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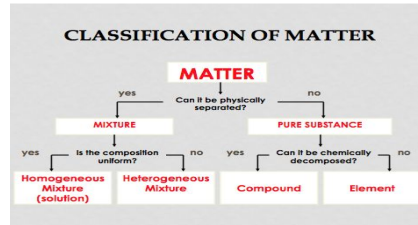
Chemistry Review Notes Units 1-7

Regents Review Unit 1 & 2

Math, Measurement & Matter



Ex. NaCl (aq) Ex. soil Ex. NaCl (s) Ex. Mg or H_2

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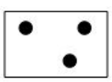
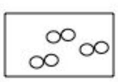
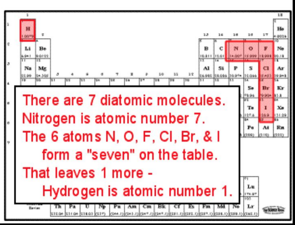
Particle Diagrams: Element



Ex. Mg



Ex. H_2



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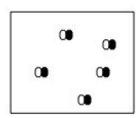
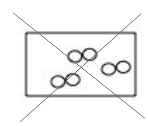
Particle Diagram: compound

Compounds contain two or more different elements chemically combined



Ex. HF



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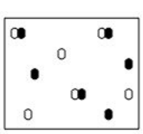
Particle Diagram: Mixture

Mixture contains two or more different substances that are not chemically combined



Mixture of 2 elements and a compound
Ex. Mg & Zn & NaCl

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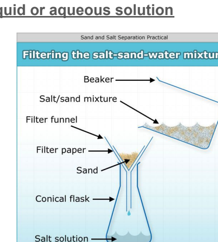
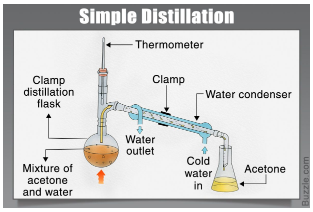
Filtration

- Physically separates **insoluble solid from a liquid or aqueous solution**
  - example: sand from salt water (NaCl aq)

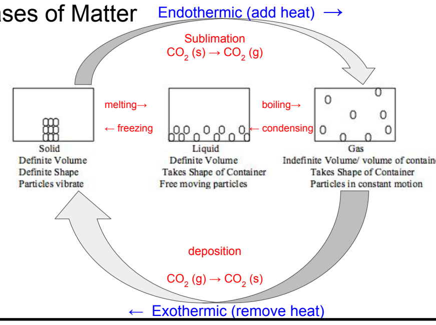


Distillation

- Physically separates two or more **liquids** based on their boiling points



Phases of Matter



Physical Change: a change that <u>does not change</u> the chemical properties of a substance:

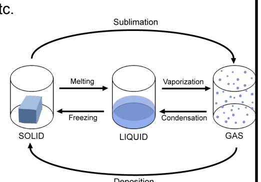
Examples:

Phase changes: boiling, melting, freezing etc.

Cutting

Dissolving

Crystallization



Chemical Change: a reaction where a <u>new substance is formed</u>

Examples:

Burning

Rusting

Reacting

Decomposing

Synthesizing

Corrosion

$N_{2(g)} + 3\,H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$

$CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$

Sig Figs: Precision of measurement

- Start counting at first nonzero digit
  - (ex. 0.000789 has 3 sig fig)
- Any zeros trapped between non zero digits are significant
  - (ex. 3006 has 4 sig fig)
- Zeros trailing after the last non zero digit are only significant if a decimal point is present in the number
  - (ex. 0.0005600 has 4 sig fig but 5600 has only 2 sig fig)

Rounding: Addition/Subtraction

- Round to the least precise place value (least number after the decimal)
  - Ex. 56.790 + 4.3 = 61.09 (round to tenths place) = 61.1
  - Ex. 56 – 45.45 = 10.5 (round to the ones place) = 11

Rounding: Multiplication/Division

- Round answer to least number of sig figs

$(13.91 g/cm^3)(23.3 cm^3) = 324.103 g$

4 SF — 3 SF — 3 SF

324 g

Unit Conversions

- Know how to convert between **milli-** and a **base unit** and vice versa
- Know how to convert between **kilo-** and a **base unit** and vice versa

Table C

Factor	Prefix	Symbol
10^3	kilo-	k
10^{-1}	deci-	d
10^{-2}	centi-	c
10^{-3}	milli-	m
10^{-6}	micro-	μ
10^{-9}	nano-	n
10^{-12}	pico-	p

(base unit*)

*Base units = grams, meters, & liters

•If a unit is getting **larger** (m → km) the number must get **smaller**.

•If the unit gets **smaller** (m→ mm) the number gets **larger**.

Examples:

1. 23.5 mm = 0.0235 m
2. 3.567 L = 3,567 mL
3. 984g = 0.984 kg

Table C

Factor	Prefix	Symbol
10^3	kilo-	k
10^{-1}	deci-	d
10^{-2}	centi-	c
10^{-3}	milli-	m
10^{-6}	micro-	μ
10^{-9}	nano-	n
10^{-12}	pico-	p

(base unit*)

*Base units = grams, meters, & liters

Temperature Conversions

Table T on reference table

Temperature	$K = °C + 273$	K = kelvin °C = degree Celsius

Memorize:

- freezing/melting pt of water = 0℃
- boiling /condensation pt of water = 100℃

Ex. What is the temperature in Celsius of an object that is 150 K?

