	Ga 70	Ge 72	As 75	Se 79·1	Br 80·0	—
X > 65	Rb 85·4	Sr 87·6	Y 89	Zr 90·6	Nb 94	Mo 96·0	—	Ru 102, Rh 103, Pd 106
—	Ag 107·9	Cd 112	In 114	Sn 118·5	Sb 120	Te 127	J 126·9	—
—	Cs 133	Ba 137·4	La 138 etc.	—	—	—	—	—
—	—	—	—	—	—	—	—	—
—	—	—	Yb 173	—	Ta 183	W 184	—	Os 191, Ir 193, Pt 195
—	Au 197·2	Hg 200·3	Tl 204·1	Pb 206·9	Bi 208	—	—	—
—	—	—	—	Th 232	—	U 240	—	—

[17] This accounts for the main group elements (the transition metals, lanthanides and actinides were not discovered until later).

4.2 The Modern Periodic Table

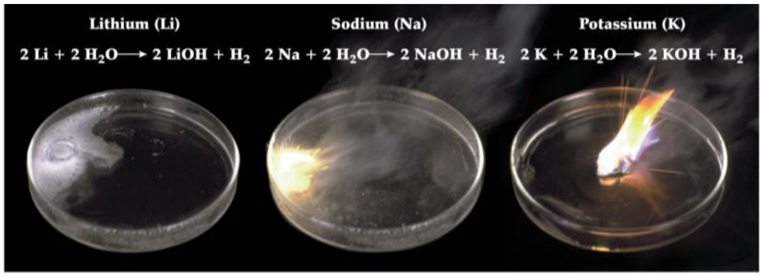
- The modern Periodic Table contains up to 118 elements[18] organized into 32 columns (eight for main group elements, ten for transition metals, 14 for lanthanides/actinides) and seven rows.
- Columns are called groups. Rows are called periods.



4.3 Periodicity

- Elements in the same group have similar chemical and physical properties.
  - ex. None of the elements in group 18 (8A) will participate in chemical reactions.
- Elements become more reactive as you move down a column.
  - ex. All the alkali (group 1) elements react with water to produce heat and energy but lithium is only mildly reactive while francium is very reactive.

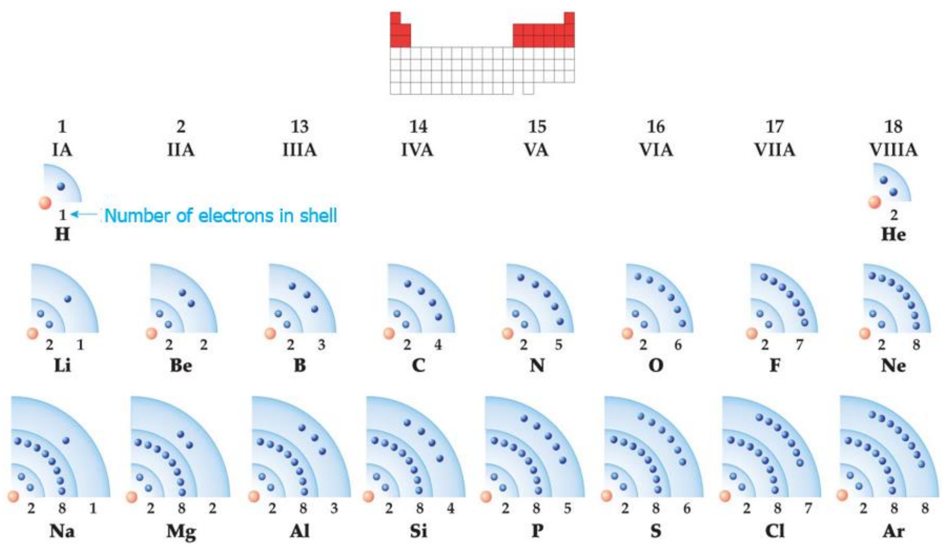
[18] There are "up to" 118 elements because elements with atomic numbers of 99 and higher have not been found on Earth. Several of the larger elements have been synthesized in a lab but very large elements (atomic number of 113+) are very unstable and so are very difficult to synthesize.



4.4 Rows & Columns

4.4.1 Columns: Valence Electrons

- The number of valence electrons in an element is equal to the column number of the element (for main group elements).
- This explains the periodicity observed in the Periodic Table because number of valence electrons determines the properties/reactivity of an element.
    - ex. All elements in column 1 have one electron in their valence shell so they all have similar properties.



4.4.2 Rows: Number of Electrons Shells

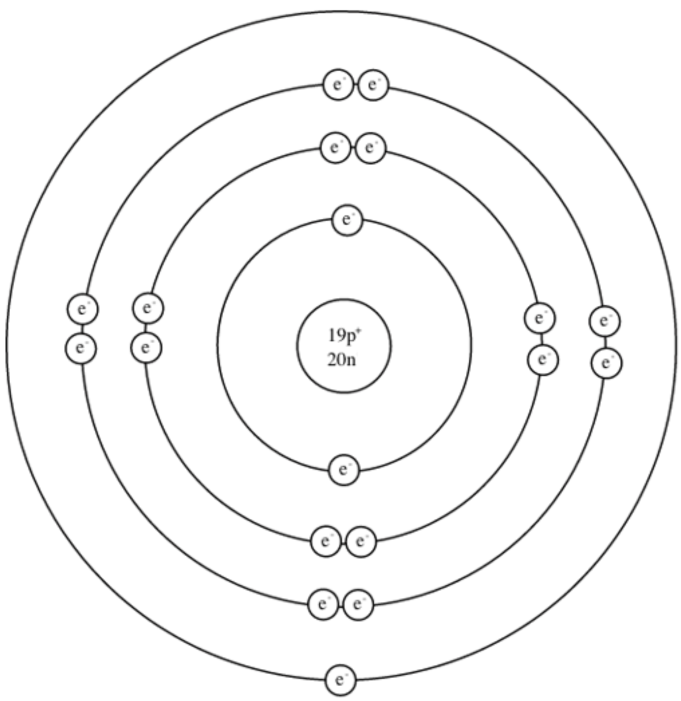
- The number of shells containing electrons can be determined from an element's row number.
    - Elements in the first row (H, He) only have electrons in the first shell.
    - Elements in the second row (Li, Be, B, C, N, O, F, Ne) have electrons in the first and second shells, etc.

Self Quiz 4.4

1. Find potassium (K) on the Periodic Table.
    a. What is its column number?
    b. What is its row number?
    c. How many electrons are potassium's valence shell?
    d. How many of potassium's shells (orbits) contain electrons?
    e. Is potassium more or less reactive than caesium?
    f. Draw the Bohr diagram of potassium.

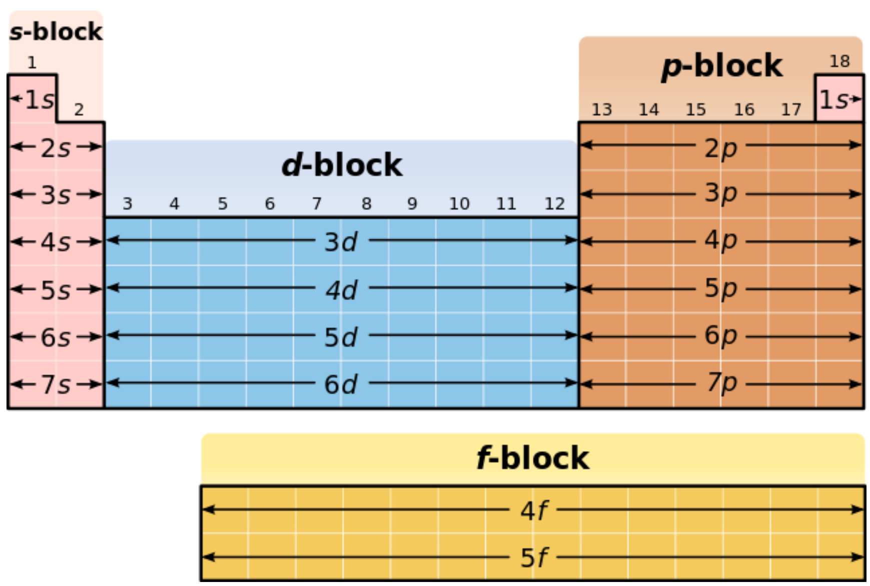
Self Quiz 4.4 – Answers

1. Find potassium (K) on the Periodic Table.
    a. What is its column number? ***one***
    b. What is its row number? ***four***
    c. How many electrons are potassium's valence shell? ***one***
    d. How many of potassium's shells (orbits) contain electrons? ***four***
    e. Is potassium more or less reactive than caesium?
    Potassium is less reactive than caesium.
    f. Draw the Bohr diagram of potassium.



4.5 Periodic Blocks

- The Periodic Table is divided into four blocks, which can be used to determine the electronic configuration of an element.
- Each of the blocks corresponds to one of the subshells (*s, p, d, f*).
- The number of columns in each block corresponds to the number of electrons that can be held in that type of subshell.
    - *s* block (main group) has two columns so *s* orbitals can hold 2 electrons
    - *p* block (main group) has six columns so *p* orbitals can hold 6 electrons
    - *d* block (transition metals) has ten columns so *d* orbitals can hold 10 electrons
    - *f* block (lanthanides/actinides) has fourteen columns so *f* orbitals can hold 14 electrons
- The periodic block shows which orbital is the last filled.
- The row number shows the *n* number of the last filled orbital.
    - ex. Carbon is in row 2 of the *p* block so the 2*p* orbital is the last orbital filled in the carbon. All the orbitals before that (1*s*, 2*s*) are full.



4.6 Determining Electronic Configuration

4.6.1 Complete Configuration

- To determine the electronic configuration of an element, start by looking at hydrogen and then move left to right across the Periodic Table until you reach the end of the block. This is the first row in the *s* block so it corresponds to the 1*s* orbital. If the element you are interested in is after this then the 1*s* orbital must be filled: $1s^2$

- Since there are only two elements in the first row, move to the second row, which exists in the *s* and *p* blocks (2*s* and 2*p*). If the element you are interested in is after this then these orbitals are full so you can add to the electronic structure: $1s^2 2s^2 2p^6$
- Now move to the third row, which exists in the *s* and *p* blocks (3*s* and 3*p*). If the element you are interested in is after this then the electronic structure becomes: $1s^2 2s^2 2p^6 3s^2 3p^6$
- Next is the fourth row, which exists in the *s, p* and *d* blocks (4*s*, 4*p*, 3*d*).
    - NOTE: The *n* number of the *d* orbital is (row number) – 1.
- If the element you are interested in is after this then the electronic structure becomes: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^6 3d^{10}$
    - NOTE: The order of orbital filling using the periodic blocks corresponds to the orbital filling tree from lecture 3, part I.
- After this is the fifth row, which exists in the *s, p* and *d* blocks (5*s*, 5*p*, 4*d*). If the element you are interested in is after this then the electronic structure becomes: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^6 3d^{10} 5s^2 5p^6 4d^{10}$
- Next is the sixth row, which exists in the *s, p, d* and *f* blocks (6*s*, 6*p*, 5*d*, 4*f*).
    - NOTE: The *n* number of the *f* orbital is (row number) – 2.
- If the element you are interested in is after this then the electronic structure becomes: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^6 3d^{10} 5s^2 5p^6 4d^{10} 6s^2 6p^6 5d^{10} 4f^{14}$
- Finally, there is the seventh row, which exists in the *s, p, d* and *f* blocks (7*s*, 7*p*, 6*d*, 5*f*). If the element you are interested in is after this then the electronic structure becomes: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^6 3d^{10} 5s^2 5p^6 4d^{10} 6s^2 6p^6 5d^{10} 4f^{14} 7s^2 7p^6 6d^{10} 5f^{14}$
- The electronic configuration above presents ununoctium (element 118).
- If you are interested in the electronic configuration of an earlier element, follow the same process but stop at the row that the element occupies. All orbitals before the orbital represented by this row are filled. The final orbital contains the number of electrons equal to the column number within the block.
    - ex. Carbon is in row 2 so the 1*s* orbital (row 1) is full. It is in the *p* block so the 2*s* orbital (first block in row 2) is full. It is in column 2 of the *p* block so the 2*p* orbital contains two electrons. So carbon's electronic configuration is $1s^2 2s^2 2p^2$

4.6.2 Short Form Configuration

- It can be very tedious to write out the entire electronic configuration of an element so there is a short form.
- Write the elemental symbol of the noble gas that precedes the element of interest in square brackets. Then include the electronic configuration of only the additional orbitals that are occupied in the element of interest.
    - ex. Titanium: [Ar] $4s^2\ 3d^2 = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$

Self Quiz 4.6

1. Write the electron configurations for the following elements (long form and short form):
   a. Sulphur (S)
   b. Iron (Fe)
   c. Selenium (Se)
2. How many valence electrons does each of these elements have?

Self Quiz 4.6 – Answers

1. Write the electron configurations for the following elements.

   Long form:
   a. Sulphur (S) = $1s^2\ 2s^2\ 2p^6$ **$3s^2\ 3p^4$**
   b. Iron (Fe) = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ **$4s^2$** $3d^6$
   c. Selenium (Se) = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ **$4s^2$** $3d^{10}$ **$4p^4$**

   Short form:
   a. Sulphur (S) = [Ne] **$3s^2\ 3p^4$**
   b. Iron (Fe) = [Ar] **$4s^2$** $3d^6$
   c. Selenium (Se) = [Ar] **$4s^2$** $3d^{10}$ **$4p^4$**

2. How many valence electrons does each of these elements have?
   a. Sulphur = 6
   b. Iron = 2
   c. Selenium = 6

4.7 Octet Rule

- Octet Rule: Elements react to form bonds/compounds in such a way as to put eight electrons in their valence shell.
- This gives elements a valence shell configuration identical to that of a noble gas and makes them very stable.
- The Octet Rule is very important for understanding chemical reactions (chapter 10).

4.8 Nobel Gases and Compounds

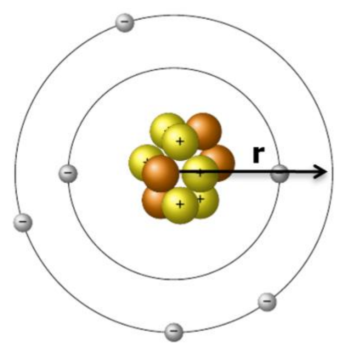
- The elements in column 8 have eight electrons in their valence shell. Since their valence shell contains only *s* and *p* orbitals, it is full with 8 electrons.[19]

[19] Helium has only two electrons in its valence shell because it is so small that it only has electrons in the *n* = 1 orbit, which only has one *s* orbital so it has a full valence shell with two electrons.

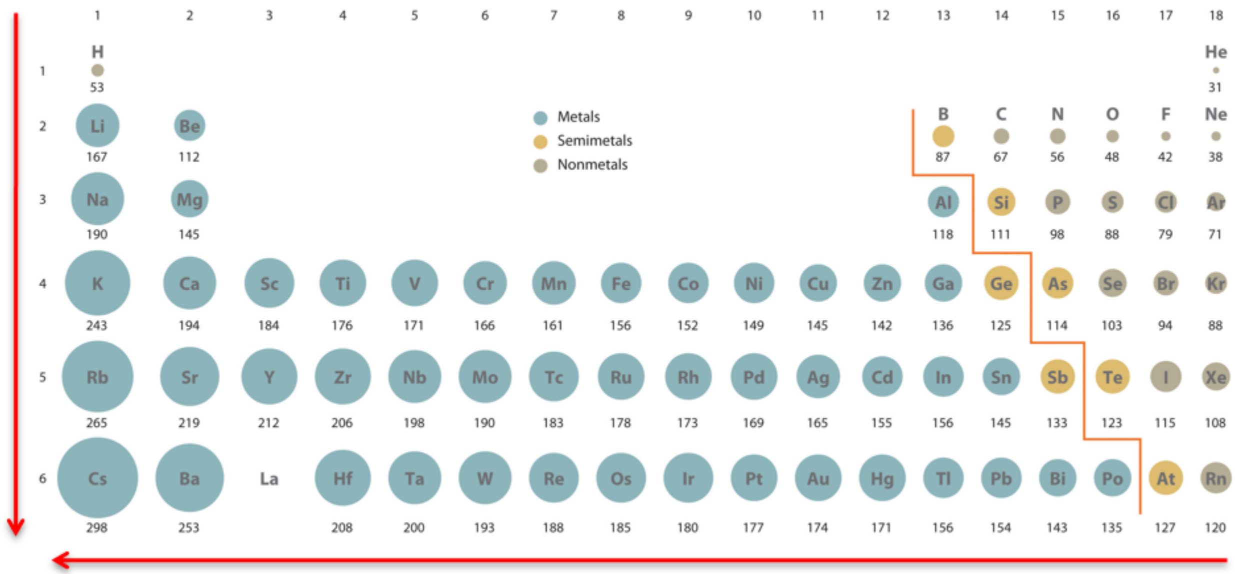
    - Atoms react in order to obtain a full valence shell but since noble gases already have a full valence shell, they have no reason to react.
    - So the noble gases are very stable (unreactive).
- All other elements must bond to achieve a full valence shell. They participate in one of two types of substances.
    - Elemental substance: A substance composed of only one type of element.
        - ex. O_2 (oxygen gas) = two oxygen atoms bound together
        - ex. Ne (neon gas) = a single neon atom
    - Compound: A collection of two or more different types of elements bound together.
        - ex. H_2O (water) = two atoms of hydrogen bound to one oxygen atom
        - ex. CH_4 (methane) = one atom of carbon bound to four hydrogen atoms

4.9 Trends in the Periodic Table

4.9.1 Atomic Radius



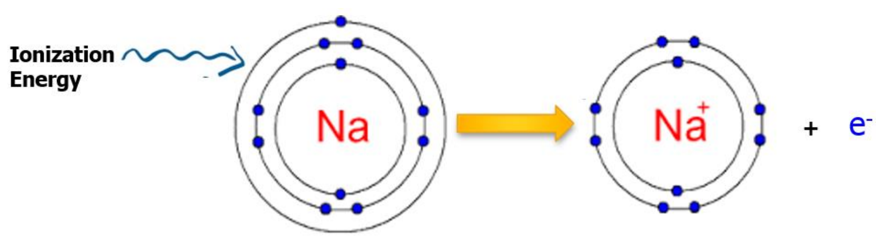
- Atomic radius is the distance from the centre of the nucleus to the boundary of the surrounding electron cloud.
- It is measured in angstroms, Å, where 1 angstrom = 10^{-10} meters.
- Atomic radius increases as you go down a column in the Periodic Table.
    - As you go down a column, the number of occupied orbits increases so the atom becomes larger.
- Atomic radius decreases as you move from left to right across a row.
    - As you move across a row from left to right, electrons are added to the valence shell and protons are added to the nucleus.
    - Since protons are positively charged, each additional proton increases the pulling force of the nucleus on the electrons. This causes the orbits to move closer to the nucleus so the atomic radius decreases.



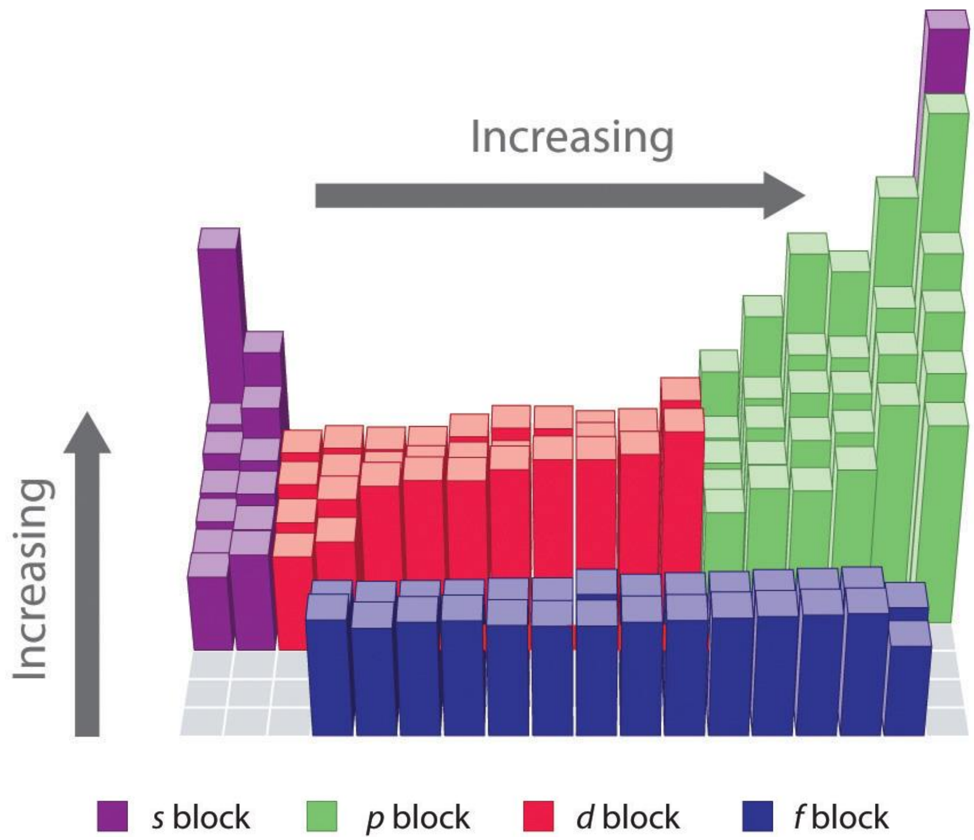
4.9.2 Ionization Energy

4.9.2.1 Trends in Ionization Energy

- Ionization energy is the energy required to remove a valence electron from an atom to form a cation.



- The trend for ionization energy is opposite that of atom radius.
- Ionization energy decreases down a column.
    - As the number of occupied orbits increases, the valence electrons are farther from the nucleus so there is a weaker attractive force between the valence electrons and the nucleus so less energy is required to remove an electron.
- Ionization energy increases left to right across a row.
    - As the atomic radius decreases, there is a stronger attractive force between the valence electrons and the nucleus so more energy is required to remove an electron.
    - Similarly, as you move left to right across a row, the elements become closer to having a full valence shell so elements on the far right really want to gain (not lose) electrons, while elements on the far left are less attached to their electrons.



4.9.2.2 Multiple Ionization Energies

- We usually talk about the first ionization energy (the energy required to remove one electron from a neutral atom).
- The second ionization energy is the energy required to remove a second electron (ie. remove an electron from a +1 cation).
    - The second ionization energy is much larger than the first because there are more protons than electrons in the +1 cation so there is a strong attractive force between the nucleus and the valence electrons.
- The third ionization energy is much larger than the second.

4.9.2.3 Ionization Energy of Noble Gases

- The noble gases have exceptionally high ionization energies because losing an electron will ruin their full octet.
- Similarly, column 1 metals have a very high second ionization energy because they have a full octet when they are +1 cations.
- Column 2 metals have very high third ionization energies because they have a full octet when they are +2 cations.

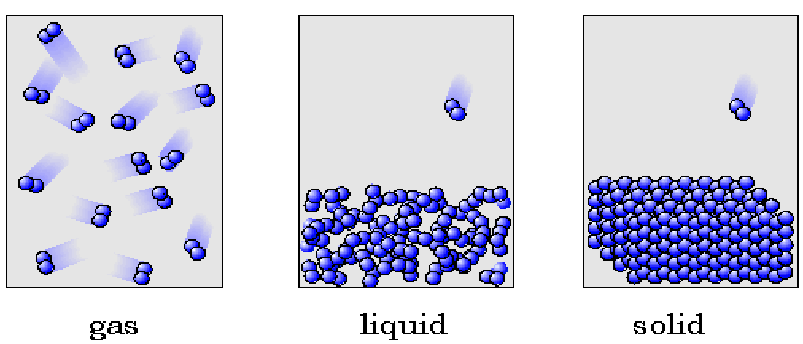
CHAPTER 5 – STATES OF MATTER AND GASES

5.1 Kinetic Energy

- Kinetic energy is the energy that a molecule processes by virtue of being in motion.
- Any molecule at a temperature above zero Kelvin has kinetic energy.
- Temperature is a representation of the amount of kinetic energy in a substance.
- Molecules at a higher temperature move faster and have more kinetic energy.

5.2 States of Matter

- The state of a molecular compound is determined by the amount of kinetic energy it contains relative to the strength of its intermolecular forces.
    - Gas: kinetic energy is stronger than intermolecular forces
    - Liquid: kinetic energy is comparable to intermolecular forces
    - Solid: intermolecular forces are stronger than kinetic energy[20]
- Intermolecular forces as forces that attract molecules to each other. They will be discussed in detail in section 9.3.
    - Intermolecular forces are only present in compounds that contain molecules (distinguishable groups of atoms). All the compounds discussed in this chapter are molecular.
- When a compound has strong intermolecular forces, more kinetic energy is needed to convert it to a less ordered state (solid to liquid, liquid to gas).
- Compounds with stronger intermolecular forces have higher melting/boiling points.
    - ex. water, mp = 0 °C, has stronger intermolecular forces than ammonia, mp = -78 °C



[20] Only gases are discussed in this chapter. Liquids and solids rely on intermolecular forces (section 9.3). Gases can be discussed at this point because their intermolecular forces are negligible.

5.3 Gases

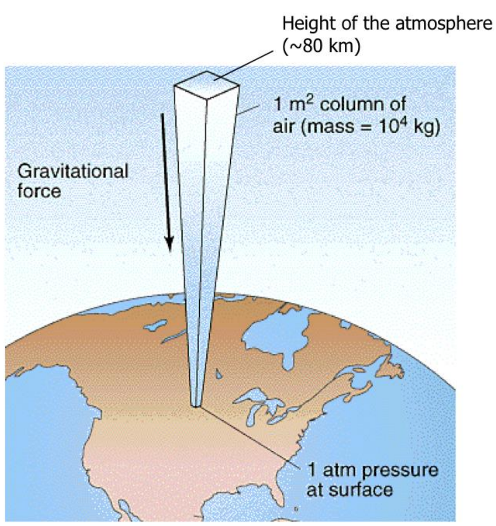
5.3.1 Introduction to Gases

- Over the last few centuries, a lot of technical innovations have involved mastering of properties of gases:
    - Gun powder (firearms, fireworks, rockets)
    - Engines (steam, combustion, aircrafts)
    - Cooling/ventilation (refrigerators, HVAC systems)
    - Energy (windmills)
- Since the kinetic energy in gases is much stronger than the intermolecular forces, we can neglect intermolecular forces when analyzing gases.
- This means that all gases behave very similarly (like an ideal gas).

