Avogadros # = 6.023 x 10^{23}
Amphoteric = substances can be either an acid or a base. ie water.
Anhydrous = Without water; especially without water of crystallization [ant: hydrated]
Electron Pair Geometry =
Endothermic = Reaction absorbing heat. Ie increase in temperature
Add Kj on the LHS(table 16.2 p503 Zumdahl)
Exothermic = a reaction that produces heat. Reduction in temperature
Add Kj on the RHS (table 16.2 p503 Zumdahl)
Ion = An atom that is charged because it looses or gains an electron.
Ionisation Energy = The quantity of energy required for an electron to be removed
from a gaseous atom.
Molecular Geometry =
Monatomic ions =
Percentage Yield = The real amount of product as a percentage of what the calculated amount should be.
R = 0.08206 L atm K-1 mol-1
Significant Figures is using the smallest number of decimals in the question to the
problem.

Example

$$\overset{Mass\#\mathbf{A}}{Atomic\#\mathbf{Z}}\mathbf{X} \qquad \overset{Mass\#\mathbf{13}}{Atomic\#\mathbf{6}}\mathbf{C}$$

$$\#Protons = 6$$
$$\#Neutrons = (A - Z) = (13 - 6) = 7$$
$$\#Electrons = Protons \cdot in \cdot neutral \cdot atom = 6$$

From Block8Lecture3LiquidsSolidsSolutionsAfterB.pdf
and n = cV

$$c = \frac{n}{V}$$

c = M = Solution Concentration
n = mol of Solute
V = Litres of Solvent

C = M = mol L^{-1} or mol/L or molar
Mass of a liquid
Volume x Density = Mass
V x D = m
Solution Dilution.
$n_1 = n_2$
where n_1 is the # moles in the solution
and n_2 is the # mol in the diluted solution.
so n = cV
therefore
$c_1 \times V_1 = c_2 \times V_2$

p429 Finding the Mass of a Solute when given as a percentage composition

$$MassPercent = \frac{gramsOfSolute}{GramsOfSolvent + gramsOfSolute} \times 100\%$$

STP means Standard Temperature and Pressure.

- Standard temperature is equal to 0 °C, which is 273.15 K.

- Standard pressure is equal to 1 atm.
R = 0.08206 L atm K-1 mol-1
- R is calculated for 1 mole of ideal gas at standard temperature and pressure (STP)
STP for gases is 0 °C (273.15 K) and 1 atm
Molar volume - at STP 1 mole of ideal gas occupies 22.4L

Gases
α = Proportional to
k = constant
Boyles

$$V\alpha \frac{1}{P}$$

P1V1 = P2V2

Charles V α T
Avogadros V α n

All together makes

$$V\alpha \frac{nT}{P}$$

Description of gas pressure (P) need this for the exam

$$V = R\frac{nT}{P}$$

R = 0.08206 L atm K-1 mol-1
rearranges to

$$R = \frac{PV}{nT} \qquad PV = nRT$$

but if n is constant use

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \qquad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Daltons Law of Partial pressures
$P_T = P_1 + P_2 + P_3 + ...$

$$P_{Total} = n_{Total}\frac{RT}{V}$$

Chemical Bonding
Dispersion < Dipole < Hydrogen << Ionic<<Covalent
COVALENT BOND
Bond formed by sharing of electrons by the nuclei of two atoms
Molar heat of fusion is the energy required to melt 1 mol of a substance.
Ie ice is 6.02kj/mol.
Molar heat of vaporisation is the energy required to turn 1 mol of a substance into gas.
Water is 40.6kj/Mol at 100°C
When atm pressure = vapour pressure the water will boil.
INTRAmolecular – bond between atoms
>> is greater than
INTERmolecular – bond between molecules
Valence Electron Orbital Boxes

1 box	**S**
3 boxes	**SP**
5 boxes	**SPD**
7 boxes	**SPDF**

Dipole – Dipole forces are only about 1% as strong as covalent or ionic bonds.
Hydrogen bonding to F, O, or N is higher than other dipole bonding because the electro negativity is greater.
Acids and Bases
Bases = Alkalis produce hydroxide Ions H+(aq) when dissolved in water or H_3O
HCl
Acids produce hydrogen Ions
NaOH