150K = ℃ + 273

150K -273 = ℃

= -123 C

Density

Located on reference table T

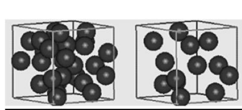
Density	$d = \frac{m}{V}$	d = density m = mass V = volume

Example:

What is the mass of an object with a density of 0.3456 g/cm^3 and a volume of 112.4cm^3?

$\frac{0.3456 g/cm^3}{1} = \frac{x}{112.4 cm^3}$

x = 38.85 g



Example:

- The volume of an aluminum sample is 251 cm^3. What is the mass of the sample?
- *The density of aluminum on Table S is 2.70g/cm^3*

$D = \frac{m}{v}$

$2.70 g/cm^3 = \frac{m}{251 cm^3}$

m = 678 g

Table S
Properties of Selected Elements

Atomic Number	Symbol	Name	First Ionization Energy (kJ/mol)	Electro-negativity	Melting Point (K)	Boiling* Point (K)	Density** (g/cm³)	Atomic Radius (pm)
1	H	hydrogen	1312	2.2	14	20.	0.000082	32
2	He	helium	2372	—	—	4	0.000164	37
3	Li	lithium	520.	1.0	454	1615	0.534	130.
4	Be	beryllium	900.	1.6	1560.	2744	1.85	99
5	B	boron	801	2.0	2348	4273	2.34	84
6	C	carbon	1086	2.6	—	—	—	75
7	N	nitrogen	1402	3.0	63	77	0.001145	71
8	O	oxygen	1314	3.4	54	90.	0.001308	64
9	F	fluorine	1681	4.0	53	85	0.001553	60.
10	Ne	neon	2081	—	24	27	0.000825	62
11	Na	sodium	496	0.9	371	1156	0.97	160.
12	Mg	magnesium	738	1.3	923	1363	1.74	140.
13	Al	aluminum	578	1.6	933	2792	2.70	124

Percent Error

$$\% \text{ error} = \frac{\text{measured value} - \text{accepted value}}{\text{accepted value}} \times 100$$

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Located on Table T of reference Table

A student measures the mass and volume of a sample of copper at room temperature and 101.3 kPa. The mass is 48.9 grams and the volume is 5.00 cubic centimeters. The student calculates the density of the sample. What is the percent error of the student's calculated density?

$D = \frac{m}{v}$

$D = \frac{48.9\text{g}}{5.00\text{cm}^3}$

$D = 9.18\ \text{g/cm}^3$

$\% \text{ error} = \frac{9.18\ \text{g/cm}^3 - 8.96\ \text{g/cm}^3}{8.96\ \text{g/cm}^3} \text{x}100$

% error = 2.46

26	Fe	iron	762	1.8	1811	3134	7.87	124
27	Co	cobalt	760.	1.9	1768	3200.	8.86	118
28	Ni	nickel	737	1.9	1728	3186	8.90	117
29	Cu	copper	745	1.9	1358	2835	8.96	122
30	Zn	zinc	906	1.7	693	1180.	7.134	120.

Calculating with Sci Notation using a Calculator

Ex. $(5.44 \times 10^7\ \text{g}) \div (8.10 \times 10^4\ \text{mol}) =$

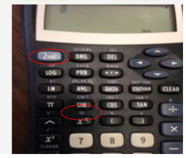
Type on your calculator:

5.44 [2nd / EE] 7 [÷] 8.10 [2nd / EE] 4 [EXE / ENTER]

= 671.60493

= 672 g/mol

$= 6.72 \times 10^2$ g/mol



Regents Review Unit 3

Moles, Stoichiometry, Naming and Formula Writing

Chemical Formulas

Counting Atoms: $4Ca(OH)_2$

Coefficient: tells moles of the compound

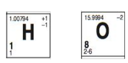
Subscript: tells moles of each type of atom within the compound

Ex. There are 4 calcium atoms, 8 oxygen atoms and 8 hydrogen atoms in $4Ca(OH)_2$

Molar Mass

- The mass of 1 mole of a substance (element or compound)
- AKA the gram formula mass (GFM)

Ex. H_2O = 2(1.0) + 16.0 =18.0g/mol



Calculating Moles

Table T: Moles = $\frac{\text{given mass (g)}}{\text{gfm}}$

gfm is the mass of the element or formula in grams Ex. the gfm of NaCl = 23 + 35.5 = 58.5 g/mol (always round to nearest tenth)

Mole Problem: What is the mass of 0.500 mol of NaCl?

$\frac{0.500 \text{ mol}}{1} = \frac{x}{58.5 \text{ g/mol}}$ x = 29.3 g

Types of Reactions

Synthesis:	**A + B → AB**	*** only one product**
Decomposition:	**BA → A + B**	*** only one reactant**
Single Replacement:	**A + BC → B + AC**	
Double Replacement:	**AB + CD → AD + CB**	

Synthesis: Forming 1 product

Ex. $2N_2 + 3H_2 \rightleftharpoons 2NH_3$

Decomposition: 1 reactant

Ex. $2NH_3 \rightleftharpoons 2N_2 + 3H_2$

Single Replacement: (always redox rx's)

Ex.

$Zn_{(s)} + H_2SO_{4\,(aq)} \rightarrow ZnSO_{4\,(aq)} + H_{2\,(g)}$

Double Replacement: (never redox) (neutralization rx's are also DR)

Ex.

