http://upload.wikimedia.org/wikipedia/commons/c/c7/Explosions.jpg

BOOMZ BOOMZ BOOMZ

Raffles Institution Secondary Three Chemistry
End of Year Examination 2009 Notes v1

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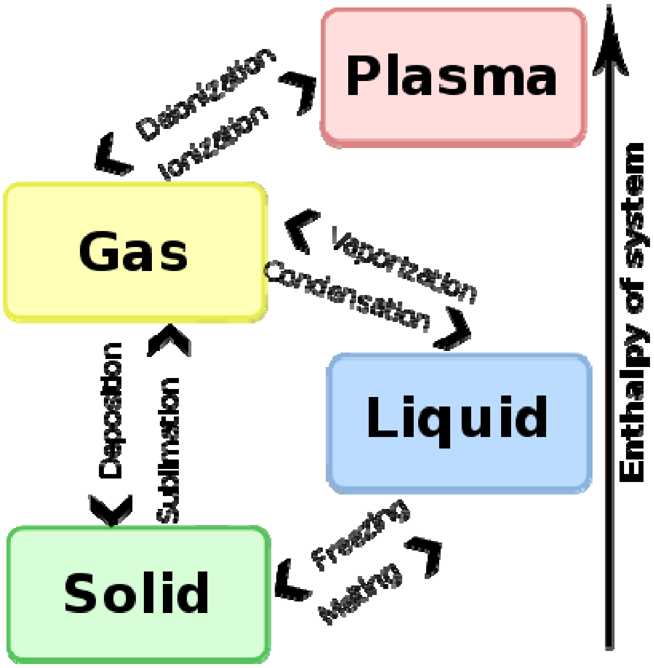
1 Particulate Nature of Matter

1a Changes of States (melting, boiling, sublimation, etc.)



Taken from http://en.wikipedia.org/wiki/File:Phase_change_-_en.svg
Everyone loves Wikipedia!

1b How to prove that matter is made up of particles

This was stolen from "Chemistry KEY.Points". Apparently my mother's colleague's daughter doesn't need it anymore, so it's mine now o.O

- The idea that matter is made up of particles is evident through diffusion.
  - Diffusion of bromine in a gas jar (when the lower bromine-containing half is separated from the upper half by a partition and the partition is removed, bromine diffuses to fill the upper half as well)
- Diffusion of a substance is understood as the movement of particles through a region of higher concentration of the substance to a lower concentration of the same substance down a concentration gradient.
- Speed of diffusion is affected by two factors:
  - A_r or M_r of particle (see 6d)
  - Temperature (high → faster diffusion)

2 Elements, Compounds and Mixtures

2a Differentiate between elements and compounds

Elements	Compounds
Made up of particles (atoms) which are the "building blocks"	
Pure substances	
Fixed melting and boiling points	
Made up of only one type of atom	Made up of more than one type of atom
Metallic or non-metallic	Can be formed by combining a metallic element with a non-metallic element, or several non-metallic elements
Cannot be chemically split into simpler substances	Can be decomposed into its elements/simpler compounds
A substance that is made up of only one type of atom. It cannot be chemically split into simpler substances.	A substance that contains two or more elements chemically joined together

