Chemistry 11 – Final Exam Study Guide

Unit 1: Safety in the Chemical Laboratory

- ensure that you understand all the basic safety rules to follow in the chemistry laboratory
- there will be a few safety questions on the final exam

Unit 2: Introduction to Chemistry

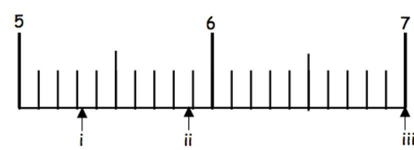
Prefix	Abbreviation	Exponent
giga	G	10^{9}
mega	M	10^{6}
kilo	k	10^{3}
hecto	h	10^{2}
deca	da	10^{1}
deci	d	10^{-1}
centi	c	10^{-2}
milli	m	10^{-3}
micro	µ	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}

UNIT CONVERSIONS

- you should know **kilo**, **centi**, **milli** and **micro**
- use conversion factors
- stop off at base unit (metre, litre, gram)
- always show your work
- always include the units
- (unknown) = (initial) X (conversion factor)
- DENSITY $d = \frac{m}{V}$

ACCURACY & PRECISION

- **significant figure** = a measured or meaningful digit; includes all certain digits and one uncertain
- all non-zero numbers significant
- leading zeros not significant
- trailing zeros significant if there is a decimal
- multiplication/division - round to lowest number of s.f.
- addition/subtraction - round to lowest number of decimal places
- hold all numbers in your calculator and round at the end
- **accurate** = close to correct or accepted value
- **precise** = more significant digits



Unit 3: The Physical Properties & Physical Changes of Substances

DEFINITIONS

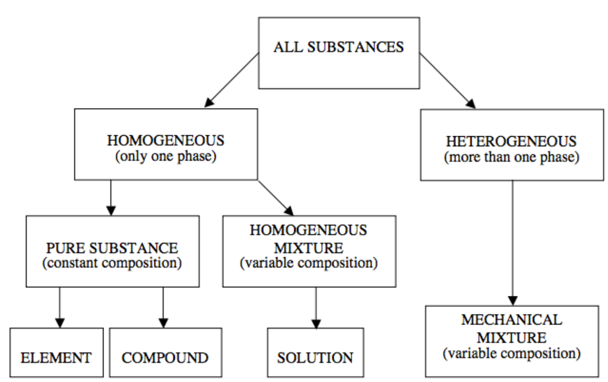
qualitative, quantitative, observation, interpretation, data, experiment, hypothesis, theory, law

PROPERTIES OF MATTER

matter, physical property, chemical property, solid, liquid, gas, aqueous, malleability, ductility, lustre, viscosity,

CLASSIFICATION OF MATTER

pure substance, mixture, homogeneous, heterogeneous, solution, solute, solvent, element, atom, molecule, ion, particle, compound



Unit 4: Inorganic Nomenclature

- metals, nonmetals, semiconductors/semi-metals/metalloids
- ions, anions, cations, polyatomic ions

IONIC COMPOUNDS

- metal and nonmetal
- criss-cross combining capacities, simplify; keep polyatomic ions together
- examples: potassium oxide, calcium phosphide, tin (IV) sulphate, iron (II) phosphate
- name metal and add “ide” to end of nonmetal
- Roman numerals for metals with >1 possible charge
- examples: CaF_2, Cu_2O, Ag_2SO_4, $Pb(SO_4)_2$

COVALENT COMPOUNDS

- two nonmetals
- use prefixes
- examples: CO, P_2O_5, SiF_6
- examples: phosphorus trichloride, tetrasulphide dinitride, disilicon hexaiodide

Common Prefixes Use in Chemical Nomenclature

Prefix	Meaning
Mono-	1
Di-	2
Tri	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Nona-	9
Deca-	10

PURE ELEMENTS

- diatomic elements end in “gen" including halogens

HYDRATES

- molecules that include water molecules in their crystal structure - ex. $CuSO_4 \bullet 5H_2O$
- use prefix and “hydrate” to indicate number of water molecules
- examples: $CoCl_2 \bullet 4H_2O$, $Al_2O_3 \bullet 3H_2O$

ACIDS

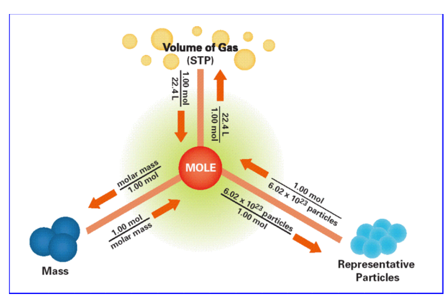
- start with “H”
- examples: HF, HNO_3, $HClO_2$

anion ending	acid ending	example
“ide"	hydro _____ ic	HCl = hydrochloric acid
“ate"	______ ic	H_2SO_4 = sulphuric acid
“ite"	______ ous	HNO_2 = nitrous acid