5.3.2 Properties of Gases

5.3.2.1 Pressure

- Pressure is force per unit area.
- Your body is very sensitive to pressure.
    - This is why your ears pop when you change elevation quickly (your body is trying to equalize the pressure in your body with that in the atmosphere).
    - If you were brought to an environment without atmospheric pressure, such as the vacuum of outer space, your body would bulge and you would die instantly.
- Atmospheric pressure is the force exerted on the surface of the Earth by air molecules colliding with the Earth's surface.
    - The air exerts a lot of pressure on Earth because air molecules are attracted to the Earth's surface by gravity.
    - Normal atmospheric pressure is 1 atm (or 760 Torr, 760 mm Hg).



5.3.2.2 Volume

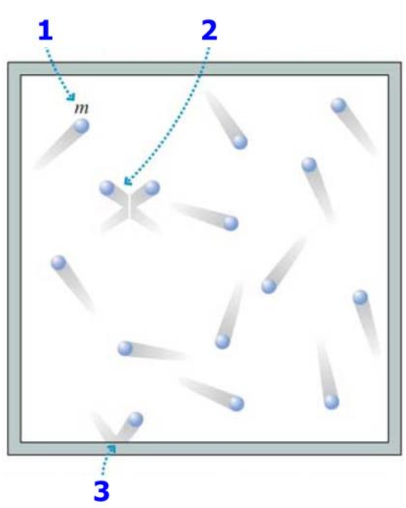
- Because of their high velocity, gases automatically fill any volume they are put in.
    - Gases can be compressed into a small space or expanded into a large space.
- So to measure the volume of gas, simply measure the dimension of its storage container.

5.3.2.3 Temperature

- As with all matter, the temperature of a gas is a reflection of its kinetic energy content.
    - If you heat a gas, you increase its kinetic energy and the speed of the gas molecules increases.
    - If you cool a gas, the kinetic energy is reduced and the speed of the gas molecules decreases.
- If you cool a gas enough, the kinetic energy will have a similar force to that of the intermolecular forces so the molecules will be attracted to each other and the gas will condense to a liquid.

5.3.3 Ideal Gases

- An ideal gas can be described by three main features:
    1. The gas is made of a large number of molecules (*N*), each with mass (*m*), moving randomly.
    2. The molecules in the gas are far from each other and interact only rarely when they collide.
    3. The collisions of the molecules with each other and the walls of the container are elastic, meaning that no energy is lost during these collisions.



5.3.4 Ideal Gas Law

5.3.4.1 The Law

- There are four properties of an ideal gas (their units are shown in brackets)[21]:
  - Pressure of the gas, *P* (atmospheres, atm)
  - Volume of space taken up by the gas, *V* (litres, L)
  - Temperature of the gas, *T* (Kelvin, K)
  - Amount of gas, *n* (moles, mol)[22]
- These properties are connected by the ideal gas law equation: $PV = nRT$ where *R* is a constant, $r = 0.08206 \frac{L \cdot atm}{K \cdot mol}$.
  - If you know or can find/measure three of these properties, you can calculate the fourth.
  - If one of the properties changes, you can recalculate the others.

[21] You must use these units so that the units of *R* cancel. If you measure gas properties in another unit, you will need to use conversion factors (section 0.6.1). Alternatively, you can use a different version of *R* but that is not recommended because it can get very confusing.

Value of *R*	Units ($V\,P\,T^{-1}\,n^{-1}$)
8.3144621	$J\,K^{-1}\,mol^{-1}$
8.31446	$VC\,K^{-1}\,mol^{-1}$
5.189×10^{19}	$eV\,K^{-1}\,mol^{-1}$
0.08205736	$L\,atm\,K^{-1}\,mol^{-1}$
1.9872041	$cal\,K^{-1}\,mol^{-1}$
1.9872041×10^{-3}	$kcal\,K^{-1}\,mol^{-1}$
8.3144621×10^{7}	$erg\,K^{-1}\,mol^{-1}$
8.3144621×10^{-3}	$amu\,(km/s)^2\,K^{-1}$
8.3144621	$L\,kPa\,K^{-1}\,mol^{-1}$
8.3144621×10^{3}	$cm^3\,kPa\,K^{-1}\,mol^{-1}$
8.3144621	$m^3\,Pa\,K^{-1}\,mol^{-1}$
8.3144621	$cm^3\,MPa\,K^{-1}\,mol^{-1}$
8.3144621×10^{-5}	$m^3\,bar\,K^{-1}\,mol^{-1}$
8.205736×10^{-5}	$m^3\,atm\,K^{-1}\,mol^{-1}$
82.05736	$cm^3\,atm\,K^{-1}\,mol^{-1}$
84.78402×10^{-6}	$m^3\,kgf/cm^2\,K^{-1}\,mol^{-1}$
8.3144621×10^{-2}	$L\,bar\,K^{-1}\,mol^{-1}$
62.36367×10^{-3}	$m^3\,mmHg\,K^{-1}\,mol^{-1}$
62.36367	$L\,mmHg\,K^{-1}\,mol^{-1}$
62.36367	$L\,Torr\,K^{-1}\,mol^{-1}$
6.132440	$ft\,lbf\,K^{-1}\,g\text{-}mol^{-1}$
1545.34896	$ft\,lbf\,R^{-1}\,lb\text{-}mol^{-1}$
10.73159	$ft^3\,psi\,R^{-1}\,lb\text{-}mol^{-1}$
0.7302413	$ft^3\,atm\,R^{-1}\,lb\text{-}mol^{-1}$
1.31443	$ft^3\,atm\,K^{-1}\,lb\text{-}mol^{-1}$
998.9701	$ft^3\,mmHg\,K^{-1}\,lb\text{-}mol^{-1}$
1.986	$Btu\,lb\text{-}mol^{-1}\,R^{-1}$

[22] Moles will be described in detail in section 6.1. For now, you only need to know how to calculate number of moles from the Ideal Gas Law.

Self Quiz 5.3.4A

1. Rearrange the ideal gas law (*PV* = *nRT*) to solve for *P, V, T* and *n*. You will need principles from section 0.7 to do this.

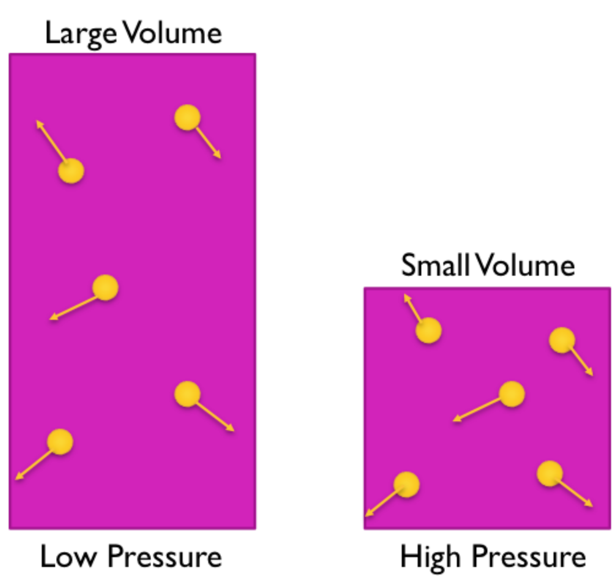
Self Quiz 5.3.4A – Answers

1. Rearrange the ideal gas law (*PV* = *nRT*) to solve for *P, V, T* and *n*. You will need principles from section 0.7 to do this.

$$P = nRT / V$$
$$V = nRT / P$$
$$n = PV / RT$$
$$T = PV / nR$$

5.3.4.2 Effect of Volume

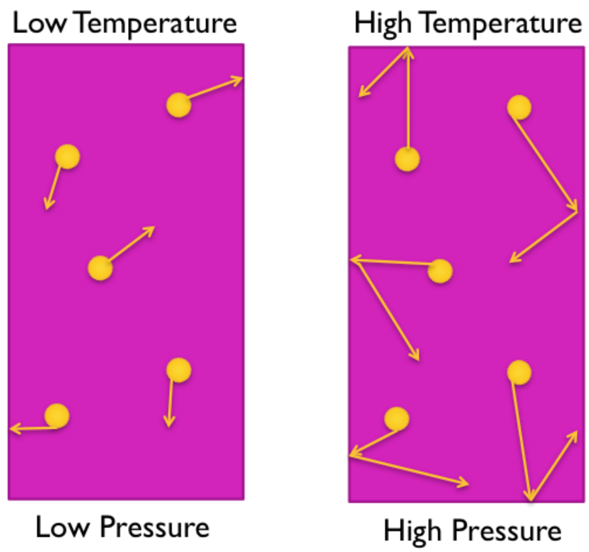
- If you reduce the volume of a container holding a gas, the gas molecules collide with the sides of the container more often than they did in the large container.
- This causes an increase in pressure.
- So pressure is inversely proportional to volume (ie. when the volume of the container decreases, the pressure increases and visa versa).
- $P \propto 1/V$



5.3.4.3 Effect of Temperature

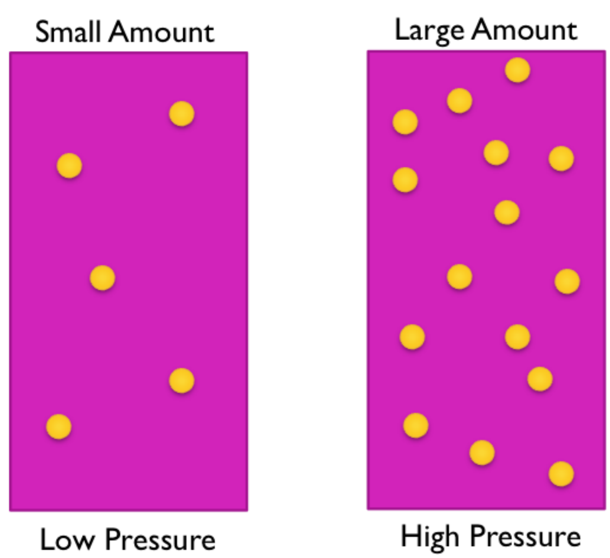
- If you increase the temperature of the gas molecules in a container, the molecules move faster.

- This causes them to collide with the container walls more often, which increases pressure.
- So pressure is proportional to temperature (ie. when temperature increases, pressure increases and visa versa).
- P α T



5.3.4.4 Effect of Amount

- If you increase the amount of gas molecules in a container, the molecules collide with the container walls more frequently.
- This causes an increase in pressure.
- So pressure is proportional to number of molecules of gas (ie. when the amount of gas, pressure increases and visa versa).
- P α n



5.3.4.5 Initial & Final States

- The Ideal Gas Law can also be used to calculate the initial state of a gaseous system from the final state and visa versa.
- This equation is developed because the Ideal Gas Law (for the initial or finial state) can be rearranged to solve for R:

$$R = \frac{P_i V_i}{n_i T_i}, \qquad R = \frac{P_f V_f}{n_f T_f}$$

- *R* is a constant so these equations are equal to each other.
- This produces a useful form of the Ideal Gas Law that can be used when dealing with a gaseous system with different initial and final states:

$$\frac{P_i V_i}{n_i T_i} = \frac{P_f V_f}{n_f T_f}$$

Self Quiz 5.3.4B

1. The gas inside a balloon has a pressure of 745.5 mm Hg, a volume of 250.0 mL and a temperature of 25.5 °C. What is the number of moles of gas in the balloon (1 atm = 760 mm Hg)?
2. What must the temperature be (in °C) if 2.0 moles of a gas in a 4.0 L steel container has a measured pressure of 100 atm?
3. A gas initially at P = 2.00 atm, T = 200.0 K and V = 50.0 mL is in a cylinder that has a moveable piston. The cylinder is heated to T = 220 K and then the piston is pushed down so that the volume of the gas decreases to V = 25.0 mL. Calculate the final pressure in litres.

Self Quiz 5.3.4B – Answers

1. The gas inside a balloon has a pressure of 745.5 mm Hg, a volume of 250.0 mL and a temperature of 25.5 °C. What is the number of moles of gas in the balloon (1 atm = 760 mm Hg)?

   P = 745.5 mm Hg x (1 atm / 760 mm Hg) = 0.981 atm

   V = 250 mL x (1 L / 1000 mL) = 0.250 L

   T = 25.5 °C + 273.15 = 298.65 K

$$n = \frac{PV}{RT} = \frac{(0.981\ atm)(0.250\ L)}{(0.0821\ ^{L\,atm}/_{K\,mol})(298.65\ K)} = 0.0100\ mol$$

2. What must the temperature be (in °C) if 2.0 moles of a gas in a 4.0 L steel container has a measured pressure of 100 atm?

$$T = \frac{PV}{nR} = \frac{(100\ atm)(4.0\ L)}{(2.0\ mol)(0.08206\ \frac{L \cdot atm}{K \cdot mol})} = 2437\ K$$

T = 2437 K – 273.15 = 2164 °C

3. A gas initially at P = 2.00 atm, T = 200.0 K and V = 50.0 mL is in a cylinder that has a moveable piston. The cylinder is heated to T = 220 K and then the piston is pushed down so that the volume of the gas decreases to V = 25.0 mL. Calculate the final pressure in litres.

Initial State	***Final State***
P_i = 2.00 atm	***P_f = ?***
T_i = 200.0 K	***T_f = 220.0 K***
V_i = 0.0500 L	***V_f = 0.0250 L***

$n_i = n_f$

The equation to compare initial and final states of a gas is: $\frac{P_i V_i}{n_i T_i} = \frac{P_f V_f}{n_f T_f}$

But since the number of moles of gas are the same at the start and end of the reaction, the n term will be the same on both sides of the equation. So it can be eliminated by multiplying both sides of the equation by n. Recall that an equation can be manipulated as long as any operations performed to the right side are performed to the left side as well:

$$\cancel{n}\left(\frac{P_i V_i}{\cancel{n}_i T_i}\right) = \cancel{n}\left(\frac{P_f V_f}{\cancel{n}_f T_f}\right) \Rightarrow \frac{P_i V_i}{T_i} = \frac{P_f V_f}{T_f}$$

The equation can now be rearranged to solve for P_f:

$$\frac{T_f}{V_f}\left(\frac{P_i V_i}{T_i}\right) = \frac{\cancel{T_f}}{\cancel{V_f}}\left(\frac{P_f \cancel{V_f}}{\cancel{T_f}}\right) \Rightarrow P_f = \frac{P_i V_i T_f}{V_f T_i}$$

Now we can solve for P_f:

$$P_f = \frac{P_i V_i T_f}{V_f T_i} = \frac{(2.00\ atm)(0.0500\ L)(220.0\ K)}{(200.0\ K)(0.0250\ L)} = 4.40\ atm$$

SECTION II – CHEMICAL COMPOUNDS

CHAPTER 6 – MOLES & MOLAR MASS

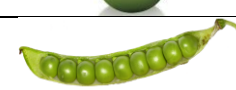
6.1 Moles

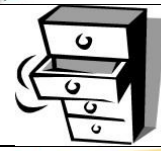
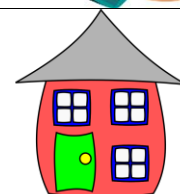
6.1.1 Definition

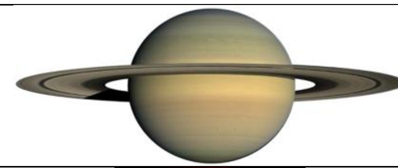
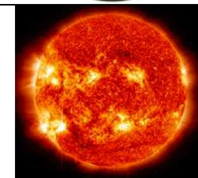
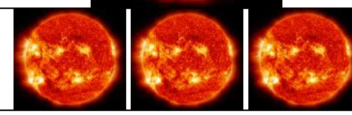
- Atoms and molecules are <u>very</u> small so we need a very large unit to count them.
    - ex. There are 8,000,000,000,000,000,000,000,000 molecules of water in 1 cup.
- The unit is called a mole (symbol: mol) where 1 mole = 6.022 x 10^{23} items
    - 6.022 x 10^{23} is called Avogadro's number (N_A)
- Counting atoms or molecules in units of moles makes the number much more manageable.
    - ex. There are 13 moles of water molecules in 1 cup.
- A mole is just a unit of measure so you can have a mole of different things just like you can have a dozen of different things.
    - ex. There are a dozen humans in a room. There are 20 dozen poppy seeds on the bagels I bought this morning.
    - ex. There is a mole of water molecules in my glass. There are 6 moles of carbon atoms in the glucose in my sugar packet.
- A mole of molecules can contain several moles of atoms (just like a dozen bagels can contain 10 dozen poppy seeds).
    - ex. One mole of water (H_2O) molecules contains two moles of hydrogen atoms and one mole of oxygen atoms.

6.1.2 How Big is a Mole?

- Humans cannot really comprehend a number as large as a mole so to work you up to it, try using the pea analogy.
- The pea analogy calculates the surface area that would be covered by 1 mole of peas.

Exponent	Value	Analogy	Picture
10^0	1	One pea	
10^1	10	10 peas = peapod	
10^2	100	10 peapods = spoonful of peas	
10^3	1000	10 spoonfuls = bowlful of peas	

10^4	10,000	10 bowlfuls = bag of peas	
10^5	100,000	10 bags = drawer filled with peas	
10^6	1,000,000	10 drawers full = bedroom covered with peas	
10^7	10,000,000	10 rooms = house covered with peas	
10^9	1,000,000,000	100 houses = a village covered with peas	
10^{11}	100,000,000,000	100 villages = a small city covered with peas	
10^{13}	10,000,000,000,000	100 small cities = GTA covered with peas	
10^{15}	1,000,000,000,000,000	100 metropolitan cities = New York State covered with peas	
10^{17}	100,000,000,000,000,000	100 small states = Canada covered with peas	
10^{19}	10,000,000,000,000,000,000	100 large countries = the surface of the Earth covered with peas	

10^{21}	1,000,000,000,000,000,000,000	100 small planets = Saturn covered with peas	
10^{23}	100,000,000,000,000,000,000,000	100 large planets = ½ the surface of the Sun covered with peas	
6×10^{23}	600,000,000,000,000,000,000,000	3 suns covered with peas[23]	

Self Quiz 6.1

1. How many moles of nitrogen atoms are in 1 mole of N_2 gas?
2. How many N_2 molecules are in 1 mole of nitrogen gas?
3. How many moles are in 1.8×10^{24} molecules of glucose ($C_6H_{12}O_6$)?
4. How many moles of carbon atoms are in 1.8×10^{24} molecules of glucose ($C_6H_{12}O_6$)?

Self Quiz 6.1 – Answers

1. How many moles of nitrogen atoms are in 1 mole of N_2 gas?

   # atoms = # molecules x (atoms/molecule)
   = 1 mole x (2 N / N_2)
   = 2 moles

2. How many N_2 molecules are in 1 mole of nitrogen gas?

   # molecules = n x N_A
   = 1 mol x (6×10^{23} molecules/mol)
   = 6×10^{23} molecules N_2

3. How many moles are in 1.8×10^{24} molecules of glucose ($C_6H_{12}O_6$)?

   n = (# molecules) / N_A
   = (1.8×10^{24} molecules) / (6×10^{23} molecules/mol)

[23] This can also be calculated algebraically (but it is well beyond the scope of this course). If you are interested, we can treat the mole of peas as a flat plane of hexagonal packed circles. This means that the πr^2 area of the peas will have to be divided by 0.9069 to give the area covered. For peas of diameter of 5 mm, this represents 1.3038×10^{13} square kilometers, which is 2.14x the surface area of the sun.

= 3 mol glucose

4. How many moles of carbon atoms are in 1.8×10^{24} molecules of glucose ($C_6H_{12}O_6$)?

   # atoms = # molecules x (atoms/molecule)
   = 3 mole x (6 C / $C_6H_{12}O_6$)
   = 18 moles

6.2 Molar Mass

6.2.1 Definition

- Molar mass, M, is the mass of 1 mole of a compound.
- It has units of grams/mol (g/mol).
- The molar mass of an element is found on the Periodic Table.
  - ex. The molar mass of hydrogen is 1.008 g/mol so 1 mol of hydrogen atoms weighs 1.008 grams or 6×10^{23} atoms of hydrogen weigh 1.008 grams.
- The mass of a mole of atoms (or molecules) has units of g/mol. The mass of a single atom or (or molecule) has units of amu.
  - ex. One hydrogen-1 atom has a mass of 1 amu. The mass of 1 mole of hydrogen atoms is larger than 1 gram because it considers the natural abundance of all hydrogen isotopes.

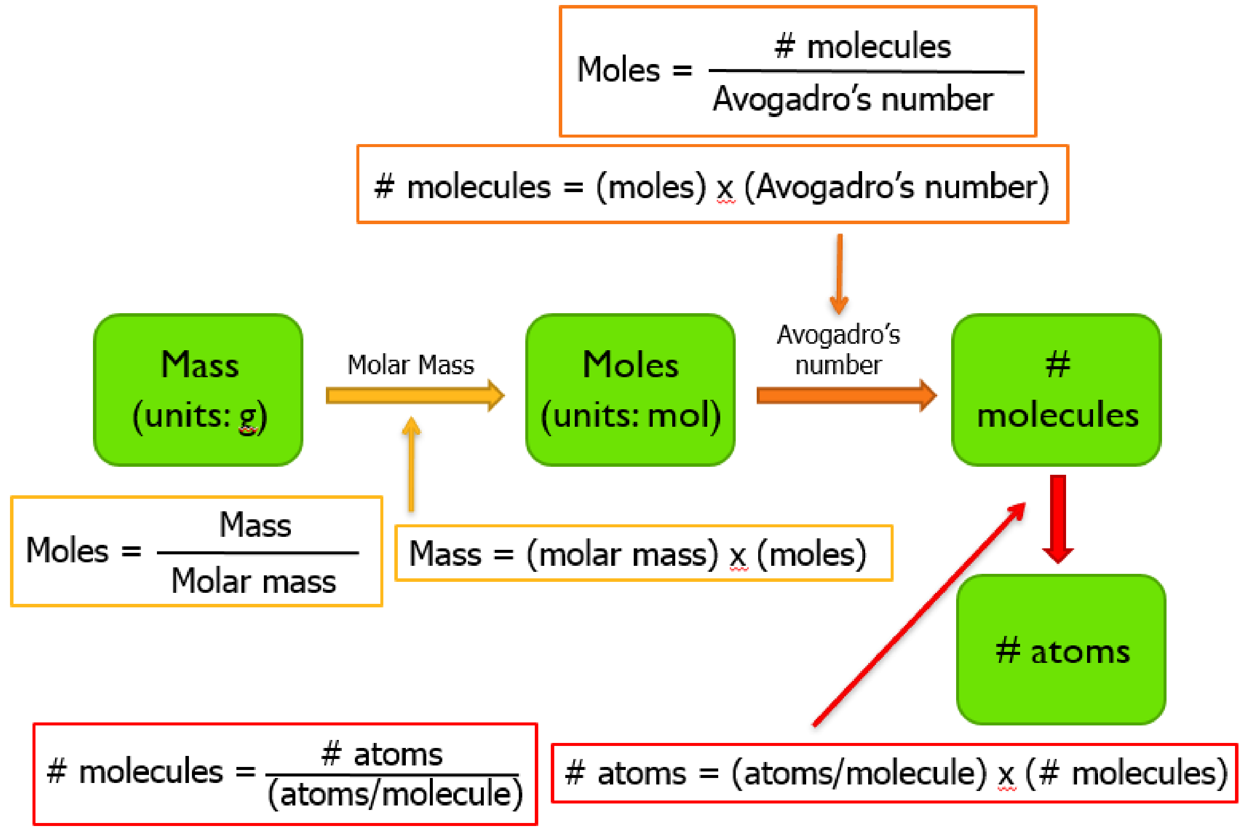
6.2.2 Calculating Molar Mass

- The molar mass of a molecule is equal to the sum of the molar masses of all its components.
- Example: $M(H_2O) = [2 \times M(H)] + M(O)$
  $= (2 \times 1.008\ g/mol) + 15.999\ g/mol$
  $= 18.015\ g/mol$

  So 1 mole of H_2O molecules weighs 18.015 grams (or 6×10^{23} H_2O molecules weigh 18.015 grams).

6.3 Converting between Mass, Moles and Number of Molecules

- Avogadro's number allows you to convert between moles and number of molecules.
- Molar mass allows you to convert between mass and moles.



Self Quiz 6.3

You will need to use knowledge from sections 0.6 and 5.3.4 in addition to 6.3 to answer these questions.

1. Answer the following questions about magnesium difluoride (MgF_2):
    a. What is the molar mass of magnesium difluoride?
    b. If you have a sealed bottle of MgF_2 that contains 35.6 grams of the gas, how many moles of MgF_2 are there?
    c. How many molecules of MgF_2?
    d. How many atoms of fluorine?
2. A car tire at has an internal volume of 20.0 L and is at a temperature of 22 °C. It is filled with air to a total pressure of 30 psi. You will need the following conversion factors:

    1 atm = 14.696 psi
    1 lb = 453.6 g

    a. What is the amount in moles of air in the tire?
    b. If the air is entirely nitrogen (N_2), how many grams of it are in the tire?
    c. How many pounds are in the tire?

3. Suppose you have two identical balloons, one filled with 1.00 mole of carbon monoxide (CO) and the other filled 1.00 mole of carbon dioxide (CO_2). Both gases are at 1.00 atm and 25.0 ^oC.
    a. What is the volume in liters of each balloon?
    b. Calculate the molar masses of CO and CO_2. What does this tell you about the relative weights of the two balloons? Does the molar mass effect the volume of gas?