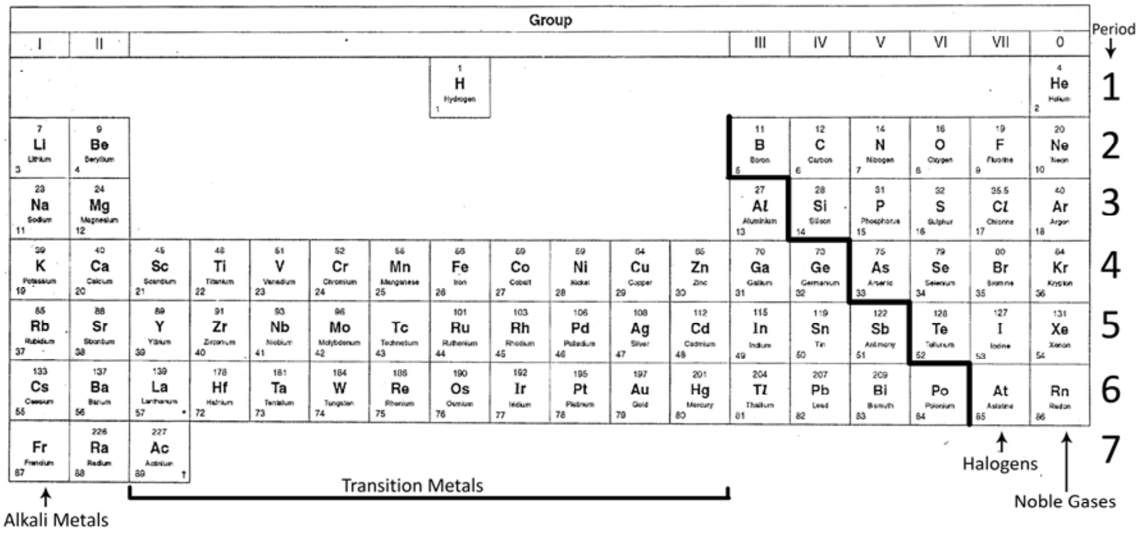
2b Differentiate between compounds and mixtures

Compounds	Mixtures
Made up of more than one type of atom	
Can only be separated into its elements by chemical methods	Can be separated into its components by physical methods
Has fixed melting/boiling points	Does not have fixed melting/boiling points
Chemical change takes place during formation	No chemical change takes place during formation
Elements are always combined in a fixed proportion by mass	Proportion of constituents can vary
Properties of a compound are different from those of its elements	Properties of a mixture are the same as those of its constituents (really…)
A substance that contains two or more elements chemically joined together	A substance that contains two or more substances physically together but not reacted with one another chemically

2c Differentiate between metals and non-metals

Metals	Non-metals
Conducts electricity	Does not conduct electricity (save for graphite)
High melting and boiling points (with the exception of Hg and Group I metals)	Generally low melting and boiling points
Malleable and ductile	Not malleable or ductile
Groups I, II and III	Groups IV to VIII
Lustrous and sonorous	(usually) not lustrous and sonorous

Be careful of metalloids. Those are sneaky.
Take note of the line that divides metals and non-metals as well.

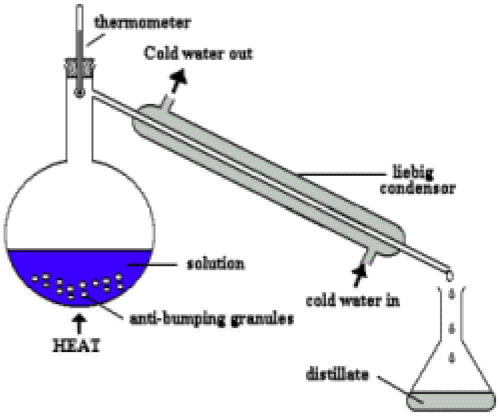
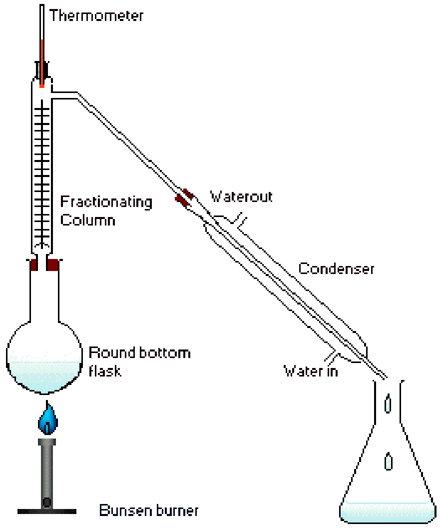