Unit 5: The Mole Concept

- **Avogadro’s Hypothesis** = equal volumes of different gases, at the same temperature and pressure, contain the same number of particles
- **mole** = the number of carbon atoms in exactly 12 g of carbon
- **molar mass** = the mass of one mole of particles
- molar mass of each element given in grams on periodic table
- for molar mass of a compound, add molar masses of atoms that make up compound
- **molar volume** = volume occupied by one mole of gas at STP = 22.7 L/mol
- **STP** = standard temperature & pressure = 0°C & 101.3 kPa
- **Avogadro’s number** = 6.02×10^{23} particles per mole

CONVERSION	CONVERSION FACTOR
MOLES ←→ NUMBER OF PARTICLES	$\frac{6.02 \times 10^{23}\ \text{particles}}{1\ \text{mol}}$ or $\frac{1\ \text{mol}}{6.02 \times 10^{23}\ \text{particles}}$
MOLES ←→ MASS	$\frac{\textit{(molar mass)}\ \text{g}}{1\ \text{mol}}$ or $\frac{1\ \text{mol}}{\textit{(molar mass)}}$
MOLES ←→ VOLUME (gases @ STP)	$\frac{22.4\ \text{L}}{1\ \text{mol}}$ or $\frac{1\ \text{mol}}{22.4\ \text{L}}$
MOLECULES ←→ ATOMS	$\frac{\textit{(atom count)}\ \text{atoms}}{1\ \text{mol}}$ or $\frac{1\ \text{mol}}{\textit{(atom count)}\ \text{atoms}}$



PERCENTAGE COMPOSITION

- assume 1 mole of compound
- determine mass of each and take % of total

EMPIRICAL FORMULA

- empirical = simplest formula
- assume 100 g; determine mass of each element; find moles of each; divide by smallest number

MOLECULAR FORMULA

- divide molar mass of molecular formula by molar mass of empirical formula

MOLAR CONCENTRATION - Solutions

- solution vocab: solution, solute, solvent, concentration, saturated, unsaturated, solubility
- **molarity** = molar concentration = moles/L = M = [solution]
- **dissociation** = separating previously existing ions in an ionic solid
  - ex. $NaCl(s) \longrightarrow Na^{+}(aq) + Cl^{-}(aq)$
- calculating [ions] in solution
  - write dissociation equation
  - use mole ratio

DILUTION CALCULATIONS

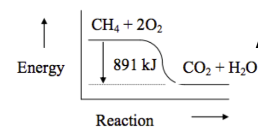
- mixing two solutions together, the concentration will change
- C1V1 = C2V2 where C = concentration and V = volume

Unit 6: Chemical Reactions

- signs: temperature change, change in colour, new phases formed
- open vs. closed system
- Law of Conservation of Mass
- balancing - number of each atom must be the same on each side
- writing formula equations - don't forget diatomic elements ('gen')

An **EXOTHERMIC** reaction gives off heat to the surroundings. (Heat **EX**its from the reaction.)

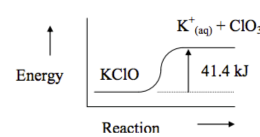
$CH_4 + 2O_2 \rightarrow CO_2 + H_2O + 891\ kJ$



ΔH = -891 kJ

An **ENDOTHERMIC** reaction absorbs heat from the surroundings. (Heat **EN**ters the reaction.)

$KClO_{3(s)} + 41.4\ kJ \rightarrow K^{+}_{(aq)} + ClO_3^{-}{}_{(aq)}$

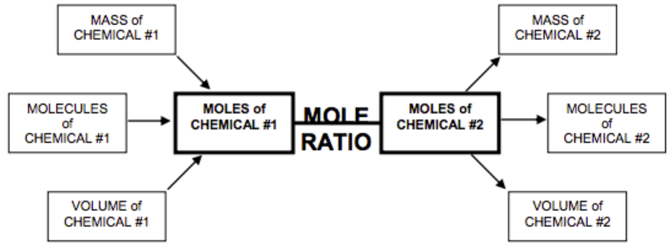


ΔH = 41.4 kJ

TYPE	HOW TO RECOGNIZE REACTANTS	HOW TO PREDICT PRODUCTS
Synthesis or Combination	2 elements	combine elements into one compound
Decomposition	1 compound	Break compound down into its elements
Single Replacement	Element + Compound	Interchange metals (or nonmetals) present
Double Replacement	Compound + Compound	Interchange positive ions in compounds
Neutralization	Acid + Base	Water is one product; remaining ions combine to form a salt
Combustion of Hydrocarbon	Compound starting with "C" + O_2	CO_2 + H_2O (if H present) + SO_2 (if S present)