Self Quiz 6.3 – Answers

You will need to use knowledge from sections 0.6 and 5.3.4 in addition to 6.3 to answer these questions.

1. Answer the following questions about magnesium difluoride (MgF_2):

    a. What is the molar mass of magnesium difluoride?

    $M(MgF_2)$ = M(Mg) + [2 x M(F)]
    = 24.305 g/mol + (2 x 18.998 g/mol)
    = 62.302 g/mol

    b. If you have a sealed bottle of MgF_2 that contains 35.6 grams of the gas, how many moles of MgF_2 are there?

    n = m/M
    = 35.6 g / 62.305 g/mol
    = 0.571 mol

    c. How many molecules of MgF_2?

    # molecules = n x N_A
    = 0.571 mol x (6 x 10^{23} molecules/mol)
    = 3.43 x 10^{23} molecules MgF_2

    d. How many atoms of fluorine?

    # atoms = (# molecules) x (atoms/molecule)
    = (3.43 x 10^{23} molecules MgF_2) x (2 F / MgF_2)
    = 6.85 x 10^{23} atoms F

2. A car tire at has an internal volume of 20.0 L and is at a temperature of 22 °C. It is filled with air to a total pressure of 30 psi. You will need the following conversion factors:

    1 atm = 14.696 psi
    1 lb = 453.6 g

a. What is the amount in moles of air in the tire?

T = 22 °C + 273.15 = 294.15 K

P = 30 psi x (1 atm / 14.696 psi) = 2.0414 atm

$$n = \frac{PV}{RT} = \frac{(2.0414\,atm)(20.0\,L)}{(0.08206\,\frac{L\,atm}{K\,mol})(294.15\,K)} = 1.69\,mol$$

b. If the air is entirely nitrogen (N_2), how many grams of it are in the tire?

$M(N_2)$ = 2 x M(N) = 2 x 14.0067 g/mol = 28.0134 g/mol

$m(N_2)$ = n x M = 1.69 mol x 28.0134 g/mol = 47.3 g

c. How many pounds are in the tire?

m_{lbs} = 47.3 g x (1 lb / 453.6 g) = 0.104 lbs

3. Suppose you have two identical balloons, one filled with 1.00 mole of carbon monoxide (CO) and the other filled 1.00 mole of carbon dioxide (CO_2). Both gases are at 1.00 atm and 25.0 °C.
   a. What is the volume in liters of each balloon?

T = 25.0 °C + 273.15 = 298.15 K

$$V_{CO} = \frac{nRT}{P} = \frac{(1.00\,mol)(298.15\,K)(0.08206\,\frac{L\,atm}{K\,mol})}{1.00\,atm} = 24.5\,L$$

$$V_{CO_2} = \frac{nRT}{P} = \frac{(1.00\,mol)(298.15\,K)(0.08206\,\frac{L\,atm}{K\,mol})}{1.00\,atm} = 24.5\,L$$

The volume of both balloons is 24.5 litres.

b. Calculate the molar masses of CO and CO_2. What does this tell you about the relative weights of the two balloons? Does the molar mass effect the volume of gas?

M(CO) = M(C) + M(O) = 12.0107 g/mol + 15.9994 g/mol = 28.0101 g/mol

$M(CO_2)$ = M(C) + {2 x M(O)} = 12.0107 g/mol + (2 x 15.9994 g/mol) = 44.0095 g/mol

Carbon dioxide (CO_2) is heavier than carbon monoxide (CO) so the balloon containing CO_2 will be heavier than the balloon containing CO since there is equal amount of gas molecules in each balloon.

The molar mass of the gases does not affect the volume of gas because the Ideal Gas Law is dependent on number of moles, not weight, of the gas.

CHAPTER 7 – CHEMICAL NOMENCLATURE

7.1 Introduction to Naming

- Chemical compounds are named using the International Union for Pure and Applied Chemistry (IUPAC) system.
- This is a very detailed system. In this course, you only need to know how to name binary compounds (substances composed of 2 elements).
- Ionic and covalent compounds have different naming conventions so you must first identify if the compound is ionic (metal + non-metal) or covalent (two non-metals) before naming.

7.2 Ionic Compounds

- The name of the metal is unchanged.
- The suffix "ide" replaces the ending of the non-metal.
  - ex. "oxygen" become "oxide"
  - ex. "sulphur" becomes "sulphide"
  - ex. "chlorine" becomes "chloride"
- Examples:
  - $NaCl$ = sodium chloride
  - K_2S = potassium sulphide
  - MgF_2 = magnesium fluoride

7.3 Covalent Compounds

- Similar to ionic compounds, the name of the less electronegative element is unchanged and it is written first in the formula.
- Also similarly, the suffix "ide" replaces the ending of the more electronegative element, which is written second in the formula.
- Unlike ionic compounds, Greek prefixes are added to show any many atoms of each element are present.

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

- By convention, the prefix "mono" is not included for the first element.
    - ex. CO is called carbon monoxide instead of monocarbon monoxide.
- Examples:
    - C_2H_6 = dicarbon hexahydride
    - PCl_3 = phosphorous trichloride
    - N_2O = dinitrogen monoxide
    - NO = nitrogen monoxide

7.4 Comparing Ionic and Covalent Naming

- Prefixes are needed for naming covalent compounds but not ionic compounds because the amount of each element in an ionic compound can be determined using the ions' charges.
    - ex. In magnesium oxide, magnesium will form a +2 ion to achieve a full valence shell; oxygen will form a -2 ion. So to balance the charges, magnesium and oxygen must be present in a 1:1 ratio so the formula is MgO.
- In a covalent compound, different combinations of elements can be achieved because atoms can bond by single, double or triple covalent bonds and can form linear, branching or cyclic structures.
    - ex. Carbon and oxygen can combine as CO or CO_2.

7.5 Compounds with Polyatomic Ions

- A list of the names of common polyatomic ions is given below.
- There is only one common polyatomic cation; the rest are anions.
- Most polyatomic anions contain oxygen.
    - The endings "ite" and "ate" and the prefixes "hypo" and "per" are used to describe how many oxygen atoms are in the ion. You do not need to understand the details of this convention for this course.
- When polyatomic ions are present in an ionic compound, the naming convention described in section 7.2 still hold.
    - ex. K(ClO) = potassium hypochlorite[24]
    - ex. $(NH_4)Cl$ = ammonium chloride

[24] The polyatomic ion is not usually surrounded by brackets if the formula contains only one (ex. this compound is usually written as KClO) but using the brackets helps to identify the polyatomic ions until you become familiar with them. Including brackets is not incorrect; it is just not the standard convention.

CATION		IUPAC NAME	
NH_4^+		ammonium ion	
ANION	**IUPAC NAME**	**ANION**	**IUPAC NAME**
$C_2H_3O_2^-$	acetate ion	OH^-	hydroxide ion*
CO_3^{2-}	carbonate ion	ClO–	hypochlorite io
ClO_3^-	chlorate ion	NO_3^-	nitrate ion
ClO_2^-	chlorite ion	NO_2^-	nitrite ion
CrO_4^{2-}	chromate ion	ClO_4^-	perchlorate ion
CN^-	cyanide ion*	MnO_4^-	permanganate ion
$Cr_2O_7^{2-}$	dichromate i ı	PO_4^{3-}	phosphate ion
HCO_3^-	hydrogen carbonate ion	SO_4^{2-}	sulfate ion
HSO_4^-	hydrogen sulfate ion	SO_3^{2-}	sulfite ion

Self Quiz 7A

1. Give the name or chemical formula for each of the following compounds:
   a. ammonium oxide
   b. NaF
   c. CO_2
   d. MgS
   e. $Na_2(SO_4)$
   f. dihydrogen monoxide
   g. SiO_2
   h. rubidium oxide
   i. carbon tetrahydride
   j. Na(CN)

Self Quiz 7A – Answers

1. Give the name or chemical formula for each of the following compounds:
   a. ammonium oxide = ***$(NH_4)_2O$*** ←[25]
   b. NaF = ***sodium fluoride***
   c. CO_2 = ***carbon dioxide***
   d. MgS = ***magnesium sulphide***
   e. $Na_2(SO_4)$ = ***sodium sulphate***
   f. dihydrogen monoxide = ***H_2O***
   g. SiO_2 = ***silicon dioxide***
   h. rubidium oxide = ***Rb_2O***
   i. carbon tetrahydride = ***CH_4***
   j. Na(CN) = ***sodium cyanide***

[25] Here, the polyatomic ion must be surrounded by brackets with the subscript “2” exterior to the brackets to indicate that there are two groups of NH_4^+.

7.6 Compounds with Transition Metals

7.6.1 Current Convention

- When ionic compounds involve a transition metal, the naming is slightly more complicated because transition metals often have multiple oxidations states (section 3.7).
- In this case, the rules from section 7.2 hold true but a Roman numeral is added to indicate the metal's oxidation state.
- This Roman numeral is put in brackets after the metal's name.
- Example:

  $FeCl_2$ = Iron (II) chloride
  $FeCl_3$ = Iron (III) chloride

7.6.2 Old Convention

- In the old convention, the charge of the metal ion was indicated by a suffix.
- When a metal has two common oxidation states, the ending "-ous" was used for the ion of lower charge and "-ic" was used for the ion of higher charge.
- These ending were added to the metal's Latin name. The Latin names of some common metals (transition and main group) are listed on below.

Element Name	Symbol	Latin Name
Antimony	Sb	Stibium
Copper	Cu	Cuprum
Gold	Au	Aurum
Iron	Fe	Ferrum
Mercury	Hg	Hydrargyrum
Potassium	K	Kalium
Sodium	Na	Natrium
Tin	Sn	Stannum

- Example:

  $FeCl_2$ = Ferrous chloride
  $FeCl_3$ = Ferric chloride

Self Quiz 7B

1. Give the name or chemical formula for each of the following compounds. For compounds involving transition metals, use the current naming system.
   a. $Sr(MnO_4)_2$
   b. Iron (III) nitrate

c. As_2O_5
d. Calcium iodide
e. $Fe(ClO_4)_3$
f. N_2O
g. Lead (II) nitrate
h. Aluminum sulphate

Self Quiz 7B – Answers

1. Give the name or chemical formula for each of the following compounds. For compounds involving transition metals, use the current naming system.
    a. $Sr(MnO_4)_2$ = ***Strontium permanganate***
    b. Iron (III) nitrate = ***$Fe(NO_3)_3$***
    c. As_2O_5 = ***Diarsenic pentaoxide***
    d. Calcium iodide = ***CaI_2***
    e. $Fe(ClO_4)_3$ = ***Iron (III) perchlorate***
    f. N_2O = ***Dinitrogen oxide***
    g. Lead (II) nitrate = ***$Pb(NH_3)_2$***
    h. Aluminum sulphate = ***$Al_2(SO_4)_3$***

7.7 Common Names

- Many common compounds have non-standard names that were assigned to them before the IUPAC naming system was developed.
- These common names are often so prevalent that they are used in scientific literature.
- Examples:
    - H_2O = water
    - NH_3 = ammonia

7.8 Acids

7.8.1 Acid Naming Systems

- Acids will be discussed in detail in a later course. For now, think of an acid as a compound that dissociates to produce H^+ ions when dissolved in water.
- Acids are named with two different systems, depending on whether or not they contain oxygen.
- Note that by definition, acids are dissolved in water so if a compound that could be an acid exists as an undissolved solid, it is named with the previously described systems (sections 7.2 – 7.6).

7.8.2 Acids without Oxygen

- Begin with the name of the non-hydrogen element. Remove the ending then add the prefix “hydro-” and the suffix “-ic acid”.
- Examples:
    - HF = hydrofluoric acid
    - HCN = hydrocyanic acid
    - H_2S = hydrosulphuric acid

7.8.3 Acids with Oxygen

- Acids with oxygen are polyatomic ionic compounds.
- If the name of the anion ends with the suffix “–ite”, replace the suffix with “–ous acid”.
    - ex. H_2SO_3 = sulphurous acid
    - ex. HNO_2 = nitrous acid
- If the name of the anion ends with the suffix “–ate,”, replace the suffix with “–ic acid.”
    - ex. H_2SO_4 = sulphuric acid
    - ex. HNO_3 = nitric acid
- Just like the endings “–ite” and “–ate” for polyatomic ions, the endings “–ous acid” and “–ic acid” for acids indicate how many oxygen atoms are present in the compound.

Self Quiz 7.8

1. Give the name or chemical formula for each of the following acids:
    a. HCl
    b. Hydriodic acid
    c. H_2CO_3
    d. Perchloric acid

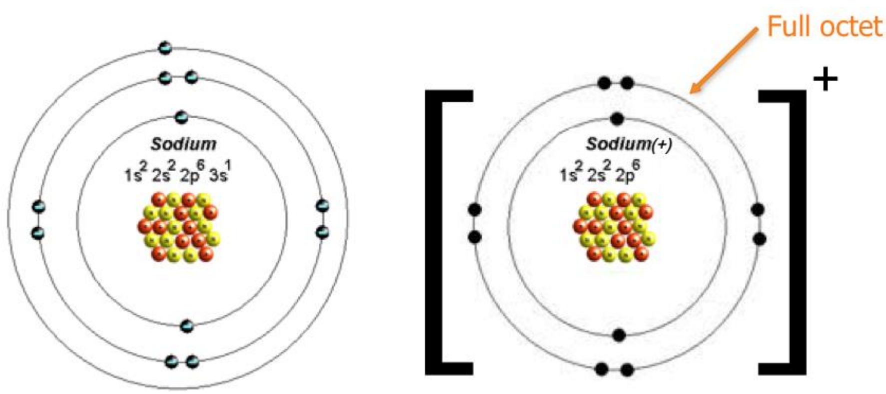
Self Quiz 7.8 – Answers

2. Give the name or chemical formula for each of the following acids:
    e. HCl = ***Hydrochloric acid***
    f. Hydriodic acid = ***HI***
    g. H_2CO_3 = ***Carbonic acid***
    h. Perchloric acid = ***$HClO_4$***

CHAPTER 8 – CHEMICAL BONDING

8.1 Octet Rule in Bonding

- Recall from section 4.7 that the Octet Rule states that elements react to form bonds (compounds) in such a way as to put eight electrons in their valence shell.
- ex. Sodium has one electron in its valence shell so it likes to form ionic bonds in which it donates an electron to a non-metal. This makes it a sodium +1 ion (Na^+), which has an identical electronic configuration as neon with eight electrons in its valence shell.
- The Octet Rule governs all chemical bonding.



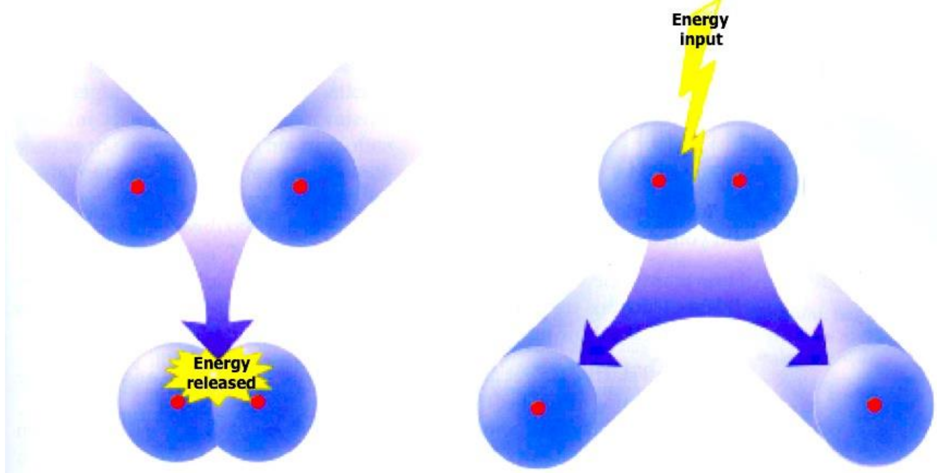
8.2 Chemical Bonds

8.2.1 Types of Bonds

- A chemical bond is the attraction between two or more atoms that allows the formation of chemical substances.
- The types of bonds that atoms form depend on the nature of the elements involved.
- They all involve electrons from the valence shell of participating atoms being shared or transferred to allow the atoms to obtain full valence shells.
- There are three types of bonds:
    - Covalent: Bond between two non-metals; involves sharing of electrons.
    - Ionic: Bond between a metal and a non-metal; involves transfer of electrons from the metal to the non-metal.
    - Metallic: Bond between two metals; involves sharing of electrons in an electron cloud.

8.2.2 Energy in Bonding

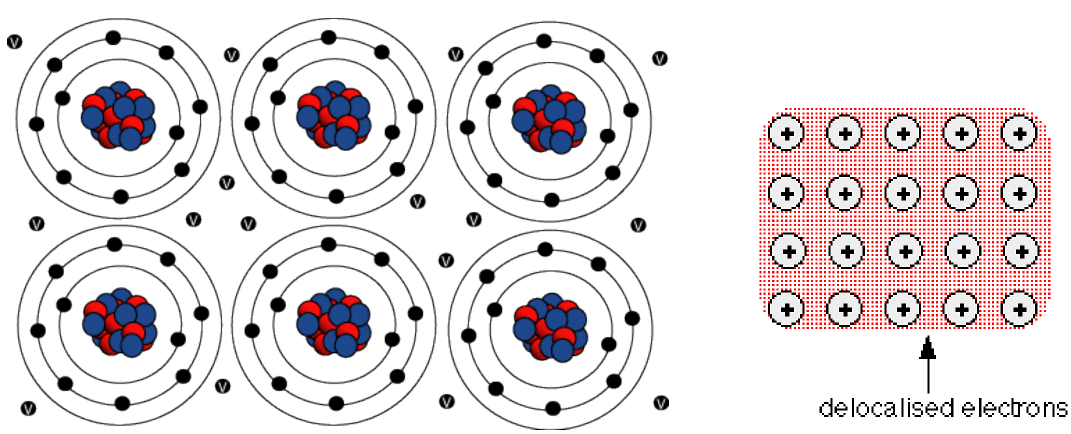
- Atoms are more stable in a covalent bond than they are individually (because formation of the bond gives them a full octet).
- Stable systems are at a lower energy than less stable systems.
- So the formation of a covalent bond is accompanied by an energy release.
- And breaking a covalent bond requires an input of energy.



8.3 Metallic Bonds

8.3.1 The Bonds

- Metallic bonds occur between multiple metal atoms when no non-metals are present.
- Metallic bonds are composed of delocalized electrons that are contained in an 'electron cloud' around all the metals' positively charged nuclei.

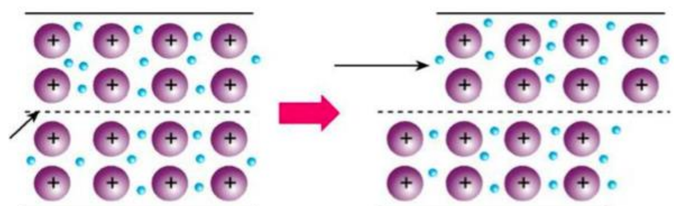
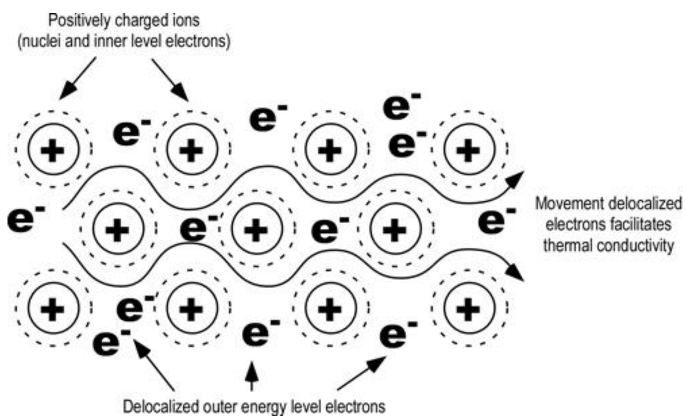
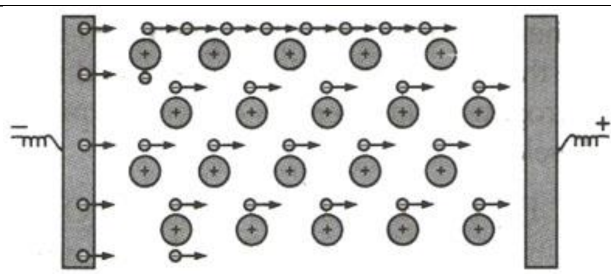


- The electrons exist simultaneously over all atoms to neutralize the nuclear charges (no ions are created).

- So there is no such thing as a single metal bond.[26]
- Metals need to bond by an electron cloud with a large number of atoms involved because metals have only a few electrons in their valence shell (usually 1 – 3 electrons) so a lot of metals need to share their electrons for the atoms to obtain a full valence shell.

8.3.2 Properties

- Metallic bonding is responsible for all of the characteristic properties of metals.

Property	Explanation	Figure
High malleability & ductility	Atoms are not rigidly bonded to each other so they can shift easily, making the material is easy to manipulate.	
High thermal conductivity	Valence electrons are not specifically associated with an atom so they can vibrate rapidly throughout the material, carrying heat.	
High electric conductivity	Valence electrons are not specifically associated with an atom so they can move freely throughout the material, carrying electric current.	
High Luster	There are valence electrons at almost every possible energy level so it is very likely that there will be an electron transition capable of adsorbing every wavelength of light that hits that metal. When these electrons drop back down to their ground state, they release the light, making the metal look shiny. Recall electron transitions from section 2.7.	

[26] It might appear like there is a single metal bond in mercury because mercury exists as Hg_2 but mercury is actually the only metal that does not form metallic bonds; it forms a covalent bond. This is why mercury is the only metal that is liquid at room temperature.

8.4 Electronegativity

- Electronegativity is important to understand the difference between ionic, covalent and polar covalent bonds.
- It is the tendency for an atom to attract a bonding electron pair.
- Electronegativity is usually inversely proportional to size (so smaller atoms have higher electronegativities) because the positively charged nucleus has a stronger pull on the valence electrons.
    - So fluorine has the highest electronegativity on the Periodic Table; caesium has the lowest.
- Electronegativity is usually measured with the Pauling scale, which has values between 0.7 and 4.0.
- The difference in the electronegativities of two elements participating in a chemical bond (ΔEN) provides insight about the nature of the bond.
    - If ΔEN is very small (ΔEN < 0.4), then the bonding electrons are equally shared and the bond is covalent.
    - If ΔEN is large (ΔEN > 1.7), then the more electronegative atom steals electron(s) from the less electronegative atom and the bond is ionic.
    - If ΔEN is between 0.4 and 1.7, then the bond has both covalent and ionic character. This is called polar covalent; it involves unequal sharing of electrons in a covalent bond.

H 2.20																
Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98
Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16
K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96
Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66
Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus
*	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27	
**	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr	

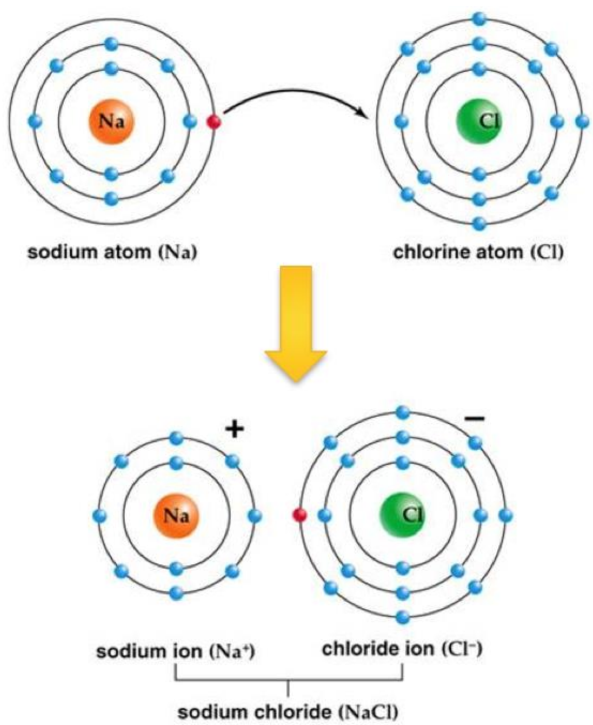
8.5 Ionic Bonds and Compounds

8.5.1 Ionic Bonds

- Recall from section 3.6 that ions are atoms with a charge.
    - Cations are positively charged (fewer electrons than protons).
    - Anions are negatively charged (more electrons than protons).
- An ionic bond is an attraction between an anion and cation.
- They form between metals and non-metals.

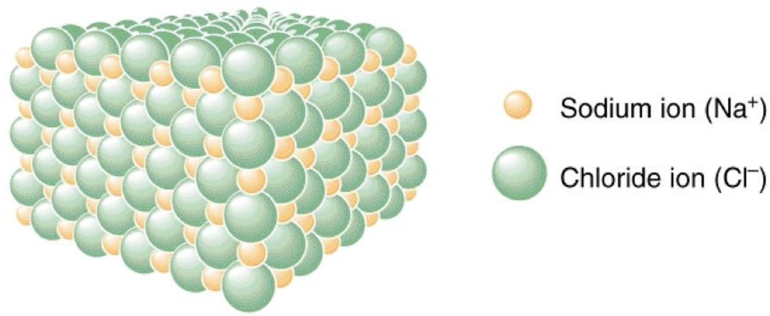
- Metals have few electrons in their valence shell (<4) so they give up electrons to achieve a full valence shell (becoming cations).
- Non-metals have a lot of electrons in their valence shell (>4) so they gain electrons to achieve a full valence shell (becoming anions).
- Semi-metals have a middle number of electrons (~4). They don't usually participate in ionic bonding but when they do, they give up or accept electrons (becoming cations or anions respectively).