$BaCl_{2\,(aq)} + Na_2SO_{4\,(aq)} \rightarrow BaSO_{4(s)} + 2NaCl_{(aq)}$

$HBr + KOH \rightarrow H_2O + KBr$

Table F
Solubility Guidelines for Aqueous Solutions

Ions That Form *Soluble* Compounds	Exceptions
Group 1 ions (Li^+, Na^+, etc.)	
ammonium (NH_4^+)	
nitrate (NO_3^-)	
acetate ($C_2H_3O_2^-$ or CH_3COO^-)	
hydrogen carbonate (HCO_3^-)	
chlorate (ClO_3^-)	
perchlorate (ClO_4^-)	
halides (Cl^-, Br^-, I^-)	when combined with Ag^+, Pb^{2+}, and Hg_2^{2+}
sulfates (SO_4^{2-})	when combined with Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Pb^{2+}

Ions That Form *Insoluble* Compounds	Exceptions
carbonate (CO_3^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
chromate (CrO_4^{2-})	when combined with Group 1 ions, Ca^{2+}, Mg^{2+}, or ammonium (NH_4^+)
phosphate (PO_4^{3-})	when combined with Group 1 ions or ammonium (NH_4^+)
sulfide (S^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
hydroxide (OH^-)	when combined with Group 1 ions, Ca^{2+}, Ba^{2+}, Sr^{2+}, or ammonium (NH_4^+)

Table J
Activity Series**

Metals (Most → Least)	Nonmetals (Most → Least)
Li	F_2
Rb	Cl_2
K	Br_2
Cs	I_2
Ba	
Sr	
Ca	
Na	
Mg	
Al	
Ti	
Mn	
Zn	
Cr	
Fe	
Co	
Ni	
Sn	
Pb	
**H_2	
Cu	
Ag	
Au	

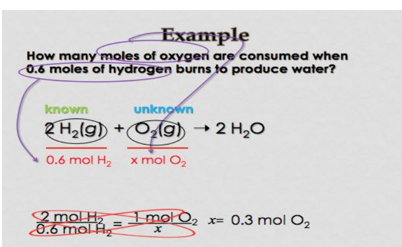
**Activity Series based on hydrogen standard
Note: H_2 is *not* a metal

Mole Ratios

Comparing moles to moles in a question

Ex. How many moles of oxygen are consumed when 0.6 moles of hydrogen burns to produce water?

$2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O$



Conservation of Mass

- Mass of reactants equals mass of product

Ex. If 50 grams of H_2 reacts with oxygen to produce 150 grams of H_2O how much O_2 reacted?

$2H_2 + O_2 \rightarrow 2H_2O$

50g + xg = 150g

50g + **100g** = 150g

Balancing Reactions

Ex.

2Al + **3**$CuCl_2$ → **3**Cu + **2**$AlCl_3$

**Keep polyatomics together...no need to balance each element individually

$BaCl_2$ + **2**$AgNO_3$ → $Ba(NO_3)_2$ + **2**AgCl

Molecular to Empirical Formula

Molecular formula = C_4H_{10}

Divide by 2 *(greatest common factor)*

C_2H_5

Empirical Formula to Molecular Formula

•The empirical formula for ethylene is CH_2. Find the molecular formula if the molecular mass is 28.1 g/mol?

empirical mass CH_2 =
(1 C X 12.0 g/mol) + (2 H X 1.0 g/mol) = 14.0 g/mol

$$\frac{28.1\ g/mol}{14.0\ g/mol} = 2.00$$

$$(CH_2) \times 2 \Rightarrow C_2H_4$$

Percent Composition

Formula on Table T

$$\%\ composition = \frac{mass\ of\ part}{mass\ of\ whole} \times 100$$

Ex.

What is the percentage by mass of carbon in CO_2?

C=12.0g x 1 = 12.0g
O=16.0g x 2 = 32.0g
+ = 44.0g/mol

$$\frac{12.0g}{44.0g} \times 100$$

= 27.3%

% Composition of a hydrate

$$\%\ composition = \frac{mass\ of\ part}{mass\ of\ whole} \times 100$$

What is the percent by mass of water in $BaCl_2 \bullet 2H_2O$?

Ba= 137.3 x 1=137.3g
Cl = 35.5 x 2 = 71.0g
H = 1.0 x 4 = 4.0g
O = 16.0 x 2 = 32.0g
+ =244.3g/mol

$$\frac{36.0g}{244.3g} \times 100$$

= 14.7%

Naming Binary Compounds

•Write the **complete name of the first element.**

•The **second element** should then be named using the **first syllable with the ending**
"-ide."