Amphoteric substances can be either an acid or a base. ie water.
Acids Neutralise bases
•Acid: proton (H+) donor (proton because a H+ ion is just a proton)
•Base: proton acceptor
Acids are <7
Bases are >7
acid + base → salt + water
HCl(aq) + NaOH(aq) → NaCl(aq) + H2O(l)
Thus
H^+(aq) + OH^-(aq) → H2O
Conjugate Acid Base Pairs
Am acidic molecule will lose one of its hydrogen Protons and thus become its Conjugate base. Eg. HCl and Cl.
To find the pH from the hydrogen ion concentration of a solution

$$pH = -\log[H+]$$

To find the hydrogen ion concentration from the pH

$$[H^+] = inv\log[-pH]$$

To find the OH
pOH = -log10[OH–]
pH + pOH = 14
To find pH from pOH
14 – pOH = pH
Pg320 Electronegativity table
Increasing to the right →
Decreasing down.
Lewis Structures
Figure out VE(Valence Electrons) and TBH (To Be Happy) amount of electrons.
H needs only 2 TBH the rest are 8 e TBH
To find No of bonds, minus the totals and divide by 2

$$\#ofBonds = \frac{TBH - VE}{2}$$

The spare Electron can be put outside the brackets
p491 **Equilibrium Constant**
eg. $H_{2(g)} + F_{2(g)} \leftrightarrows 2HF_{(g)}$

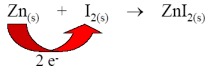
K= The Equilibrium Constant
Where the number of mol is the power The

$$K = \frac{[HF]^2}{[H_2][F_2]} \qquad K = \frac{[Product]}{[Reactant]}$$

substances in the square brackets are the concentration (mol L^{-1}) This a temperature dependent reaction.
Liquid and **solid** states are not included in this equation.
Homogeneous equilibria are when all substances are the same state
Heterogeneous equilibria are when substances may have different states
When using this for acids.

PEC140 :
Dion Patelis – 30505542
dion@7rocks.com

If K > 1 it is a forward reaction and there is more product produced If K < 1 it is a backward reaction and there is more reactant kept. K basically says what rates solutes dissolve in solutions.
K_{sp} = the solubility product constant
Le Châtelier's Principle p503
From pec140AlanKnightSB9L3aChateliersPrinciple.pdf
In order for K to remain constant the levels of the products or reactants will change.
Thus adding more F in the above reaction will mean that HF will go up or H will go down to keep K constant.
Pressure – When the pressure decreases there are less gas molecules. Thus a forward reaction is favoured.
For Example (p500) :
- When the volume decreases the pressure increases.
- Thus because there is more pressure the particle collisions are more frequent.
- Because particle collisions are more frequent molecules will bond more often.
- The volume of a bonded molecule is usually less than the individual atoms (sometimes molecules) therefore it's pressure would be less.
- Inturn the reaction favours a move to the side of the equation with less volume
Temperature -
Enthalpy – the energy stored in chemical bonds.
Exothermic – Release of Heat – forward reaction
Endothermic – Absorption of Heat – reverse reaction
Redox reactions.
OILRIG – Oxidation is Loss, Reduction is Gain
are simultaneous reactions. One can not occur without the other.
Oxidisation is the loss of usually hydrogen (or oxygen) atoms. (ie loss of electrons)
Reduction is the gain of hydrogen atoms (ie, gain of electrons)



•Zn as a reductant (or reducing agent) because it has caused the I2 to be reduced, and the
•I2 as an oxidant (or oxidising agent) because it has caused the Zn to be oxidised.

P524 Oxidisation State (Number) is the charge of a molecule. Pure Elements are neutral.

Rules for Assigning Oxidation States

1. For elements in their elemental state: oxidation number = 0
   Eg. Cl2: oxid'n no. of Cl = 0
2. For monatomic ions: oxidation number = charge on the ion.
   Eg. $CuBr_2$: In this compound, Cu exists as Cu^{2+} and Br exists as Br^-, therefore oxid'n no. of Cu = +2, Br = -1
3. For combined O: oxidation number = -2 (except for peroxides where it is -1 and in F O where it is +2). Eg. CO_2: oxid'n no. of O = -2
4. For combined H: oxidation number = +1 (except in metal hydrides where it is -1). Eg. CH_4: oxid'n no. of H = +1.
5. For polyatomic species: sum of the oxidation numbers of all the atoms equals the charge on the species. Eg. CH_4: oxid'n no. of H = +1, C = -4
   NO^{3-}: oxid'n no. for O = -2, N = +5
6. For binary species, the element with the higher electronegativity is assigned an oxidation number equal to its charge as an anion.
   Eg. CF_4: oxid'n no. of F = -1, C = +4
7. For an ionic species, the sum of the oxidization states must equal the overall charge. Eg CO_3^{2-} +4 + 3(-2) = -2

1. elements = 0
2. monatomic = charge on ion
3. O = -2 except peroxides
4. H = +1 except in metallic hydrides
5. polyatomic and ionic = must equal the overall charge
6. binary = highest negative charge is equal to anion charge.

