Chemistry 103, Dr. Hamers
REVIEW EXAM I, Chapters 1-4
Tony Jacob

CHAPTER 1 and *Let's Review*
SI UNIT CONVERSIONS

1ml = $1cm^3$	c = centi = 1×10^{-2}
1000L = $1m^3$	m = milli = 1×10^{-3}
T = tera = 1×10^{12}	μ = micro = 1×10^{-6}
G = giga = 1×10^{9}	n = nano = 1×10^{-9}
M = mega = 1×10^{6}	p = pico = 1×10^{-12}
k = kilo = 1×10^{3}	f = femto = 1×10^{-15}

1×10^{10}Å = 1m   or   1Å = 1×10^{-10}m = 1×10^{-8}cm = 100pm

TEMPERATURE CONVERSIONS: **K = 273 + ˚C**   [˚F = $^9/_5$˚C + 32]

DENSITY = $\frac{\textbf{MASS}}{\textbf{VOLUME}}$ units: g/ml (liquids), g/cm^3 (solids), or g/L (gases)

TERMS:

Matter, Phases of Matter (solid, liquid, gas)
Pure Substance & Mixture
Homogeneous Mixture & Heterogeneous Mixture
Element: a material that can't be decomposed into 2 or more substances; one of about 106 unique atoms on the periodic table (e.g., Ag, Fe, or Br_2)
Compound: 2 or more elements bonded together in fixed proportions (e.g., $KAl(SO_4)_2$ or NO); a Pure Substance that contains 2 or more elements
Atom: the smallest chemical substance that has the properties of that element
Molecule: smallest discrete element or compound (e.g., O_2(g) or H_2O(g)) that maintains the chemical characteristics of that species
Chemical Change & Physical Change (e.g., boiling)
Physical properties (intrinsic = independent on quantity: boiling pt, D; extrinsic = dependent on quantity: mass)

KINETIC MOLECULAR THEORY: All matter consists of tiny particles in constant motion; *Solids*: the particles are packed in close array; particles vibrate but do not move; *Liquids and Gases*: particles are arranged randomly and can move past one another; Liquids - particles are close to one another so they can't be compressed; Gases - particles are far apart so they can be compressed; always fill the container they're in.

DIMENSIONAL ANALYSIS (UNIT CONVERSIONS)

SIG FIGS

CHAPTER 2
ATOMIC STRUCTURE

Electrons (-1), protons (+1), neutrons (0); Relative masses of atomic particles (n ≈ p^+; e^- ≈ $^1/_{2000}p^+$ or n);
nucleus = p^+ + n; mass n = 1.675×10^{-24}g > mass p^+ = 1.673×10^{-24}g >> mass e^- = 9.109×10^{-28}g;
size of atom = size of electron cloud = ~1Å or ~100pm or ~1×10^{-10}m or ~1×10^{-8}cm

Elements: <u>*defined*</u> by the #<u>*protons*</u> they contain; **Charge** determined by comparing the #e^- to the #p^+;
Mass determined by number of n + p^+ (e^- mass is too small)

PERIODIC TABLE: families: alkali metals, alkaline earth metals, noble gases, halogens, chalcogens; metals, nonmetals, metalloids; periods = rows; families or groups = columns;

Allotropes: the naturally occurring forms of an element; e.g., C: diamond, graphite, Buckyballs; O: O_2 and O_3;

Diatomic Elements: F_2, Cl_2, Br_2, I_2, N_2, O_2, H_2

ISOTOPES: $^{A}_{Z}\mathbf{Symbol}$: where Symbol = atomic symbol, A = mass number (#n + $\#p^+$), and Z = atomic number ($\#p^+$); know how to determine #n, $\#p^+$, and $\#e^-$ in an isotope

$1\ amu \equiv \dfrac{mass\ of\ 1\ atom\ of\ ^{12}C}{12}$; 12 amu ≡ mass of 1 atom of ^{12}C; 1 amu = 1.66×10^{-27}kg

H isotopes: $^{1}_{1}H$, $^{2}_{1}H = ^{2}_{1}D = D$, $^{3}_{1}H = ^{3}_{1}T = T$; D = deuterium; T = tritium

AVERAGE ATOMIC WEIGHT (AAW) = (M1)(RA1) + (M2)(RA2) + …

M = mass, RA = relative abundance (decimal); for elements with only 2 isotopes: **RA1 + RA2 = 1**

When 2 masses and AAW are given and solving for RA1 and RA2, then let RA1 = x and RA2 = 1-**x**