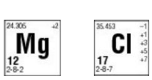
$NaCl$ sodium chloride
KI potassium iodide
$MgCl_2$ magnesium chloride
Ca_3N_2 calcium nitride

Writing chemical formulas for binary compounds

Use the drop and swap (drop the + and - signs and swap the numbers as subscripts)

Remember the whole compound must equal zero

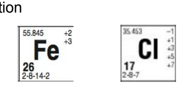
$$Mg^{+2}\ Cl^{-1} \longrightarrow MgCl_2$$



Naming binary compounds w/ multiple charges

If the first element has more than one charge listed (**Some transition metals** and **all nonmetals compounds**)

1. Find the charge of the anion *(the most electronegative element which is usually listed last in the formula)* It will always be a negative charge
2. Set the whole thing equal to zero *(since compounds are always neutral)*
3. Then solve for the the charge of the positive cation



Example: $Fe^{+3}Cl^{-1}_3$

Iron (III) chloride

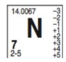
Writing chemical formulas with multiple charges

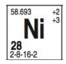
The roman numeral is the charge of the first substance

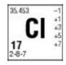
1. Write the charges above the element symbols then drop and swap

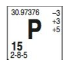
Ex. nickel (II) chloride $\overset{+2}{Ni}\overset{-1}{Cl}_2$

nitrogen (II) phosphide $\overset{+2}{N}_3\overset{-3}{P}_2$









Naming Tertiary compounds

More than two elements (capital letters)

Use table E

1. Name the same as other compounds except use name listed on table E for that substance

Ex. $CaCO_3$

Calcium Carbonate

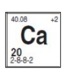
Formula	Name	Formula	Name
H_3O^+	hydronium	CrO_4^{2-}	chromate
Hg_2^{2+}	mercury(I)	$Cr_2O_7^{2-}$	dichromate
NH_4^+	ammonium	MnO_4^-	permanganate
$C_2H_3O_2^-$ / CH_3COO^-	acetate	NO_2^-	nitrite
CN^-	cyanide	NO_3^-	nitrate
CO_3^{2-}	carbonate	O_2^{2-}	peroxide
HCO_3^-	hydrogen carbonate	OH^-	hydroxide
$C_2O_4^{2-}$	oxalate	PO_4^{3-}	phosphate
ClO^-	hypochlorite	SCN^-	thiocyanate
ClO_2^-	chlorite	SO_3^{2-}	sulfite
ClO_3^-	chlorate	SO_4^{2-}	sulfate
ClO_4^-	perchlorate	HSO_4^-	hydrogen sulfate
		$S_2O_3^{2-}$	thiosulfate

Writing Chemical Formulas for Tertiary Compounds

- Same as Binary Compounds
  - **Never** change the polyatomic ion subscripts
  - If more than 1 polyatomic ion put in parenthesis followed by subscript

$Ca^{+2}\ PO_4^{-3} \longrightarrow Ca_3(PO_4)_2$

Formula	Name	Formula	Name
H_3O^+	hydronium	CrO_4^{2-}	chromate
Hg_2^{2+}	mercury(I)	$Cr_2O_7^{2-}$	dichromate
NH_4^+	ammonium	MnO_4^-	permanganate
$C_2H_3O_2^-$ / CH_3COO^-	acetate	NO_2^-	nitrite
CN^-	cyanide	NO_3^-	nitrate
CO_3^{2-}	carbonate	O_2^{2-}	peroxide
HCO_3^-	hydrogen carbonate	OH^-	hydroxide
$C_2O_4^{2-}$	oxalate	PO_4^{3-}	phosphate
ClO^-	hypochlorite	SCN^-	thiocyanate
ClO_2^-	chlorite	SO_3^{2-}	sulfite
ClO_3^-	chlorate	SO_4^{2-}	sulfate
ClO_4^-	perchlorate	HSO_4^-	hydrogen sulfate
		$S_2O_3^{2-}$	thiosulfate

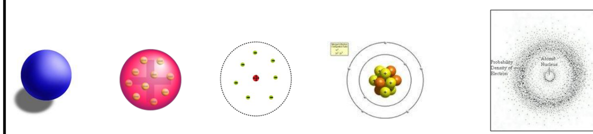


Regents Review Unit 4

Atomics

Atomic Theory

1. Atom is a solid sphere
2. Atom consists of a uniform positive charge with electrons embedded in it
3. Atom consists of a small positive nucleus and is mostly empty space
4. Electrons orbit around the nucleus in energy levels
5. There is a high probability of finding electrons in orbitals

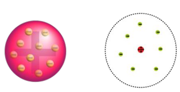


Thomson vs Rutherford

Thomson discovered electrons

Similarities: Both agreed that there were negative electrons and that the atom was neutral

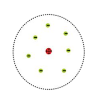
Differences: Rutherford said protons (+ charges) were in nucleus and electrons were outside the nucleus

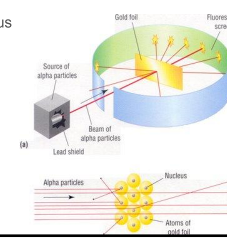


Rutherford's Gold Foil Experiment

2 conclusions:

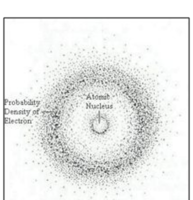
1. Atom consists of a small dense positive nucleus
2. Most of the atom is empty space



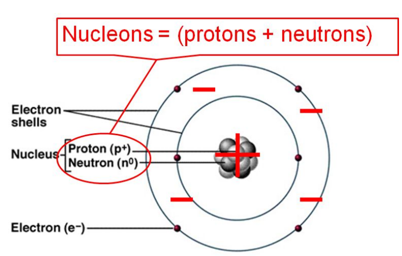


Current Model: Wave mechanical model

There is a high probability of finding electrons in regions called orbitals



Parts of the atom: Subatomic Particles



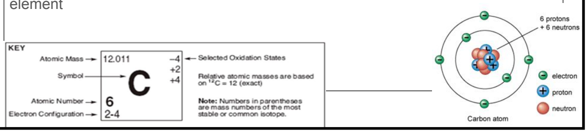
Calculating # of subatomic particles

Atomic number = the # of protons

Protons = electrons in a neutral atom

Neutrons = mass # - protons

Atomic mass: the weighted average of all the naturally occurring isotopes of that element

