February 12 2020

Scientific Notation

FOR LARGE NUMBERS

124000000

↳ 1.24000000

↳ 1.24×10^{8}

FOR SMALL NUMBERS

.000000001 24

↳ 000000001.24

↳ 1.24×10^{-9}

1. 10,000
↳ 1×10^{4} ✓
2. .0001
↳ 1×10^{-4} ✓
3. 10,000,000,000
↳ 1×10^{10} ✓
4. 50,000
↳ 5×10^{4} ✓
5. 2,000,000,000
↳ 2×10^{9} ✓
6. .000004
↳ 4×10^{-6} ✓
7. .0003
↳ 3×10^{-4}
8. 790,000
↳ 7.9×10^{5} ✓
9. 26,000,000,000,000
↳ 2.6×10^{13}
10. .000000045
↳ 4.5×10^{-8}
11. 48,000,000
↳ 4.8×10^{7}
12. .0000000000000000013
↳ 1.3×10^{-18}
13. .0175
↳ 1.75×10^{-2}
14. .0000462
↳ 4.62×10^{-5}

CALCULATOR
X10– or EXP or EE or ee

Significant Digits

February 12 2020

certain certian uncertain
7.92
c c u

Minor graduation → small units, which some measuring divies don't have so a gues without measuring these units → guess is an uncertian value → last digit

1. All non-zero digits are significant.
2. zeros between significant figure is significant.
3. zeros that are just placeholders are not significant
4. Final zeros after a significant figure, decimal are significant.

Certian vs significant

Certian vs uncertian → is the last digit
any number before the last digit.
significant vs unsignificant → a zero at the end of a number
↳ an uncertian value is significant

c c c c u
28760 4 significant.
s s s s u

c c c c u
28064 5 significant.
s s s s s

c c c c u
.00218 3 significant
u u s s s

c c c c u
218.00 5 significant
s s s s s

1. 307.0 cm 4 significant
2. 61 m/s 2 significant
3. 0.03 m 1 significant
4. .5060 km 4 significant
5. 3.00×10^8 m/s 3 significant

Counts are infinitly significant → Person, Protons ect.

Multiplying & Dividing.

GIVE THE ANSWER IN THE LOWEST NUMBER OF SIGNIFICANT FIGURES.

19.57(3.1)
= 60.667
= 61

$\frac{268.46}{5.00001}$ = 53.69189262 = 53.692

of significant digits the final answer is the same amnt as the lower of the numbers.

ONLY ROUND FINAL ANSWERS

Energy & Matter

February 12 2020

TEMPERATURE

Celsius & Kelvin

Celcius

Boiling point: 100c

Freezing point: 0c

Kelvin

Boiling point: 373K

Freezing point: 273K

0 celsius is 273 kelvin

Temperature is Average Kenetic Energy

*Hot is fast } realitve terms
*cold is slow

Kenetic Energy (KE) = $\frac{1}{2} m v^2$

*vector = gives direction & magnitude

Thermal Expansion

most objects expand when heated

Large structures must be built to leave room for thermal expansion

All features expand together

COLD • HOT ●

Hot will be absorbed by cold.

GALILEAN THERMOMETER

*consistent mass

*volumes change → densities

↳ sink + floating = temp

CONSERVATION OF MATTER

reactants $\xrightarrow{\text{yeild}}$ products.

DENSITY

*Cubic measurment.

*Mass

↳ comparing amount/weight to the amont of space

Density → how much matter is in a given amount of space.

Mass, Density, Volume

$D = \frac{M}{V}$ $M = D \times V$ $V = \frac{M}{D}$

WATER DENSITY = 1 gram / cm³

equal volumes with unequal masses.

ARCHIMEDES PRINCIPAL

Submerging an object in water will give you its volume.

↳ subtracting origional & final = object.

Development of the Atom

February 12 2020

DEMOCRITUS

- Greek Philosopher
- Idea of "democracy"
- Idea of Atoms
  - Atoms = invisible
  - Atom is derived
- Continuous vs discontinuous
- no protons, natrons, electron

FOUNDATIONS of ATOMIC THEORY

LAW of Conservation of Mass

Mass is neither distroyed nor created during ordinary chemical reactions

* Conservation of Atoms → what goes in must come out on opp side

LAW of DEFINITE PROPERTIES

The fact that a chemical compound contains the same elements in exact the same proportions by mass reguardless of size of the sample or source of the compound.

LAW of MULTIPLE PROTONS

2 or more different compounds are composed of the same 2 elements.
ratio of small whole numbers = 1st mass & 2nd mass

DALTON

- All matter is tiny particles.
- Atoms of 1 element can neither be divided or subdivided.
- Atoms cannot be created or destroyed.
- All atoms of the same element are identical, mass, size, ect.
- Atoms of one element differ in mass & other properties from atoms of other elements.
- Atoms of 1 element differ in mass

THOMAS

- RASIN Bread model → bread = nucleus, rasins = electrons.
- Rasin bread aswas plumb pudding.

RUTERFORD

- Gold foil expirment
  ↳ determined: atom → most electricty could pass - atom empty space would deflect, defining its space by it's electrons around it.

BHOR RUTHERFORD

- Core: positive , electron: negitive.

orbits { electrons can exist only in certian places.
↳ electrons cannot absob certian light wave lengths.

Orbits 2 , 8 , 8 , 8 , 8/16

Bhor Rutherford doesn't really work past 1st 20 elements.

PERIODIC TABLE

MASS		
PROTON	1	+1
NUETRON	1	0
ELECTRON	.0001	-1

ATOMIC NUMBER

6
C
CARBON
12.01

MASS

PROTONS
6p = Atomic Mass

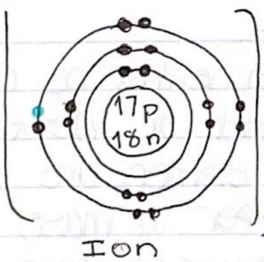
NUETRONS
12 (mass) - 6 (atomic number)
= 6n.

ELECTRONS 6

Atom → pure unalterd form of a meterial
Ion → version of an atom; has more electrons than it should.

17
Cl
CHLORINE
35.45

atom
17p
18n
17e



• → added to make an ion.

Ion ↵
↳ 17+
18-
1- → Charge

Isotope → different number of neutrons than normal
↳ neutron: additon → 1+ or more
deficit → -1 or more

Balencing Chemical equation.

Balancing Chemical Equations

Adding a multiplication in front of an Element

$$\underset{2}{2Al} + \underset{3}{1.5O_2} \longrightarrow \underset{2}{Al_2}\ \underset{3}{O_3}$$

* Balancing Al (aluminum) → multiply by 2
* Balancing O (oxygen) → multiply by 1.5 ?
  ↳ cannot use decimals (would be partial molecule)

$$\underset{4}{4Al} + 3O_2 \longrightarrow \underset{4}{2Al_2}\ O_3$$

mutiply the whole equation by 2 (Aluminums)
mutipy the oxygens by 3 on left

the whole right side was mutiplied by 2
the Al on left is by 4 so 2 × 2 on right is equivilent to 4
the O on left is by 3 so 3 × 2 on right is equavlent to 3 × 2 on left.

1.1 The Nature of Chemistry

SUBATOMIC PARTICLE	SYMBOL	LOCATION IN THE ATOM	Charge	APPROXIMATE MASS (kg)
electron	e^-	IN energy levels outside nucleus	−1	9.11×10^{-31}
proton	p^+	IN nucleus	+1	1.67×10^{-27}
neutron	n^0	IN nucleus	0	1.67×10^{-27}

Atomic Mass Unit → AMU → 1/1000 of an electron/proton/neutron

Atomic Structure

MASS NUMBER

* mass # = protons + neutrons.
* always a whole number
* not on the periodic table

Isotopes

* atoms of the same element with different mass #s
* nuclear symbol → Carbon 12 → mass # 6 → atomic #
  has an atom → 6p, 6n, 6e
* sometimes may have more neutrons → isotope
* Protons → determine the element
* neutrons → determine isotope or Atom / neutrons in isotope have extra weight

Average Atomic Masses

* weighted average of all isotopes
* round to 2 decimal places

* 2 variable percents should add to ~~be~~ 100%

$$\text{Avg Atomic Mass} = \frac{\text{mass}(\%) + \text{mass}(\%)}{100}$$

* Calculate avg atomic mass of oxygen if it's abundance in nature is 99.76% ^{16}O, 0.04% ^{17}O, and 0.20% ^{18}O.

$$\text{Avg} = \frac{16(99.76) + (17)(.04) + 18(.20)}{100} = 16.00 \text{ amu}$$

Lewis Dot Diagrams

→ Amount of electrons in final valent around an element name

1st coloum
H
Li
Na } Alkali metals
↳ All have 1 in last valence shells.

$:\ddot{\underset{..}{Ne}}:$

PERIODOTIC TABLE

Mendeleeu
- organized elements by increasing atomic mass.
- Elements with similar properties were grouped together.

Mosley
- organized by increasing atomic number.
- resolved discrepincies in mendeleve's arrangement

Metallic Character
- Metals
- Non-Metals
- Metaloids

Blocks
- main group elements
  ↳ 1, 2, 13, 14, 15, 16, 17, 18
- Transition Metals → 3-12
- Inner transition metals
  ↳ bottom

PERIODOTIC LAW

* when elements are arranged in order of increasing atomic number, elements with similar properties appear at regular intervals

Families → similar valence e- within a group result in similar properties.

Atomic Radius → size of atom

1st Ionization → Energy required to remove 1 e- from a neutral atom.
↳ ex-hard to remove e- from a noble gas (18)
↳ Energy → increases up & right

- more electrons removed → exponential energy release
  ↳ closer to core

Melting & Boiling Point
- medial & outward to the periodic table.
  ↳ least reactive → centre

Ionic Radius

CATION → lose electrons, smaller

ANION → gain electrons, larger

adding electrons → 1- → 1 unbalanced (no pair)

Chemical Bonding

February 24 2020

introduction to bonding

CHEMICAL BOND

- attractive force between atoms or ions that binds together as a unit.
- bonds form in order to...
  - ↳ decrease potential energy (PE) / stored energy
  - ↳ Increase stability.