- When a metal and non-metal come into contact, the metal valence electrons transfer to the non-metal, causing the atoms to become oppositely charged.
- The oppositely charged ions are attracted to each other so they form an ionic bond (ex. NaCl).
- An ionic bond is not a true bond (like covalent bonds); it is a strong attraction between oppositely charged ions.



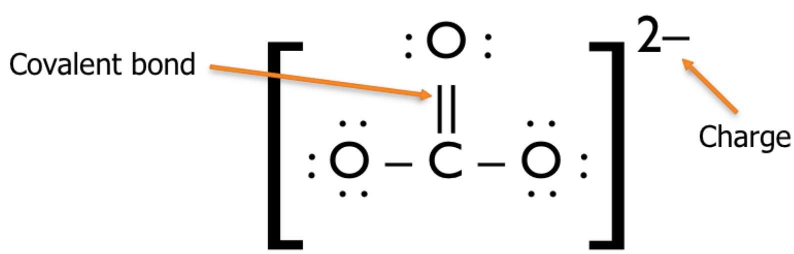
8.5.2 Ionic Compounds

- Ionic compounds do not exist as molecules. Instead, they exist as a large, ordered, 3D lattices of positively and negatively charged ions called an ionic lattice.
- An ionic lattice is arranged so that positive ions are adjacent to negative ions (because like charges attract).



8.5.3 Polyatomic Ions

- Polyatomic ions are ionized (charged) molecules, ie. ions that contain two or more atoms covalently bonded together.
  - There are covalent bonds between the atoms within a polyatomic ion.
  - There are ionic bonds between positive and negative polyatomic ions (or between a polyatomic ion and a single atom ion).
- They form ionic compounds because the main bonds holding together the compound are ionic (covalent bonds are present within the ions only).
- ex. CO_3^{2-}

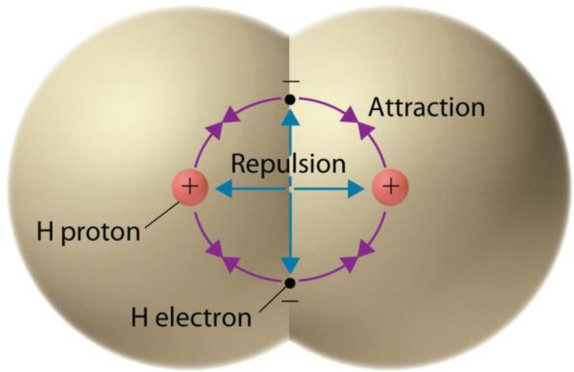


8.6 Covalent Bonds and Compounds

8.6.1 Covalent Bonds

- Covalent bonds are considered the only "true" chemical bonds.
- They involve sharing of valence electrons between two atoms.
- The name describes the bonds accurately:
  - "Co" means partnership, collaboration, cooperation (aka. sharing);
  - "Valent" refers to the valence electrons.
- Atoms can share two (single bond), four (double bond) or six (triple bond) electrons.
- All shared electrons are counted as simultaneously being in the valence shell of both atoms involved in the bond.
- Covalent bonds hold two atoms together because the nucleus of one atom (positive) attracts the electrons of the other atom (negative).

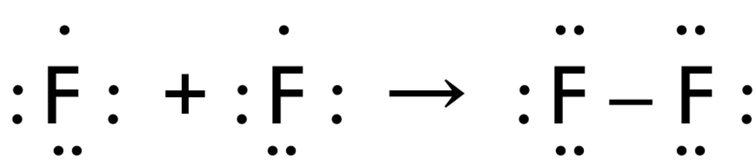
- But the atoms do not fuse because the nuclei of the two atoms repel each other (both positively charged) and the electrons of the two atoms repel each other (all negatively charged).
- So there is an equilibrium between attractive and repulsive forces that determines how far apart the atoms in the bond are from one another.
- ex. H_2



8.6.2 Representing Covalent Bonds

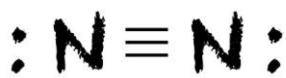
- Covalent bonds can be represented in three ways:
    - Two horizontal dots;
    - Two vertical dots;
    - A single vertical line.
- Each dot represents one electron; each line represents two electrons.
- A double bond will have either four dots or two lines and a triple bond will have six dots or three lines.

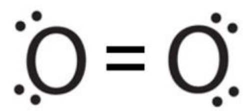
- Electrons in a covalent bond are double counted because the shared electrons are considered to belong to both atoms simultaneously.
    - ex. Two fluorine atoms join with a single covalent bond to form F_2. Each fluorine atom has seven electrons in its valence shell so it needs one more for a full octet. When two fluorine atoms bond, they both share one of their valence electrons. These two electrons are counted as being in the valence shell of both fluorine atoms so both have a full octet.



8.6.3 Multiple Covalent Bonds

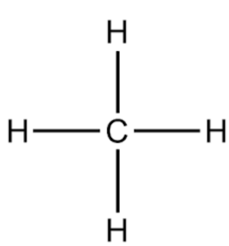
- If an atom needs more than one additional electron to achieve a full octet, it can bond with multiple atoms and/or form multiple covalent bonds.
- Multiple covalent bonds are bonds in which four (ex. O_2) or six electrons (ex. N_2) are shared. They are called a double and triple covalent bond respectively.
- When more electrons are shared between atoms, the bond is shorter and stronger.
    - A double bond is shorter and stronger than a single bond.
    - A triple bond is shorter and stronger than a double bond.

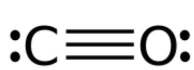


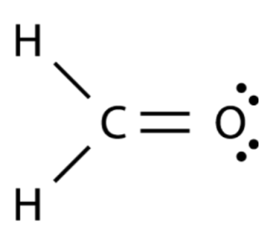


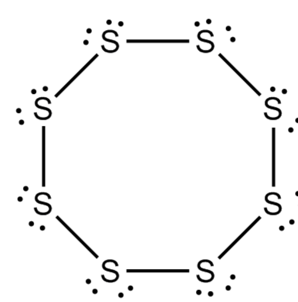
8.6.4 Forming Covalent Compounds

- Atoms will often need to form covalent bonds with more than one atom to achieve a full octet.
- These bonds can be single (ex. CH_4) or multiple (ex. CO) covalent bonds or a combination (ex. H_2CO).
- Some elements form predictable molecules.
    - ex. All elements in column 7 (halogens) form single-bonded diatomic (two atom) molecules with other elements from their same column (F_2, Cl_2, Br_2, I_2).
- The structure of some elements must be determined experimentally.
    - ex. Sulphur should theoretically be stable as S_2 (like O_2) but it actually exists as S_8.





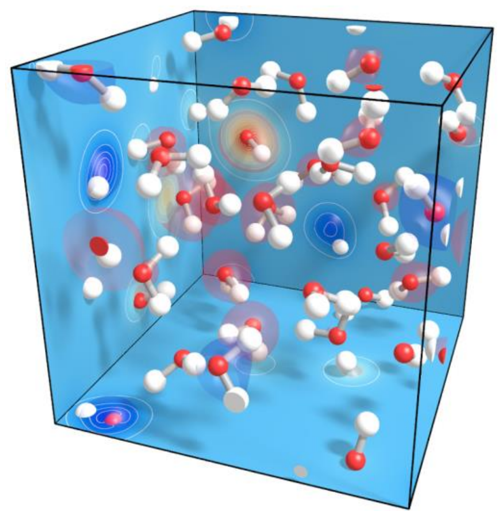




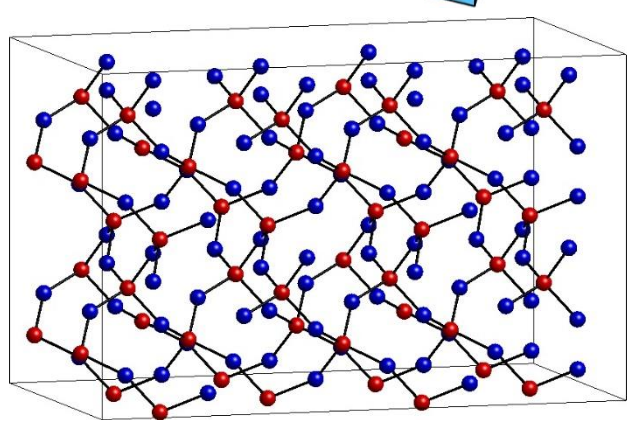
8.6.5 Types of Covalent Compounds

- Covalent bonds either form molecules (molecular substance) or network solids.
- Molecules are distinguishable groups of atoms.
    - ex. H_2O -- a glass of water contains millions of millions H_2O molecules; each molecule is distinguishable from the others
- Network solids are large 3-dimensional structures of atoms joined by covalent bonds.
- Network solids are exceptionally strong with very high melting and boiling points (ex. diamond, quartz).

A block of water showing the individual H_2O molecules…



A block of quartz showing the continuous SiO_2 network…



8.7 Polar Covalent Bonds and Compounds

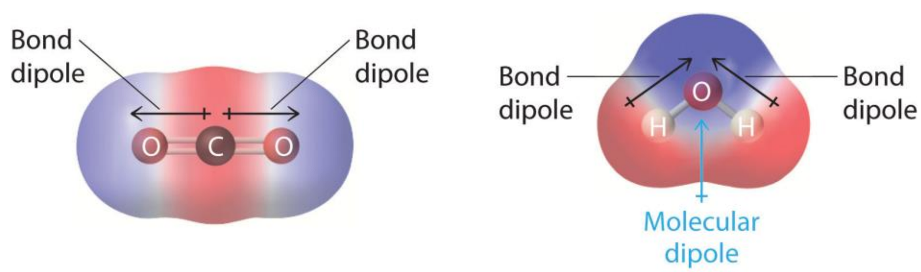
8.7.1 Polar Covalent Bonds

- Polar covalent bonds are thought of as having both ionic and covalent character.

- Polar covalent bonds involve an unequal sharing of electrons between two atoms, ie. they are covalent bonds where the electrons are attracted more strongly by one atom in the bond.
- One atom acquires a partial positive charge (δ^+) while the other atom acquires an equivalent partial negative charge (δ^-).
    - δ is the symbol for 'partial'

8.7.2 Polar Covalent Compounds

- Determining if a compound is polar covalent can be tricky because not all molecules that contains polar bonds are polar.
    - If a molecule contains all non-polar bonds, it is non-polar.
    - If a molecule containing polar covalent bonds is symmetrical, the partial charges cancel and the overall molecule is non-polar.
    - If a molecule containing polar covalent bonds is asymmetrical, the overall molecule is polar.



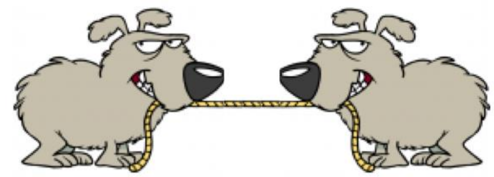
- So to determine if a molecular is polar, you must know its shape.
- It is important to know if a molecular is polar because this has a large effect on the intermolecular forces (section 9.3).

8.8 Determining if a Bond is Ionic, Covalent or Polar Covalent

- There are some general rules to determine if a bond is ionic, polar or non-polar.
    - A metal bonded to a non-metal is ionic (ex. NaCl).
    - A bond between two atoms of the same non-metal element is non-polar (ex. Cl_2).
    - Hydrocarbons (compounds containing only carbon and hydrogen) are non-polar (ex. CH_4).
    - Most other bonds between non-metals are polar (ex. H_2O, HF).
- If the bond type is not obvious, you can use electronegativity.
- Electronegativity can be thought of like a game of tug-o-war where the rope represents the electrons involved in the bond.
    - When two opponents are evenly matching, the rope will not move = covalent bond.

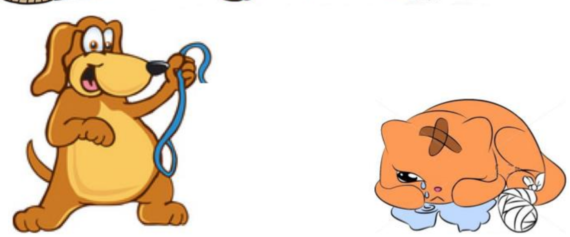
- When there is competition with one opponent is stronger, the stronger opponent has more rope but does not steal it entirely = polar covalent bond.
- When one opponent is substantially stronger, he takes the rope completely from the weaker opponent = ionic bond.

Covalent bond (non-polar)…



Polar covalent bond
(cat is less electronegative)…

Ionic bond
(cat is much less electronegative)…



8.9 VSEPR Theory

8.9.1 The Theory

- Every molecule has a distinct 3-dimensional shape, which can be predicted by valence shell electron pair repulsion (VSEPR) theory.
- VSEPR theory is based on the idea that the valence electrons (bonding electrons and lone pairs) around different atoms in a molecule repel each other.
- So the molecule arranges itself in order to put electron pairs as far from each other as possible.
- The ideal VSEPR shape for a molecule minimizes the repulsion between electron pairs (and so maximizes stability of the molecule).

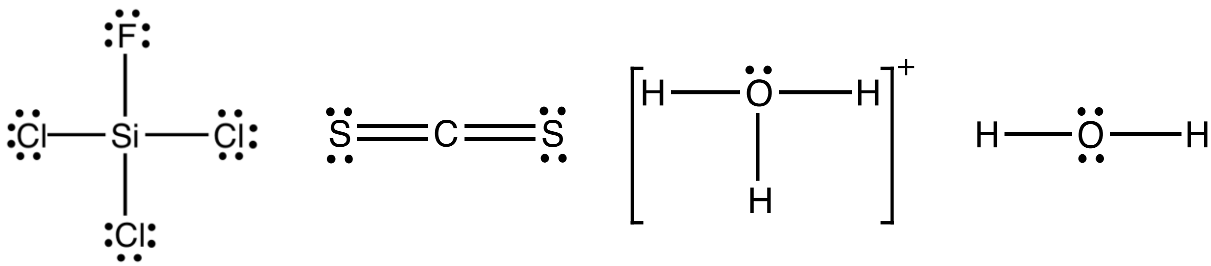
8.9.2 Electron Groups

- A non-bonding electron pair, a single bond and a multiple bond are each treated as a single electron group.
- Non-bonding electron pairs influence the shape of a molecule (and are actually more repulsive than bonding electron pairs) but they are not

included when describing the shape of a molecule because the name of the molecular shape indicates the arrangement of atoms only.
- The number of electron groups around a central atom is the steric number of that atom.
    - ex. methane, CH_4: steric number = 4
      formaldehyde, H_2CO: steric number = 3
      carbon dioxide, CO_2: steric number = 2

Self Quiz 8.9.2

1. Calculate the steric number of the central atom in the following molecules.

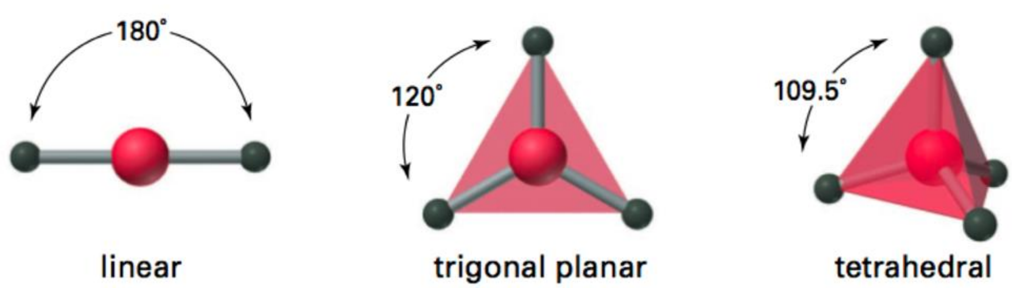


Self Quiz 8.9.2 – Answers

1. Calculate the steric number of the central atom in the following molecules.
    a. CS_2 -- ***C = 2***
    b. $SiCl_3F$ -- ***Si = 3***
    c. H_2O -- ***O = 4***
    d. H_3O^+ -- ***O = 4***

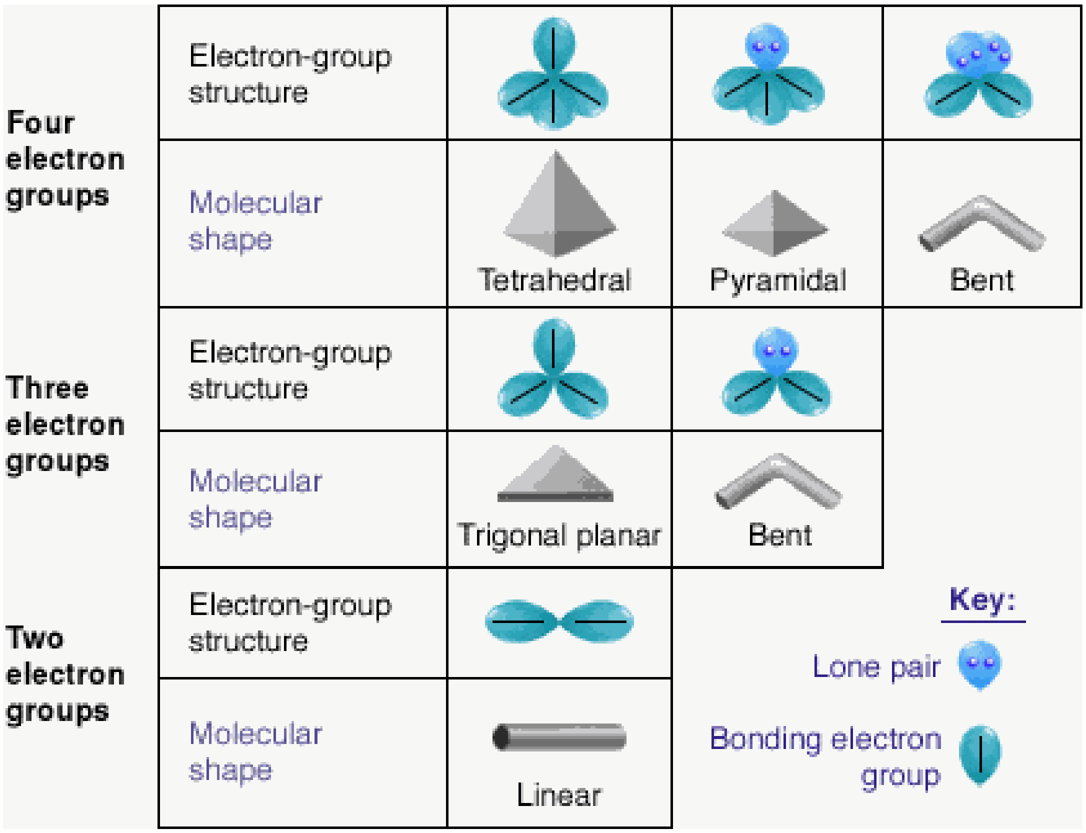
8.9.3 VSEPR Geometries

- The main VSEPR geometries are linear, trigonal planar and tetrahedral.[27]



[27] There are other important geometries for compounds involving transition metals because they can have steric numbers greater than 4 but they are not covered in this course.

- Each of these geometries is further sub-divided depending on whether the surrounding electron groups are bonding (attached to atoms) or non-bonding (lone pairs).



Self Quiz 8.9.3

1. What is the geometry of each molecule shown in Self Quiz 8.7.2?
    a. CS_2
    b. $SiCl_3F$
    c. H_2O
    d. H_3O^+
2. Remembering that molecules must have asymmetrical polar bonds to be polar overall, which of the molecules are polar?

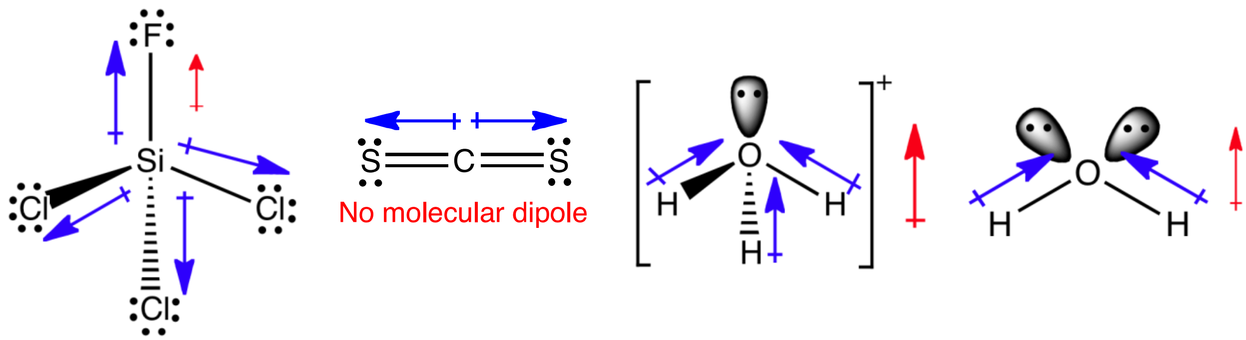
Self Quiz 8.9.3 – Answers

1. What is the geometry of each molecule shown in Self Quiz 8.7.2?
    a. CS_2 = ***Linear***
    b. $SiCl_3F$ = ***Tetrahedral***
    c. H_2O = ***Bent***

d. H_3O^+ = ***Trigonal pyramidal***

2. Remembering that molecules must have asymmetrical polar bonds to be polar overall, which of the molecules are polar?

   The molecules are shown below with the correct shape. Bond dipoles[28] are shown in blue; molecular dipoles are shown in red where applicable.



   CS_2 has no molecular dipole because it has symmetrical polar bonds.

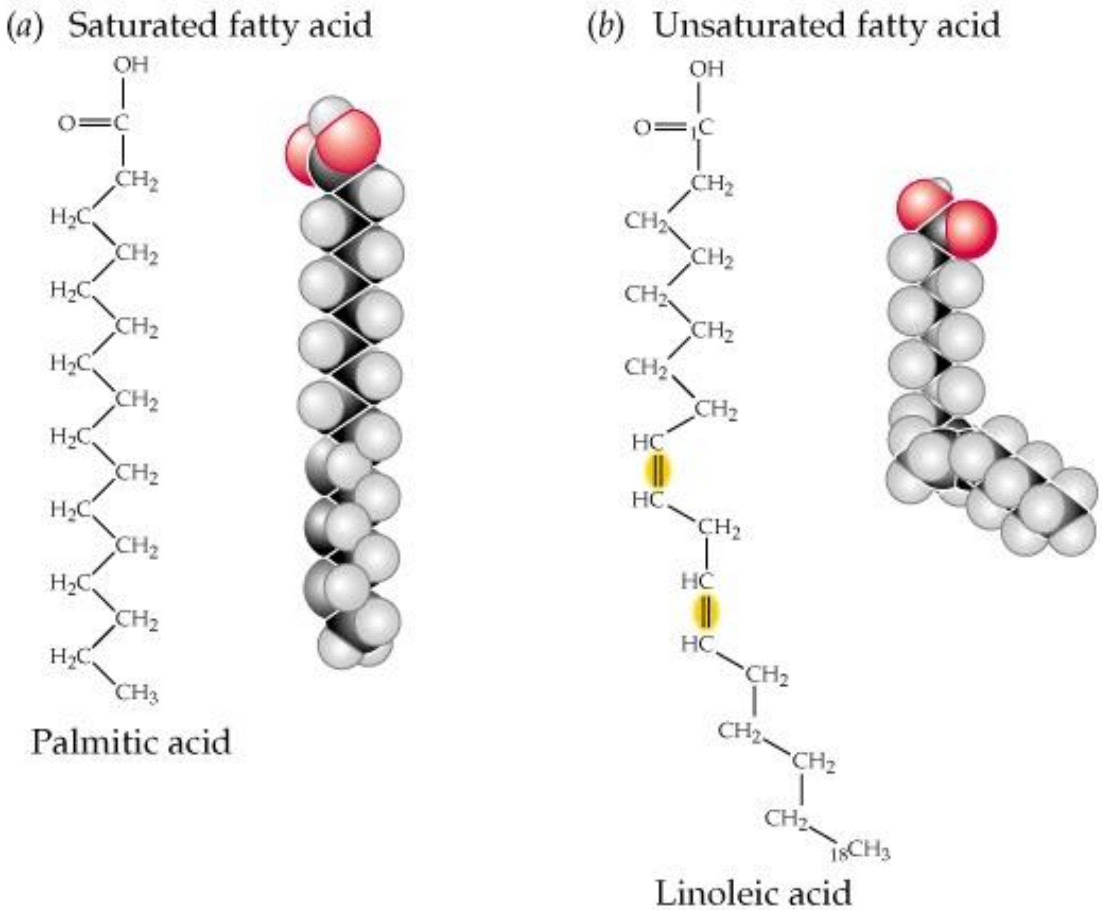
   H_2O and H_3O+ have molecular dipoles because they do not have atoms located symmetrically around the central atom.

   $SiCl_3F$ has a molecular dipole because, even though it has atoms located symmetrically around the central atom, the identity of the exterior atoms varies (there is one F atom and three Cl atoms). Fluorine is more electronegative than chlorine so it will pull slightly more strongly on the electrons, producing a molecular dipole.

8.9.4 Importance of Shape

- Shape is important for more than just determining if a molecular is polar.
- The shape of a molecule can have substantial effects on many of its properties.
- ex. Saturated fats are linear and so stack easily in the body to clog arteries. In contrast, unsaturated fats have bends so they do not stack as easily and are less solid so the body eliminates them more easily.
  - Saturated fats are linear because all of the carbon atoms have a steric number of 4.
  - Unsaturated fats are bent because they have some C = C (double bonds). The participating carbon atoms have a steric number of 3.

[28] The words polar and dipole are used to describe the same thing. A bond is polar if it has a dipole (dipole = two different poles), ex. H_2O has two polar H – O bonds; H – O bonds have a bond dipole; H_2O is polar molecule; H_2O has a molecular dipole.



8.10 Differences between Covalent and Ionic Compounds

- Covalent and ionic compounds have very different properties because (a) covalent bonds are slightly stronger than ionic bonds; (b) ionic bonds are interactions between ions while covalent bonds are "true" bonds.