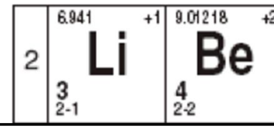


Example: Li

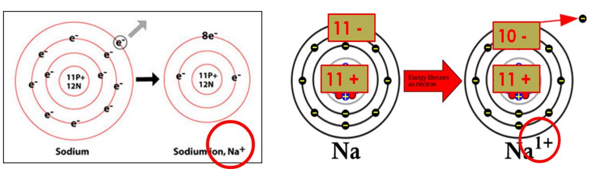
Number of protons = 3

Number of electrons = 3

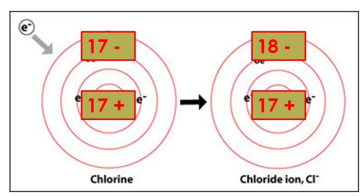
Number of neutrons: 7-3 = 4



Ions: cations (+)



Ions: anions (-)



Isotopes

- **Same** number of <u>protons</u> **different** number of <u>neutrons</u>

- Ex: Determine the average atomic mass of Boron using the information below

Boron-10	19.78%	10.013 amu
Boron-11	80.22%	11.009 amu

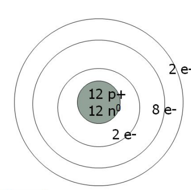
Multiply the mass of each isotope by its percent abundance then divide by 100

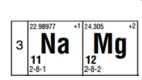
<u>(19.78 x 10.013) + (80.22 x 11.009)</u>

100

=10.812 amu

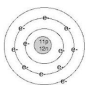
Bohr Diagrams (they are Bohr-ring)





Excited vs ground state

Na (ground state)	Na (possible excited state)	Na (another possible excited state)	Na (another possible excited state)
2-8-1	2-7-2	2-6-3	2-5-4



*****Remember when in excited state the total # of electrons DOES NOT change**

How to determine if it is in an excited state

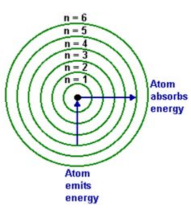
- Add up total # of electrons in configuration
- Determine element
- If it **matches** element configuration on periodic table **= GROUND**
- If it **doesn't match = EXCITED**

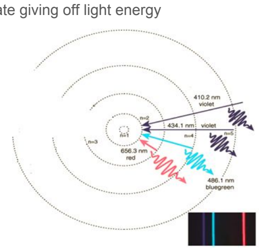
Example: Identify the electron configuration as being ground state or excited state: 2-6-1

18.998403 −1
F
9
2-7

How is light produced

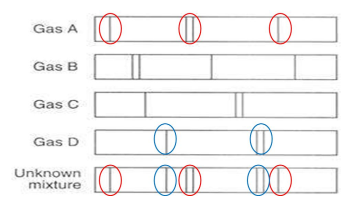
Electrons go from excited state to ground state giving off light energy





Example

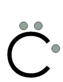
What gases are present in the unknown mixture? Gases A and D



Lewis dot diagrams: atoms

Locate the last number (valence number) in the electron configuration.

Each dot represents 1 valence electron (no more than 2 dots per side of symbol)

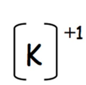


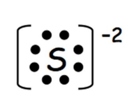
12.0111 −4 +2 +4
C
6
2-4

Lewis dot diagrams: ions

- Metals lose electrons to form positive ions
  - No dots in diagram, use brackets and charge of ion
- Non metals gain electrons to form negative ions
  - 8 dots in diagram, use brackets and charge

Example Draw the lewis diagram for K^{+1} and S^{-2}





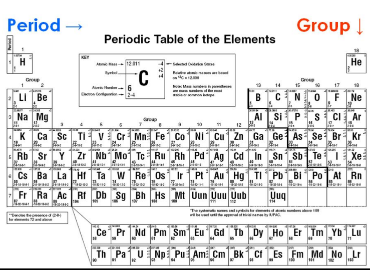
Regents Review Unit 5

Periodic Table

Periodic Table Organization

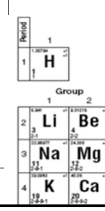
Mendeleev: organized my mass

Mosley: current table organized by atomic number (number of protons)



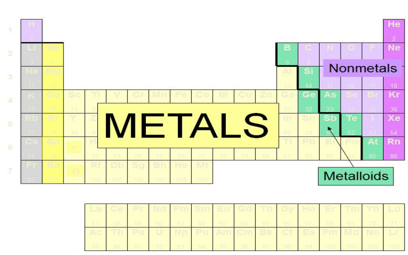
Periods: Same number of electron shells

Groups: Same number of valence electrons...have similar properties



Example: Which two elements have similar chemical properties and why? Na, K, Li, Be

Na and K because they are in the same group so they have the same number of valence electrons



Group 1 Alkali Metals: most reactive metal Fr

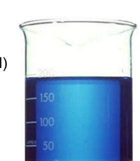
Group 2: Alkaline earth metals

Groups 3-12: Transition metals: tend to form colored solutions

Group 17: Halogens: most reactive nonmetal F

Group 18: Noble Gases: inert (unreactive due to full valence shell)