CHEMICAL FORMULA

Ionic	Covalent
↳ formula unit	↳ molecular formula
↳ NaCl	↳ CO_2
↳ can be continously added to & joined to the first ion.	↳ when adding to it will create a new molecule not onto old molecules.

Ion

↳ monoatomic → 1 atom

↳ Polyatomic → 2+ atoms → have charges (covalent)

	IONIC	COVALENT
Bond Formation	e- are transfered from metal to non-metal	e- are shared between non-metals
type of Structure	crystal lattice	true molecules
Physical State	solid	solid / liquid / gas
Melting point	high	low
Solubility in water	yes	usually not
Electrical Conductivity	yes (soloution/liquid)	no

Metalic

↳ e- are demoulecatized.

Bond Polarity

- most bonds are a blend of Ionic & covalent characteristics
- difference in electronegitivity determines bond type.
- anything over 0 - .4 pure covalent

  .4 - 1.7 polar covalent

  1.7 + Ionic

Al + P

$|1.5 - 2.1| = .6$ → covalent bond, not ionic

Electronegitivity

- attraction an atom has for a shared pair of electrons
- higher e-neg atom ⇒ $\delta-$
- lower e-neg atom ⇒ $\delta+$
- flows up & to the right.

Non-Polar Covalent

- e- shared equally
- symmetrical e- density
- usually identical atoms.

Polar Covalent Bond

- e- are shared unequally
- asymmetrical e- density
- results in partial charge (dipole)

  $\delta+$ & $\delta-$ ($\delta+$ $\delta-$)

Non-Polar		EN < .5
Polar	$\delta+$ $\delta-$	EN .5 - 1.7
Ionic	⊕ ⊖	EN > 1.7

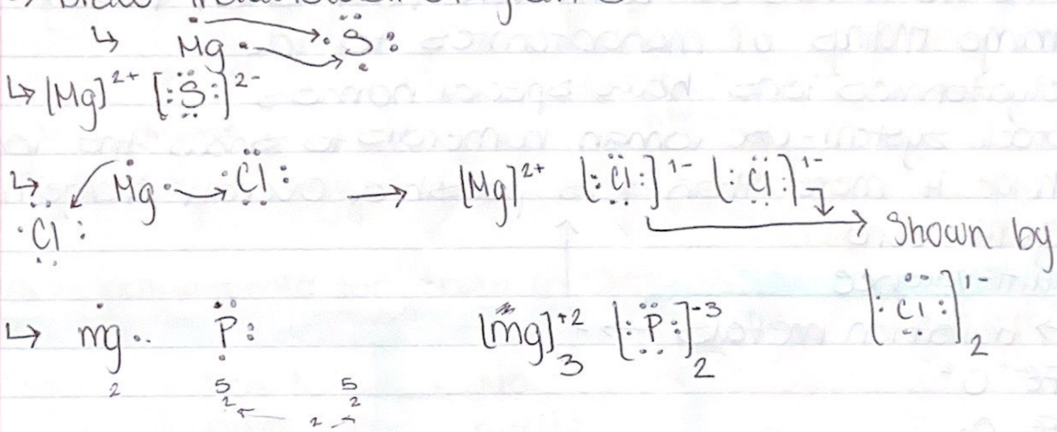
Lattice Energy

- a measure of the strength of the bonds in a ionic compound

Lewis Dot Diagram

↳ coloum # of main group metals will determine outter dot #s

↳ Draw individual diagrams



1 ionic → high melting point, electrons fully transfer, solid

2 electrolyte → dislaves in water, producing a conducting

ionic bond → the jump of an electron to complete the valence ring.

formula unit → smallest repeating unit in a crystal

PG 60 1-10 hw

6 Ionic compound shatter → their planes / molecules are on a straight grid.

Intermolecular force → attraction between molecules

Intramoleculer fore →

Ionic Formulas

- write each ion, cation first. Don't ~~show~~ charges in final.
  ↳ only when bonding
- overall charge of 0
  ↳ if charges cancel write symbols
  ↳ if not: subscripts to balance

- use parentheses to show more than one polyatomic ion.

Ionic Names

Base elements
- Write the names of both ions, cat-ions, first.
- Change endings of monoatomics to -ide
- Polyatomics ions have special names
- Stock system - use roman numerals to show the ions charge if more than 1 is possible. Overall charge must equal zero.

Multi-valence

↳ transition metals →

Fe^{2+} O^{2-}

Fe_2O_2

FeO Iron (II) oxide

Does it contain a polyatomic ion?

↳ -ide, 2 elements → no

↳ -ate, 3+ elements → yes

Does it contain a Roman numeral?

↳ Check the table for metals not in group 1 or 2.

NO PREFIXES

Covalent → PREFIXES → mono, di, tri, tetra, penta (etc).

Ionic Nomenclature

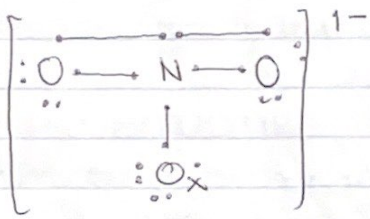
1+ 2+ $\xrightarrow[\text{transitions}]{\text{skip}}$ 3+ 4+/- 3- 2- 1- 0

Nick the Camel ate a Calm Supper in Phoenix

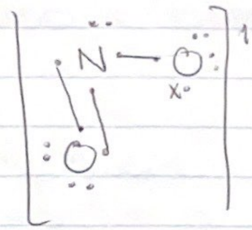
↓ ↓ ↓ ↓ ↓

NO_3^{-1}	CO_3^{-2}	ClO_3^{-1}	SO_4^{-2}	PO_4^{-3}
Nitrate	Carbonate	Chlorate	Sulfate	Phosphate

vowels = charge
constenants = subscript



NO_3^{1-}
Nitrate



NO_2^{1-}
Nitrite.

remove any 1 oxygen from -ates → -ite
↳ remove 1 more
↳ hypo — ite

For any polyatomic ion ending in -ate

Per- ate (1 more)	-ate (standard)	- ite (1 less)	Hypo- -ite. (2 less)
	NO_3^{-1} nitrate	→ NO_2^{-1} nitrite	
	SO_4^{-2} sulfate	→ SO_3^{-2} sulfite	
ClO_4^{-1} perchlorate	ClO_3^{-1} Chlorate	→ ClO_2^{-1} chlorite	→ ClO^{-1} hypochlorite

Charges never Change
↳ only oxygen quantity

81 1-5 7 & 8

1. a) Li

1.1 The Nature of Chemistry

matter - anything that has mass & takes up space.

ATOMIC STRUCTURE

atom - smallest particle or an element.

Chronologic order of atom advances
- Dalton
- Thompson
- Rutheford
- Bhor-Rutherford

Chadwick } neutron discovery

1.2 ATOMIC STRUCTURE

Proton - Positive, in nucleus.

Neutron - Neutral, in nucleus.

Energy Level - theroetical sphere around an atom where electrons exist in orbits.

Valance shell - orbit of an atom or ion.

Valance electron - electron in the outermost energy level or orbit.

Atomic Number - number of Protons top right/~~left~~

Mass Number - sum of protons (Atomic) & neutrons

Atomic Mass Unit - very small unit of mass defined as 1/12 of a carbon-12 atom; unit symbol u.

1.3 Ions & Octet Rule

full/stable Octet - when the outer most valance is full usually 8 but can be 2 for oxygen or carbon.

ion - Charged entity formed when an atom gains or loses 1+ electrons

cation - positivly charged ion formed by the removal of one or more electrons from, the valance shell of a neutral atom

Valance - the charge of an ion; determined by electron loss, addition or share.

1.4 Isotopes, Radioisotopes & Atomic Mass

Isotope form of an element in which the atoms have the same number of protons as all its other forms but it has a different amount of neutrons.

Isotopic Aubundance percentage of an isotope in a sample of an element.

Mass Spectromoter measuring instrament used to determine the mass & abundance of isotopes.

there's many high energy ways that are as a result of differences in electrons

1.5 Periodic Table & Perriodic Law

using groups we know electron numbers of valance outermost shell for the dot diagrams.
↳ not for transition metals.

Period → row
group → column

+1 +2 +3 +4 +5 +6 +7 +8

1.7 Periodic Trends in Atomic Properties

Atomic radius → measurment of a size of an atom.
Ionic radius → measurment of a size of an ion

2.1 Ionic Compounds

electrolyte - compound that dissolves in water
ionic Bond - Electrostatic force of attraction between a positive & negative ions.

2.2 Molecular Compounds & Elements

the sharing of electrons to complete a valance shell
↳ non & non → compound
↳ 2 atoms (any)

Pure substance composed of molecules made of 2+ non-metallic elements.

A pure substance composed of molecules made of 2+ ~~non-metallic~~ elements.

2.3 Chemical Bonding & Electromagnetivity

the ability an atom has to attract bonding electrons to its self.

2.4 Chemical Formulas & Nomenclature

Ionic Compounds
(IUPAC Name) ↳ (I), (II), (III) ⟶ when writing out names
↳ $CuCl_2$ → copper Cu^{2+} & Chlorine Cl^-

would be found @ cross down

$Cu^2 \cancel{Cl}^1$
$CuCl_2$

I Brought Clay For Our New House
Iodine Bromine Chlorine Florine Oxygen Nitrogen Hydrogen

Exterior elements need to be totalled.