8.10.1 Strength and Melting/Boiling Points

- Covalent network solids are the strongest substances on Earth with the highest melting and boiling points because they are composed entirely of strong covalent bonds, ex. diamonds.
- Ionic compounds are almost as strong with almost as high melting and boiling points because they are composed entirely of strong ionic bonds, ex. salt crystals.
- Covalent molecular substances are usually weak with low melting and boiling points because, although they contain strong covalent bonds

between atoms, the forces between molecules are weak and easily broken, ex. oxygen gas.

8.10.2 Solubility

- Ionic compounds dissolve easily in water because the charges on the ions allow them to form ion-dipole forces with water (section 13.2), ex. salt water.
- It is possible to dissolve some covalent molecular substances in water, but not as much as ionic compounds, ex. sugar water.
- Network solids do not dissolve, ex. diamonds.

8.11 Lewis Dot Diagrams

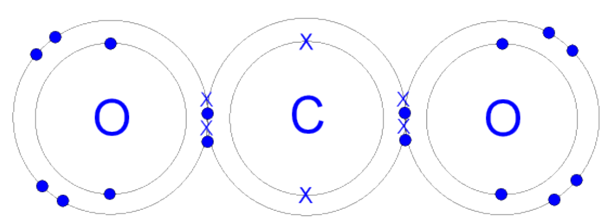
8.11.1 The Diagrams

- The Bohr diagram (section 2.3) of atoms and molecules is very informative but it is very time consuming to draw.
- Since only valence electrons are involved in bonding, Gilbert Lewis developed a shorthand drawing method in 1916 called the Lewis Dot Diagram.
- In Lewis Dot Diagrams, the symbol of the element is surrounded by dots and lines.
    - Each valence electron is shown as a dot. If there are more than 4 valence electrons, they are drawn in pairs.[29]
    - Each bonding pair is shown as a line.
    - No lower level electrons are drawn.

8.11.2 Examples

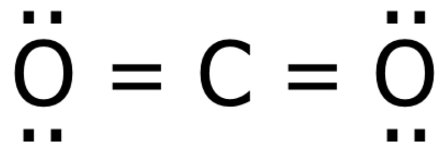
Covalent compound: CO_2

Bohr diagram…



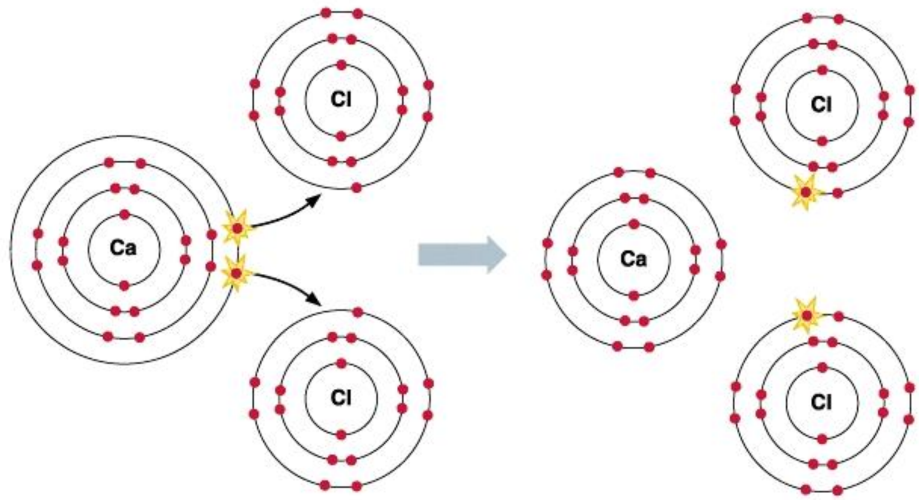
[29] Chemists sometimes get lazy and do not draw the non-bonding electrons because they know the electrons are present without drawing them (because of the Octet Rule). You cannot use that shortcut.

Lewis Dot diagram…



Ionic compound: $CaCl_2$

Bohr diagram…



Lewis Dot diagram…

8.11.3 Steps for Drawing a Lewis Dot Diagram

1. Determine the type of bonding (ionic, covalent, metallic). If it metallic, you cannot draw a Lewis structure. If it is ionic or covalent, continue to step #2.
2. Determine the total number of valence electrons on all the atoms in the molecule. If the bonding is covalent, continue to step #3. If the bonding is ionic, continue to step #9. For compounds involving polyatomic ions, first use steps #3-8 to arrange the covalently bound atoms within the ion then use steps #9-14 to arrange the ions.

3. Draw the atoms in the correct orientation without bonds. The least electronegative atom is usually the central atom.
4. Connect the atoms with single bonds.
5. Calculate how many valence electrons you have not yet placed (remember that every bond uses two electrons).

6. Add the remaining electrons two at a time in lone pairs around the outside atoms until every atom has eight valence electrons (except hydrogen, which only needs two) or until you run out of electrons.
7. If you run out of electrons and any of the atoms do not yet have full octets, move some of the electrons in lone pairs to bonding pairs, making multiple bonds. Remember that you cannot form multiple bonds with hydrogen because it is too small.
8. Continue to step #15.
9. If the chemical formula is not given, determine the correct ratio of metals to non-metals so that the metal(s) can get rid of all their valence electrons and the non-metal(s) can accept enough valence electrons to achieve a full valence shell.
10. Calculate the charges on all ions.
11. Draw all atoms in the compound, showing the full valence shell for non-metals and the empty valence shell for metals.
12. Put square brackets around the ions, showing their charges.
13. Arrange the ions so that they alternate anions and cations (since opposite charge attract).
14. Continue to step #15.

15. Check that three criteria are met: (1) the number of valence electrons total in the molecule is correct; (2) each atom has eight valence electrons (except hydrogen, which only needs two); (3) for ionic compounds, the charges on all ions add to zero.

Self Quiz 8.11

1. Draw a Lewis structure for cyanogen, C_2N_2, a poisonous gas used as a fumigant and rocket propellant. *Hint*: It is linear.
2. Draw the Lewis structure for the polyatomic ion nitronium, NO_2^+.
3. Draw the Lewis structure of potassium nitride, K_3N.

Self Quiz 8.11 – Answers

1. Draw a Lewis structure for cyanogen, C_2N_2, a poisonous gas used as a fumigant and rocket propellant. *Hint*: It is linear.

   Both carbon and nitrogen are non-metals so the bonding will be covalent.

   Valence electrons (C_2N_2) = 2(C) + 2(N) = (2 x 4) + (2 x 5) = 18

   EN(C) = 2.5,   EN(N) = 3.0   so the carbon atoms are likely in the centre

$$N-C-C-N$$

Remaining electrons = 18 – 2(3) = 12

$$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{N}}-C-C-\overset{\cdot\cdot}{\underset{\cdot\cdot}{N}}:$$

Remaining electrons = 12 – 2(6) = 0

Do all the atoms have eight electrons in their valence shell? NO! The nitrogen atoms have eight electrons in their valence shell but the carbon atoms only have four.

So move some lone electron pairs to bonding pairs, making double bonds:

$$:\overset{\cdot\cdot}{N}=C-C=\overset{\cdot\cdot}{N}:$$

Do all the atoms have eight electrons in their valence shell? NO! The nitrogen atoms have eight electrons in their valence shell but the carbon atoms only have six. We're moving in the right direction though.

So move some lone electron pairs to bonding pairs, making triple bonds:

$$:N\equiv C-C\equiv N:$$

Do all the atoms have eight electrons in their valence shell? YES!

Is the number of valence electrons total in the molecule correct? YES!

2. Draw the Lewis structure for the polyatomic ion nitronium, NO_2^+.

   Both nitrogen and oxygen are non-metals so the bonding will be covalent.

   Valence electrons (NO_2^+) = N + 2(O) – charge = 5 + (2 x 6) – 1 = 16

   EN(O) = 3.5,   EN(N) = 3.0   so nitrogen likely in the centre

$$[O-N-O]^+$$

   Remaining electrons = 16 – 2(2) = 12

..   ..
[:N – O – N:]$^+$
..   ..

Remaining electrons = 12 – 2(6) = 0

Do all the atoms have eight electrons in their valence shell? NO! The nitrogen atoms have eight electrons in their valence shell but the oxygen only has four.

So move some lone electron pairs to bonding pairs, making double bonds:

..   ..
[:N = O = N:]$^+$

Do all the atoms have eight electrons in their valence shell? YES!

Is the number of valence electrons total in the molecule correct? YES!

3. Draw the Lewis structure of potassium nitride, K_3N.

Potassium is a metal and nitrogen is a non-metal so the bonding will be ionic.

Valence electrons (K_3N) = 3(K) + N = (3 x 1) + 5 = 8

Potassium needs to lose one electron to achieve a full valence shell so it will become K^+

Nitrogen needs to gain three electrons to achieve a full valence shell so it will become N^{3-}

..
[K]$^+$ [:N:]$^{3-}$ [K]$^+$
..
[K]$^+$

Do all the atoms have eight electrons in their valence shell? YES!

Is the number of valence electrons total in the molecule correct? YES!

Do all the charges on the ions add to zero? YES!

CHAPTER 9 – FORCES

9.1 Types of Forces

- Intramolecular forces: "Intra" means "on the inside" or "within" so these are forces that occur within a compound.
- Intermolecular forces: "Inter" means "between" or "among" so these are the forces that occur between molecules.

9.2 Intramolecular Forces

- Intramolecular forces are attractive forces that hold together the atoms in a compound.
- All substances contain intramolecular forces (except noble gases because they are stable as individual non-bonded atoms).
- The intramolecular forces were described in section 9.2. They include:
    - Covalent bonds
    - Ionic bonds
    - Metallic bonds
- Intramolecular forces are much stronger (>10x) than intermolecular forces.

9.3 Intermolecular Forces

9.3.1 Definition

- Intermolecular forces are attractive forces that hold molecules together in condensed phases (liquid and solid).
- They are only present in molecular compounds.
    - Non-molecular compounds cannot have intermolecular forces because intermolecular means "in between molecules". Individual molecules cannot be identified in non-molecular compounds so there is no opportunity for intermolecular forces.

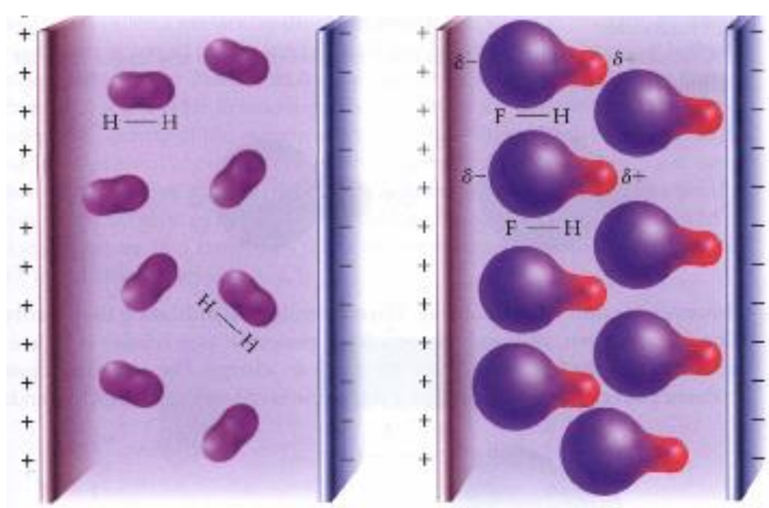
- Intermolecular forces are relatively weak and easily broken compared to intramolecular forces.
- There are three main types of intermolecular forces in molecular compounds. From weakest to strongest they are:
    1. London dispersion forces
    2. Dipole-dipole forces
    3. Hydrogen bonding
- Ion-dipole forces are also an intermolecular force (section 13.2). They are only present when ions are dissolved in a liquid (usually water); they do not occur in pure substances.

9.3.2 Effect on State

- Intermolecular forces determine the state of a compound at room temperature (solid, liquid or gas).
- Compounds with stronger intermolecular forces usually exist in a more condensed state (liquid or solid) because the molecules are attracted to each other and so are located in close proximity to each other.
- Intermolecular forces must be broken when a solid is melted to a liquid or when a liquid is vapourized to a gas.

9.3.3 Dipole-Dipole Forces

- Dipole-dipole forces occur between two polar molecules.
- Recall that polar molecules have partial negative and partial positive ends. Since unlike charges attract, the partial negative end of one polar molecule is attracted to the partial positive end of another polar molecule (and visa versa).
- This attractive force is called a dipole-dipole force.

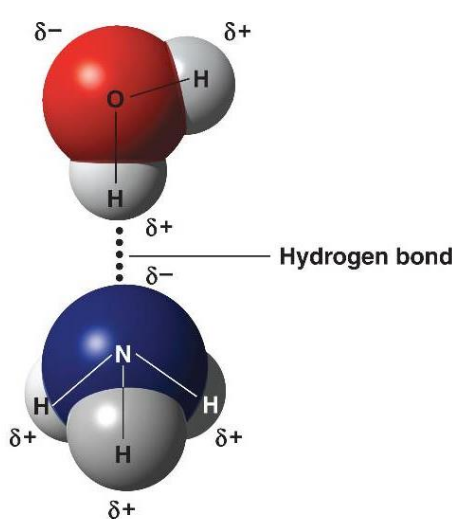


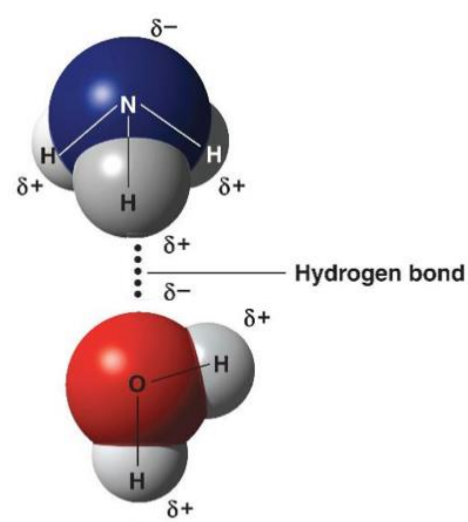
Non-polar molecule **Polar molecule**

9.3.4 Hydrogen Bonding

9.3.4.1 Forces

- Hydrogen bonding (H-bonding) is a particular type of very strong dipole-dipole force.
- Since hydrogen only has one electron, when hydrogen is engaged in a covalent bond, all of its electrons are in that bond.
- If hydrogen is bonded to a very electronegative atom (nitrogen, oxygen or fluorine), its only electron is pulled strongly away from it so the hydrogen nucleus acquires a very large partial positive charge, δ^+.
- When hydrogen has a very large δ^+ charge, it experiences a very large attractive force with the lone electron pair on very electronegative atoms (nitrogen, oxygen or fluorine) in neighbouring molecules.
- This strong intermolecular attraction is called is a hydrogen bond (even though it is not a true bond).
- Hydrogen bonds only occur between hydrogen bonded to nitrogen, oxygen or fluorine and a nitrogen, oxygen of fluorine atom on another molecule.
    - o Hydrogen will have strong dipole-dipole forces with other elements but the force will not be strong enough to classify it as a hydrogen bond.
- ex. H_2O and NH_3

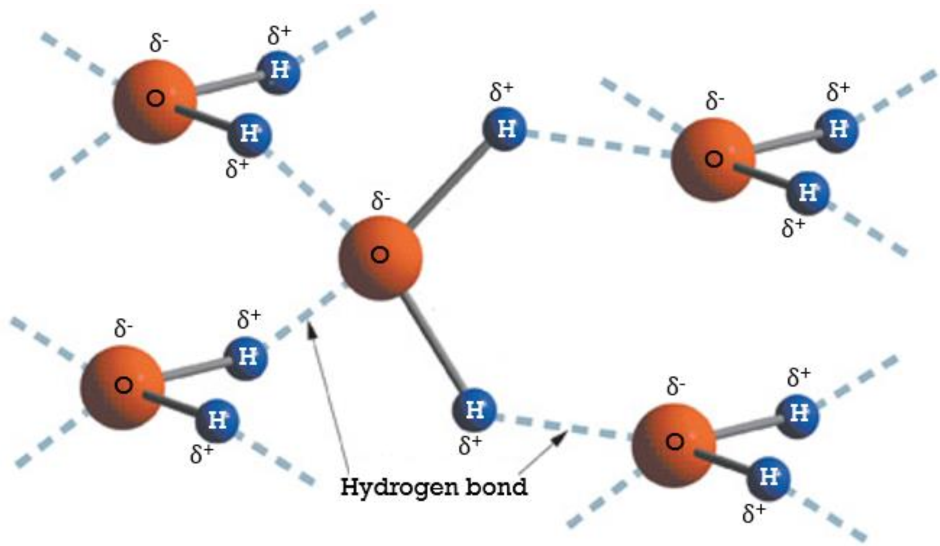




9.3.4.2 Importance

- Hydrogen bonds are integral to human life.

- They are responsible for the relatively high boiling point of water compared to other small molecules (ex. methane (CH_4), carbon dioxide (CO_2)), which allows water to be liquid at room temperature instead of a gas.[30]
- Hydrogen bonds are also responsible for holding two strands of DNA together. The H-bonds are strong enough to attract complimentary base pairs during replication but they are weak enough that the strands can separate when required.



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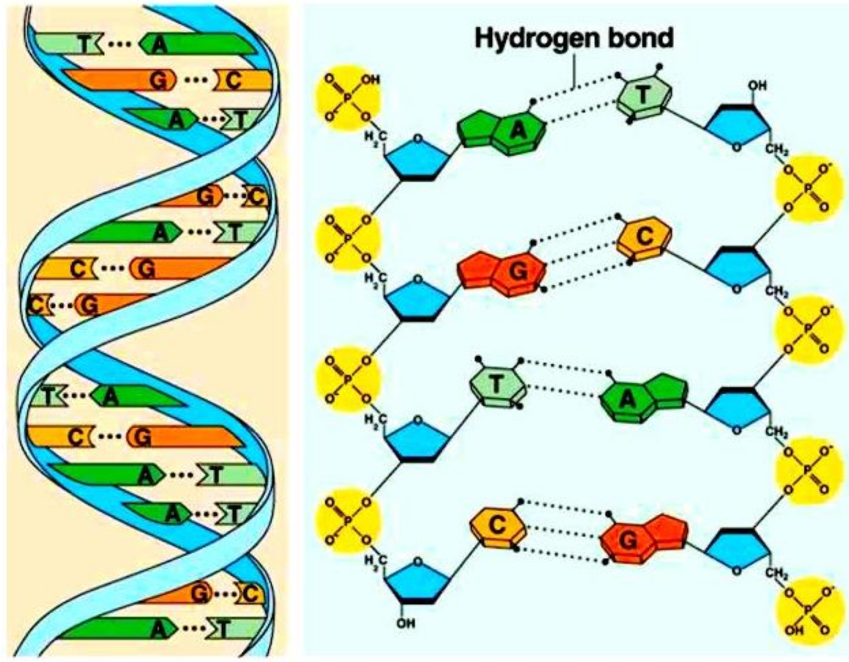
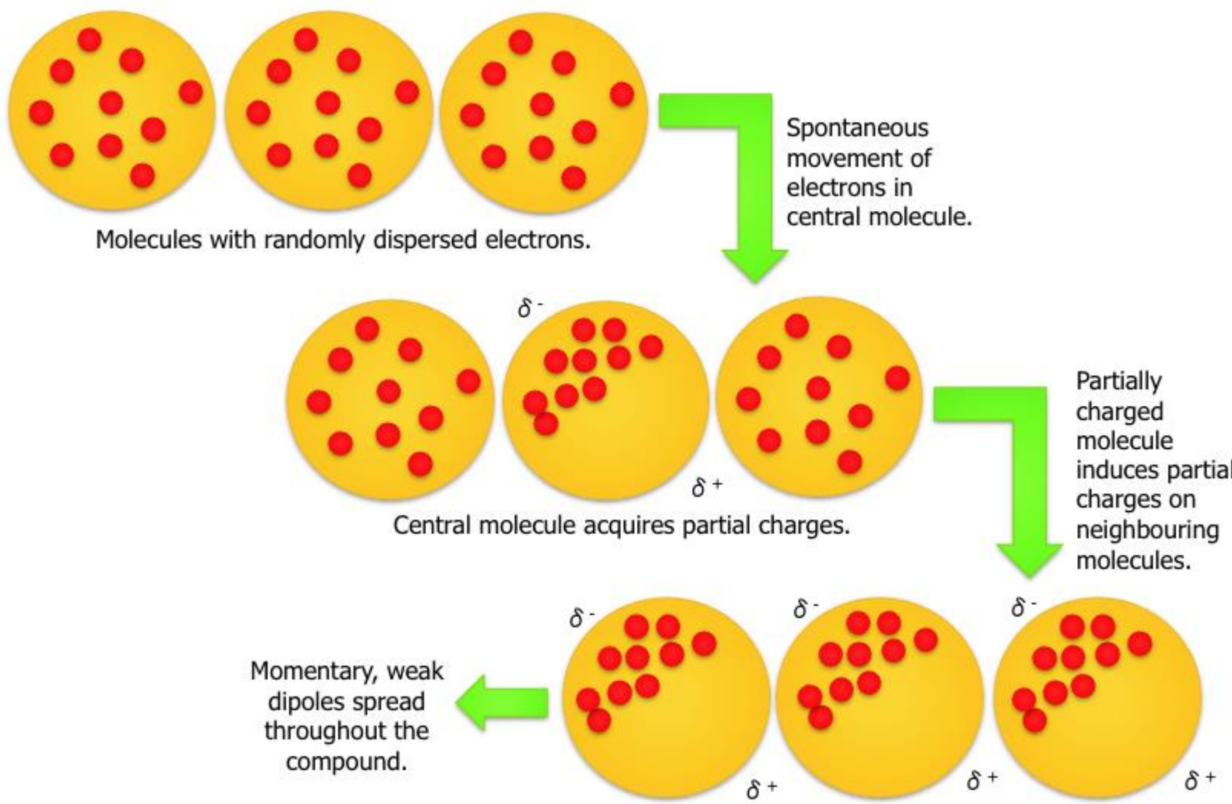
[30] Hydrogen bonding also explains why solid H_2O is less dense than liquid H_2O (usually solids are more dense than liquids) but that is beyond the scope of this course.



9.3.5 London Dispersion Forces

9.3.5.1 Forces

- London dispersion forces are the weakest intermolecular force.
- They occur in <u>every</u> molecular substance but they are only relevant in non-polar compounds because non-polar molecules do not experience any other intermolecular force.
- Dispersion forces result from random, momentary, very weak dipoles created in non-polar molecules due to the random shifting of electrons.
- When this happens, the small, momentary dipole in one molecule induces a small, momentary dipole in its neighbouring molecule, which then spreads throughout the substance like the wave at a baseball game.



9.3.5.2 Strength

- Although dispersion forces are weak, they can become very significant in large molecules due to the large number of electrons participating.
    - ex. CH_4 (natural gas) → gas a room temperature
      C_8H_{18} (gasoline) → liquid at room temperature
      $C_{20}H_{42}$ (candle wax) → solid at room temperature
- The strongest London dispersion forces are stronger than hydrogen bonds.
    - ex. H_2O contains hydrogen bonds and is a liquid at room temperature; $C_{20}H_{44}$ (candle wax) contains only London forces and is solid at room temperature.

SECTION III – REACTIONS

CHAPTER 10 – CHEMICAL REACTIONS

10.1 What is a Chemical Reaction?

- A chemical reaction occurs when one or more substances are converted into new substances that have compositions and properties different from those of the starting substances.
- A chemical reaction either produces or consumes energy (heat, electricity, light or mechanical energy).
- The starting substance(s) are called reactants; the new substance(s) are called products.
- A chemical reaction is represented by an equation:

$$\text{Reactants} \rightarrow \text{Products}$$

10.2 Energy

- To transform reactants into products, we must break chemical bonds in the reactants and form new chemical bonds in the products.
- So energy is at the heart of all chemical reactions.
- A chemical reaction is a two step process:
    1. Breaking the bonds of the reactants.
    2. Forming new bonds in the products.
- Whether an overall reaction produces or consumes energy depends on the balance between the energy needed to break bonds in the reactants and the energy released by forming new bonds in the products.
    - Recall from section 8.2.2 that the formation of bonds between two atoms increases their overall stability because the bonded atoms have a full octet and a more stable system is at a point of lower energy. So energy is always required to break a chemical bond and energy is always released when new bonds form.

10.3 Coefficients

- Chemical reactions have to respect the Law of Conservation of Matter so the number of atoms of each element must be equal on both sides of the equation.
- Numerical coefficients ensure this when a chemical reaction is balanced.
- Example:

  - $H_2 + O_2 \rightarrow H_2O$ is unbalanced because there are two atoms of oxygen on the left side and only one atom of oxygen on the right side
  - $2\ H_2 + O_2 \rightarrow 2\ H_2O$ is the balanced equation where 2 is a coefficient (the coefficient '1' is not written by convention)
- It is a convention to use the smallest whole numbers possible as coefficients.
- Examples:
  - $H_2 + \frac{1}{2}\ O_2 \rightarrow H_2O$ is correctly balanced but it does not use the standard conventions for coefficients because there is a fraction coefficient.
  - $4\ H_2 + 2\ O_2 \rightarrow 4\ H_2O$ is also correctly balanced but it does not use the standard conventions for coefficients because the coefficients are larger than necessary (by a factor of two).
- Typically, fraction coefficients are accepted (ie. considered good enough) but unnecessarily large coefficients are not.
  - The important thing to remember when using fraction coefficients is that these coefficients represent a correct ratio but you cannot have half a molecule of O_2.
- You <u>cannot</u> change the molecular formula of the reactants or products to balance an equation.
  - ex. $H_2 + O \rightarrow H_2O$ is incorrect because the molecular formula of O_2 has been changed to O. Coefficients show only the ratio of molecules to one another; the molecular formula cannot be changed.
- The coefficients refer to ratios of number of molecules or moles of the reactants and products, <u>not mass or volume</u>.