- Exception is He which has 2 valence electrons
- monoatomic

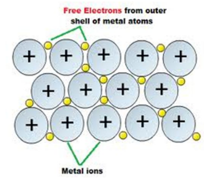


Properties of Metals

- **Malleable** (can be hammered or rolled into thin sheets)
- **Ductile** (can be drawn into a wire)
- **Excellent conductors** of heat and electricity
- **Luster** (shiny)
- Lose electrons to form **cations**
- Solid @ STP (except Hg)

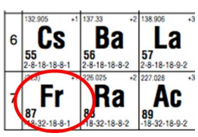
Sea of mobile electrons

Why metals conduct electricity in the solid state



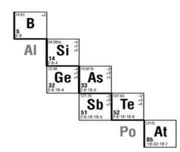
Metallic Character

- How much “like a metal” an element is
- **Franicium** is most metallic
- Closer to Fr more metallic…further from Fr least metallic



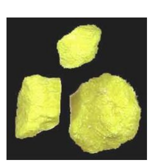
Properties of Metaloids

- **Semiconductors** (Good/moderate conductor)
- **Luster** (like metals) and **Brittle** (like nonmetals)
- Used for making computer microchips



Properties of Nonmetals

- **Poor conductors** of heat and electricity
- **Brittle** (shatter when struck)
- **Dull**
- Tend to gain electrons to form **anions**



	Metals	Metalloids	Nonmetals
Phys. prop.	• malleable • ductile • shiny • excellent conductors (heat, electricity... MOBILE e-'s)	in-between	• brittle • dull • poor conductors (heat, electricity)
Chem. prop.	• lose e-'s • form + ions • low E.N. • low I.E.	B, Si, Ge, As, Sb, Te	• gain e-'s • form - ions • high E.N. • high I.E.

Period Trends

- Atomic radius
- Electronegativity
- Ionizationation energy

Table S
Properties of Selected Elements

Atomic Number	Symbol	Name	First Ionization Energy (kJ/mol)	Electro-negativity	Melting Point (K)	Boiling* Point (K)	Density** (g/cm³)	Atomic Radius (pm)

For all trends look up two elements in either the group or period and locate the values. If they are increasing the trend is increasing and vice versa.

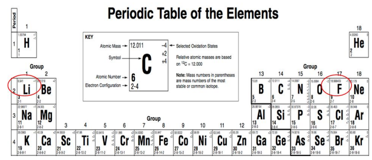
The reasons:

- Going across a period is due to increased proton pull
- Going down a group is due to more electron shells

Atomic Radius: size of atom

Example: what is the trend in atomic radius across a period?

Decreasing due to increased nuclear pull



Atomic Number	Symbol	Name	Atomic Radius (pm)
1	H	hydrogen	32
2	He	helium	37
3	Li	lithium	130.
4	Be	beryllium	99
5	B	boron	84
6	C	carbon	75
7	N	nitrogen	71
8	O	oxygen	64
9	F	fluorine	60.
10	Ne	neon	62
11	Na	sodium	160.
12	Mg	magnesium	140.
13	Al	aluminum	124
14	Si	silicon	114
15	P	phosphorus (	109
16	S	sulfur (mono	104
17	Cl	chlorine	100.
18	Ar	argon	101
19	K	potassium	200.
20	Ca	calcium	174
21	Sc	scandium	159
22	Ti	titanium	148
23	V	vanadium	144
24	Cr	chromium	130.
25	Mn	manganese	129

Electronegativity

•Measure of the **ATTRACTION** for electrons

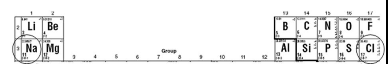
•**FLUORINE** most electronegative (4.0)

•The closer an atom is to Fluorine the **HIGHER** the electronegativity

•Scale of 0 - 4

Example: What is the trend in electronegativity going across a period?

Increasing due to more proton pull



Atomic Number	Symbol	Name	First Ionization Energy (kJ/mol)	Electro-negativity	Melting Point (K)	Boiling* Point (K)	Density** (g/cm³)	Atomic Radius (pm)
1	H	hydrogen	1312	2.2	14	20.	0.000082	32
2	He	helium	2372	—	—	4	0.000164	37
3	Li	lithium	520.	1.0	454	1615	0.534	130.
4	Be	beryllium	900.	1.6	1560.	2744	1.85	99
5	B	boron	801	2.0	2348	4273	2.34	84
6	C	carbon	1086	2.6	—	—	—	75
7	N	nitrogen	1402	3.0	63	77	0.001145	71
8	O	oxygen	1314	3.4	54	90.	0.001308	64
9	F	fluorine	1681	4.0	53	85	0.001553	60.
10	Ne	neon	2081	—	24	27	0.000825	62
11	Na	sodium	496	0.9	371	1156	0.97	160.
12	Mg	magnesium	738	1.3	923	1363	1.74	140.
13	Al	aluminum	578	1.6	933	2792	2.70	124
14	Si	silicon	787	1.9	1687	3538	2.3296	114
15	P	phosphorus (white)	1012	2.2	317	554	1.823	109
16	S	sulfur (monoclinic)	1000.	2.6	388	718	2.00	104
17	Cl	chlorine	1251	3.2	172	239	0.002898	100.