Chemical Bonding

March 2 2020

ENERGY OF BOND FORMATION

Potential Energy

- Based on position of an object
- low PE - high density

LEWIS STRUCTURE

Electron Dot Diagrams

- Show valence e- as dots
- Place on electron on each side; then pair

OCTET RULE

- most atoms form bonds in order to get a full 8 valence
- Full energy level stability - Nobel Gas.

COVALENT

- share electrons; to complete 8, insted of a jump.

Non-Polar Covalent - no charges

:Cl: + .Cl: ⟹ :Cl:Cl: (chlorine molecule)

(chlorine atom)

Polar Covalent - Partial Charges

oxygen atom

H. H. + :O: ⟶ H:O: H

2 hydrogen atoms

WATER MOLECULE

Only up to 3 bonds w/1 of same elements no 4

Molecular Nomenclature *PREFIX SYSTEM*

- Less e-neg atom comes first.
- Add prefixes indicating number of atoms.
  - ↳ omit mono- on the first element
- Add ide- to the ending of the second element.

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

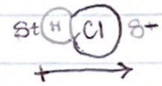
- CCl_4 ~~mono~~ carbon tetrachloride
- N_2O dinitrogen mon~~o~~oxide

Molecular Structure

MOLECULAR POLARITY + BOND POLARITY

Dipole Moment

- Direction of the polar bond in a molecule.
- Arrow points toward the more e-neg atom.



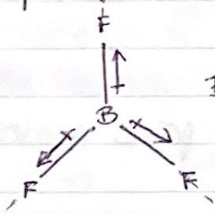
$C^{\delta+} - F^{\delta-}$

DETERMINING MOLECULAR POLARITY

- Dipole moments & - molecular shape.

NONPOLAR MOLECULE

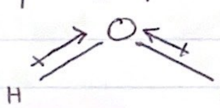
- Dipole moments are symetrically & can cancle out.



BF_3 - Symetrical pulling → cance out
↳ no movement

POLAR MOLECULES

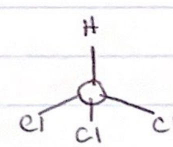
- Dipolar momements are asymetrical; don't cancel.



H_2O

- therefore, polar molecules have ...
  - Asymetrical shape (lone pairs) OR
  - Asymetrical atoms

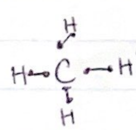
$CHCl_3$



Bond polarity

	Non-Polar	Polar
Non-Polar	CH_4	PH_3
Polar	CF_4	CH_2F_2

CH_4

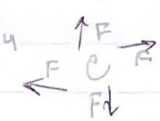


C - H
2.5 2.1 = .4
non-polar

PH_3

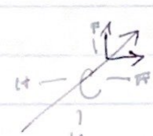
P - H
2.2 - 2.1 = .1
non-polar bond

CF_4



C - F
2.5 - 4 = 1.5 → polar

CH_2F_2



C - H .4 non
C - 1.5 polar

Pull toward 1 end.

SAME ATOMS
↳ side by side

Molecular Structure

VSEPR Theory

- Valance Shell Electron Pair Repulsion Theory.
- Electron Pairs orient them selves to minimize repulsive forces.

TYPES OF e⁻ PAIRS

- Bonding Pairs – form bonds
- Lone Pairs – nonbonding e⁻
  - ↳ higher e⁻ negitivity.

Lone pair repel more strongly than bonding pairs

Draw lewis Diagrams

Tally up e⁻ pairs on central atom,

↳ double/triple bonds = 1 pair

Shape is determined by the # of bonding pairs & lone pairs.

8 common shapes & bond angels.

*Electronegitivities.
larger-smaller.

everything equally speced.

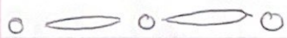
lone pairs can only come from the central/larger electron amount.

Common Molecular Shapes

Linear

2 total 180°

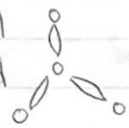
2 bond



Trigonal Planar

3 total 120°

3 bond

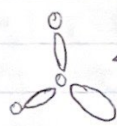


Bent I

3 total

2 bond

1 lone

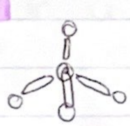


∠120°

Tetrahedral

4 total

4 bond



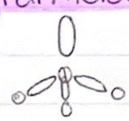
109.5°

Trigonal Pyramidal

4 total

3 bond

1 lone



107°

Bent II

4 total

2 bond

2 lone

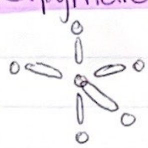
104.5°

Trigonal Bipymardial

5 total

5 bond

0 lone

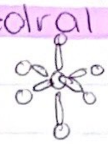


120°/90°

Octaheolral

6 total

6 bond



90°

3.3 Bond Polarity & Polar Molecules

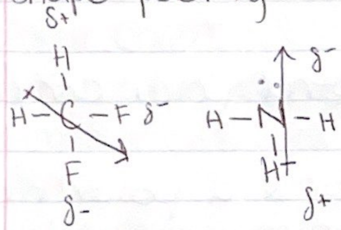
March 9 2020

bond polarity X-Y

|electromagnetivity difference|

ionic covalent
non polar

shape polarity



Intra → between atoms covalent + ionic
Inter → between molecules

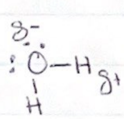
Intermolecular forces

intramolecular forces: forces that bond the atoms to eachother within a molecule.

Intermolecular forces: forces affecting the relationship between molecules.
- london forces
- dipole-dipole
- Ion dipole
- hydrogen.

London Dispersion forces
- weakest Intermolecular force
- all types of molecules
- tempory, when electrons in 2 adjacent atoms ocupy positions making atoms form temporary dipoles.

gas

liquid 

Diff in energy & attractive force.

solid

$\delta^{+/-}$ location
↳ + at end w/smaller e- amt
↳ - at end w/larger e- amnt.

Dipole - Dipole
- polar - molecule
- change their direction so their + & - end line up with eachothers.
- be w/ any compound
- polar compounds can dislove other polar compounds
ion Dipole Forces
- Force of attraction between an ion & a polar molecule
- NaCl breaks up because the ion dipole with water is stronger than the attraction of Na^+ to Cl^-

Hydrogen - Bonding
stronger than normal dipole forces. (strongest)
- not as strong as covalent 1/10.
- strong enough for water to be different.

N - H
O - H
F - H

March 11 2020

Chemical Reactions

A. Signs of a Chemical Reaction

* evoloution of heat & light
* formation of gas
* formation of precipate
* colour change

Atleast 2/4 but all must affect chemical structure

B. Law of Conservation of Mass

* mass is neither created nor deystroyed in a chemical reaction (balancing chem equations)
* total mass stays the same

C. Chemical Equations

A + B ⟶ C + D

REACTANTS PRODUCTS

cannot change subscripts

SYMBOL	MEANING
⟶	produces, forms
+	plus, and
(s)	solid
(l)	liquid
(g)	gas
(aq)	aqueous (solid disloved in water)
$\xrightarrow{\Delta}$	reactants are heated

D. Writing Equations

$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

* Identify substances involved
* use symbols to show
  * How many? coefficent
  * Of what? chemical formula
  * In what state? Physical State

*Rember diatomic elements.

Word equations; asume normal state if not given

2 atoms of aluminum react w/ 3 units of copper (II) chloride to produce 3 atoms of copper & 2 units of aqueus Aluminum Chloride

$2Al(s) + 3CuCl_2(aq) \longrightarrow 3Cu(s) + 2AlCl_3(aq)$

E. Describing Equations

* Describing Coeffients
  * Induvidual - atom
  * covalent substance — molecule
  * Ionic substance – unit

$3CO_2$ 3 molecules of carbon dioxide

2Mg 2 atoms of magnesium

4MgO 4 units of magnesium oxide

TYPES OF REACTIONS

A. Decomposition

* a compound breaks down into 2+ simpler substances.
* Only 1 reactant

AB ⟶ A + B //

* Products
  * Binary - break into elements
  * Others hard to tell

$2KBr(l) \longrightarrow 2K(s) + Br_2(l)$

B. Synthesis

* Combination of 2+ substances to form a compound
* Only 1 product

A + B ⟶ AB //

* Products
  * Ionic - cross charges
  * Covalent - hard to tell.

$Al(s) + Cl_2(g) \longrightarrow AlCl_3(s)$

C. Single Replacement (dis)

* 1 element replaces another in a compound
  * metal replaces metal (+)
  * non-metal replaces non-metal (-)

A + BC ⟶ B + AC

C. Single Replacement

* products
  * metal → metal (+)
  * non-metal → non-metal (-)
  * free element must be more active (activity series)

$Fe(s) + CuSO_4(aq) \longrightarrow Cu(s) + FeSO_4(aq)$

~~AB + CD~~

D. Double Displacement

* Ions in 2 compounds "change partners"
* cations can switch with anion of the other but never with 2 anion or 2 cations

$AB + CD \longrightarrow AD + CB$ // 

* switch negitive ions
* 1 product must be insauble (gas or water)

$NaNO_3(aq) + KI(aq) \longrightarrow$ N.R

$Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow PbI_2(s) + 2KNO_3(aq)$

E. Combustion

* the burning of any substance in O_2 to produce heat

$A + O_2 \longrightarrow B$

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$ } elements become oxides.

* Complete combustion
  * contains oxygen
  * hycarbonates form $CO_2 + H_2O$

$4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$

$C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

March 11 2020

Chemical Reactions

IV Reaction Energy

A. Reaction Pathway

* shows the change in energy during a chemical energy.

C. Endothermic Reaction (rare 1/2)

* reaction that absorbs energy.
* reactants have lower PE than products

$2Al_2O_3$ + energy $\longrightarrow$ $4Al + 3O_2$

B. Exothermic Reaction (giving energy common)

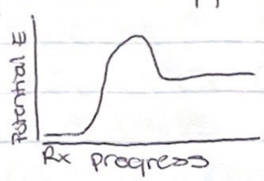
* reaction that releases energy.
* products have lower PE than reactants

$2H_2(l) + O_2(l) \rightarrow 2H_2O(g)$ + energy

April 17
2020

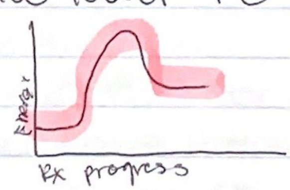
Chapter 17 Chemical Reactions Reaction Energy

A reaction pathway
-shows change in energy during chemical reaction.