Using oxidation numbers to identify redox reactions
- SB10L2Redox.pdf page 22

• oxidation involves an increase in oxidation number, and
• reduction involves a decrease in oxidation number.

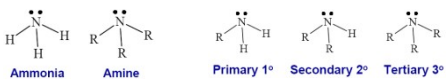
Steps to balancing redox reactions by half-equations -
SB10L3Redox.pdf

1. Separate the reaction into an oxidation half-equation and a reduction half-equation.
2. For each half-equation:
   a) Balance all the atoms except O and H.
   b) balance O by adding H_2O to the appropriate side.
   c) balance H by adding H+ to the appropriate side.
   d) balance the charge by putting electrons on the appropriate side. (The charge is the total charge on all the elements on one side.)
3. If needed, multiply one or both of the half-equations by an integer to give equal numbers of electrons in the two half-equations.
4. Add the half-equations and cancel common species on opposite sides.
5. Ensure the atoms and charges balance.

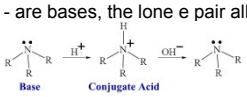
Organic Chemistry – remember the order by the 4th letter in the Alk…
Alkane – all single bonds
Alkene – a double bond somewhere in the chain
Alkyne – a triple bond somewhere in the chain

1.	Meth	6.	Hex
2.	Eth	7.	Hept
3.	Prop	8.	Oct
4.	But	9.	Non
5.	Pent	10.	Deca

R – Can represent any molecule. Used to represent a standard group to simplify expressions.
Functional Groups
Amines – suffix –'mine' organic derivative of Ammonia



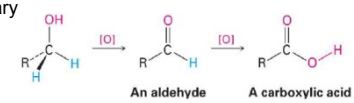
-Amines are available for hydrogen bonding
-fishy smell
- lower bp than alcohol as N is less electronegative than O.
- are bases, the lone e pair allows a covalent bond with a H



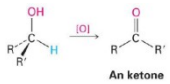
Alcohols –OH suffix 'ol'
Added to a standard Carbon chain CH_3-OH
Oxidisation of Alcohols –
- reagent potassium permanganate $KmnO_4$
- catalyst requied such as - sulfuric acid H_2SO_4

privilege

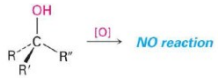
Primary



Secondary



Tertiary



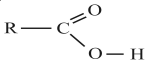
Carbonyl – 2 R groups single bonded to a C then double bonded to an O.

P602 **Keytones** suffix 'one' - RCORcarbonyl group appear in the middle of a carbon chain.

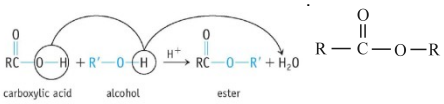
P602 **Aldehydes** suffix 'al' –COH Same as Keytones but the carbonyl group appears at the end of the chain.

Carboxyl group - -COOH – naming of this group must contain the carboxyl.

P605 **Carboxlic Acids** suffix 'oic' - RCOOH where R represents the Hydrocarbon **'diene'** group – what is this group???? Does it mean 2 sets of double bonds somewhere in the chain???

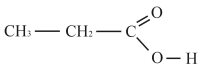


Ester – suffix – 'ate' or 'oate' – (also known as acid derivatives) Made by combining a Carbolic acid and an Alcohol. It needs an acid catalyst for the reaction

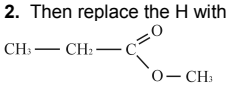


1. eg Methyl Propanoate.
Take your Carboxylic Acid (the Propanoate part) and put in the right amount of carbon atoms by replacing R.

PEC140 :
Dion Patelis – 30505542
dion@7rocks.com



2. Then replace the H with the (Methyl) part.



pg 137 Zumdahl **Table 5.4 Names of Common Polyatomic Ions –** With Wikipedia links

NH_4^+	Ammonium	CO_3^{2-}	Carbonate
NO_2^-	Nitrite	HCO_3^-	Hydrogen carbonate (bicarbonate)
NO_3^-	Nitrate	ClO^-	Hypochlorite
SO_3^{2-}	Sulfite	ClO_2^-	Chlorite
SO_4^{2-}	Sulfate	ClO_3^-	Chlorate
HSO_4^-	Hydrogen sulfate (bisulfate)	ClO_4^-	Perchlorate
OH^-	Hydroxide	CH_3COO^- or $C_2H_3O_2^-$	Acetate
CN^-	Cyanide	MnO_4^-	Permanganate
PO_4^{3-}	Phosphate	$Cr_2O_7^{2-}$	Dichromate
HPO_4^{2-}	Hydrogen phosphate	CrO_4^{2-}	Chromate
$H_2PO_4^-$	Dihydrogen phosphate	O_2^{2-}	Peroxide