TYPES OF COMPOUNDS

Ionic Compounds: metal + nonmetal (can sub polyatomic ion); *electrons transferred*; held together by electrostatic forces (+ and -); solids; high bp/mp;

Polyatomic Ions (memorize): these compounds should **never** be broken up;

NH_4^+, NO_3^-, SO_4^{-2}, OH^-, PO_4^{-3}, CO_3^{-2}, $C_2H_3O_2^-$, etc. (see Table 2.4 + pg 77, or see last page of handout)

Molecular Compounds: contain 2 nonmetal elements; examples: CO, CO_2, N_2O_4; *electrons shared in bonds*; low bp/mp

COULOMB'S LAW: Describes the attractive force between a positive (cation) and negative (anion) ion – that is, in ionic compounds

$$\textbf{Force of Attraction} = \mathbf{k}\frac{(\mathbf{Q_1})(\mathbf{Q_2})}{\mathbf{r^2}}$$

where Q_1 and Q_2 are the charges on the ions, and r is the distance between the centers of the nuclei

ionic charges ↑ or ionic radii ↓ ⇒ Force ↑ ⇒ mp/bp ↑

NAMING COMPOUNDS

Ionic Compounds – metal + nonmetal (can substitute polyatomic ion for one or both)

- Metal & Nonmetal: Metal name & Nonmetal stem+"ide"
- Metal & Polyatomic Ion: Metal name & Polyatomic ion name
- Transition Metal: Metal(charge in Roman numerals) & Nonmetal stem+"ide" (or polyatomic ion name)

Molecular Compounds – 2 nonmetals

- prefix(only if > 1)+1st nonmetal name & prefix(always)+2nd nonmetal stem+"ide"

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

WRITING FORMULAS FROM NAMES

Ionic Compounds – balance anion and cation charges

1. Metal name or polyatomic ion ⇒ write formula including charge
   - With Roman numeral ⇒ charge from numeral (iron(III) = Fe^{+3})
   - Without Roman numeral ⇒ charge for polyatomic ion *memorized* (sulfate = SO_4^{-2}); use column in PT for charge of non-transition metal (calcium = Ca^{+2}, +2 because in column IIA)

   Group IA: +1 Group IIA: +2 Group IIIA: +3 Group VA: -3 Group VIA: -2 Group VIIA: -1
2. Nonmetal name or polyatomic ion ⇒ write formula including charge
   - Name ends with "ide" ⇒ *single* atom anion (sulf*ide* = S^{-2}; except: hydroxide (OH^-), cyanide (CN^-))
   - Name ends with "ite" or "ate" ⇒ *memorized* polyatomic ion (carbon*ate* = CO_3^{-2})
3. Balance charges within formula (calcium phosphate = Ca^{+2} and PO_4^{-3}; balance charge: $Ca_3(PO_4)_2$)

Molecular Compounds: translate name with prefixes; e.g., dinitrogen trioxide → N_2O_3

MOLE: 1 mole = 6.023×10^{23} particles (Avogadro's number, N_A)

MOLAR MASS = mass of 1 mol of any substance; sum of the individual average atomic weights; also called molecular weight, MW

MASS PERCENT

$$\textbf{mass\%A} = \frac{\textbf{massA}}{\textbf{total mass}} \textbf{ x 100\%} \text{ and } \textbf{mass\%A} = \frac{\textbf{\#molA x AW}_A}{\textbf{molar mass compound}} \textbf{ x 100\%} \text{ (AW = atomic weight)}$$

EMPIRICAL FORMULA (EF): simplest ratio of atoms in a chemical formula; ***determine EF from mass%***:

1. Assume 100g
2. % → g
3. g → mol (if g originally given and not mass% → start at step 3!)
4. write chemical formula; divide by smallest number of moles
5. fractions: $^1/_2$ (0.5) → x2; $^1/_3$ or $^2/_3$ (0.33, 0.66) → x 3; $^1/_4$ or $^3/_4$ (0.25, 0.75) → x 4

MOLECULAR FORMULA (MF): molecular formula obtained by multiplying EF by ratio of molar masses:

$$\frac{\textbf{molar mass}_{\textbf{MF}}}{\textbf{molar mass}_{\textbf{EF}}} = \textbf{ratio}; \quad \textbf{EF x ratio} = \textbf{MF}$$ where EF = empirical formula and MF = molecular formula