B endothermic reaction
-releases energy
-products have lower PE than reactants.

C Endothermic reaction
-reaction that absorbs energy
-reactants have lower PE than products



April 17
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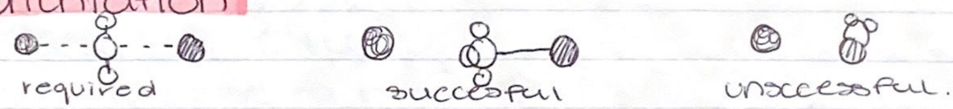
Chemical Reactions — Reaction Rates

A collision theory
- reaction rate depends on collisions between reacting particles.
- successful collission occurs
  - collide w/eachother
  - have correct orrientation
  - enough kenetic energy.

orientation



Activation energy
- minimum energy required for a reaction to occur
- depends on reactants.
- low Ea = fast rxn rate.

B Factors Affecting Rxn Rate.

Surface area
- high SA = fast rxn rate
- more opportunities for reaction.
- increase surface Area by...
  - using small particles — dissloving water

concentration
- high conc = fast rxn rate
- more opportunities for colision.

temperature
- high temp = fast rxn rate
- high KE • fast moving particles • mor likley to achieve

Catalyst
- substance that increases rxn rate without being used
- lowers activation energy.

Molar Conversions.

A What is a mole?

- A counting number (ex dozen)
- Avogadro's number (N_A)
- 1 mol = 6.02×10^{23} (atoms, molecules, particles)
  very large.

Origin of the mole

- 1 proton = 1.66×10^{-24} g
- how many protons are there in 1.00 grams of protons?
  ↳ $1.00g / 1.66 \times 10^{-24} = 6.02 \times 10^{23}$
  ↳ moles =
- based on carbon.
- size is close to planets

B Molar Mass

- mass of 1 mole of an element or compound.
- Atomic mass tells the ...
  - atomic mass units per atom (amu)
  - gram per mole (g/mol)
    - ex 1 atom of Carbon (c) = 12 amu = 1 mole = 12 g
    - round 2 decimal places.

ex C = 12.01 g/mol   al = 26.98 g/mol   Zn = 65.39 g/mol
 └ g/mol is similar/same as mass on table

ex sodium hydrogen carbonate ($NaHCO_3$)
 ↳ 22.99 + 1.01 + 12.01 + 3(16.00) = 84.01 g/mol

C Molar Conversions

$$\text{Mass in Grams} \xrightarrow{\text{molar mass}} \text{moles} \xrightarrow{6.02 \times 10^{23}} \text{Number of Particles}$$

$$\text{Mass in Grams} \xleftarrow[\text{(g/mol)}]{} \text{moles} \xleftarrow[\text{(particles/mol)}]{} \text{Number of Particles}$$

C Molar Conversion Examples

how many moles of carbon are in 26 g of carbon.

$$\frac{26\,\cancel{g}\,C}{} \bigg| \frac{1\ mol\ C}{12.01\,\cancel{g}\,C} = 2.2\ mol\ C$$

how many molecules are in 2.50 moles of $C_{12}H_{22}O_{11}$

$$\frac{2.50\,\cancel{mol}}{} \bigg| \frac{6.02 \times 10^{23}\ molecules}{1\ \cancel{mol}} = 1.51 \times 10^{24}\ molecules\ C_{12}H_{22}O_{11}$$

find the mass of 2.1×10^{24} molecule $NaHCO_3$.

$$\frac{2.1 \times 10^{24}\ \cancel{molecules}}{} \bigg| \frac{1\ \cancel{mol}}{6.02 \times 10^{23}\ \cancel{molecules}} \bigg| \frac{84.01\,g}{1\ \cancel{mol}} = 290\,g\ NaHCO_3.$$

April 20/20 # FORMULA CALCULATIONS formula calculation

Law of Definite proportions:
- elements in a chemical compound are always present in the same proportions by mass
- by Joseph Louis Proust in late 18^{th} century

Percentage Composition
- the percentage by mass of each element in a compound.

formula ⟶ $\% \text{ Composition} = \frac{\text{mass of element}}{\text{total mass}} \times 100$

Cu_2S

$\% Cu = \frac{127.10g\ Cu}{159.17\ g\ Cu_2S} \times 100 = 79.85\% Cu$ $\quad \% S = \frac{32.07gS}{159.17g\ Cu_2S} \times 100 = 20.15\% S$

$28g\ Fe$ & $8g\ O$

$\% Fe = \frac{28g}{36g} \times 100 = 78\% Fe$ $\quad \% O = \frac{8.0g}{36g} \times 100 = 22\% O$

copper in 38.0 g sample of Cu_2S

$Cu_2S = 79.85\% Cu$

$(38.0g\ Cu_2S)(0.7985) = 30.3g\ Cu$

$CaCl_2 \cdot 2H_2O$

$\% H_2O = \frac{36.04g}{147.02g} \times 100 = 24.51\% H_2O$

Empirical Formula

- smallest whole number ratio of atoms in a compound $C_2H_6 \rightarrow CH_3$

1. Find mass or % of each element
2. Find moles of each element
3. Divide moles by smallest # to find subscripts.
4. When necessary multiply subscripts by 2, 3 or 4 to get whole numbers

- Percent

25.9% N & 74.1% O

$$N \quad \frac{25.9\,g}{} \cdot \frac{1\,mol}{14.01\,g} = \frac{1.85\,mol\,N}{1.85\,mol} = 1\,N$$

$$O \quad \frac{74.1\,g}{} \cdot \frac{1\,mol}{16.00\,g} = \frac{4.63\,mol\,O}{1.85\,mol} = 2.50$$

$N_1O_{2.5} \times 2$

N_2O_5

Molecular Formula

- true formula, the actual number of atoms in a compound. $CH_3 \rightarrow C_2H_6$

1. Find the empirical formula
2. Find the empirical formula mass
3. Divide the molecular mass by the empirical mass
4. Multiply each subscript by the answer from 3.

$$\frac{MF\ mass}{EF\ mass} = n \qquad (EF)_n$$

CH_2 molecular mass 28.1 g/mol
emperical mass = 14.03 g/mol

$$\frac{28.1\,g/mol}{14.03\,g/mol} = 2.00 \qquad (CH_2)_2 \Rightarrow C_2H_4$$

molar mass 159.2 composition 20.2 S and
79.8 % Cu what's the formula?

$(20.2\%)(159.2g) = 32.16g\ S$

$32.16g \times 1mol = 1\ mol\ S$

$79.8\%Cu\ (159.2g) = 127.0g\ Cu$

$127.0g \times 1mol = 2$ mol copper

$\therefore\ Cu_2S$

April 23 2020

STOCHIOMETRY

STOICHIOMETRIC CALCULATIONS

(a) Proportional Relationships

2 1/4 c. flour
1 tsp baking soda
1 tsp salt
1 c. butter
3/4 c sugar
3/4 brown sugar
1 tsp vanilla
2 eggs
2 c chocolate chips
makes 5 dozen

I have 5 eggs. How many cookies can I make?

$\frac{5 \text{ eggs}}{} \Big| \frac{5 \text{ doz}}{2 \text{ eggs}} = 12.5$ dozen cookies ← ratio of eggs : cookies

- mass relationships between substances in a chemical reaction
- based on the mole ratio
  - ↳ indicated by coefficients in a balanced equation

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$$

(b) Mole Ratios

$C_{(s)} + O_{2(g)} \longrightarrow CO_2$ can be read as . . .

1 atom carbon + 1 molecule oxygen ⟶ 1 molecule carbon dioxide
1 mole carbon atom + 1 mole oxygen molecule → 1 mole carbon dioxide

$$2Al_{(s)} + 6H_2O_{(l)} \longrightarrow 3H_{2(g)} + 2Al(OH)_{3(s)}$$

If you have 3.2 mole Al, how many moles of H_2 gas will be produced

2 mol Al : 3 mol H (from balanced equ)
3.2 mol Al : x mol H_2 (from question)
$2x = 9.6$
$x = 4.8$ mol H_2 produced

(c) Box Method

How many grams of MgO will be produced from 9.7g of Mg?

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$$

	Mg	MgO
m	9.7g	16g
mm	24.31 g/mol	40.31 g/mol
n	0.40 mol	0.40 mol
R	2	2

How many grams of $KClO_3$ must decompose in order to produce 9.0 g of oxygen gas?

$$2KClO_{3(s)} \longrightarrow 2KCl_{(s)} + 3O_{2(g)}$$

	$2KClO_3$	$3O_2$
m	23 g	9.0 g
mm	122.55 g/mol	32.00 g/mol
n	0.19 moles	.28 moles
R	2 moles	3 moles

$3x = .56 \quad x = 0.19$ moles KClO

② Unit Analysis

How many grams of silver will be formed from 12.0 copper?

$$Cu_{(s)} + 2AgNO_{3\,(aq)} \rightarrow 2Ag_{(s)} + Cu(NO_3)_{2\,(aq)}$$

12.0 g ? g

$$\frac{12.0\,\cancel{g\,Cu}}{} \left| \frac{1\,\cancel{mol\,Cu}}{63.55\,\cancel{g\,Cu}} \right| \frac{2\,\cancel{mol\,Ag}}{1\,\cancel{mol\,Cu}} \left| \frac{107.87\,g\,Ag}{1\,\cancel{mol\,Ag}} \right. = 40.7\,g\,Ag$$

How many grams of Cu are required to react with 1.5 L of .10 M $AgNO_3$? $Cu_{(s)} + 2AgNO_{3\,(aq)} \longrightarrow 2Ag_{(s)} + Cu(NO_3)_{2\,(aq)}$

? g 1.5 L .10 M

$$\frac{1.5\,\cancel{L}}{} \left| \frac{.10\,\cancel{mol\,AgNO_3}}{1\,\cancel{L}} \right| \frac{1\,\cancel{mol}\,Cu}{2\,\cancel{mol}\,AgNO_3} \left| \frac{63.55\,g\,Cu}{1\,\cancel{mol\,Cu}} \right. = 4.8\,g\,Cu$$

CHEMISTRY 04/30/20

STICHIOMETRY IN THE REAL WORLD

April 30 2020

A LIMITING REACTIONS.

avaliable ingredients • 1/2 jar of jelly
• 4 slices of bread • 1 jar of peanut butter
- limiting reactant → bread - excess reactants → peant butter & jelly.