Group																		Period
I	II											III	IV	V	VI	VII	0	
							1 H Hydrogen 1										4 He Helium 2	1
7 Li Lithium 3	9 Be Beryllium 4											11 B Boron 5	12 C Carbon 6	14 N Nitrogen 7	16 O Oxygen 8	19 F Fluorine 9	20 Ne Neon 10	2
23 Na Sodium 11	24 Mg Magnesium 12											27 Al Aluminium 13	28 Si Silicon 14	31 P Phosphorus 15	32 S Sulphur 16	35.5 Cl Chlorine 17	40 Ar Argon 18	3
39 K Potassium 19	40 Ca Calcium 20	45 Sc Scandium 21	48 Ti Titanium 22	51 V Vanadium 23	52 Cr Chromium 24	55 Mn Manganese 25	56 Fe Iron 26	59 Co Cobalt 27	59 Ni Nickel 28	64 Cu Copper 29	65 Zn Zinc 30	70 Ga Gallium 31	73 Ge Germanium 32	75 As Arsenic 33	79 Se Selenium 34	80 Br Bromine 35	84 Kr Krypton 36	4
85 Rb Rubidium 37	88 Sr Strontium 38	89 Y Yttrium 39	91 Zr Zirconium 40	93 Nb Niobium 41	96 Mo Molybdenum 42	Tc Technetium 43	101 Ru Ruthenium 44	103 Rh Rhodium 45	106 Pd Palladium 46	108 Ag Silver 47	112 Cd Cadmium 48	115 In Indium 49	119 Sn Tin 50	122 Sb Antimony 51	128 Te Tellurium 52	127 I Iodine 53	131 Xe Xenon 54	5
133 Cs Caesium 55	137 Ba Barium 56	139 La Lanthanum 57 *	178 Hf Hafnium 72	181 Ta Tantalum 73	184 W Tungsten 74	186 Re Rhenium 75	190 Os Osmium 76	192 Ir Iridium 77	195 Pt Platinum 78	197 Au Gold 79	201 Hg Mercury 80	204 Tl Thallium 81	207 Pb Lead 82	209 Bi Bismuth 83	Po Polonium 84	At Astatine 85	Rn Radon 86	6
Fr Francium 87	226 Ra Radium 88	227 Ac Actinium 89 †																7

This annotated table was dug up from my S2 EOY notes. Whee!

3 Experimental Design and Purification Techniques

3a Simple distillation versus Fractional distillation

 http://www.docbrown.info/page12/gifs/distill.gif	 http://www.biocrawler.com/w/images/1/16/Fractional_distillation_lab_apparatus.png
Simple Distillation	**Fractional Distillation**
Collection of a solvent from a solution	Separation of miscible liquids with (widely) differing boiling points (Note: liquid with lower bp boils off first and is at the top of the column)

Condenser's inlet is at the bottom end so as to ensure complete condensation of distillate (since the bottom part of the Liebig condenser is consequently colder than the top part) and minimise loss of distillate in vapour form.

3b Chromatography

TENETS OF CHROMATOGRAPHY

Chromatography is defined as the process of separating the various components of a mixture by selective removal of these components from a moving fluid. Oh wow.

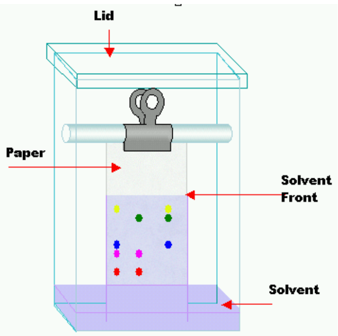
- **Selective solubility**
  - Solutes that are more soluble in the solvent travel faster
- **Selective adsorption**
  - Solutes that are more adsorbed to the solid (e.g. paper, in paper chromatography) travel slower

HOW CHROMATOGRAPHY WORKS

Mobile phase containing mixture of compounds to be analysed is allowed to pass over a stationary phase.

Take, for example, paper chromatography.