- energy change can be determined using mole ratios

Unit 7: Stoichiometry



- **stoichiometry** = the relationship between the amounts of reactants used in a chemical reaction and the amounts of products produced by the reaction
- **coefficients** in a balanced equation give us the **MOLE RATIO**
- Step 1 - change to moles; Step 2 - mole ratio; Step 3 - change to desired units

TITRATIONS
- a measured amount of a solution is reacted with a known volume of another solution of unknown concentration
- **equivalence point** = the point in a titration where the ratio of the moles of each species involved exactly equals the a ratio of the coefficients of the species in the balanced equation
- same as any stoichiometry question

STOICHIOMETRY OF EXCESS QUANTITIES
- **limiting reactant** = reactant that sets a limit on the amount of product that can be formed (completely used up in a chemical reaction)
- **excess reactant** = reactant that is not completely used up in a chemical reaction
- ALWAYS use limiting reactant to determine other amounts of products formed or reactants used

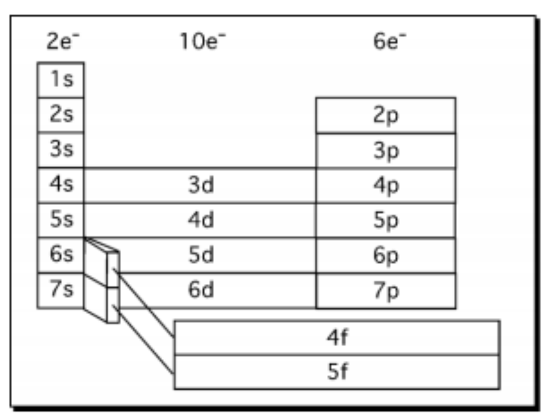
PERCENT YIELD
- amount of product actually obtained as a percentage of the expected amount
- theoretical yield - comes from stoichiometry calculation

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \text{ percent}$$

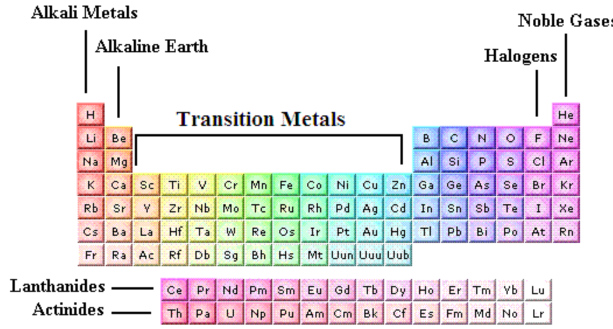
Unit 8: Atoms & the Periodic Table

ATOMIC THEORY
- John **Dalton** - solid sphere; atoms of different elements have different properties
- J.J. **Thomson** - plum pudding model; ball of + charge with random - charges throughout
- Sir E. **Rutherford** - small + nucleus surrounded by cloud of - charged electrons
- Niels **Bohr** - electrons in fixed energy orbits
- **quantum mechanics** - e- location based on mathematical calculation of probability
- **orbital** = actual region of space occupied by an electron in a particular energy level
- **electron configuration** = describes how electrons arranged in an atom

ATOMIC NUMBER & ATOMIC MASS
- **atomic number** = number of protons in the nucleus
- **mass number** = number of protons + neutrons
- **isotopes** = atoms with same atomic number but different masses (different number of neutrons)
- **atomic mass** = weighted average of isotopes of an element
- protons = atomic number
- neutrons = mass number - protons
- electrons = protons - ionic charge



THE PERIODIC TABLE

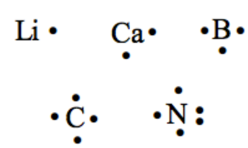


- created by Dmitri Mendeleev in 1869
- organized by atomic number
- **metals**: reflective, opaque, good conductors, ductile, malleable, solid at room T
- **nonmetals**: gases, liquids or brittle solids at room T, poor conductors,
- **semiconductor**: nonmetal having electrical conductivity which increases with T

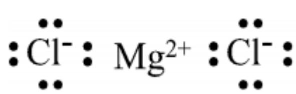
PERIODIC TRENDS
- **metallic character** - increases right to left and top to bottom
- **atomic radius** - decreases left to right; increases top to bottom
- **ionization energy** = energy to remove an electron from outermost shell of neutral atom; increases left to right, decreases top to bottom
- **electronegativity** = tendency of an atom to attract electrons from neighbouring atom; increases left to right, decreases top to bottom