HYDRATED COMPOUNDS: compounds that have waters attached; include water in molar mass

$$\textbf{\#H}_2\textbf{O} = \frac{\textbf{mol H}_2\textbf{O}}{\textbf{mol compd}_{\textbf{no H}_2\textbf{O}}}$$ where

$$\text{mol } H_2O = \frac{\text{mass}_{\text{hydrated compd}} - \text{mass}_{\text{anhydrous compd}}}{\text{molar mass } H_2O} \quad \text{and} \quad \text{mol compd}_{\text{no } H_2O} = \frac{\text{mass compd}_{\text{no } H_2O}}{\text{molar mass compd}_{\text{no } H_2O}}$$

CHAPTER 3
EQUILIBRIUM

Product-favored: most of the reactants form product; little of the reactants remain
Reactant-favored: products not readily formed; reactants react only a little
Dynamic equilibrium: macroscopically no visible change, microscopically chemicals are reacting forming products and products are reacting to re-form reactants; *reaction is going in both directions*

Hints (in general; there may be exceptions!)

- Reaction contains strong acid or strong base: Other side is favored
- Reaction contains weak acid or weak base (not both): Side containing weak acid or weak base is favored
- Reaction forms a precipitate: Side with the precipitate (solid) is favored
- Combustion reaction: Side with CO_2 and H_2O is favored

TYPES OF REACTIONS

Acid/Base (A/B): acid(aq) + base(aq) → salt(aq) + H_2O(l)
Precipitation (ppt): MA(aq) + NB(aq) → MB(aq) + NA(s)   (2 aqueous solutions form a solid)
Gas forming: HCO_3^-(aq) or CO_3^{-2}(aq) + H^+(acid(aq)) → H_2CO_3(aq) → H_2O(l) + CO_2(g)
Combustion: $C_xH_y(O_z)$ + O_2(g) → CO_2(g) + H_2O(g)

WRITING REACTIONS

1. Molecular Reaction: Compounds in molecular form; no ions. Steps for A/B, ppt, and gas forming reactions.

a. Names → formulas
b. Assign charges and balance reactant formulas
c. Switch partners; use only 1 of each ion initially
d. Balance product formulas
e. Balance reaction

2. Complete Ionic Reaction: Break chemicals into ions: Strong acids, Strong bases, Soluble ionic compds
Exception: H_2CO_3(aq) → H_2O(l) + CO_2(g); NH_4OH and $HC_2H_3O_2$ – don't break up

Break up

1. strong acids (HCl, etc.)
2. strong bases (NaOH, etc.)
3. soluble ionic compounds ($NaNO_3$, etc.)

Don't break up

1. weak acids ($HC_2H_3O_2$, etc.)
2. weak bases (NH_3, etc.)
3. insoluble ionic compounds (AgCl, etc.)
4. molecular compounds (CO_2, etc.)

3. Net Ionic Reaction: Cross out ions that appear on both sides of the reaction; these are called Spectator Ions
If everything cancels out → No reaction!; don't forget phases: (s), (l), (g), (aq)

Strong acid + Strong base net ionic reaction: H^+(aq) + OH^-(aq) → H_2O(l) (exception: H_2SO_4 + $Ba(OH)_2$)

Strong Acids (memorize): HCl (hydrochloric acid), HBr (hydrobromic acid), HI (hydroiodic acid), HNO_3 (nitric acid), H_2SO_4 (sulfuric acid), $HClO_4$ (perchloric acid)

Weak Acids (memorize acetic acid): some common: $HC_2H_3O_2$ (acetic acid) = CH_3COOH = CH_3CO_2H, H_3PO_4 (phosphoric acid), HF (hydrofluoric acid), HCN (hydrocyanic acid), H_2CO_3 (carbonic acid)

Strong Bases (memorize): LiOH, NaOH, KOH, $Ba(OH)_2$ (sometimes: $Sr(OH)_2$)

Weak Bases (memorize ammonia): NH_3 or NH_4OH = ammonium hydroxide; this is NH_3 in water

Solubility Rules: Soluble: Na^+; K^+; NH_4^+ (not: NH_4OH); NO_3^-; $C_2H_3O_2^-$ (not: $HC_2H_3O_2$); ClO_3^-; ClO_4^-; SO_4^{-2} (not: Ca^{+2}, Sr^{+2}, Ba^{+2}, Pb^{+2}); Cl^-, Br^-, I^- (not: Ag^+, Hg_2^{+2}, Pb^{+2})

Polyatomic ions (memorize): NH_4^+, NO_3^-, SO_4^{-2}, OH^-, PO_4^{-3}, CO_3^{-2}, HCO_3^-, $C_2H_3O_2^-$, etc. (see text)