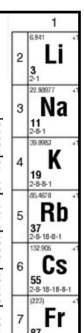
Ionization Energy

How much energy is required to remove a valence electron

Ex. What is the trend in ionization energy going down a group?

Decreasing due to more electron shells

10	Ne	neon	2081	—
11	Na	sodium	496	0.9
12	Mg	magnesium	738	1.3
13	Al	aluminum	578	1.6
14	Si	silicon	787	1.9
15	P	phosphorus	1012	2.2
16	S	sulfur	1000	2.6
17	Cl	chlorine	1251	3.2
18	Ar	argon	1521	—
19	K	potassium	419	0.8
20	Ca	calcium	590	1.0
21	Sc	scandium	633	1.4
22	Ti	titanium	659	1.5
23	V	vanadium	651	1.6
24	Cr	chromium	653	1.7
25	Mn	manganese	717	1.6
26	Fe	iron	762	1.8
27	Co	cobalt	760	1.9
28	Ni	nickel	737	1.9
29	Cu	copper	745	1.9
30	Zn	zinc	906	1.7
31	Ga	gallium	579	1.8
32	Ge	germanium	762	2.0
33	As	arsenic	944	2.2
34	Se	selenium	941	2.6
35	Br	bromine	1140	3.0
36	Kr	krypton	1351	—
37	Rb	rubidium	403	0.8
38	Sr	strontium	549	1.0
39	Y	yttrium	600	1.2
40	Zr	zirconium	640	1.3

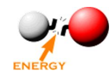


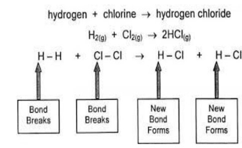
Regents Review Unit 6

Bonding

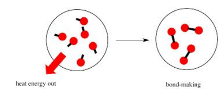
Bond breaking and formation (BARF)

- Breaking Bonds takes (absorbs) energy





- Making a bond releases energy



Octet Rule

- Atoms will bond to obtain a lower energy state
- This is done by obtaining a full valence shell
  - 8 valence electrons (stable octet)
  - 2 valence electrons for Hydrogen (stable duet)

Types of Bonds

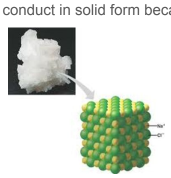
Ionic	Covalent	Metallic
Metal and nonmetal	Nonmetal and nonmetal	Metals
Transfer of electrons	Sharing of electrons	Sea of mobile electrons
-hard - good conductor as liquid or (aq) ONLY because of **free moving ions** - high melting and boiling point	- **S**oft - **P**oor conductors of heat and electricity because no charged mobile particles - **L**ow melting and boiling point	hard - always good conductors, because of **sea of mobile valence electrons** - high melting point - malleable (made into sheets) - ductile (made into wires)

Electrolytes

Substances that conduct electricity when dissolved in water

Ex. NaCl (aq)

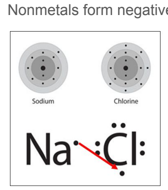
Doesn't conduct in solid form because there are not free moving ions.

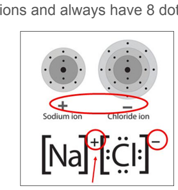


Lewis Dot Diagrams: Ionic

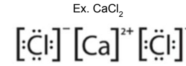
Metals form positive ions and always have no dots in the brackets

Nonmetals form negative ions and always have 8 dots in the brackets





Ex. $CaCl_2$



Lewis Dot Diagrams: Covalent

H_2 Cl_2 CH_4 H_2O NH_3

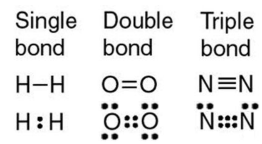
CO_2 O_2 N_2

---

2 electrons (1 Pair) shared in a single bond

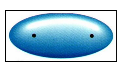
4 electrons (2 Pair) shared in a single bond

6 electrons (3 Pair) shared in a single bond

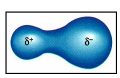


Types of Covalent/ Bonds: (Bond Polarity)

- Nonpolar covalent bond: Equal sharing of electrons
  - Electronegativity difference (E.N.D.) 0-0.4
  - Ex. H_2 CH_4

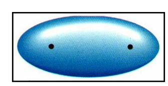


- Polar covalent bond: Unequal sharing of electrons
  - Electronegativity difference (E.N.D.) greater than 0.4
  - Ex. H_2O HF

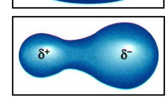


Comparing Ionic to Covalent

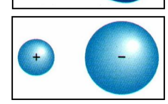
Nonpolar covalent

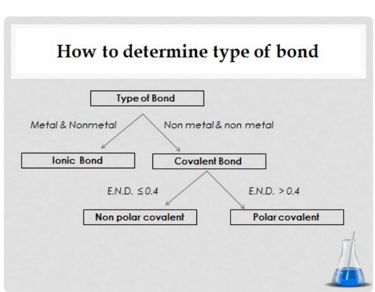


polar covalent



Ionic



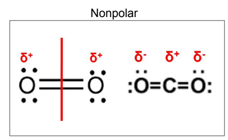


Types of Molecules (Molecular Polarity)

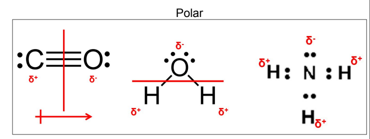
- Nonpolar molecules have a symmetrical distribution of charge
- Polar molecules have an asymmetrical distribution of charge