LIMITING REACTANT
• used up in a reaction • determines amount of product.

EXCESS REACTANT
• added to ensure that the other reactant is used up
• cheaper & easier to recycle.

1. write a balanced equation
2. for each reactant calculate the amount of product formed.
3. smaller answer indicates
   - limiting reactant
   - amount of product.

79.1g of zinc react with .90L of 2.5M HCl. Identify the limiting excess reactants. How many gs of hydrogen are formed?

$Zn + 2HCl \rightarrow ZnCl_2 + H_2$

79.1g .90L 2.5M ?g

$$\frac{79.1\text{ g Zn}}{} \times \frac{1\text{ mol Zn}}{65.39\text{ g Zn}} \times \frac{1\text{ mol } H_2}{1\text{ mol Zn}} \times \frac{2.02\text{ g } H_2}{1\text{ mol } H_2} = 2.44\text{ g } H_2$$

$$\frac{.90\text{ L}}{} \times \frac{2.5\text{ mol HCl}}{1\text{ L}} \times \frac{1\text{ mol } H_2}{2\text{ mol HCl}} \times \frac{2.02\text{ g } H_2}{1\text{ mol } H_2} = 2.3\text{ g } H_2$$

Limiting reactant HCl

Excess Reactant Zn

Product formed 2.3g H_2

B Percent Yeild

$$\%\text{ yeild} = \frac{\text{actual yeild}}{\text{theoretical yeild}} \times 100$$

actual yeild: measured in lab

theoretical yeild: calculated with schitometry.

When 45.8g of K_2CO_3 react with excess HCl, 46.3g of KCl are formed. Calculate the theroetical & % yeild of KCl.

$K_2CO_3 + 2HCl \rightarrow 2KCl + H_2O + CO_2$ Actual 46.3g

45.8g ?g ACTUAL 46.3g

THEROETICAL YEILD:

$$\frac{45.8\text{ g } K_2CO_3}{} \times \frac{1\text{ mol } K_2CO_3}{138.21\text{ g } K_2CO_3} \times \frac{2\text{ mol KCl}}{1\text{ mol } K_2CO_3} \times \frac{74.55\text{ g KCl}}{1\text{ mol KCl}} = 49.4\text{ g KCl}$$

$$\%\text{ yeild} = \frac{46.3\text{ g}}{49.4\text{ g}} = 100 = 93.7\%$$

SOLOUTIONS

Chemistry
SCH3U
April 31 20

THE NATURE OF SOLOUTIONS 1/3

A Definitions.

solution a homogeneous mixture.
solute a substance being disloved.
solvent present in a greater amount.

B Solvation.

solvation the process of dissloving.
FIRST solute particles are surrounded by solvent particles
THEN solute particles are seperated & pulled into the soloution
non-electrolyte solute exists as molecules only
weak-electrolyte solute exists as molecules and ions.
strong-electrolyte solute exsists as ions only.
branches into <u>dissociation</u> and <u>ionization</u>.

dissociation seperation of an Ionic solid into aqueous ions
($NaCl_{(s)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)}$)

ionization breaking apart of some polar molecules into aqueous ions. ($HNO_{3(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_3^-_{(aq)}$)

molecular solvation molecules stay in tact.
($C_6H_{12}O_{6(s)} \longrightarrow C_6H_{12}O_{6(aq)}$)

like disloves like Polar + Polar & Non-polar + Non-Polar

C solubility

unsaturated solution more solute dissloves
saturated solution no more solute dissloves.
supersaturated solution unstable, crystals form.

* maximum grams of solute that will dissloue in 100 g of solvent at a given temperature.
  - varies in temp
  - based on a saturated soln.
* Solubility Curve - shows dependence of solubility on tem
* Solids are more soluable at high temperatures.
* Gases are more soluable at -low temperature & -high pressure (Henry's law) -ex nitrogen narcosis the "bends" soda

April 31 2020
CHEMISTRY SCH3U

LIQUIDS + SOLIDS

INTERMOLECULAR FORCES

A Definition of Intermolecular Forces.

* Attractive forces between molecules.
  - ↳ weaker than chemical bonds within molecules
  - ↳ aka vanderwaals forces.

B Types of Intermolecular Forces.

ion ion interactions

* very strong bond between positive & negitive ions.
  - ↳ not really an intermolecular force.

ion dipole interactions

* Attraction between a charged ion & the oppositley charged end of a polar molecule.
  - ↳ ex bond between positive end of polar water molecule & negitive chlorie ion.

dipole dipole forces/Keesom forces

* Attraction between oppositley charged ends of permanent. dipole/polar molecules.
  - ↳ can be very strong.

dipole induced forces/debye force.

* A polar molecule can torarily induce a dipole in a non-polar molecule.
  - ↳ negitive end of a diple can repel electrons in a non-polar molecule, causing a slight positive end.

induced/induced interactions/london dispersion forces.

* occour between all molecules, espically non-polar.
  - ↳ as electrons move around, slightly +/- regions occour
    - ↳ very weak attraction.

hydrogen bonding.

* Strongest intermolecular force
* Bond between sigre hydrogen nulcus & electrons on 2 adjacent molecule
* forms & breaks quickly
* only H-F, H-O & H-N bonds.

- Dertermining Intermolecular Forces.
  - NCl3
    - Polar = dipole-dipole / keesom, london disperson, ion-dipole
  - CH4
    - Non-polar = london dispersion.
  - HF
    - Polar = dipole / dipole / keesom, london elisperson, iondipole, hydrogen bonding.

Chemistry May 7 2020

8.5 SOLUBILITY & SATURATION

Degrees of Saturation

depending on the quantity of solute they hold at a given temp

saturated soloution contains maximum quantity of solute at a given temperature & pressure.

unsaturated soloution more solute can disolve at a given temperature & pressure.

supersaturated soloution contains more than the maximum quantity of solute at a given temperature & pressure.

Solubility Curve.

a graph of the solubility of a substance over a range of temperatures

Solubility Curves of Ionic Compounds.

most soubilities of ionic compounds increase with temperature. They're affected differently but some rise more sharply than others.

Solubility Curves of Gases

unlike solids, solubility decreases as temperature rises.

thermal pollution increase of water in temperature, result of warm water being added to an aquatic ecosystem.

pressure (P) force per unit area: unit is the pascal, unit symbol Pa: 1 Pa = 1 N/m^2

Pressure is an excess force applied to the unit area.

KEY POINTS

* Solubility of a soloution is expressed as the mass of solute required to form a saturated soloution in 100 g of water at a given temperature.

Name: ______________________ Date: __________ Class__________

Solubility Curve Practice Problems Worksheet

You'll notice that for most substances, solubility increases as temperature increases. As discussed earlier in solutions involving liquids and solids typically more solute can be dissolved at higher temperatures. Can you find any exceptions on the graph? NH_3, $Ce_2(SO_4)_3$

Here's an example of how to read the graph. Find the curve for $KClO_3$.

At 30°C approximately 10g of $KClO_3$ will dissolve in 100g of water. If the temperature is increased to 80°C, approximately 40 g of the substance will dissolve in 100g (or 100mL) of water.

Directions: Use the graph to answer the following questions. REMEMBER UNITS!

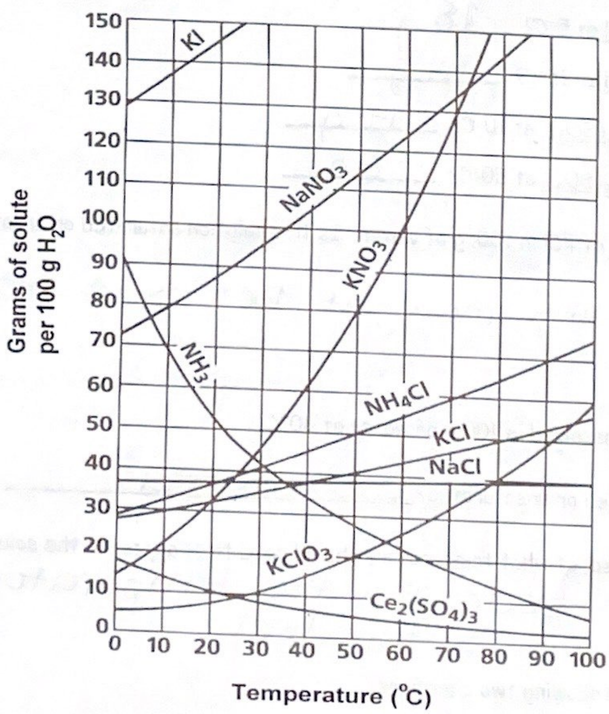
1) What mass of solute will dissolve in 100mL of water at the following temperatures?

a. KNO_3 at 70°C = 130 grams

b. NaCl at 100°C= 40 grams

c. NH_4Cl at 90°C= 70 grams

d. Which of the **above** three substances is most soluble in water at 15°C. = KNO_3



2) Types of Solutions

On a solubility curve, the lines indicate the concentration of a saturated **solution** - the maximum amount of solute that will dissolve at that specific temperature.