ELECTROLYTES – chemicals that produce ions in solution

Strong – strong acids, strong bases, soluble ionic compounds
Weak – weak acids, weak bases
Non – insoluble ionic compounds, molecular compounds

OTHER REACTIONS

Oxides of nonmetals + $H_2O(l) \rightarrow$ Acid Example: $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

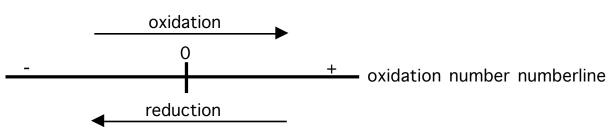
Oxides of metals + $H_2O(l) \rightarrow$ Base Example: $MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$

REDOX

Assigning Oxidation Numbers

1. Elements in elemental form 0
2. In a *compound*:
    - a. Group 1A (Li,Na,...) +1
    - b. Group 2A (Be, Mg,...) +2
    - c. F -1
    - d. H +1 (can be –1 in MH_x compds; e.g., NaH)
    - e. O -2 (can be –1 in O_2^{-2} compds; can be $-{}^1/_2$ in O_2^- compds; e.g., H_2O_2, KO_2)
3. Sum Rule: Sum of all the oxidation numbers = total charge on compound

Oxidized, reduced, oxidizing agent, reducing agent (reactants only)
Redox reaction occurs when oxidation numbers change in a reaction



CHAPTER 4
BALANCE REACTIONS

STOICHIOMETRIC CALCULATIONS

LIMITING REAGENTS: the chemical that runs out first; limiting reagent problems can be identified when 2 reactant quantities are given; use each reactant to find out how much product can be made; the smaller amount is the amount that can be made and the reactant that produces the smaller amount is the limiting reagent

PERCENT YIELD: $\%\text{yield} = \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$ (actual yield is given in the problem)

EMPIRICAL FORMULA determined from mass CO_2 and H_2O

A. Compound contains C and H only

1. g CO_2 → mol CO_2 → **mol C**
2. g H_2O → mol H_2O → **mol H**
3. Write formula and divide by smallest #moles
4. If needed, fractions: $^1/_2$ (0.5) → x 2; $^1/_3$ or $^2/_3$ (0.33, 0.66) → x 3; $^1/_4$ or $^3/_4$ (0.25, 0.75) → x 4

B. Compound contains C, H, and O

1. g CO_2 → mol CO_2 → **mol C** → **gC** (need both mol C and g C)
2. g H_2O → mol H_2O → **mol H → gH** (need both mol H and g H)
3. g O from: total g sample = g C + g H + g O (g O = total g sample - g C - g H)
4. g O → **mol O**
5. Write formula and divide by smallest #moles
6. If needed, fractions: $^1/_2$ (0.5) → x 2; $^1/_3$ or $^2/_3$ (0.33, 0.66) → x 3; $^1/_4$ or $^3/_4$ (0.25, 0.75) → x 4

EMPIRICAL FORMULA determined from other reactions

MOLARITY

- *M calculations*
- *Dilutions*
- *Stoichiometry*

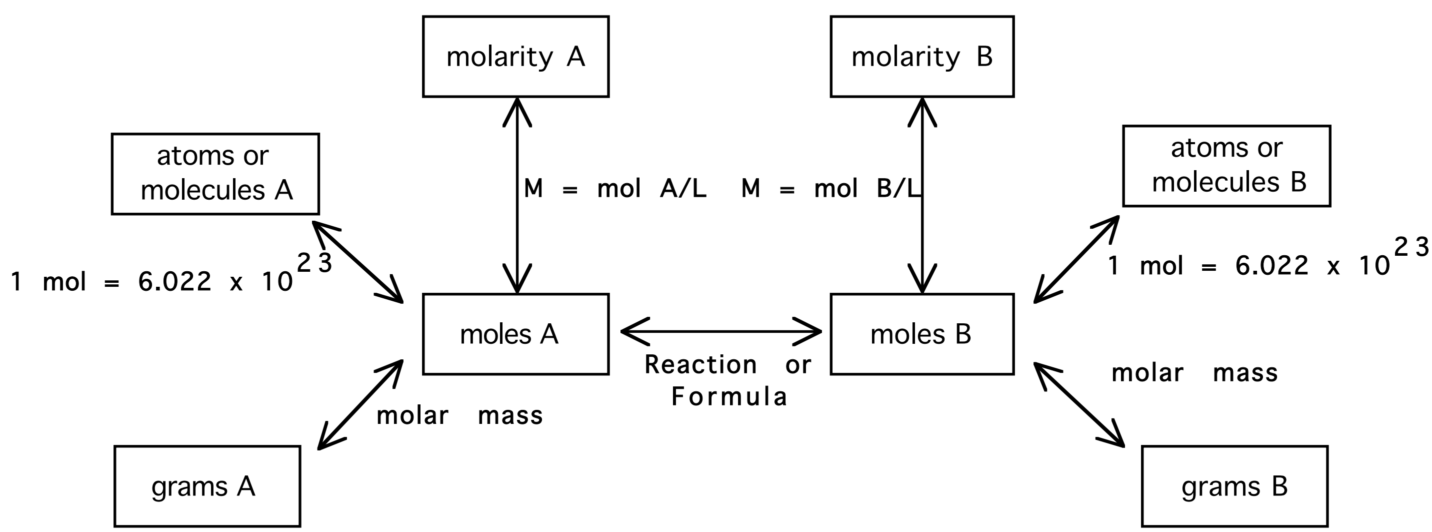
Solute: The chemical that is dissolved into the solvent
Solvent: The liquid that the solute is dissolved into
Solution: Solvent + Solute