******Remember SNAP (symmetrical non polar, asymmetrical polar

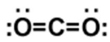
Nonpolar



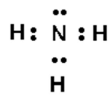
Polar



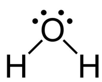
Molecular Geometry



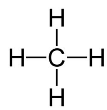
Linear



pyramidal



Bent



tetrahedral

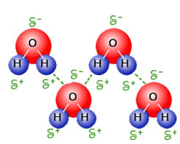
Like Dissolves Like

Polar molecules dissolve in other polar molecules

Nonpolar molecules dissolve in nonpolar molecules

Intermolecular Forces of Attraction:

- Weak Bond between molecules
- The strongest IMF is **Hydrogen bonding**
  - Occurs between H and FON
  - Remember H bonding is FON
- The stronger the IMF the higher the melting and Boiling Pt
  - Responsible for thee high melting and boiling pt of H_2O, NH_3 and HF



Regents Review Unit 7

Heat

Heat

- **Heat** flows from hot to cold
- **Average kinetic energy** = temperature

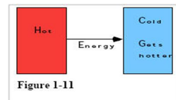


Figure 1-11

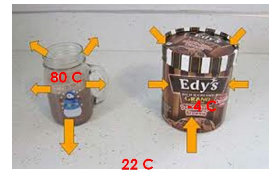


Table I
Heats of Reaction at 101.3 kPa and 298 K

Reaction	ΔH (kJ)*
$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(\ell)$	−890.4

*The ΔH values are based on molar quantities represented in the equations. A minus sign indicates an exothermic reaction.

- **Endothermic:** absorb heat; heat on left side of reaction (reactant side) +Δ**H**

$2C + H_2 + 52.4kJ \rightarrow C_2H_4$

$2C(s) + 2H_2(g) \longrightarrow C_2H_4(g)$	+52.4

- **Exothermic:** release heat; heat on right side of reaction (product side) -Δ**H**

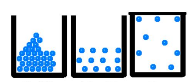
$C + O_2 \rightarrow CO_2 + 393.5kJ$

$C(s) + O_2(g) \longrightarrow CO_2(g)$	−393.5

•Solid – definite shape and volume. STRONGEST force of attraction between particles

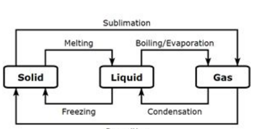
•Liquids – definite volume and takes shape of container. MODERATE force of attraction between particles

•Gases – not definite shape or volume...fills a container. WEAKEST force of attraction between particles



Solid Liquid Gas

Endothermic →



← Exothermic

Specific Heat

- Amount of heat needed to raise the temperature of 1 gram of a substance 1 degree Celsius
  - Specific Heat of water listed on table B
  - The higher the specific heat, the longer it takes to heat up and cool down
  - Metals have low specific heat, water has a high specific heat

Table B
Physical Constants for Water

Heat of Fusion	334 J/g
Heat of Vaporization	2260 J/g
Specific Heat Capacity of $H_2O(\ell)$	4.18 J/g•K

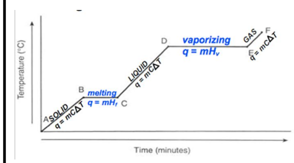
Heat of fusion: amount of heat required to turn a solid into a liquid

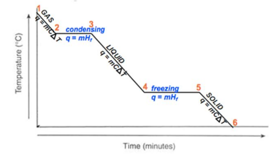
Heat of vaporization: amount of heat required to turn a liquid into a gas

Table B
Physical Constants for Water

Heat of Fusion	334 J/g
Heat of Vaporization	2260 J/g
Specific Heat Capacity of $H_2O(\ell)$	4.18 J/g•K

Heating vs cooling curve

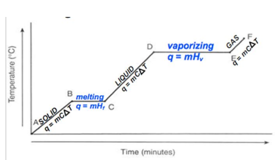




Kinetic energy increases during Solid, liquid, gas phases. Potential stays the same.
Potential energy increases during phase change. Kinetic stays the same

Calculating heat

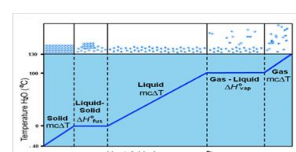
- In the question, if there is a change in temperature use q=mcΔT
- If no change in temperature:
  - At 0°C or freezing/melting use $q=mH_f$
  - At 100°C or boiling/vaporizing/condensing use $q=mH_v$



Calculating heat

How much heat is needed to raise the temperature of 500. g of water from 5°C to 20°C ?

- **Step1:** List the known variables
  m = 500. g
  C = 4.18 J/g°C (from Table B)
  ΔT = 15°C
- **Step 2:** Determine the product
  q = mCΔT
  q = (500. g)(4.18 J/g°C)(15°C)
  q = 31,500 J = 32,000 J

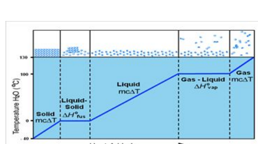


How many joules does it take to melt a 16.00 gram sample of water at 0 C?

$q= mH_f$

q = (16.00g)(334J/g)

q= 5344 J



Calculating heat

How many joules does it take to boil a 250. gram sample of water at 100 C?

$q= mH_v$

q = (250.g)(2,260J/g)

q= 565,000 J