- A solvent (mobile phase) is allowed to travel along paper (stationary phase)
- If there are solutes on the solvent they will be carried by the solvent and thus also travel along the paper



This is descending chromatography by the way.

http://upload.wikimedia.org/wikipedia/commons/8/8a/Cromatography_tank.png

EVERYTHING ELSE YOU NEED TO KNOW ABOUT CHROMATOGRAPHY

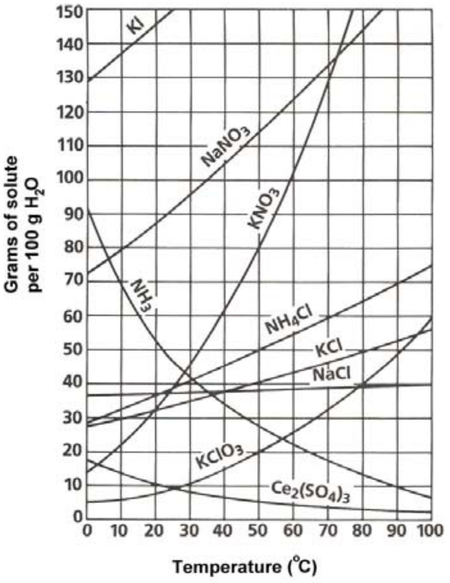
- Pure substances will give only one spot on a chromatogram when it is subjected to chromatography. I think.
- R_f = distance travelled by solute / distance travelled by solvent (leave in 2dp?)
- Different solvents give different R_f (retention factor) values (duh) and offer different results
- The lid in that picture prevents the loss of solvent through evaporation (it DOES NOT prevent evaporation)
- Draw start line with pencil so as to avoid interference with results (if start line is drawn with ink the dyes might also travel with the solvent and mess things up)
- Solvent level should be BELOW start line for ascending chromatography so that the solute spots don't just dissolve in the solvent
- Both ascending and descending chromatography involve capillary action (of the solvent up/down the paper), but the latter involves gravitational pull as well → faster separation
- If the solutes are colourless, use a locating agent! (e.g. ninhydrin for amino acids)
- Allow solvent front to travel as far as possible to ensure best separation results (you can use a longer piece of paper, for example)

3c Crystallisation and solubility curves

The figure on the right shows solubility curves of a few soluble substances.

For most soluble substances, solubility increases with temperature (see KNO3). Some exceptions are NaCl (doesn't change much) and $Ce_2(SO_4)_3$ (solubility actually decreases with increased temperature).

Crystallisation is what happens what a saturated solution (one that holds the maximum amount of solute) cools down – since temperature drops, the maximum amount of solute that the solvent can hold drops, and the remaining solute crystallises out.



Crystallisation at room temperature forms better, larger crystals than freezing because the atoms have more time to arrange themselves in the lattices which characterise their solid state.

Taken from https://www.glastonburyus.org/staff/BREINANH/regular/worksheets/Documents/solubility%20curve.jpg

3d Other methods (e.g. filtration, evaporation, reverse osmosis, etc.)

Here's a lowdown on everything for easy reference.

Dissolving, filtering and evaporating	Two solids, one soluble and one insoluble
Centrifuging/Decanting	Insoluble solid suspended in liquid
Crystallisation	Heat-sensitive (thermally unstable) solid dissolved in liquid (if you try to use DF&E you're going to destroy the solid)
Simple Distillation	Collection of a solvent from a solution
Fractional Distillation	Miscible liquids with (widely) differing boiling points
Separating Funnel	Two immiscible liquids
Sublimation	Mixture of two solids, one of them being able to sublimate
Reverse Osmosis	Collection of water from a solution

REVERSE OSMOSIS (continued from previous page)

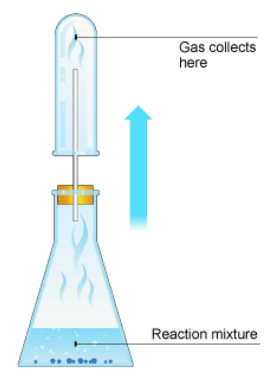
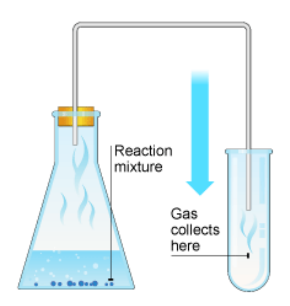
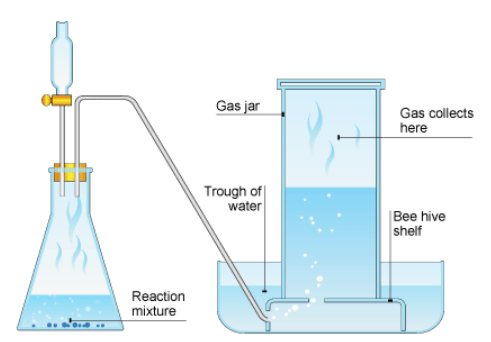
- Usually, in a container where a region of high solute concentration and a region of low solute concentration are separated by a semi-permeable membrane, the solute will diffuse from the HSC region to the LSC region via osmotic pressure
- In RO, pressure greater than the osmotic pressure is applied to the HSC region to force the solvent (water) across the semi-permeable membrane (which, in water purification, only allows water to pass through) to the LSC region (which supposedly contains pure water)