Values on the graph beneath a curve represent **unsaturated solutions** - more solute could be dissolved at that temperature.

Label the following solutions as saturated or unsaturated. If unsaturated, write how much more solute can be dissolved in the solution.

Solution	Saturated or Unsaturated?	If unsaturated: How much more solute can dissolve in the solution?
a solution that contains 70g of $NaNO_3$ at 30°C (in 100 mL H_2O)	unsaturated	20 g +
a solution that contains 50g of NH_4Cl at 50°C (in 100 mL H_2O)	saturated	
a solution that contains 20g of $KClO_3$ at 50°C (in 100 mL H_2O)	saturated	
a solution that contains 70g of KI at 0°C (in 100 mL H_2O)	unsaturated	60 g+

1. a. What is the solubility of KCl at 5°C? aprox 18 g
   b. What is the solubility of KCl at 25°C? 34 g
   c. What is the solubility of $Ce_2(SO_4)_3$ at 10°C? 14 g
   d. What is the solubility of $Ce_2(SO_4)_3$ at 50°C? 5 g

2. a. At 90°C, you dissolved 10 g of KCl in 100. g of water. Is this solution saturated or unsaturated?
   b. How do you know?
   unsaturated, its max at 100° is ≈ 55 g

3. A mass of 100 g of $NaNO_3$ is dissolved in 100 g of water at 80°C.
   a) Is the solution saturated or unsaturated? unsaturated
   b) As the solution is cooled, at what temperature should solid first appear in the solution? Explain.
   at ≈ 35° because the temperature is wher 100g is saturated.

4. Use the graph to answer the following two questions:
   Which compound is most soluble at 20 °C? KI
   Which is the least soluble at 40 °C? $Ce_2(SO_4)_3$

5. Which substance on the graph is **least** soluble at 10°C? $KClO_3$

6. A mass of 80 g of KNO_3 is dissolved in 100 g of water at 50 °C. The solution is heated to 70°C. How many more grams of potassium nitrate must be added to make the solution saturated? Explain your reasoning

$\boxed{50g}$ max g at → 50° = 80 g ∴ 130g - 80g.
↳ 70° = 130 g = 50g.

7. Elements review: Fill in the chart below for some of the compounds on the graph:

Formula	# of atoms in formula	If the following amounts of solute are dissolved in 100 mL of water: Is the solution SATURATED OR UNSATURATED
Example: NaCl	*Na* = 1 Cl = 1	*3 grams dissolved at 0°C* unsaturated
KI	K = 1 I = 1	*120 grams dissolved at 0°C* unsaturated
$Ce(SO_4)_3$	Ce = 1 S = 3 O = 12	*7.2 grams dissolved at 70°C* saturated
NH_4Cl	N = 1 H = 4 Cl = 1	*11 grams dissolved at 46.7°C* unsaturated

Radius → increases down + left
1st Ionization → increases up + right

Pt A 7/9 = 77
Know 25

St. Aloysius Gonzaga S.S.
SCH3U Unit 1 Test – Matter, Chemical Trends and Chemical Bonding

PART A: Multiple Choice – *Answer all questions on the SCANTRON provided. Do NOT write on the question sheet!*

1. Elements in the same chemical family tend to have the same
   (a) atomic numbers
   (b) numbers of valence electrons
   (c) atomic radii
   (d) ionization energy

2. An element has a low first ionization energy and a low electron affinity. What is it most likely to be?
   (a) a halogen
   (b) a noble gas
   (c) an alkali metal
   (d) a Group VI (16) element

3. Which of the ions would not have a noble gas electron configuration?
   (a) Li^+ (1 loss) (b) Cl^- (1 gain) (c) O^{2-} (2 gain) (d) Mg^{2+} (2 loss) (e) Sn^{2+} (2 loss)

b 4. In which pair of elements is the element with the smaller radius listed first?
   (a) potassium, calcium (smaller, 194)
   (b) oxygen, sulfur (smaller; 48, 88)
   (c) aluminum, silicon (118 smaller, 111)
   (d) iodine, bromine (larger, smaller; 115, 94)

d 5. In which pair of elements is the element with the smaller electron affinity listed first?
   (a) fluorine, oxygen (16, 15)
   (b) sodium, magnesium (22, 24; smaller)
   (c) beryllium, lithium (10, 6)
   (d) helium, neon (4, 20)

6. What is the ionic charge of the iron in the compound FeN?
   (a) -3 (b) +1 (c) +3 (d) +2

   Fe: N· +2 -2 +3 -3

7. Which of the following correctly represents a stable ion of aluminum?
   (a) Al: (b) Al^{3+} (c) Al^{3-} (d) Al

a 8. A metallic element X forms a chlorate with the formula $X(ClO_3)_2$. Which element can represent X?
   (a) Ca (b) Na (c) ~~B~~ (d) S

★9. Which element will form a polar covalent bond with nitrogen?
   (a) F (b) Li (c) K (d) O

   5 6 7

b 10. Consider the equation $X_{(g)}$ + energy → $X^+ + e^-$. What does the "energy" term in the equation represent?
   (a) electron affinity
   (b) ionization energy
   (c) electronegativity
   (d) none of the above

11. Why does atomic radius increase from top to bottom in a chemical family?
   (a) Ionic charge increases from top to bottom in a chemical family.
   (b) Ionic charge decreases from top to bottom in a chemical family.
   (c) The number of orbits around the nucleus increases from top to bottom in chemical family.
   (d) The number of orbits around the nucleus decreases from top to bottom in a chemical family.

15 Com. $\frac{4}{8}$ = 50 = 7.5 = 32.?

4 12. If the classical name of an acidic substance has the ending "...ic acid", then the formula for the "per...ic acid" of the same family is obtained by
(a) removing two oxygen atoms (c) subtracting one oxygen atom
(b) adding one hydrogen atoms (d) adding one oxygen atom

13. Which of the following elements requires the least amount of energy to remove an electron from an atom to form an ion?
(a) O (b) He (c) K (d) H

b 14. Which of the following is an element consisting of diatomic molecules:
(a) copper (b) hydrogen (c) helium (d) silicon

*15. Which bond is most polar?
(a) H-O ~~(b) I-Br~~ ~~(c) F-Cl~~ (d) O-S

a 16. Which list correctly gives the chlorates of the alkaline earth metals?
(a) $Mg(ClO_3)_2$, $Ca(ClO_3)_2$, $Sr(ClO_3)_2$
~~(b)~~ $NaClO_3$, $KClO_3$, $RbClO_3$
(c) $MgCl_2$, $CaCl_2$, $SrCl_2$
~~(d)~~ NaCl, KCl, RbCl

d 17. Metals tend to :
~~a) share electrons~~
b) gain electrons
c) form positive ions
d) form negative ions
e) all of the above

*18. Which molecule or ion has the same shape as CH_4?
(a) PCl_3 ~~(b) CO_2~~ (c) NH_4^+ (d) NO_3^-

19. Which species has the same total number of valence electrons as a carbon monoxide, CO, molecule?
(a) O_2 (b) NO 11 (c) NO^- 5+6 12 (d) CN^- 9 +1 → 10 10

b 20. Which statement about atomic and ionic radii would be false?
(a) Na has a larger atomic radius than the ionic radius of Na+ 1 lost smaller/more e-
(b) Mg 2+ has a smaller ionic radius than the atomic radius of Mg w/o 2 e- 1 less V.
(c) Br has a larger atomic radius than the ionic radius of Br- +1 e cant be it has 1+
(d) Cl – has a larger ionic radius than the atomic radius of Cl +1 e

25 Application 25/130 = 76

19

Version A Part B: Short Answer (36 marks) Name: Alexis Williams

1. Choose 5 of the following compounds and name them. (5 marks)

(a) PO_5 (b) $CuBrO_3$

(c) $Li_2SO_4 \cdot 8H_2O$ (d) $HI_{(aq)}$

(e) $KMnO_4$ (f) $Pb(CH_3COO)_2$

4/5

x 2

a) Phosporus pentaoxide

b) copper (I) Bromate

c) dilithium sulfate ~~octahydrate~~ droxide

e) Potasium ~~monomanganese (IV) tetraoxide~~ Permangate

f) lead (II) acetate

8/10

85.82

2. Choose 5 of the following and write the chemical formula for each. (5 marks)

(a) tin (ii) sulfate (b) calcium hypofluorite

(c) lithium hydrogen phosphate (d) sodium persulfate

(e) iron (III) cyanide (f) potassium chromate

4.5/5

x 2

9/10

a) $Sn(SO_4)_2$

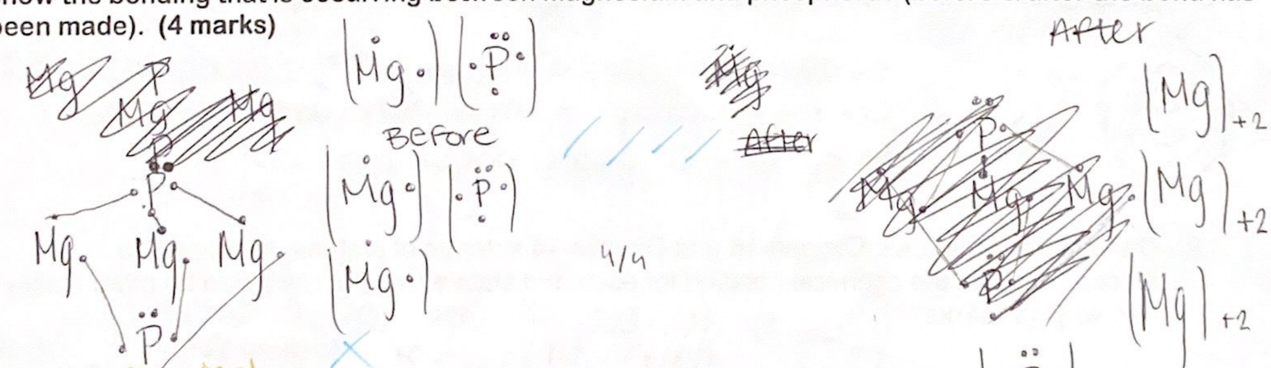
c) LiH_2PHO_4

d) $NaPO_5$

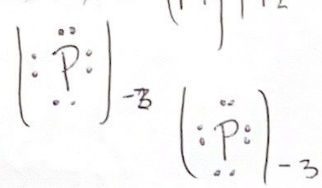
f) $KCrO_4$

e) $Fe(CN^-)_3$

3. Magnesium combines with phosphorus to form magnesium phosphide. Use Lewis-Dot diagrams to show the bonding that is occurring between magnesium and phosphorus (before & after the bond has been made). (4 marks)



4/4



1. An ~~ionic compound~~ chemical is a compound of two or more elements.
2. The Lewis Dot Diagram for calcium is Ca• .
3. A carbon dioxide molecule has a ~~straight? triangular~~ linear shape.
4. Sodium is isoelectronic with ~~Chlorine~~ neon.
5. A(n) Polar - Covalent bond is formed when the difference in electronegativity is less than 1.67.
6. Beryllium will have a total of 2 valence electrons when it is covalently bonded with other atoms.