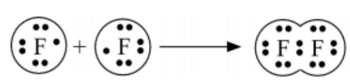
BONDING
- **ionic bond** - electron transfer; attraction between positive and negative ions; between metal and nonmetal; very strong; high melting points
- **covalent bond** - electrons shared equally; between nonmetals
- **octet rule** = atoms in in groups 14-17 form covalent bonds to have 8 electrons in outer shells
- **double bond** = sharing 4 electrons
- **triple bond** = sharing 6 electrons
- **polar covalent bond** = electrons shared unequally; one end slightly more negative (δ-) and one end slightly more positive (δ+)
- valence electron = electrons in outer s and p orbitals
- valence of an atom = number of unpaired electrons

LEWIS STRUCTURES



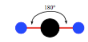
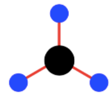
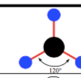
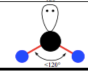
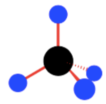
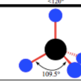
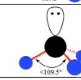
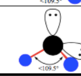
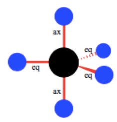
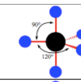
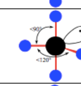
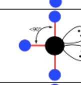
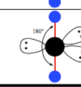
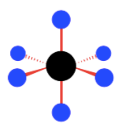
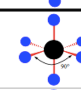
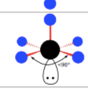
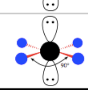
- electrons in outermost *s* and *p* orbitals shown with dots





- **intramolecular forces** - covalent bonds between atoms in a molecule; STRONG!
- **intermolecular forces = van der Waals forces** - hold molecules next to other molecules; WEAK!
- **dipole-dipole forces** - due to dipole on polar molecules; ex. **hydrogen bond**
- **London forces** - weakest intermolecular force; due to momentary dipoles
- **dipole** = partial separation of charge
  - a molecule is **polar** if is partial + charge at one end and partial - charge at other end
  - a molecule that is not polar is **nonpolar**
  - **polar** molecules must be **asymmetrical** (uneven)
  - a molecule that is **symmetrical** is **nonpolar**
- **like dissolves like**
  - polar/ionic solutes dissolve in polar solvents
  - nonpolar solutes dissolve in nonpolar solvents

- **VSEPR = Valence Shell Electron Pair Repulsion** - predicts molecular geometry by examining bonding and non-bonding electron pairs of electrons on a molecule

Electron Groups on central atom[1]	Electron-Group Shape	Bonds[2]	Lone Pairs	AX_mE_n[3]	Molecular Shape	Bond angles	Polarity	Hybrid -ization	Appearance
2	 Linear	2	0	AX_2	linear	180°	nonpolar[4]	sp	
3	 Trigonal Planar	3	0	AX_3	trigonal planar	120°	nonpolar[4]	sp^2	
		2	1	AX_2E	bent	<120°[5]	polar	sp^2	
4	 Tetrahedral	4	0	AX_4	tetrahedral	109.5°	nonpolar[4]	sp^3	
		3	1	AX_3E	trigonal pyramidal	<109.5°	polar	sp^3	
		2	2	AX_2E_2	bent	<109.5°	polar	sp^3	
5	 eq = equatorial ax = axial Trigonal Bipyramidal	5	0	AX_5	trigonal bipyramidal	120° eq 90° ax	nonpolar[4]	sp^3d	
		4	1	AX_4E	seesaw	<120° eq <90° ax	polar	sp^3d	
		3	2	AX_3E_2	T-shaped	<90°	polar	sp^3d	
		2	3	AX_2E_3	linear	180°	nonpolar[4]	sp^3d	
6	 Octahedral	6	0	AX_6	octahedral	90°	nonpolar[4]	sp^3d^2	
		5	1	AX_5E	square pyramidal	<90°	polar	sp^3d^2	
		4	2	AX_4E_2	square planar	90°	nonpolar[4]	sp^3d^2	

Unit 9: Organic Chemistry

- chemistry of **carbon** compounds
- **hydrocarbon** = compound containing only H and C
- each carbon forms four bonds
- **structural isomers** = compounds which have the same molecular formula but a different arrangement of atoms
- **saturated** = contains maximum number of H-atoms
- **unsaturated** = contain some double/triple bonds
- **cis-trans isomerization** = arrangement around double bond
- **'cis' isomer** = 2 groups on same side of double bond
- **'trans' isomer** = 2 groups on opposed sides of double bond
- **resonance** structure = differ in placement of double bonds
- **aromatic** molecule = contains one or more benzene rings

NAME	ENDING	FUNCTIONAL GROUP	NAME	ENDING	FUNCTIONAL GROUP
alkene	–ene	$C = C$	ether	–oxy	—O—
alkyne	–yne	$C \equiv C$	amine	amino	$—NH_2$
halide		–F, –Cl, –Br, –I	amide	–amide	$—CONH_2$
alcohol	–ol	—OH	carboxyllic acid	–oic acid	—COOH
aldehyde	–al	—CHO	ester	–oate	—COO—
ketone	–one	—CO—	aromatic ring	phenyl benzene