3e Collection of gases; drying agents

COLLECTION OF GASES

Gases can be collected in 3 ways:

- Upward delivery (gases less dense than air)
- Downward delivery (gases more dense than air)
- Gas jar (gases that do not readily dissolve in water)

		
Upward Delivery	**Downward Delivery**	**Gas Jar**

The average M_r of dry air is 29.0 (3sf) as determined empirically (so says http://www.engineeringtoolbox.com/molecular-mass-air-d_679.html). This is NOT in the syllabus, but it does come in handy when hacking…

Gases such as H_2 ($M_r = 2$) and NH_3 ($M_r = 17$) are usually collected via upward delivery. Gases such as CO_2 ($M_r = 44$) are usually collected via downward delivery.

DRYING AGENTS

- $CaCl_2$ (s)
- Concentrated H_2SO_4 (aq)
- Anhydrous CaO (s)
- Silica gel

What else are there? Na_2SO_4, $MgSO_4$ and $CaSO_4$ are used in organic chemistry…blah I don't know. Alternatively, use http://en.wikipedia.org/wiki/List_of_desiccants

3f Identify gases (O_2, H_2, CO_2, NH_3, SO_2, Cl_2)

Gas	Test and Results
NH_3	Turns damp red litmus paper blue
CO_2	Forms white ppt $CaCO_3$ when bubbled into limewater (prolonged bubbling makes the white ppt disappear)
Cl_2	Bleaches damp blue/red litmus paper (which turns the former red first, see below)
H_2	Pops with lighted splint
O_2	Relights glowing splint
SO_2	When bubbled into acidified $K_2Cr_2O_7$ (aq), $K_2Cr_2O_7$ (aq) turns from orange to green because SO_2 is a reducing agent

Mr Ong says that one of these is going to come out. Everyone loves hints!

NH_3

NH_3 (g) + H_2O (l) $\rightleftharpoons$ NH_4^+ (aq) + OH^- (aq)

Aqueous ammonia is alkaline. Whee!

CO_2

CO_2 (g) + $Ca(OH)_2$ (aq) → $CaCO_3$ (s) + H_2O (l)

With prolonged bubbling: CO_2 (g) + $CaCO_3$ (s) + H_2O (l) → $Ca(HCO_3)_2$ (aq)

Calcium hydrogencarbonate is not a compound that you can actually extract by evaporation. It's just free-floating Ca^{2+} (aq), CO_2 (aq), CO_3^{2-} (aq) and HCO_3^- (aq). Either way, it's colourless, so prolonged bubbling makes the white ppt ($CaCO_3$) disappear.

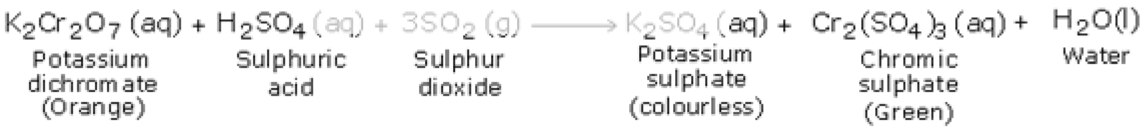
Cl_2

Cl_2 + H_2O → HCl + HClO (this explains why damp blue litmus paper turns red first)

Cl_2 is a very strong oxidising agent and thus oxidises the litmus dyes in litmus paper, hence effectively bleaching it.