10.4 Balancing Chemical Reactions

10.4.1 Tips

- When balancing chemical equations, deal with the pure elemental substances last (H_2, O_2, Fe, etc.) because these substances are the easiest to balance because they contain only one element so adjusting their coefficient will not affect multiple elements.
- Similarly, try not to change the amount of the most complicated compound because adjusting its coefficient will affect many elements

10.4.2 Example

- $C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$
- Remembering the tips, we balance oxygen last because of the presence of O_2. Only put a coefficient in front of C_8H_{18} if necessary.

Carbon:

- Starting from the left, there are 8 carbon atoms in the reactants so there must be 8 carbon atoms in the products.
- Since there is one carbon atom in every CO_2 molecule, we need 8 CO_2 molecules.
- $C_8H_{18} + O_2 \rightarrow 8\ CO_2 + H_2O$

Hydrogen:

- There are 18 hydrogen atoms in the reactants so we need 18 hydrogen atoms in the products.
- Since there are 2 hydrogen atoms in every water molecule, we need 18/2 = 9 water molecules.
- $C_8H_{18} + O_2 \rightarrow 8\ CO_2 + 9\ H_2O$

Oxygen:

- Since we want to balance the pure elemental substances last, we count the number of oxygen atoms in the products and then adjust the O_2 coefficient accordingly (instead of counting the oxygen atoms in the reactants and adjusting the product coefficients).
- There are 8 x 2 = 16 oxygen atoms from CO_2 and 9 x 1 = 9 oxygen atoms from H_2O so there are 16 + 9 = 25 oxygen atoms total in the products.
- Since there are two oxygen atoms in every molecule of O_2, we need 25/2 O_2 molecules.
- $C_8H_{18} + 25/2\ O_2 \rightarrow 8\ CO_2 + 9\ H_2O$

Whole number coefficients:

- By convention, we want whole number coefficients, so we must multiply all coefficients by two to remove the fraction.
- $2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$

Self Quiz 10.4

1. Balance the following chemical reactions:
    a. $C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$
    b. $HCl + Na_2CO_3 \rightarrow NaCl + CO_2 + H_2O$
    c. $CaH_2 + H_2O \rightarrow Ca(OH)_2 + H_2$
    d. $Al_2Cl_6 + H_2O \rightarrow Al(OH)_3 + HCl$
    e. $KClO_3 \rightarrow KCl + O_2$

Self Quiz 10.4 – Answers

1. Balance the following chemical reactions:
    a. $2\ C_6H_{14} + 19\ O_2 \rightarrow 12\ CO_2 + 14\ H_2O$
    b. $2\ HCl + Na_2CO_3 \rightarrow 2\ NaCl + CO_2 + H_2O$
    c. $CaH_2 + 2\ H_2O \rightarrow Ca(OH)_2 + 2\ H_2$
    d. $Al_2Cl_6 + 2\ H_2O \rightarrow 2\ Al(OH)_3 + 6\ HCl$
    e. $2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$

10.5 Types of Chemical Reactions

10.5.1 Overview

- There are six main types of chemical reactions.
    - Single displacement (single replacement)
    - Double displacement (double replacement)
    - Synthesis (combination)
    - Decomposition
    - Dissociation & precipitation
    - Neutralization
- One reaction can be classified as more than one type (ex. a neutralization reaction can also be a double displacement).

10.5.2 Simple Reactions

- Single-replacement: Two compounds exchange one element (or polyatomic ion) by decomposing one reactant and producing two distinct products.

$$2\ Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + 2\ Fe_{(\ell)}$$

- Double-replacement: Two compounds exchange two elements (or polyatomic ions) by decomposing both reactants and producing two distinct products.

$$PCl_{3(\ell)} + 3\,AgF_{(s)} \rightarrow PF_{3(g)} + 3\,AgCl_{(s)}$$

- Decomposition: One compound breaks down (decomposes) into two or more simpler products.

$$2\,HgO_{(s)} \rightarrow 2\,Hg_{(\ell)} + O_{2(g)}$$

- Synthesis: Two or more substances combine to produce one, more complex product.

$$4\,Ni_{(s)} + 4\,Al_{(s)} + S_{8(s)} \rightarrow 2\,Ni_2Al_2S_{4(s)}$$

Self Quiz 10.5.2

1. Balance the following chemical equations and assign the reaction type:
    a. $N_2O_{5\,(g)} \rightarrow NO_{2\,(g)} + O_{2\,(g)}$
    b. $Li_{\,(s)} + N_{2\,(g)} \rightarrow Li_3N_{(s)}$
    c. $Fe(NO_3)_{3\,(aq)} + Na_2S_{\,(aq)} \rightarrow Fe_2S_{3\,(s)} + NaNO_{3\,(aq)}$
    d. $CO_{\,(g)} + NO_{\,(g)} \rightarrow CO_{2\,(g)} + N_{2\,(g)}$
    e. $SO_{2\,(g)} + O_{2\,(g)} \rightarrow SO_{3\,(g)}$

Self Quiz 10.5.2 – Answers

1. Balance the following chemical equations and assign the reaction type:
    a. $2\,N_2O_{5\,(g)} \rightarrow 4\,NO_{2\,(g)} + O_{2\,(g)}$ (decomposition)
    b. $6\,Li_{\,(s)} + N_{2\,(g)} \rightarrow 2\,Li_3N_{(s)}$ (combination)
    c. $2\,Fe(NO_3)_{3\,(aq)} + 3\,Na_2S_{\,(aq)} \rightarrow Fe_2S_{3\,(s)} + 6\,NaNO_{3\,(aq)}$ (double replacement)
    d. $2\,CO_{\,(g)} + 2\,NO_{\,(g)} \rightarrow 2\,CO_{2\,(g)} + N_{2\,(g)}$ (single replacement)
    e. $2\,SO_{2\,(g)} + O_{2\,(g)} \rightarrow 2\,SO_{3\,(g)}$ (combination)

10.5.3 Dissociation and Precipitation Reactions

10.5.3.1 Dissociation

- When an acid or ionic compound is dissolved in water, it is labeled with the subscript '(*aq*)', meaning 'aqueous' (or 'dissolved in water').
    - ex. solid table salt: $NaCl_{(s)}$
      table salt dissolved in water: $NaCl_{(aq)}$
- When the compound dissolves, it forms ions.
    - ex. $NaCl_{(aq)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$

10.5.3.2 Solubility Limits

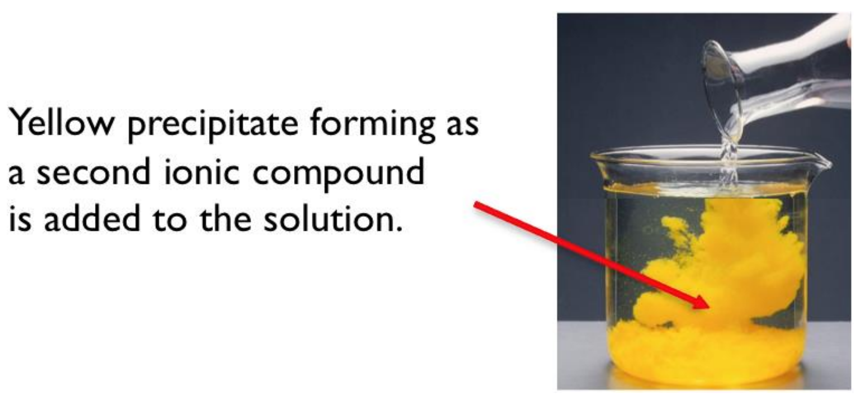
- The general solubility rules are shown below.
- Most ionic compounds are soluble but there are some exceptions.

Compounds	Exceptions
Soluble in Water	
Sodium, potassium and ammonium salts	None
Acetates and nitrates	None
Halides (chlorides, bromides, iodides)	Halides of lead (II), sliver (I) and mercury (I)
Sulphates	Sulphates of calcium, barium, lead (II) and strontium
Insoluble in Water	
Phosphates, carbonates and sulphides	Sodium, potassium and ammonium salts; calcium sulphide
Hydroxides	Sodium, potassium, calcium and barium hydroxides

- Even soluble compounds have a solubility limit, after which no more substance will dissolve.
    - ex. You can dissolve a lot of table salt in water but eventually it will stop dissolving and pool in the bottom of the glass as a solid (or form large salt crystals, such as in the Dead Sea).
- The solubility of a compound is higher in warmer water. This is why more salt dissolves in warm water than cold water.

10.5.3.3 Precipitates

- A precipitate is a solid that forms in an aqueous solution when one of three things occur:
    1. The amount of compound added to the solution exceeds the solubility limit of that compound.
    2. Two of the ionic compounds added to a solution produce ions that interact with each other to form an insoluble compound.
    3. The temperature of a saturated solution is lowered, which lowers the solubility limit, so the amount of compound in the solution now exceeds the solubility limit at the new temperature.



10.5.3.4 Rules for Identifying a Precipitate

1. Identify all ions present when the starting solutions are dissolved.
2. Identify all new possible cation–anion combinations. These are your potential precipitates.
3. Use the solubility rules (section 10.5.3.2) to determine if any of the new cation–anion combinations are insoluble salts. If no combinations represent insoluble salts, there is no precipitation reaction, just a dissociation reaction.
4. If a precipitation reaction occurs, write the net ionic equation. Make sure the formula for the precipitate is electrically neutral (ie. the charges on the cations and the charges on the anions add to zero).

10.5.3.5 Ionic Equations

- A complete ionic equation shows all ions and precipitates present in a solution after ionic compounds have been dissolved.
- A net ionic equation omits the spectator ions (any ions that are not part of the precipitate and so appear on both sides of the reaction).
- Example:
  - Precipitation reaction:
    $2\ Na_3PO_{4(aq)} + 3\ CaCl_{2(aq)} \rightarrow 6\ NaCl_{(aq)} + Ca_3(PO_4)_{2(s)}$
  - Complete ionic equation:
    $6\ Na^+_{(aq)} + 2\ PO_4^{3-}{}_{(aq)} + 3\ Ca^{2+}{}_{(aq)} + 6\ Cl^-{}_{(aq)} \rightarrow 6\ Na^+_{(aq)} + 6\ Cl^-{}_{(aq)} + Ca_3(PO_4)_{2(s)}$
  - Net ionic equation:
    $2\ PO_4^{3-}{}_{(aq)} + 3\ Ca^{2+}{}_{(aq)} \rightarrow Ca_3(PO_4)_{2(s)}$

Self Quiz 10.5.3

1. Lead (II) iodide (PbI_2) is a bright yellow powder used as a colouring pigment in paint. PbI_2 can be made by mixing a solution of lead (II) nitrate ($Pb(NO_3)_2$) and sodium iodide (NaI), both white powders. Write the resulting precipitation reaction and the net ionic equation. Can these reactions be classified as any other reaction type?

2. Consider three separate aqueous solutions of sodium nitrate ($NaNO_3$), aluminum sulphate ($Al_2(SO_4)_3$) and sodium hydroxide (NaOH). Will there be a precipitate when the three solutions are combined? Write the net ionic equations for all precipitates that form in the beaker if applicable.

Self Quiz 10.5.3 – Answers

1. Lead (II) iodide (PbI_2) is a bright yellow powder used as a colouring pigment in paint. PbI_2 can be made by mixing a solution of lead (II) nitrate ($Pb(NO_3)_2$) and sodium iodide (NaI), both white powders. Write the resulting precipitation reaction and the net ionic equation. Can these reactions be classified as any other reaction type?

$Pb(NO_3)_{2(s)}$ and $NaI_{(s)}$ are both soluble in water so the first step is to dissolve them:

$$Pb(NO_3)_{2(s)} + NaI_{(s)} \xrightarrow{H_2O} Pb(NO_3)_{2(aq)} + NaI_{(aq)}$$

Now you can write all the ions that are produced when these compounds dissolve:

$$Pb(NO_3)_{2(aq)} + NaI_{(aq)} \rightarrow Pb^{2+}_{(aq)} + 2\ NO_3^-{}_{(aq)} + Na^+_{(aq)} + I^-_{(aq)}$$

The new possible ion combinations are PbI_2 and $NaNO_3$. Using the solubility chart, we see that halides are normally soluble but not with lead (II) so PbI_2 is insoluble and will form a precipitate. We also see that sodium salts are always soluble so $NaNO_3$ is soluble and will remain in solution and Na^+ and NO_3^- ions.

We can now write the complete ionic equation:

$$Pb^{2+}_{(aq)} + 2\ NO_3^-{}_{(aq)} + 2\ Na^+_{(aq)} + 2\ I^-_{(aq)} \rightarrow PbI_{2(s)} + 2\ Na^+_{(aq)} + 2\ NO_3^-{}_{(aq)}$$

Note that in order to balance the above reaction, the starting materials have to be mixed in a ratio of 1 $Pb(NO_3)_2$: 2 NaI

The net ionic equation is:

$$Pb^{2+}_{(aq)} + 2\ I^-_{(aq)} \rightarrow PbI_{2(s)}$$

This can also be classified as a synthesis.

The complete reaction is:

$$Pb(NO_3)_{2(aq)} + 2\ NaI_{(aq)} \rightarrow PbI_{2(s)} + 2\ NaNO_{3(aq)}$$

This can also be classified as a double displacement reaction.

2. Consider three separate aqueous solutions of sodium nitrate ($NaNO_3$), aluminum sulphate ($Al_2(SO_4)_3$) and sodium hydroxide ($NaOH$). Will there be a precipitate when the three solutions are combined? Write the net ionic equations for all precipitates that form in the beaker if applicable.

$NaNO_{3(s)}$, $Al_2(SO_4)_{3(s)}$ and $NaOH_{(s)}$ are all soluble in water so the first step is to dissolve them:

$$NaNO_{3(s)} + Al_2(SO_4)_{3(s)} + NaOH_{(s)} \xrightarrow{H_2O} NaNO_{3(aq)} + Al_2(SO_4)_{3(aq)} + NaOH_{(aq)}$$

Now you can write all the ions that are produced when these compounds dissolve:

$$NaNO_{3(aq)} + Al_2(SO_4)_{3(aq)} + NaOH_{(aq)} \rightarrow 2\ Na^+_{(aq)} + NO_3^-{}_{(aq)} + Al^{3+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} + OH^-{}_{(aq)}$$

The new possible ion combinations are Na_2SO_4, $Al(NO_3)_3$ and $Al(OH)_3$. Using the solubility chart, we see that all sodium salts are soluble so Na_2SO_4 is soluble and will exist in the solution as ions; all nitrates are soluble so $Al(NO_3)_3$ is soluble and will exist in the solution as ions; and, most hydroxides are insoluble ($Al(OH)_3$ is not an exception) so $Al(OH)_3$ is insoluble and will form a precipitate.

Therefore, yes there will be a precipitate when all three solutions are combined.

The complete ionic equation is:

$$4\ Na^+_{(aq)} + NO_3^-{}_{(aq)} + Al^{3+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} + 3\ OH^-{}_{(aq)} \rightarrow Al(OH)_{3(s)} + 4\ Na^+_{(aq)} + NO_3^-{}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Note that in order to balance the above reaction, the starting materials have to be mixed in a ratio of 1 $NaNO_3$: 1 $Al_2(SO_4)_3$: 3 $NaOH$

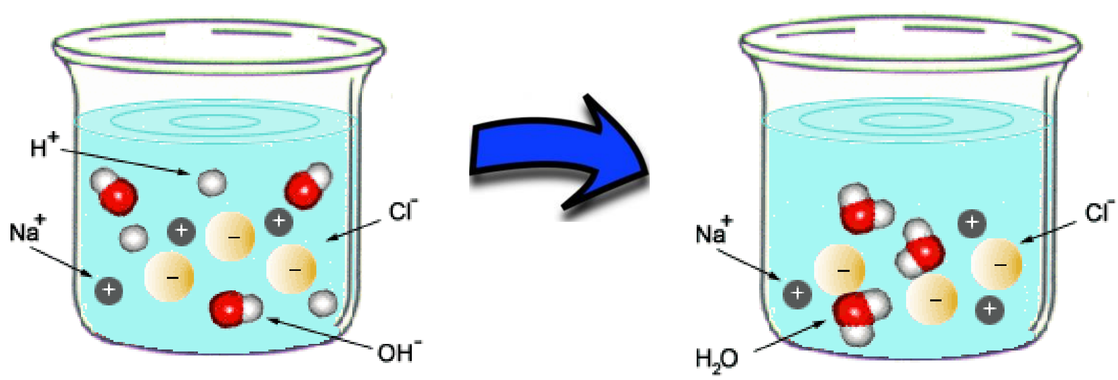
The net ionic equation is:

$$Al^{3+}{}_{(aq)} + 3\ OH^-{}_{(aq)} \rightarrow Al(OH)_{3(s)}$$

10.5.4 Neutralization Reactions

- Acids and bases are discussed in a later course. For now, like in section 7.8, think of an acid as any compound that produces $H^+_{(aq)}$ ions when dissolved in water.
  - ex. HCl, HNO_3, HF

- Think of a base as any compound that produces $OH^-_{(aq)}$ ions (hydroxide ions) when dissolved in water.
    - ex. NaOH, KOH, $Ca(OH)_2$
- When an acid and base are combined in aqueous solution, a double replacement reaction. This reaction is also a neutralization reaction that produces water.
    - The acid neutralizes the base by removing the OH^- ions.
    - The base neutralizes the acid by removing the H^+ ions.
    - The counterions (the cation paired with OH^- and the anion paired with H^+) are spectator ions.
- Example: $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$
    - The Na^+ and Cl^- ions stay dissolved in the solution because NaCl is soluble.
    - But the H^+ and OH^- ions will combine to produce a new molecule, H_2O.
    - So the net ionic equation is: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$



Self Quiz 10.5.4

1. An aqueous solution of hydrosulphuric acid, $H_2S_{(aq)}$, reacts with an aqueous solution of sodium hydroxide, $NaOH_{(aq)}$.
    a. Write the balanced complete ionic equation for the acid-base neutralization.
    b. What salt will remain behind if you evaporate the water from the neutralized solution?

Self Quiz 10.5.4 – Answers

1. An aqueous solution of hydrosulphuric acid, $H_2S_{(aq)}$, reacts with an aqueous solution of sodium hydroxide, $NaOH_{(aq)}$.
    a. Write the balanced complete ionic equation for the acid-base neutralization.

    When the compounds are first dissolved, they produce ions:

    $H_2S_{(aq)} + NaOH_{(aq)} \rightarrow 2\ H^+_{(aq)} + S^{2-}_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)}$

    The H^+ and OH^- ions then react to produce the following reaction:

    $H_2S_{(aq)} + 2\ NaOH_{(aq)} \rightarrow 2H_2O_{(l)} + S^{2-}_{(aq)} + 2Na^+_{(aq)}$

    Note that in order to balance the above reaction, the reactants must combine in a ratio of 1 H_2S : 2 NaOH

    b. What salt will remain behind if you evaporate the water from the neutralized solution?

    Na_2S

CHAPTER 11 – NUCLEAR REACTIONS

11.1 Introduction

- Normal chemical reactions occur between the electrons of atoms so the only way that the nucleus is affected is through a nuclear reaction.
- Nuclear reactions are very uncommon in nature.
- There are two types of nuclear reactions:
    - Fission: Breaks one large atom into two smaller atoms.
    - Fusion: Joins two small atoms into one larger atom.

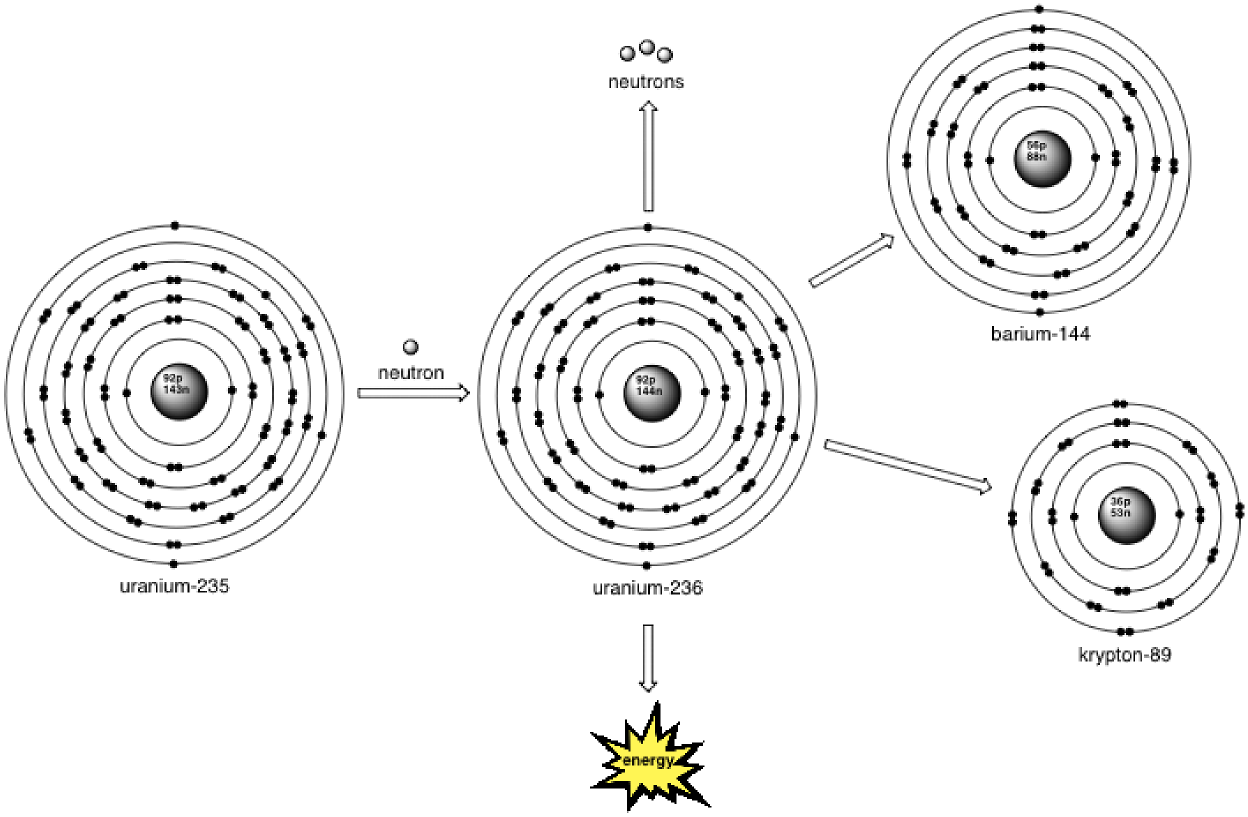
11.2 Fission

11.2.1 The Reaction

- Fission occurs only in radioactive isotopes through natural or human-induced nuclear reactions.
- The most common fission reaction occurs with uranium-235, which is naturally unstable so it breaks down into smaller elements.
- This happens naturally every few billion years.
- Humans speed up the process in nuclear reactors.

11.2.2 Nuclear Reactors

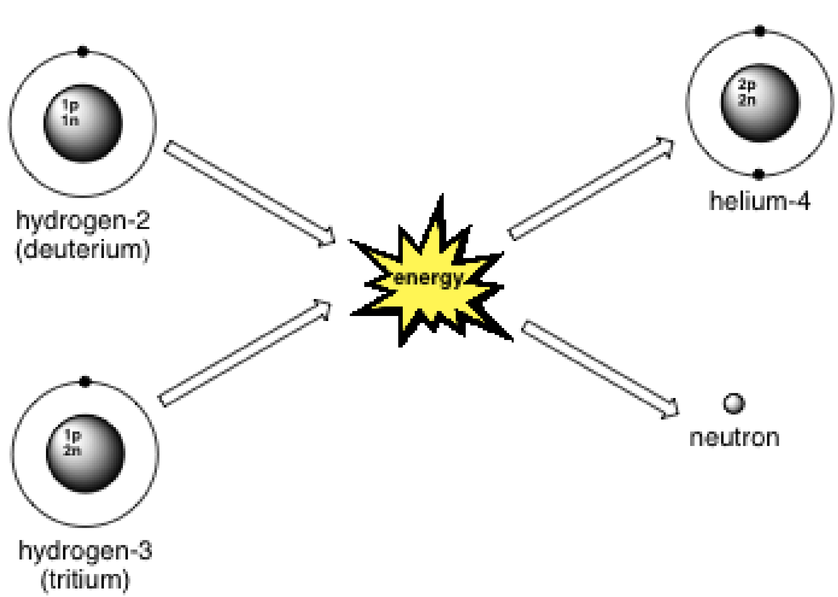
- First, a neutron is thrown at uranium-235 at very high speeds to produce uranium-236.
- Uranium-236 is very unstable so it breaks down into barium-144, krypton-89 and three extra neutrons and releases a lot of energy.
- The created neutrons collide with more uranium-235 to produce more uranium-236, which breaks down to produce more energy.
- The reaction eventually becomes self sustaining (it will continue without any assistance) and continues to produce energy.
- This reaction is shown below. Note that electrons do not participate in nuclear reactions. They are shown in the figure for illustrative purposes only.



11.3 Fusion

11.3.1 The Reaction

- Fusion occurs between small and somewhat unstable atoms (such as deuterium and tritium).
- If these atoms collide at very high speeds, they can combine to form a more stable element.
    - ex. Hydrogen-2 + Hydrogen-3 → Helium-4 + neutron
        - ↻This reaction is shown below. Again, electrons do not participate.
- Fusion does not occur naturally on Earth, but is constantly occurring in the sun (and every other star in the universe).
- Fusion reactions produce so much energy that the reactions occurring in the sun 150 million kilometers are felt on Earth and can even cause burns.
- Fusion is the theoretical answer to unlimited power because it does not produce a radioactive byproduct but humans cannot properly control it.



11.3.2 H-Bombs

- The only fusion reaction humans have produced is in an H-bomb.
- In the bomb, a small fission reaction produces the power to initiate a hydrogen fusion reaction.
- This produces a huge amount of energy but it cannot be contained.