$M = \frac{\text{moles}}{L}$ moles = MxL $L = \frac{\text{moles}}{M}$

Dilution - this equation is only used with dilutions, not reactions!

$\mathbf{M_1 \times V_1 = M_2 \times V_2}$ (1 = initial, 2 = final; V_2 = total volume = V_1 + water added)

Neutralization: acid completely reacts with base (referred to as a titration with non acid/base reactions)



Polyatomic Ions, Acids, Common Names

If polyatomic ion name ends with "ate" then acid (adding H^+'s until the ion is neutral) name ends with "ic acid"

If polyatomic ion name ends with "ite" then acid (adding H^+'s until the ion is neutral) name ends with "ous acid"

If ion is a single atom (Cl^-) then the ion name ends with "ide" (exceptions: CN^-, cyanide, and OH^-, hydroxide) and acid name is "hydro+stem+'ic' acid"

NH_4^+ (ammonium)

SO_4^{-2} (sulfate)	HSO_4^- (hydrogen sulfate)	H_2SO_4 (sulfuric acid)
SO_3^{-2} (sulfite)	HSO_3^- (hydrogen sulfite)	H_2SO_3 (sulfurous acid)
NO_3^- (nitrate)		HNO_3 (nitric acid)
NO_2^- (nitrite)		HNO_2 nitrous acid)
CO_3^{-2} (carbonate)	HCO_3^- (hydrogen carbonate)	H_2CO_3 (carbonic acid)
$C_2H_3O_2^-$ (acetate)		$HC_2H_3O_2$ (acetic acid)
ClO_4^- (perchlorate)		$HClO_4$ (perchloric acid)
ClO_3^- (chlorate)		$HClO_3$ (chloric acid)
ClO_2^- (chlorite)		$HClO_2$ (chlorous acid)
ClO^- (hypochlorite)		$HClO$ (hypochlorous acid)
CN^- (cyanide)		HCN (hydrocyanic acid)
OH^- (hydroxide)		
MnO_4^- (permanganate)		
CrO_4^{-2} (chromate)		
$Cr_2O_7^{-2}$ (dichromate)		

PO_4^{-3} (phosphate)	HPO_4^{-2} (monohydrogen phosphate)	$H_2PO_4^-$ (dihydrogen phosphate)	H_3PO_4 (phosphoric acid)
PO_3^{-3} (phosphite)	HPO_3^{-2} (monohydrogen phosphite)	$H_2PO_3^-$ (dihydrogen phosphite)	H_3PO_3 (phosphorous acid)

F^- (fluoride)	HF (hydrofluoric acid)
Cl^- (chloride)	HCl (hydrochloric acid)
Br^- (bromide)	HBr (hydrobromic acid)
I^- (iodide)	HI (hydroiodic acid)
O^{-2} (oxide)	
S^{-2} (sulfide)	
Se^{-2} (selenide)	
N^{-3} (nitride)	
P^{-3} (phosphide)	
As^{-3} (arsenide)	
C^{-4} (carbide)	

Common Names

NH_3 (ammonia)	CH_4 (methane)	CH_3OH (methanol)
H_2O_2 (hydrogen peroxide)	$C_2H_6 = CH_3CH_3$ (ethane)	CH_3CH_2OH (ethanol)