SO_2

I'm lazy to work it out so…



$$K_2Cr_2O_7\,(aq) + H_2SO_4\,(aq) + 3SO_2\,(g) \longrightarrow K_2SO_4\,(aq) + Cr_2(SO_4)_3\,(aq) + H_2O(l)$$

Taken from http://www.tutorvista.com/content/chemistry/chemistry-ii/sulphur/chemical-properties-sulphur-dioxide.php

H_2SO_4 (aq) comes from dissolved SO_2 (since $K_2Cr_2O_7$ (aq) contains water)

Note: Acidified $KMnO_4$ (aq) and $K_2Cr_2O_7$ (aq) test for oxidising agents. Just because a gas makes acidified $K_2Cr_2O_7$ (aq) solution turn green doesn't mean that it MUST be SO_2.

3g How to test whether a substance is pure

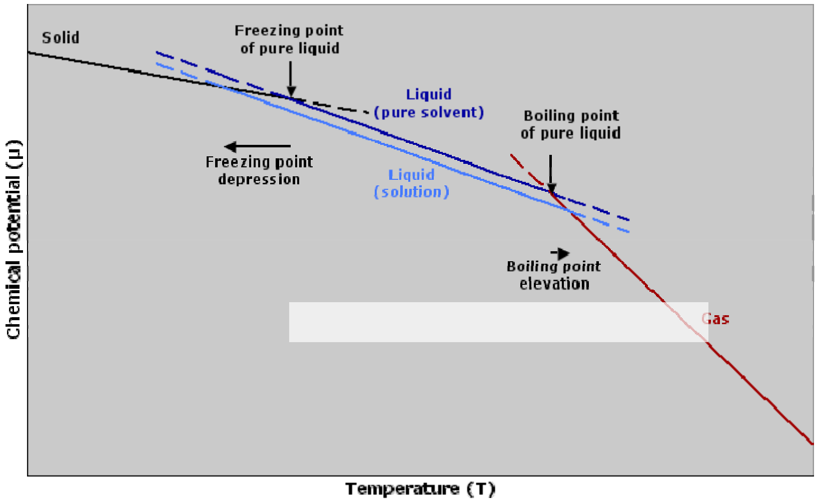
Refer to section 2b if needed for a discourse pertaining to compounds and mixtures.

FINALLY, AN ANSWER TO THE QUESTION

- A pure substance has a sharp melting and boiling point.
- A pure substance will give only one spot on a chromatogram when it is subjected to chromatography (oh really?)
- Presence of **impurity lowers m.p.** of substance and **increases b.p.** of substance - resultant mixture **melts and boils over a range of temperatures**
    - This also means that impure substances remain liquid over a range of temperatures. Whee!
    - The more impurities you have, the greater this effect is exaggerated.
    - My own (unverified) explanation: Impurities disrupt the lattice/structure of a solid substance and thus the intermolecular bonds that hold the molecules together are weaker → less thermal energy required to overcome bonds → lower m.p.
    - Try http://en.wikipedia.org/wiki/Freezing_point_depression and http://en.wikipedia.org/wiki/Boiling_point_elevation for a good discussion
- Ask the substance in a nice and non-intrusive tone (usually doesn't work)
- Stalk the substance and uncover its lineage (MIGHT work)

FREEZING-POINT DEPRESSION & BOILING POINT ELEVATION

Don't worry, this isn't tested. IT HAD BETTER NOT BE.



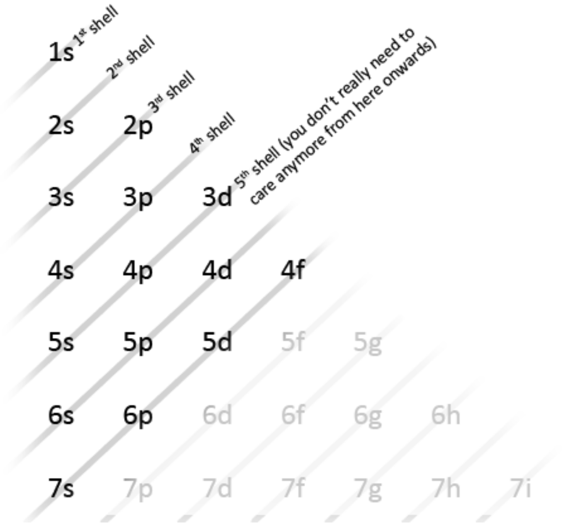
Taken from: http://en.wikipedia.org/wiki/File:Freezing_point_depression_and_boiling_point_elevation.png