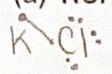
2/6

Alexis Williams

35 Thinking 16/17 = 94 32.9

5. Determine the type of bond formed in each of the following. Show your work. **(3 marks)**

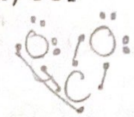
(a) KCl



ionic

metal + non-metal

(b) CO_2

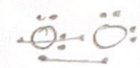


carbon dioxide

covalent

non + non

(c) O_2



dioxygen

covelent

non non.

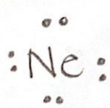
6. The following table lists some of the properties of three different compounds: glucose ($C_6H_{12}O_6$), potassium iodide (KI), and chlorine (Cl_2). Match the compound with its properties. **(3 marks)**

Substance	State at Ambient Temperature	Electrical Conductivity	Melting Point (°C)	Answer
A	gas	No	185	Cl_2
B	solid	No	115	$C_6H_{12}O_6$
C	solid	yes	968	KI

7. For nonmetals only, what is the relationship between electron affinity and reactivity? **(2 marks)**

Electron affinity requires more energy at first ionization & are least reactive because they have the most electrons out of ~~majority~~ of the metals. with more electrons they are more stable & require most energy.

8. Using a diagram explain why neon is such a non reactive element (4 marks)



since neon has a full valence shell naturally it is a non-reactive element since it's a stable nobel gas.

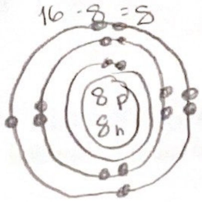
9. Discuss the 2 nuclides: Oxygen-16 and Oxygen-14 in terms of protons, neutrons and electrons. Write the chemical notation for each and state which nuclide would be more stable and why. **(5 marks)**

∴ 16 would be more stable than 14 because its shell is most full.

and 16 has a more full nucleus which is a greater

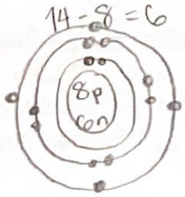
16

16 - 8 = 8



14

14 - 8 = 6



→ needs 4 more to be stable

↳ only needs 2 more to be stable.

march 6 2020

REVIEW
Chemistry
UNIT 1

Chapter 3

3.3 Polar Bonds & Polar Molecules

IONIC BONDS	COVALENT BONDS
• between ions	• between atoms
• transfer of e⁻ between atoms	• Sharing e⁻ between atoms
• large crystal lattice	• Individual molecules

use electromagnetivity numbers to find the difference & the difference determines bond type.

larger atom x − smaller atom x = EN

POLAR MOLECULE: uneven e⁻ distribution makes + at one end & − at the other

NON POLAR MOLECULE: mole-cule where e⁻ are equally distributed δ^+ & δ^-

3.4 Intermolecullar Forces

intermolecular force: Attraction between whole molecules

DIPOLE to DIPOLE FORCES: attraction at slightly positive end of a molecule to the slightly negitive end of another molecule

London dispersion force: weak attractive force between entities (non-polar, unbonded atoms) caused by temporary electron imbalanee by entities.

Vander Waals Forces: weak forces of attraction between molecules with dipole-dipole & London dispersion forces

3.5 Hydrogen Bonding & Water

special Properties of Water

- unusually high melting & Boiling points. 0°C 100°C
- Ice's lowdensity
- High surface tension
- water holds thermal energy well without temp change

Law of Definite Proportions.

1st Quiz

A .23 · 76 = 19.25

T .35 · 94 = 32.9

K .25 · 78 = 19.5

T .15 · 50 = 7.5

79.15

March 5 2020

REVIEW

Chemistry

UNIT 1

1.4 Isotopes Radioisotopes & Atomic Mass

Isotope: Atoms with same number of usual protons but different neutron number. (changing mass total#)

Isotopic abundance: percentage of isotopes in total to the other versions $\frac{mass(\%) + mass(\%)}{100}$ (formula)

Atomic Mass: Weighted average of all versions of an isotope (naturally occurring)

Radioisotopes: Isotopes that spontaneously decay & become radioactive

1.5 Periodic Table & Periodic law.

hydrogen: unique element with many phyiscal properties of a non-metal but chemical properties of a metal (s)

alkali metal: 1 soft, silver   alkali earth metals: light very reactive 2

transition metals: 3 - 12 ≈, high melting points, conductors.

nobel gases: stable, 8 (18) halogens: almost stable (7)(17).

metaloid: properties of metals & non-metals.

group: colum Period: row Periodic law: same # e^- in groups, properties

Elements can be represented by Bhor-Rutherford or Lewis

Ruther-ford → mass - Atomic = neutrons Atomic = Protons = e^-

1.7 Trends in the Periodic Table

Atomic Radius: Left + Down (increasing) 1st ionization: right + up (increasing)

electron Affinity: Change that occurs when an electron is added to a neutral atom in a gas state. (goes with/same as ionization)

1st ionization energy: quantity of energy required to remove an ~~atom~~ e^- from an atom or ion in a gaseous state.

Ionic radius: measurment of an atom after bonding; centre to outer

Chapter 2

2.1 Ionic Compounds

Chemical Bond: Force of Attraction between 2 atoms or ions.

Ionic Compounds: metalic w/ 1+ non-metalic; can conduct not as solids, high boiling

Ionic Bond: electrostatic force of attraction between + ion & - ion (chem bond)

Ionic compounds are in fixed ratios; smallest unit in crystals (formula unit)

USING LEWIS SYMBOLS → $\dot{Mg}\cdot + 2\cdot\ddot{Cl}: \longrightarrow [:\ddot{Cl}:]^- [Mg]^{2+} [:\ddot{Cl}:]^-$

2.2 Molecular Elements & Compounds

molecular Element: a pure substance 2+/more atoms of the same element

diatomic: made of 2 atoms

March 5 20

REVIEW

Chemistry

UNIT 1

Chapter 1

matter: anything that has mass and takes up space.

1.2 Atomic Structure

Dalton's theory: Particles, atoms: are all identical.

Thomson's theory: 1st electron = negitive

Rutherford's theory: Gold Experiment atom + electricity.

Chadwicks theory: neutron discovery

Bhor Rutherford: full combination; diagrams n°, e^{-}, p^{+}

Atomic Number: Protons in 1 element. unique to each one.

Mass Number: sum of everything in the nucleus (Mass - Atomic = neutro

Atomic Mass unit: $\frac{1}{12}$ a carbon atom

1.3 Ions & Octect Rule

full/stable octect: electron arrangement where valance is full at 8

octect rule: genral statement; atoms are stable @ 8 e^-

Ion: Charged entity when an atom gains or loses 1+ e^-

- cation: Positive ion; w/ loss of 1+ e^-
- anion: negitive ion; w/ gain of 1+ e^-

Valance: Charge of an ion, determined by gain, loss, share.

ex "determine the valance of NaCl."

ELEMENTS WITH MULTIVALANCE

copper	Cu	$(I)^{+1}$ $(II)^{+2}$
Iron	Fe	$(II)^{+2}$ $(III)^{+3}$
Tin	Sn	$(II)^{+2}$ $(IV)^{+4}$
Lead	Pb	$(II)^{+2}$ $(IV)^{+4}$
Manganese	Mn	$(II)^{+2}$ $(III)^{+3}$ $(IV)^{+4}$ $(VI)^{+6}$ $(VII)^{+7}$

Numerals are from IPUAC System.

↳ written as: Iron (II) Chloride

POLYATOMIC IONS

acetate	$C_2H_3O_2^-$	cromate	$Cr_2O_4^{2-}$	hydrogen Phosphite	HPO_3^{2-}		
bromate	BrO_3^-	dicromate	$Cr_2O_7^{2-}$	hydrogen Phosphate	HPO_4^{2-}		
carbonate	CO_3^{2-}	cyanide	CN^-	dihydrogen phosphite	$H_2PO_3^-$		
hydrogencarbonate	HCO_3^-	hydroxide	OH^-	dihydrogen Phosphate	$H_2PO_4^-$		
hypochlorite	ClO^-	iodate	IO_3^-	Sulfate	SO_4^{2-}	thiosulfate	S_2O_3
		permanganate	MnO_4^-	Sulfite	SO_3^{2-}	amonimmum	NH
chlorite	ClO_2^-	nitrate	NO_3^-	hydrogen Sulfide	HS^-		
Chlorate	ClO_3^-	nitrite	NO_2^-	hydrogen Sulfate	HSO_4^-		
Perchlorate	ClO_4^-	Phosphate	PO_4^{3-}	hydrogen Sulfite	HSO_3^-		

Hilary

February 28 2020

Chemistry

UNIT 1 REWIEW

TEMPERATURE

- is average keretic energy.
  - ↳ Hot is fast ↳ cold is slow
- THERMAL EXPANSION
  - ↳ Most objects expand when heated → structures built for this → All features expand

DENSITY

- Cubic measumment
- mass
  - ↳ comparing amount/weight to the amount of space
- $D = \frac{M}{V}$
- Water
  - ↳ 1 gram/cm^3

ARCHIMEDES PRINCIPAL

- submerging an ogject in water will give you it's volume.

final - origional = ob

Development of the Atom

LAW OF CONSERVATION

Mass is neither destroyed or created.