CHAPTER 12 – STOICHIOMETRY

12.1 Definition

- Stoichiometry is the study of the quantitative aspects of chemical reactions.
- It is used to calculate the amount of reactant needed to make a desired amount of product or the amount of product that will be produced from given amounts of reactants.
- A stoichiometric chemical reaction is (a) balanced and (b) respects the Law of Mass Conservation.
- Example: Assume this is a balanced, stoichiometric chemical reaction: A + B → 2 C. Acknowledging the stoichiometry, this means:
  - 1 molecule of A + 1 molecule of B → 2 molecules of C
  - Or to use a larger (more realistic) scale:
    1 mol of A + 1 mol of B → 2 mol of C
  - It is very important to remember though that this does <u>not</u> mean:
    ~~1 gram of A + 1 gram of B → 2 grams of C~~

12.2 Stoichiometric Masses

- Recall from section 6.1 that a mole is the unit used to measure the quantity of atoms or molecules.
  - 1 mole = Avogadro's number (N_A) = 6.022×10^{23} molecules
  - You can convert between mass, moles and molecules using molar mass and Avogadro's number.
- The mass of reactants required and products produced can be calculated using the reaction stoichiometry and the molar masses of the reactants/products.
- Example: Water is produced in the reaction: $2\ H_2 + O_2 \rightarrow 2\ H_2O$. If we need 100 grams of water, we can calculate the amount of H_2 and O_2 needed for the reaction.
  - First, figure out the amount of moles of water in 100 grams (because stoichiometry refers to moles <u>not mass</u>).
    - $M(H_2O) = M(O) + [2 \times M(H)] = 16.0$ g/mol + (2 x 1.01 g/mol) = 18.02 g/mol
    - $n(H_2O) = m / M = 100$ g / 18.02 g/mol = 5.55 mol
  - Second, figure out the stoichiometric quantities of H_2 and O_2 that are needed to produce 5.55 mol of H_2O.
    - $n(H_2) = n(H_2O) \times (2\ H_2 / 2\ H_2O) = 5.55$ mol x (2/2) = 5.55 mol H_2
    - $n(O_2) = n(H_2O) \times (O_2 / 2\ H_2O) = 5.55$ mol x (1/2) = 2.775 mol H_2
  - Third, calculate the masses of H_2 and O_2 based on this number of moles.

        - $M(H_2) = 2 \times M(H) = 2 \times 1.01$ g/mol = 2.02 g/mol
        - $m(H_2) = n \times M = 5.55$ mol x 2.02 g/mol = 11.2 g
        - $M(O_2) = 2 \times M(O) = 2 \times 16.0$ g/mol = 32.0 g/mol
        - $m(O_2) = n \times M = 2.775$ mol x 32.0 g/mol = 88.8 g
    - So 11.2 grams of H_2 and 88.8 grams of O_2 are required to produce 100 grams of H_2O.
    - Finally, check your answer. If you did everything correctly, the mass of the reactants will equal the mass of the products (because of the Law of Conservation of Mass).
        - m(reactants) = $m(H_2) + m(O_2)$ = 11.2 g + 88.8 g
        - m(products) = $m(H_2O)$ = 100 g
        - m(reactants) = m(products)

Self Quiz 12A

Several questions in this quiz are review from chapters 5 and 6.

1. Suppose you have 0.10 mole of uranium (U) atoms.
    a. How many uranium atoms do you have?
    b. What is the mass in grams of this much uranium?

2. How many aluminum atoms are there in 10.0 grams of aluminum oxide, Al_2O_3?

3. If you have 2.00 moles of propane (C_3H_8), a gas used in outdoor grills and industrial torches;
    a. How many molecules of propane do you have?
    b. How many hydrogen atoms do you have?
    c. What is the mass in gram of this much propane?
    d. Propane reacts with oxygen (O_2) to produce CO_2 and H_2O. Write the equation for this reaction.
    e. Balance the reaction (ie. determine the correct stoichiometry).
    f. Calculate how much mass of each reactant is needed and how much mass of each product will be produced if you start with 2.00 moles of propane.

Self Quiz 12A – Answers

Several questions in this quiz are review from chapters 5 and 6.

1. Suppose you have 0.10 mole of uranium (U) atoms.
    a. How many uranium atoms do you have?

 # atoms $= n \times N_A$

$= 0.10 \text{ mol} \times (6.022 \times 10^{23} \text{ atoms/mol})$

$= 6.022 \times 10^{22} \text{ atoms U}$

b. What is the mass in grams of this much uranium?

m $= n \times M$

$= 0.10 \text{ mol} \times 238.0289 \text{ g/mol}$

$= 23.80 \text{ g}$

2. How many aluminum atoms are there in 10.0 grams of aluminum oxide, Al_2O_3?

$M(Al_2O_3)$ $= \{2 \times M(Al)\} + \{3 \times M(O)\}$

$= (2 \times 26.9815 \text{ g/mol}) + (3 \times 15.994 \text{ g/mol})$

$= 101.9612 \text{ g/mol}$

$n(Al_2O_3)$ $= m / M$

$= 10.0 \text{ g} / (101.9612 \text{ g/mol})$

$= 0.098 \text{ mol}$

 # molecules $Al_2O_3 = n \times NA$

$= 0.0098 \text{ mol} \times (6.022 \times 10^{23} \text{ molecules/mol})$

$= 5.91 \times 10^{22} \text{ molecules}$

 # atoms Al $= \text{\# molecules } Al_2O_3 \times (2 \text{ Al} / Al_2O_3)$

$= 5.91 \times 10^{22} \text{ molecules} \times (2 \text{ Al} / Al_2O_3)$

$= 1.18 \times 10^{23} \text{ atoms Al}$

3. If you have 2.00 moles of propane (C_3H_8), a gas used in outdoor grills and industrial torches;
   a. How many molecules of propane do you have?

 # molecules $= n \times N_A$

$= 2.00 \text{ mol} \times (6.022 \times 10^{23} \text{ molecules/mol})$

$= 1.204 \times 10^{24} \text{ molecules } C_3H_8$

b. How many hydrogen atoms do you have?

 # atoms $= \text{\# molecules} \times (8 \text{ H} / C_3H_8)$

$= (1.204 \times 10^{24} \text{ molecules } C_3H_8) \times (8 \text{ H} / C_3H_8)$

$= 9.635 \times 10^{24} \text{ atoms H}$

c. What is the mass in gram of this much propane?

$M(C_3H_8)$ = [3 x M(C)] + [8 x M(H)]
= (3 x 12.0107 g/mol) + (8 x 1.00794 g/mol)
= 44.09562 g/mol

m = n x M
= 2.00 g x 44.09562 g/mol
= 88.19 g

d. Propane reacts with oxygen (O_2) to produce CO_2 and H_2O. Write the equation for this reaction.

$$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$$

e. Balance the reaction (ie. determine the correct stoichiometry).

$$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$$

f. Calculate how much mass of each reactant is needed and how much mass of each product will be produced if you start with 2.00 moles of propane.

$m\ (C_3H_8)$ = 88.2 g (from c)

$n\ (O_2)$ = $n\ (C_3H_8)$ x ($5\ O_2$ / C_3H_8)
= 2 mol C_3H_8 x ($5\ O_2$ / C_3H_8)
= 10 mol O_2

$M\ (O_2)$ = 2 x M(O)
= 2 x 16.0 g/mol
= 32.0 g/mol

$m\ (O_2)$ = n x M
= 10 mol x 32.0 g/mol
= 320 g O_2

$n\ (CO_2)$ = $n\ (C_3H_8)$ x ($3\ O_2$ / C_3H_8)
= 2 mol C_3H_8 x ($3\ O_2$ / C_3H_8)
= 6 mol O_2

$M\ (CO_2)$ = M(C) + [2 x M(O)]
= 12.0 g/mol + (2 x 16.0 g/mol)
= 44.0 g/mol

$m\ (CO_2)$ = n x M = 6 mol x 44.0 g/mol = 264 g CO_2

n (H_2O) ***= n (C_3H_8) x (4 H_2O / C_3H_8)***
= 2 mol C_3H_8 x (4 H_2O / C_3H_8)
= 8 mol O_2

M (H_2O) ***= M(O) + [2 x M(H)]***
= 16.0 g/mol + (2 x 1.01 g/mol)
= 18.02 g/mol

m (H_2O) ***= n x M = 8 mol x 18.02 g/mol = 144 g H_2O***

So 88.2 g propane burned with 320 g oxygen will produce 264 g carbon dioxide and 144 g water.

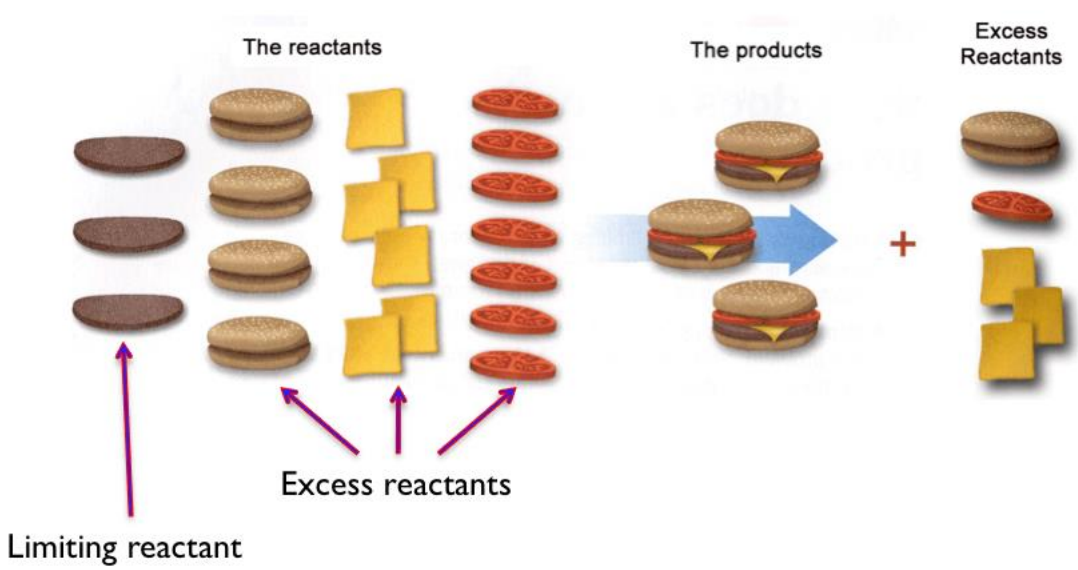
Check answer:

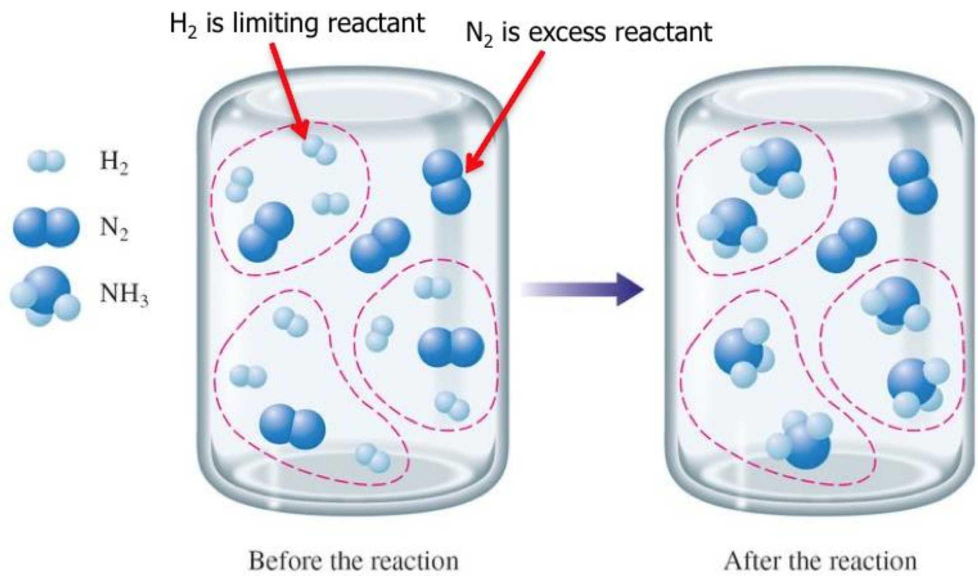
m(reactants) = m(C_3H_8) + m(O_2) = 88.2 g + 320 g = 408 g
m(products) = m(CO_2) + m(H_2O) = 264 g + 144 g = 408 g
m(reactants) = m(products) ✓

12.3 Limiting and Excess Reactants

12.3.1 Definition

- If the reactants are not present in stoichiometric quantities, there will be a limiting reactant(s) and excess reactant(s).
- The reactant(s) that is completely consumed by the reaction is the limiting reactant.
- All other reactants are in excess.
- The limiting reactant determines how much product is formed.
- In industrial processes, the limiting reactant is usually the most expensive reactant to ensure it is used completely and not wasted.





12.3.2 Example

- Returning to the example from 12.2, 2.78 mol of O_2 and 5.55 mol of H_2 are needed to produce 5.55 mol of H_2O.
- If we mix 4 mol of O_2 with 5.55 mol of H_2, only 5.55 mol of H_2O will be produced (even though there is more available O_2) because only 2.78 mol of O_2 can react with 5.55 mol of H_2. The remaining 1.22 mol of O_2 is excess (because it has nothing to react with).
- H_2 is the limiting reactant; O_2 is the excess reactant.
- Note that even though there are more moles of H_2, O_2 is the excess reactant because O_2 and H_2 react in a 1:2 ratio.
- Also note that these calculations must be done using moles, not mass!

12.3.3 Finding the Limiting Reactant

12.3.3.1 The Steps

- To determine which reactant is limiting (if any), choose one of the reactants to begin your analysis.
- Working entirely with moles:
    1. Calculate the number of moles of the first reactant.
    2. Calculate how many moles of the second reactant are required for the reactants to react in stoichiometric amounts.
    3. Calculate the moles of the second reactant actually present in the reaction.
- Alternatively, you can use the moles to find mass of the second reactant (this only changes step 3):

1. Calculate the number of moles of the first reactant.
2. Calculate how many moles of the second reactant are required for the reactants to react in stoichiometric amounts.
3. Use the molar mass of the second reactant to calculate the mass of the second reactant required to react in a stoichiometric amount with the first reactant.

- If the actual amount of the second react is equal to the calculated value then the reactants are present in stoichiometric amounts and there is no limiting reactant.
- If the actual amount of the second react is less than the calculated value then the second reactant is the limiting reactant.
- If the actual amount of the second react is more than the calculated value then the first reactant is the limiting reactant.

12.3.3.2 Example

- Carbon monoxide is produced in the balanced reaction: $2\ C + O_2 \rightarrow 2\ CO$. If you are initially given 50 g of carbon and 60 g of O_2, what is the limiting reactant?
- First, find the number of moles of one of the reactants. We will start with C (you could also start with O_2).
  - n(C) = m / M = 50 g / 12.011 g/mol = 4.16 mol
- Second, find the stoichiometric quantities of O_2 required to react with 4.16 mol C.
  - $n(O_2)$ = n(C) x (O_2 / 2 C) = 4.16 mol x (1/2) = 2.08 mol C
  - $M(O_2)$ = 2 x M(O) = 2 x 15.999 g/mol = 31.998 g/mol
  - $m(O_2)$ = n x M = 2.08 mol x 31.998 g/mol = 66.6 g
- The actual amount of O_2 is less than the calculated amount so O_2 is the limiting reactant (and C is the excess reactant).

12.4 Actual and Theoretical Yields

- The theoretical yield is the amount of product that should be produced.
- It assumes two criteria are met: (1) all of the limiting reactant reacts with a stoichiometric amount of the excess reactant(s), and (2) all of the product is recoverable.
- The actual yield is the amount of product that is actually obtained from a reaction.
- The actual yield is always less than the theoretical yield because of:
  - Non-ideal environmental conditions;
  - Impure reactants;
  - Competing side reactions.

12.5 Percent Yield

- The % yield shows how much product was obtained (actual yield) relative to what could have been in perfect conditions (theoretical yield).

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

- This calculation can be done using any quantity (mass, moles, volume, etc.) as long as you are consistent for the actual and theoretical yield.
- ex. If the reaction from section 12.2 actually produces 90 g of H_2O, what is the % yield?
  - % yield = (90 g / 100 g) x 100% = 90%

Self Quiz 12B

1. Cyclohexane (C_6H_{12}) is made in the following unbalanced chemical reaction: $C_6H_6 + H_2 \rightarrow C_6H_{12}$
   a. Balance the reaction.
   b. To produce 1 mole of C_6H_{12} from this reaction, how many grams of C_6H_6 and H_2 must you combine?
   c. What is the theoretical yield in grams of C_6H_{12}?
   d. Suppose you recover 24.0 g of C_6H_{12}. What is the percent yield?
2. Glucose ($C_6H_{12}O_6$) is burned following unbalanced chemical reaction: $C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$
   a. Balance the reaction.
   b. To produce 6 moles of H_2O from this reaction, how many grams of glucose and O_2 must you combine?
   c. What is the theoretical yield in grams of CO_2?
   d. Suppose you recover 196.0 g of CO_2. What is the percent yield?

Self Quiz 12B – Answers

1. Cyclohexane (C_6H_{12}) is made in the following unbalanced chemical reaction: $C_6H_6 + H_2 \rightarrow C_6H_{12}$
   a. Balance the reaction.

      $C_6H_6 + 3\ H_2 \rightarrow C_6H_{12}$

   b. To produce 1 mole of C_6H_{12} from this reaction, how many grams of C_6H_6 and H_2 must you combine?

      $n(H_2) = n(C_6H_{12})$ x $(3\ H_2 / C_6H_{12})$ = 1 mol x (3/1) = 3 mol

$M(H_2)$ = 2 x M(H) = 2 x 1.01 g/mol = 2.02 g/mol

$m(H_2)$ = n x M = 3 mol x 2.02 g/mol = 6.06 g

$n(C_6H_6) = n(C_6H_{12})$ x (C_6H_6 / C_6H_{12}) = 1 mol x (1/1) = 1 mol

$M(C_6H_6)$ = {6 x M(H)} + {6 x M(C)} = (6 x 1.01 g/mol) + (6 x 12.011 g/mol) = 78.126 g/mol

$m(C_6H_6)$ = n x M = 1 mol x 78.126 g/mol = 78.126 g

c. What is the theoretical yield in grams of C_6H_{12}?

$M(C_6H_{12})$ = {12 x M(H)} + {6 x M(C)} = (12 x 1.01 g/mol) + (6 x 12.011 g/mol) = 84.186 g/mol

$m(C_6H_{12})$ = n x M = 1 mol x 84.186 g/mol = 84.186 g

Check the results using the Law of Conservation of Mass:

m(reactants) = $m(H_2)$ + $m(C_6H_6)$ = 6.06 g + 78.126 g = 84.186 g

m(products) = $m(C_6H_{12})$ = 84.186 g

m(reactants) = m(products) ✔

d. Suppose you recover 24.0 g of C_6H_{12}. What is the percent yield?

% yield = {(actual yield) / (theoretical yield)} x 100%
= (24.0 g / 84.186 g) x 100%
= 28.5 %

2. Glucose ($C_6H_{12}O_6$) is burned following unbalanced chemical reaction:
$C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$
   a. Balance the reaction.

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

   b. To produce 6 moles of H_2O from this reaction, how many grams of glucose and O_2 must you combine?

   $n(O_2) = n(H_2O)$ x (6 O_2 / 6 H_2O) = 6 mol x (6/6) = 6 mol

   $M(O_2)$ = 2 x M(O) = 2 x 15.999 g/mol = 31.998 g/mol

$m(O_2)$ = n x M = 6 mol x 31.998 g/mol = 191.988 g

$n(C_6H_{12}O_6)$ = $n(H_2O)$ x ($C_6H_{12}O_6$ / 6 H_2O) = 6 mol x (1/6) = 1 mol

$M(C_6H_{12}O_6)$ = {12 x M(H)} + {6 x M(C)} + {6 x M(O)} = (12 x 1.01 g/mol) + (6 x 12.011 g/mol) + (6 x 15.999 g/mol) = 180.18 g/mol

$m(C_6H_{12}O_6)$ = n x M = 1 mol x 180.18 g/mol = 180.18 g

c. What is the theoretical yield in grams of CO_2?

$n(CO_2)$ = $n(H_2O)$ x (6 CO_2 / 6 H_2O) = 6 mol x (6/6) = 6 mol

$M(CO_2)$ = M(C) + {2 x M(O)} = 12.011 g/mol + (2 x 15.999 g/mol) = 44.009 g/mol

$m(CO_2)$ = n x M = 6 mol x 44.009 g/mol = 264.054 g

d. Suppose you recover 196.0 g of CO_2. What is the percent yield?

% yield = {(actual yield) / (theoretical yield)} x 100%
= (196.0 g / 264.054 g) x 100%
= 74.2 %

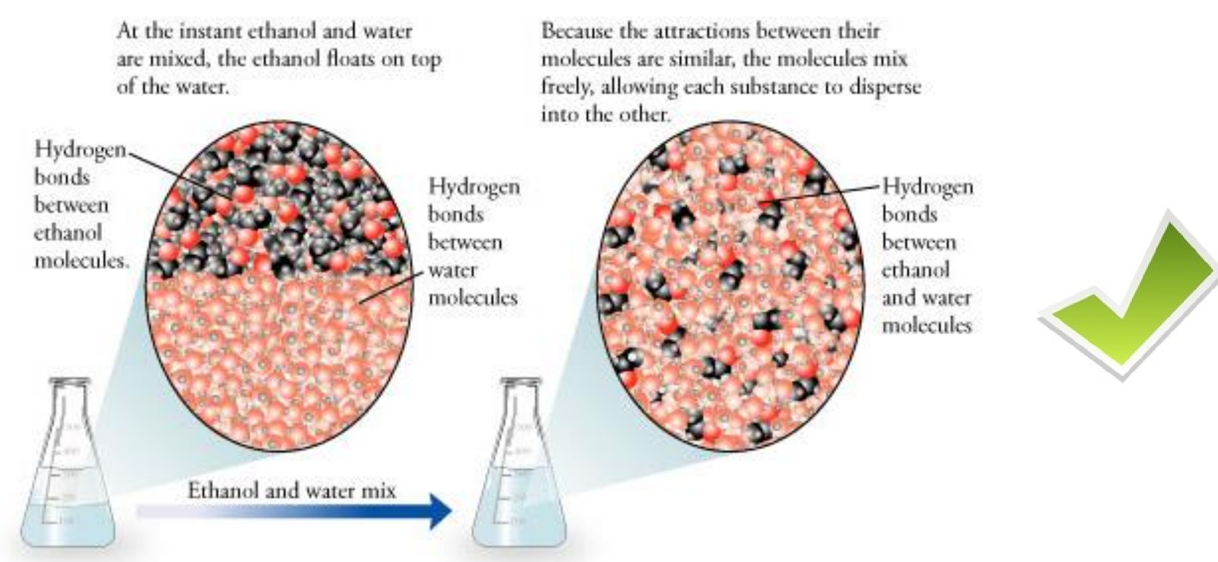
CHAPTER 13 – SOLUTIONS

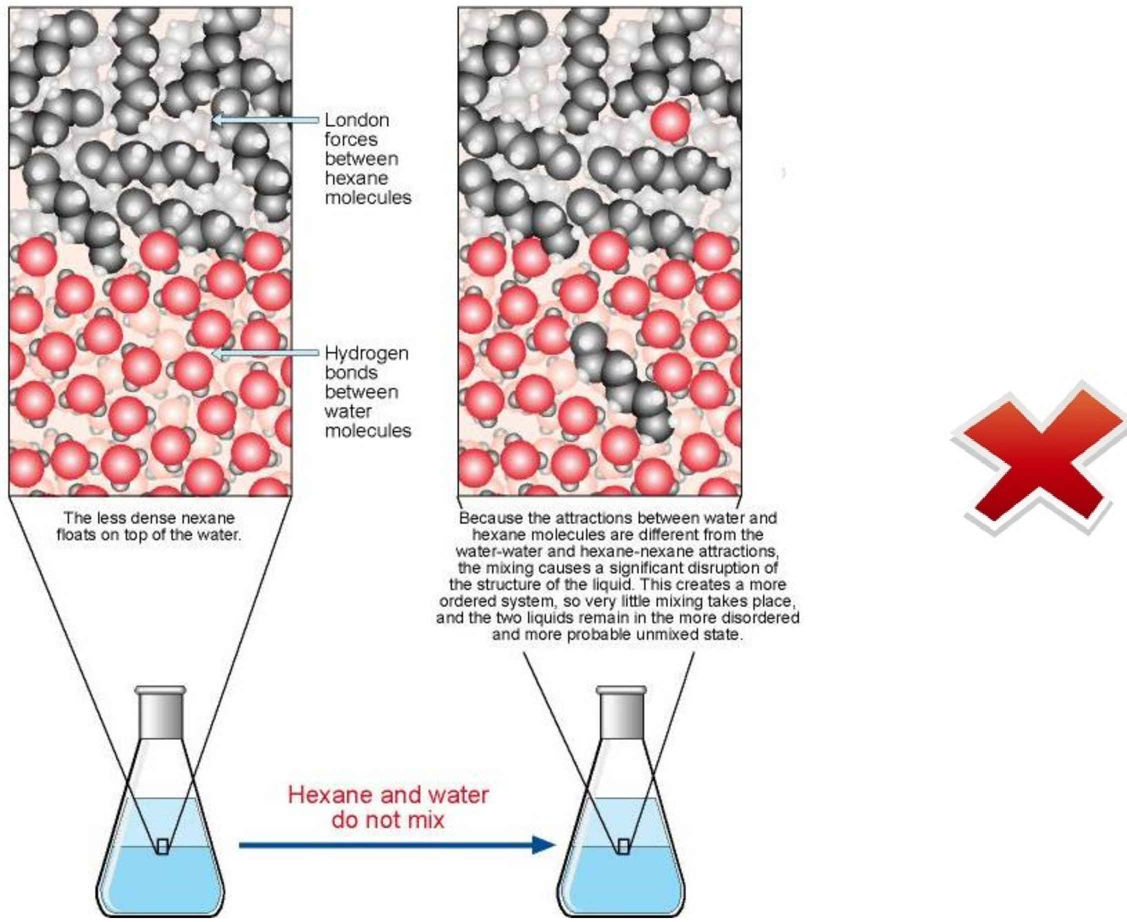
13.1 Definitions

- Solution: A homogeneous mixture of two or more substances (section 1.1). Its composition is uniform throughout the mixture so the concentration of solute in the solvent is uniform.
    - o Solutions can involve solids, liquids and/or gasses.
    - o The most common solutions involve solids dissolved in water. These are called aqueous solutions (section 1.3.3).
- Solvent: The substance present in the greatest amount (ex. in coffee, hot water is the solvent).
- Solute(s): The other substance(s) that are dissolved in the solvent (ex. in coffee, the solutes are the coffee, cream and sugar).
- Concentration: The amount of each solute in the total volume of solvent.