P178 **SOLUBILITY RULES** of Ionic Compounds (Salts) in Water at 25°C
1. Most Nitrate NO_3 –, acetate (CH_3COO–) and perchlorate (ClO_4–) salts are soluble.
2. Most alkali metal (Li, Na+,K+) and NH_4+ salts are soluble.
3. Most Chloride salts are soluble. Notable exceptions are AgCl, $PbCl_2$ and Hg_2Cl_2 (mercury(I)) .
4. Most sulfate salts are soluble. Notable exceptions are $BaSO_4$, $PbSO_4$, $CaSO_4$.
5.Most Hydroxide compounds are only slightly soluble. The import exceptions are NaOH and KOH, $Ba(OH)_2$ and $Ca(OH)_2$ and $Ca(OH)_2$ are only moderately soluble.
6. Most sulfide (S^{2-}), carbonate (CO_3^{2-}), and phosphate (PO_4^{3-}) salts are only slightly soluble.

Solubility Summary
Soluble
Sulfate, perchlorate, acetate, Nitrate - Chloride, bromides, iodides, - Sodium, Ammonium, Potassium, Lithium.

Ate,ide,ium

SulPerAcNit, ChlorBriodide, SodAmPotLith.

Insoluble
Carbonate, phosphate, (Barium, Lead & Calcium Sulfates), - Hydroxide, Sulfide, (Silver,Lead & Mercury(I) Chlorides),

BarCalCarbPhosLead, MerHydroSulfSilLead

- The **bar** (symbol: **bar**), defined as 10^5 Pa *exactly*.
- The **atmosphere** (symbol: **atm**), defined as 101,325 Pa *exactly*.
- The **torr** (symbol: **Torr**), defined as 1/760 atm *exactly*. Ie 1atm = 760Torr

Examination Study Plan

- Go through Pritams notes about the actual exam and find memorise the methods I need to know.
1.0 Memorise ionic compound table.
1.1 Memorise solubility rules
2.0 Memorise oxidisation and half equations
3.0 Memorise Le Chateliers law
4.0 Find the quickest way to find limiting reagents.
5.0 Memorise Bromine tests of alkanes and alkenes.
- Memorise the bits Pritam said to first.
- Read the organic chem. section of the text book.
- Practise organic chem. questions.
- Go through the practise tests I've done so far and figure out why I went wrong

- Whilst going through everything, transfer everything I need to remember to the flash cards.
- Redo all the exams one by one within the time limits allocated so that I get good marks in them.
- Re go through SB10L3Redox.pdf pages 7 – 11 and page 529 of Zudahl.
- Green book pg58. Repractise these organic structures.
- Read up on how Catalysts work.
- Memorise patterns and naming extensions for Carboxillic acids (oic) and Esters (oate)
- Memorise groups 1 to 8 not including the transition metals or the nuclear type ones.
- Memorise how the orbitals work.
- revise pH calcs
- Boyles law?

Test Method.
- When doing the tests, Go through the whole lot just answering the ones you can see really quickly and you know you'll find super easy.
-Then come back and do the easy calculation ones.
-Then come back and do the trickier calculation ones.

Catalysts p484

- Activation Energy (E_a) – The energy required to break bonds in a substance so that it can make new bonds with another substance. Often just the particle collision provides enough energy to allow this to happen. Sometimes you need more.
- is something that reduces the activation energy required for the bond breaking.
- Can be an added chemical that does not get involved in the reaction.
- In the human body they are called enzymes.
- Energy stored in a chemical bond is called enthalpy (H).

Testing of various functional groups.

Alcohols - Test of oxidisation with Dichromate. – will go from orage to green
Or test with a Na_2CO_3 Sodium Carbonate. – no Reaction

Ketone - Test of oxidisation with Dichromate. – nothing will happen.

Alkane - Test with bromine and ultraviolet light – gradual loss of brown colour and product of HBr gas

Alkene - Test with Bromine – rapid loss of brown colour

Aldehydes - Dichromate – orange to green

Carboxylic Acids - Dichromate – no reaction. Or test with a Na_2CO_3 Sodium Carbonate. – gives a bubbling which is CO_2forming. Red on litmus paper.

Esters - Dichromate – no reaction. No positive test for an ester
At pec140 level. Look for the positive test of the other Group.

Amines - Alan knight said we don't have to test for it specifically could use ph paper / litmus paper. Carbox red shade
Bases give blue on paper
Amine is a base so blue on Litmus paper.
Carboxylic is an acid –

PEC140 :
Dion Patelis – 30505542
dion@7rocks.com