↳ balancing chem equ.

LAW OF DEFINITE PROPER

the fact that an element has the exact same properties reguardless of amount.

Atom Theories

- DALTON → All matter is particles
- DEMOCRITUS → GREEK
- THOMAS
- RUTHERFORD.

BHOR-RUTHERFORD

- Core = Positive
- electron = negitive
  - Diagram does not work after 1st 20 elements

PERIODIC TABLE

Atomic # → 6

Ionic charges (IF APPLICABLE) ← x x

C

CARBON

12.01 ← Atomic Mass

Protons = Atomic #

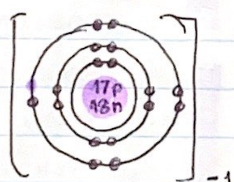
Neutrons = mass - Atomic #

Electrons = proton number

ATOM vs IONS

Atom - Pure & unaltered

Ion - atom w/more electrons



Chlorine ion

when there's additional e- there's anegitive charge.

when there's less e- there's a positive charge.

ISOTOPE

different # of neutrons than normal.

add → 1+ or more.

deficit → 1- or more

BALANCING

- Placing multiples in front of a chemical equ.
- ex $4Al + 3O_2 \longrightarrow 2Al_2\,2O_3$

4 Al on both sides & 6 O on both sides.

Chemical Bonding + Nature

SUBATOMIC PARTICLE	SYMBOL	ATOM LOCATION	CHARGE	MASS ≈ (kg)
electron	e^-	energy levels outside	-1	9.11×10^{-31}
proton	p^+	nucleus	+1	1.67×10^{-27}
neutron	n^0	nucleus	0	1.67×10^{-27}

MASS NUMBER

- Protons + neutrons
- whole number
- not on periodic table

ISOTOPES

- Atoms of 1 element w/different mass #'s
- P = element
  - n = isotope or atom.

AVERAGE ATOMIC

$$\frac{\text{Mass}(\%) + \text{mass}(\%)}{100}$$

MOLECULAR NOMENCLATURE

Non-Polar Covalent – share electrons completing 8 → no charges
Polar covalent – share electrons to complete 8 → w/charges.
Prefixes for mono, di, tri, tetra, penta, hexa, hepta, oxa, nona, decal -ide
1st is w/lest electrons.

LEWIS DOT DIAGRAMS
dots around an element
symbol → last valance
electron amount = dot #

METALIC CHARACTER
- metals - non-metals
- metaloids.
- transition 3-12

BLOCKS
- main group
- inner transition
- 12, 13, 14, 15, 16, 17, 18

PERIODIC LAW

Elements are arranged by increasing atomic number.

FAMILIES
similar e-

MELTING • BOILING POINT
- medial + outward
↳ centre = least reactive.

Atomic RADIUS
size of Atom
down + left

1st IONIZATION
energy to remove an e-
↳ increasing up & right

IONIC RADIUS
Cation – loose (smaller) Anion – gain (larger)

CHEMICAL Bonding

Attractive force between atoms or ions that bond together as a unit
Bonds are to decrease potential energy. & increase stability.

CHEMICAL FORMULA

IONIC
* formula unit (NaCl)
* can be continuously added onto the origional

COVALENT
* molecular formula (CO_2)
* additional molecules are on new molecules.

ION
* monoatomic
↳ 1 atom
* Polyatomic
↳ 2+ atoms
↓
Changes
↳ Covalent

IONIC
crystal lattice
solid high (melting point)
soluble in water
electricity conducting
e- metal to non-metal

COVALENT
e- are shared between non-metal
true molecules solid / liquid / gas
low melting point
not souble in water

BOND POLARITY & IONIC & COVALENT

Bonds clasified as ionic or covalent are determined by electromagnetivity

Ø-.4 Pure Covalent · .4 - 1.7 Polar Covalent · 1.7+ Ionic

ELECTROMEGNITIVITY
atom's attraction for an electron that's being shared.
ON PERIODIC TABLE
↳ up & right.

NON-POLAR COVALENT
e- equal sharing
symetrical

POLAR COVALENT
e- unequal sharing
Partial charge

LATTICE ENERGY measure of strength in bonds in ionic compound

When using Lewis Dot diagrams connect w/lines & arrows.

IONIC FORMULAS
① ION & Cation
② overall Ø charge
③ If cancel write symbols

MULTI-VALANCE
transition metals ions with diff possible charges

IONIC NAMING
① IONS, CATIONS
② -ide ③ Numerals (I) (II) (III)

-ate ending
1+ per- -ate
1- -ite 2- Hypo- -ite

Potential energy - object position.

Nick the Camel are a Calm Supper in Phoneix
NO_3^{-1} CO_3^{-2} ClO_3^{-1} SO_4^{-2} PO_4^{-3}

Octect Rule
↳ 8 valance

I Brought Clay For Our New House
Iodine I_2, Bromine Br_2, Chlorine Cl_2, Fluorine F_2, Oxygen, Nitrogen N_2, House – Hydrogen H_2

April 29/20
~~SCH 3U~~

STICHIOMETRY LAB

Alexis Williams

PRE-LAB QUESTIONS

1 $NaHCO_{3(s)} + HCl_{(aq)} \longrightarrow CO_{2(g)} + H_2O_{(g)} + NaCl_{(aq)}$

2 $NaHCO_{3(s)} + HCl_{(aq)} \longrightarrow CO_{2(g)} + H_2O_{(g)} + NaCl_{(aq)}$
? 82g 1.5g

$1.5g\,\cancel{NaCl} \times \frac{1\,mol\,NaCl}{58.44\,\cancel{g\,NaCl}} = .0256 \times 1.5 = .0257\ mol\ NaCl$ I forgot to multiply it

$82g\,\cancel{HCl} \times \frac{1\,mol\,HCl}{36.46\,\cancel{g\,HCl}} = 2.25\ mol\ HCl$

$n\,NaHCO_3 = .0256\ mol\,\cancel{NaHCO_3} \times \frac{1\,mol\,NaHCO_3}{1\,mol\,NaCl}$

$n\,NaHCO_3 = .0256\ mol$

$m\,NaHCO_3 = .0256\ \cancel{mol\,NaHCO_3}\left(\frac{84.01g}{1\,mol\,NaHCO_3}\right)$

$m\,NaHCO_3 = 2.156\,g$

3 During the exiprment water vapour will be formed. If additional water is added to the solid sodium hydrogen carbonate before the HCl is added then it can be heated & all moisture will leave the product. to weigh the product i would've had taken the weight of the empty beaker before & then after heating weigh the beaker with the remaining contents then subtract the beaker weight from the total weight wich leaves the weight of NaCl.

4 The H_2O & CO_2 will diffuse into the surrounding air from the reaction.

5 a) splattering of HCl with contact of $NaHCO_3$
* to prevent I will use a cover/watch glass.

b) the air/surroundingspace where the CO_2 & H_2O may go will change the weight.
* trap it with the watch glass.

c) to make sure all splattered NaCl is within the bottom of the beaker
* use water to wash excess splatter back into the bottom of the beaker.

6 a) scientific scale/balance b) beaker 2x c) watch glass d) protective equipment * eye gogles e) bunsen burner f) beaker stand for bunsen burner g) dropper. h) sodium hydrogen carbonate i) Hydrochloric Acid J) scoop k) water l) water wash bottle

7 1. put on protective equipment (goggles & apron)
2. Turn on the scale & measure the weight of the empty beaker, then remove the empty beaker & measure the weight of the watch glass & remove it. measure 82 g HCl into a different beaker
3. Put the beaker on the watch glass, zero the scale & measure the Sodium Hydrogen Carbonate to 2.156 grams.
4. Cover with the watch glass.
5. ~~measure~~ drop the 82 grams HCl with the dropper into the small hole created by the beak

7 6. when all the 82 grams of HCl are fully added into the beaker through the opening created by the watch glass & all splashing & fizzing has stopped complietley carefuly pick up the watchglass and rinse the splatters into the beaker by using the wash bottle with water.

7. when all the NaCl is rinsed off the ~~beaker~~ watch glass use the wash bottle with water again to rinse the beaker.

8. now all contents of the beaker are in the bottom place the beaker stand & beaker with its contents on the bunson burner & Place the watch glass back on.

9. once the burner is on ~~place~~ Place the temperature on low & wait until all liquid in the bottom & moisture on the watch glass is complietley gone. turn off the bunsen burner.

10. once all moistures left the beaker & watch glass and the glass ware has cooled remove it from the stand and place it on the scale. Record the weight, using the empty beaker & watch glass weight from step 1 subtract the excess weight of the equipment. This will give you the weight of the NaCl. Dispose by sink wash with soap & water.

11. Repeat steps 2 to 10 recording all weights & trial numbers.

12. once complete NaCl may be disposed of by the sink & equipment used must be washed with soap & water as well as surfaces.

POST-LAB QUESTIONS

1. The mole ratio of $NaHCO_3$ to NaCl is a 1:1 ratio because the equation is balanced.

2. $NaHCO_3$ is the limiting regagent because theres an excess of HCl.

3. No, because the small opening of the beaker where the dropper was put into may have been enough space for it to splatter out even a little bit so it may have not been a 100% yeild.