13.2 Forces in Dissolution

- Ionic and polar compounds dissolve in polar solvents because of ion-dipole and dipole-dipole forces.
- Non-polar compounds dissolve in non-polar solvents because of London dispersion forces.
- Ionic and polar compounds will not dissolve in non-polar solvents and visa versa.
    - o ex. oil (non-polar) does not mix with water (polar)
- The general principle is "like dissolves like" (polar/ionic dissolves in polar; non-polar dissolves in non-polar).





13.3 Types of Solutions

13.3.1 Examples

- Solid solutes dissolved in a liquid (ex. salt dissolved in water).
- Liquid solutes dissolved in a liquid (ex. antifreeze additive to a gas tank).
- Solid and liquid solutes dissolved in a liquid (ex. cream (liquid) and sugar (solid) in coffee).
- Gaseous solutes dissolved in a liquid (ex. O_2 and CO_2 in your blood).
- Gaseous solutes dissolved in a gas (ex. O_2, CO_2 and Ar dissolved in N_2 in air).
- Solid solutes dissolved in a solid (ex. zinc dissolved in copper in brass).

13.3.2 Solid Solutions

- Most solid-solid solutions are prepared as liquids and then cooled to solidify.

- If precipitates form as the mixture cools, it becomes a heterogeneous mixture (ex. granite). If all components stay evenly distributed, it is a solid solution (ex. brass).

Granite
(heterogeneous solid mixture)

Brass
(homogenous solid mixture)

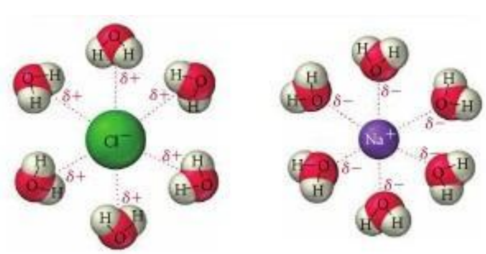
13.3.3 Aqueous Solutions

13.3.3.1 Definition

- Aqueous solutions contain a solute dissolved in water (ex. carbonated water, salt water).
- They are very common because water is a great solvent.
    - Water is often called the 'universal solvent' because of the large number of substances it can dissolve.
- Aqueous solutions (and other solutions with polar solvents) are stabilized by ion-dipole forces.

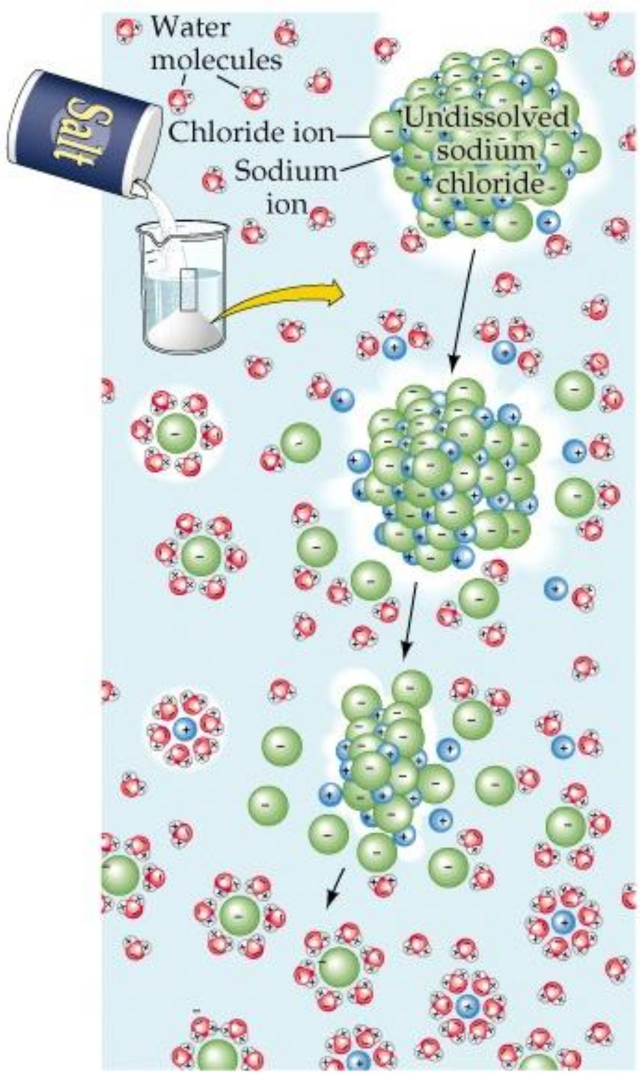
13.3.3.2 Ion-Dipole Forces

- Ion-dipole forces are an intermolecular force that occurs in solutions (not pure substances like the forces discussed in section 9.3).
- Ion-dipole forces are attractive forces between:
    - The partial negatively charged end of a polar molecule and a cation.
    - The partial positively charged end of a polar molecule and an anion.
- Because the ion has a full charge (+1, +2, +3, -1, -2, -3), many solvent molecules with partial charges surround one ion.
- The sum of all the partial charges on the solvent molecules must equal the full charge on the ion to stabilize the solution.



13.3.3.3 Steps to Form an Aqueous Solution (or Any Solution With a Polar Solvent)

1. Solute separation:
   - Solid solute: The solute salt (ex. NaCl) breaks apart into cations and anions (ex. Na^+ and Cl^-). This requires breaking of the strong intramolecular forces in the solute (ionic bonds). Requires energy.
   - Gaseous solute: If you are dissolving a gas in water (ex. CO_2 in water to make carbonated soda), no energy is required for solute separation because gas molecules are already very far apart with negligible intermolecular forces so skip directly to step #2.
2. Solvent separation: The solvent (water) molecules must separate to create spaces for the solute ions. This requires breaking the intermolecular forces (dipole-dipole or hydrogen bonds) in the solvent. Requires energy.
3. Solvation: Individual solute ions enter the spaces in the solvent. Attractive ion-dipole forces form. Releases energy.
   Note: When water is the solvent, this process is referred to as hydration.



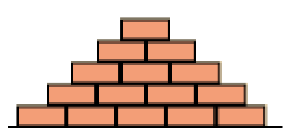
13.3.3.4 Energy of Dissolution

13.3.3.4.1 Enthalpy

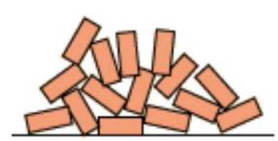
- Enthalpy, *H*, is the total heat content of a system. It accounts for the internal energy and the effects of pressure and volume.
    - o By convention, when enthalpy is added to a system, the change in enthalpy (ΔH) is positive since the system contains more enthalpy after the change than it started with.
    - o When enthalpy is released/removed from a system, ΔH is negative since the system contains less enthalpy after the change than it started with.
- The first two steps in the dissolution process (section 13.3.3.3) require an input of enthalpy; the last step releases enthalpy.
- In order for any dissolution process to occur, the sum of the overall energy changes in the system (section 13.3.3.4.3) must be negative. The change in enthalpy is a large contributor to this.
- ΔH can be negative naturally or it can be encouraged by providing an external source of energy (ex. heat, light, mechanical energy). This is why more salt dissolves in hot water or while stirring.

13.3.3.4.2 Entropy

- Entropy, *S*, is the measure of disorder of a system.
- Systems naturally want to move to a state of higher entropy (higher disorder) because:
    1. Things (objects, reactions, etc.) go in a direction that lowers their overall energy so the end up in a lower energy state (ex. a ball rolling down a hill).
    2. A system moves toward a more probable situation/arrangement. This is why if you throw a pile of bricks into a truck, they will be disordered (a disordered state is more probable than an ordered state).



Order
(improbable; requires energy)

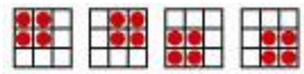


Disorder
(probable; lowest energy state)

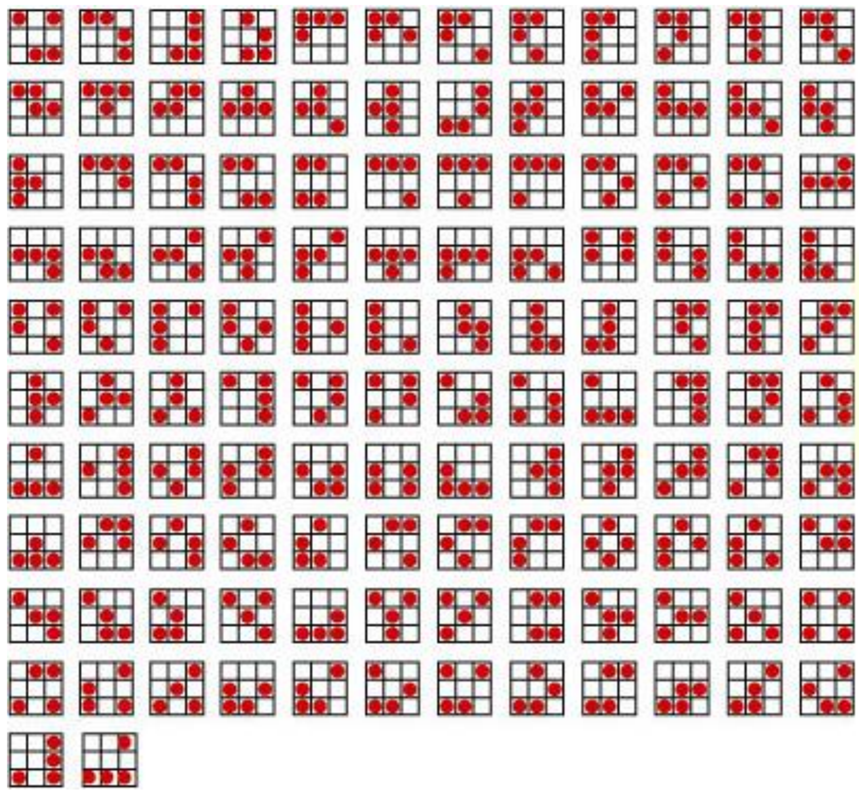
- A solute placed in a solvent will distribute randomly throughout the solution. This increases entropy so entropy favours the dissolution occurring (ie. entropy assists with dissolution).

13.3.3.4.3 Entropy Example

- If you assume that a very simple substance has only four ions and that there are nine potential spots for the ions to fill.
- In a solid there will only be four possible arrangements of the ions (because all ions must be touching in a solid):



- But in a solution, the ions don't have this constant so there are 122 possible arrangements:



- In summary, disordered ions are more probable and have higher entropy than ordered ions:

Solid
(ordered ions)

Solution
(disordered ions, increased entropy)

13.3.3.4.4 Gibbs Free Energy

- We have now talked about two types of energy associated with dissolution: enthalpy, *H,* and entropy, *S*.
- To determine if a solute will dissolve, we need to account for both of these energy changes.
- This is done using Gibbs Free Energy, *G*, which is the overall energy change of a process.

$$\Delta G = \Delta H - T\Delta S$$

- A process (reaction, dissolution, etc.) will occur only if the Gibbs energy change is negative.

Self Quiz 13.3.3.4

1. For a given solute in water, the energy changes are found to be:

   $\Delta H_{solute\ separation}$ = 835 kJ
   $\Delta H_{solvent\ separation}$ = 98 kJ
   $\Delta H_{solvation}$ = –805 kJ
   ΔS = 300 J/K

   If the system temperature is 25 °C, will this solute dissolve in water?

Self Quiz 13.3.3.4 – Answers

1. For a given solute in water, the energy changes are found to be:

   $\Delta H_{solute\ separation}$ = 835 kJ
   $\Delta H_{solvent\ separation}$ = 98 kJ
   $\Delta H_{solvation}$ = –805 kJ
   ΔS = 300 J/K

   If the system temperature is 25 °C, will this solute dissolve in water?

$$\begin{aligned}\Delta H_{total} &= \Delta H_{solute\ separation} + \Delta H_{solvent\ separation} + \Delta H_{solvation}\\ &= 735\text{ kJ} + 98\text{ kJ} - 805\text{ kJ}\\ &= 28\text{ kJ} \times (1000\text{ J / kJ})\\ &= 28{,}000\text{ J}\end{aligned}$$

$$T = 25\text{ °C} + 273\text{ K} = 298\text{ K}$$

$$\begin{aligned}\Delta G &= \Delta H_{total} - T\Delta S\\ &= 28{,}000\text{ J} - \{(298\text{ K})(300\text{ J/K})\}\\ &= -61{,}400\text{ J}\\ &= 61\text{ kJ}\end{aligned}$$

The energy change is less than zero so the solute will dissolve. Note: This calculation is specific for this temperature. At a different temperature, the answer will be different.

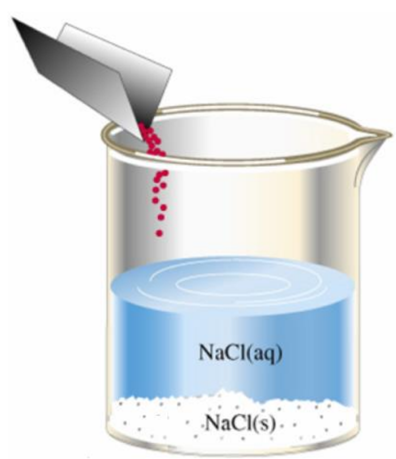
13.4 Solubility

13.4.1 Definition

- Solubility is the ability of a solute to dissolve in a given solvent.
- It is measured using a saturated solution to determine the maximum solubility.
- Units: (grams of solute) / (100 grams of solvent)

13.4.2 Saturated Solution

- A saturated solution contains the maximum amount of solute that can dissolve in a particular volume of solvent at a given temperature and pressure.
- You know you've reached the saturation point when the concentration of the solution remains constant and any additional solute added stays in its native form (ie. does not dissolve).



13.4.3 Effect of Temperature & Pressure

- Solubility is dependant on temperature and pressure.
- In general, solubility increases with increasing temperature.
- Pressure has little effect on solubility if both the solute and the solvent are liquids and/or solids.
- Pressure is very important when gasses are involved (ex. dissolving CO_2 in water to make carbonated soda).

Self Quiz 13.4

1. The solubility of sucrose at 20.0 ˚C is 203.9 g/100 g water. How many grams of sucrose are required to saturate 1 ton of water at this temperature (1 ton = 2000 lbs; 1 lb = 453.6 g)?

2. How many grams of potassium chloride, KCl, will it take to prepare a saturated solution in 500 mL of boiling water? The solubility of KCl at 100 °C is 56.7 g/100 g water. The density of water is 1.000 g/mL.

Self Quiz 13.4 – Answers

1. The solubility of sucrose at 20.0 ˚C is 203.9 g/100 g water. How many grams of sucrose are required to saturate 1 ton of water at this temperature (1 ton = 2000 lbs; 1 lb = 453.6 g)?

   The solubility in units of grams sucrose / ton of H_2O is:

   $$\frac{203.9\ g\ sucrose}{100\ g\ H_2O} \times \frac{453.6\ g}{1\ lb} \times \frac{2000\ lbs}{1\ ton} = 1.85 \times 10^6\ \text{g sucrose}/\text{ton water}$$

   So 1.85×10^6 grams of sucrose are required to saturate 1 ton of water at 20.0 ˚C.

2. How many grams of potassium chloride, KCl, will it take to prepare a saturated solution in 500 mL of boiling water? The solubility of KCl at 100 °C is 56.7 g/100 g water. The density of water is 1.000 g/mL.

   First, we need to convert 500 mL of water to a mass because the solubility of KCl is given in units of mass solute / mass solvent.

   m_{H2O} = 500 mL x (1.000 g/mL) = 500 g

   Now we can use the solubility of KCl to calculate the amount of solute needed to saturate 500 grams of H_2O.

   m_{KCl} = $\frac{56.7\ g\ KCl}{100.0\ g\ H_2O} \times 500\ g\ H_2O = 284\ g\ KCl$

   So 284 grams of KCl is needed to saturate 500 mL of boiling water.

13.5 Concentration

13.5.1 Definition

- Concentration, C, measures the amount of substrate in a solvent.
- It can be calculated from one of two equations, depending on the unit of interest:

$$C = \text{moles} / \text{volume} = n / V$$

$$C = \text{mass} / \text{volume} = m / V$$

- Concentration can be used to calculate how to make a solution starting from a solid solute or a stock solution.
    - A stock solution is a solution of known concentration that is used to make more dilute solutions of the same solute.
- Molarity, M, is the unit of concentration when concentration is calculated with moles (equation 1 above).
    - M = (moles of solute) / (litre of solvent) = mol/L
    - ex. A four molar (4.0 M) solution has a concentration of 4.0 moles of solute in 1.0 L of solvent.

13.5.2 Solving Concentration Problems

1. Calculate the moles of solute required: $n_{solute} = (V_{solution})(C_{solution})$
2. Calculate the mass of solute or volume of stock solution required to obtain this number of moles.
    - To make a solution from a solid solute, convert the number of moles to grams and add this amount to a graduated container.
    - To make a solution from stock, convert the number of moles of solute to a volume of stock solution using the dilution equation (you must know the molarity of the stock solution): $(C_{stock})(V_{stock}) = (C_{dilute})(V_{dilute})$ Add this volume to a graduated container.
3. Add water to the desired volume in the graduated container.

Self Quiz 13.5

1. What mass of solid NaCl do you need to prepare 400.0 mL of a 2.00 M NaCl solution?

2. You have 1L of a 3.00 M stock solution of NaCl and need to prepare 400.0 mL of a 2.00 M solution. Describe how you would do it.

3. How many millilitres of a 2.55 M solution of glucose, $C_6H_{12}O_6$, do you need to obtain 25.0 g of glucose?

4. How many grams of ethanol, C_2H_6O, are there in 200.0 mL of a 2.00 M aqueous ethanol solution?

Self Quiz 13.5 – Answers

1. What mass of solid NaCl do you need to prepare 400.0 mL of a 2.00 M NaCl solution?

   n_{NaCl} = C x V = 2.00 mol/L x 0.400 L = 0.800 mol

   M_{NaCl} = MNa + MCl = 22.9898 g/mol + 35.453 g/mol = 58.4428 g/mol

   m_{NaCl} = n / M = 0.800 mol / 58.4428 g/mol = 46.8 g

2. You have 1L of a 3.00 M stock solution of NaCl and need to prepare 400.0 mL of a 2.00 M solution. Describe how you would do it.

   V_{dilute} = 400.0 mL x (1 L / 1000 mL) = 0.4000 L

   C_{stock} x V_{stock} = C_{dilute} x V_{dilute}

   V_{stock} = (C_{dilute} x V_{dilute}) / C_{stock}

   = (2.00 mol/L x 0.4000 L) / 3.00 mol/L

   = 0.267 L

   = 267 mL

   So to make 400.0 mL of dilute solution with concentration 2.00 M, (i) put 267 mL of your 3.00 M stock solution in a graduated flask then (ii) fill the empty space in the flask to 400.0 mL with water.

3. How many millilitres of a 2.55 M solution of glucose, $C_6H_{12}O_6$, do you need to obtain 25.0 g of glucose?

   $M_{C6H12O6}$ = (6 x M_C) + (12 x M_H) + (6 x M_O)

   = (6 x 12.0107 g/mol) + (12 x 1.00794 g/mol) + (6 x 15.9994 g/mol)

   = 180.1559 g/mol

   $n_{C6H12O6}$ = m/M = 25.0 g / 180.11559 g/mol = 0.13877 mol

   V = n / C = 0.13877 mol / 2.55 mol/L = 0.05442 L = 54.4 mL

So you will need 54.4 mL of a 2.55 M glucose solution to obtain 25.0 grams of glucose.

4. How many grams of ethanol, C_2H_6O, are there in 200.0 mL of a 2.00 M aqueous ethanol solution?

$\boldsymbol{V}$ **= 200.0 mL x (1 L / 1000 mL) = 0.2000 L**

$\boldsymbol{n_{C2H6O}}$ **= C x V = 2.00 mol/L x 0.2000 L = 0.400 mol**

$\boldsymbol{M_{C2H6O}}$ **= (2 x M_C) + (6 x M_H) + M_O**

= (2 x 12.0107 g/mol) + (6 x 1.00794 g/mol) + 15.9994 g/mol

= 46.0684 g/mol

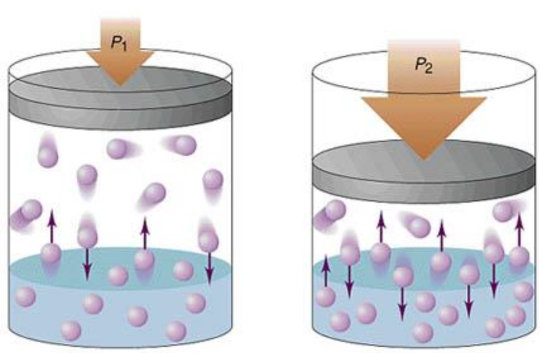
$\boldsymbol{m_{C2H6O}}$ **= n x M = 0.400 mol x 46.0684 g/mol = 18.4 g**

So there are 18.4 grams of ethanol in 200.0 mL of a 2.00 M ethanol solution.

13.6 Applications of Solutions

13.6.1 Soda

- Carbonated soda is a solution of CO_2 dissolved in water.
- It is produced by putting water in a CO_2 atmosphere (so all of the air above the water contains only CO_2).
- The pressure is then increased (using a piston), which forces CO_2 into the water.

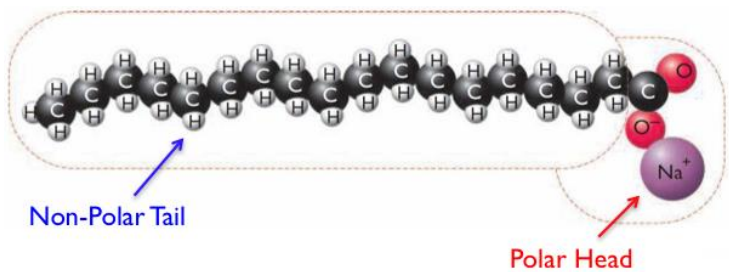


- The soda is then stored under pressure so the CO_2 remains dissolved.
- When the soda is opened, the pressure decreases so the maximum solubility of CO_2 decreases and CO_2 leaves the solution.
- This is why a pop becomes flat over time.

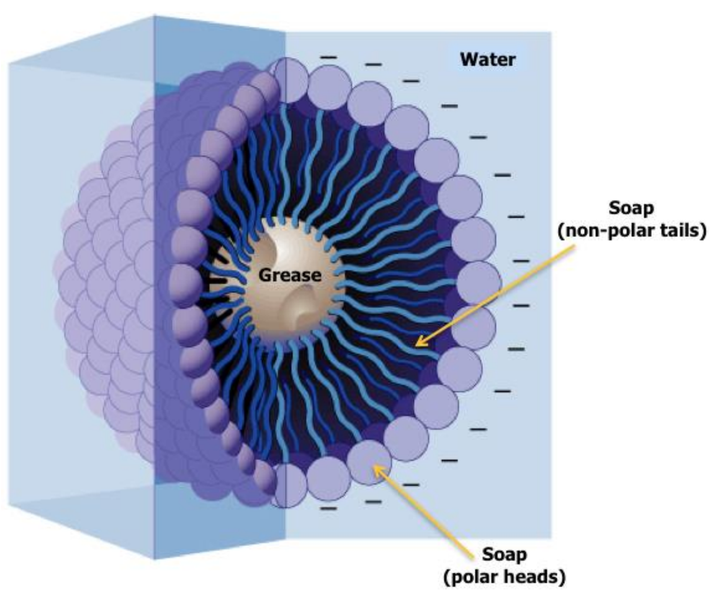
13.6.2 Soap

13.5.6.1 Soap Molecules and Micelles

- Since 'like dissolves like', water is great for removing polar and ionic stains but most tough stains are non-polar (ex. grease) so we need soap.
- Soap molecules contain a long non-polar hydrocarbon tail (hydrocarbon = hydrogen + carbon) connected to a polar head (so they have a polar and a non-polar end).



- The polar head is hydrophilic (water-loving). It will engage in intermolecular attractive forces with water molecules and so enables soap to dissolve in water.
- The non-polar tail cannot form strong intermolecular forces with water so it is hydrophobic (water-hating).
- The soap molecule tails want to interact with other non-polar molecules so when you put soap in water, the soap molecules naturally organize themselves in order to hide their non-polar tails from the water by forming a sphere called a micelle.
- In the micelle, the hydrocarbon tails are not in contact with water and they can engage in London dispersion forces with each other.
- A micelle containing a grease particle is shown below.



13.5.6.2 Mechanism of Action

- When a shirt or pot has a grease stain, you cannot clean it with water because the water molecules repel the grease because water is polar and grease is non-polar.
- But when soap comes into contact with the grease, the non-polar tails of the soap molecules form strong attractive intermolecular forces with the non-polar grease.
- This causes the soap micelles to encapsulate the non-polar grease particles within their non-polar tails.
- The grease is now detached from the shirt/pot so when you rinse your item, the soap and grease are carried away by the water.