4 Atomic Structure

4a Electron configuration (2.8.8.2 or $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$); octet state, valence electrons

AUFBAU PRINCIPLE

Electrons fill lower energy orbitals first. Be careful when filling…4s comes before 3d and so on☺ Note that chromium and copper (as well as some of their transition metal friends) do crazy things that don't follow the Principle in order to achieve lower energy configurations (e.g. chromium has $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{3d^5\ 4s^1}$ instead of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{3d^4\ 4s^2}$. How sad.).



HUND'S RULE (well, the 1st one)

If two or more energetically equivalent orbitals are available then electrons should be spread out before they are paired up. In short, fill each "box" with one electron before pairing. Look at the 2p orbital.
Box diagrams extracted from http://antoine.frostburg.edu/chem/senese/101/electrons/

	1s	2s	2p
Boron	[↓↑]	[↓↑]	[↑] [ ] [ ]
Carbon	[↓↑]	[↓↑]	[↑] [↑] [ ]
Nitrogen	[↓↑]	[↓↑]	[↑] [↑] [↑]
Oxygen	[↓↑]	[↓↑]	[↓↑] [↑] [↑]
Fluorine	[↓↑]	[↓↑]	[↓↑] [↓↑] [↑]
Neon	[↓↑]	[↓↑]	[↓↑] [↓↑] [↓↑]

OCTET STATE

Most atoms are quite happy achieving a stable octet state (outer shell filled with 8 e^-) by sharing / gaining/losing e^- and forming covalent/ionic bonds. There isn't much else to say…

VALENCE ELECTRONS

These are essentially electrons in the outer shell. Wow. I am amused.

4b | Isotopes; use isotopic masses to calculate atomic mass

DEFINITION
Isotopes are atoms of the same element with the same number of protons but a different number of neutrons → same proton number, different mass/nucleon number.

CALCULATING (RELATIVE) ATOMIC MASS

- Determine the percentage abundance in nature of the each isotope
- Determine the mass number of each isotope
- Calculate as shown below.

Chlorine-35 ($^{35}_{17}\mathrm{Cl}$) has 75% abundance in nature

Chlorine-37 ($^{37}_{17}\mathrm{Cl}$) has 25% abundance in nature

$$\left(\frac{75}{100}\times 35\right)+\left(\frac{25}{100}\times 37\right)$$
$$=26.25+9.25$$
$$=35.5$$

Moral of the story: When lazy, dig up your S2 EOY notes!

4c | Ionisation energy

DEFINITION
(The first) IE is the amount of energy needed to remove one electron from a gaseous atom (the IE value is usually for one mole of these atoms)

Jumps in IEs on an IE graph usually imply that the electrons in question are in different shells/subshells. Interpreting tables with values isn't particularly an exact science – just compare the increases. Increases that are significantly relatively larger usually imply a shell jump as well.

As if that helped. Sigh.

5 Atomic Structure

5a Ionic bonding; dot-and-cross diagrams, ionic lattice structure, strength of bond

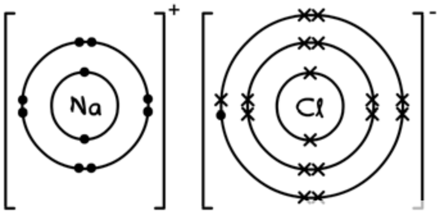
IONIC BONDING 101

- Atoms can **gain or lose valence electrons** to achieve a **stable octet structure**.
- Usually, metals lose electrons and non-metals gain electrons.
- There are no molecules in ionic bonding! You cannot say that the M_r of NaCl is 58.5 – that's the relative formula mass. Of course, you CAN have polyatomic ions as seen in $MgCO_3$ but that's beside the point.
- Ions are held in their **ionic lattice** by **strong electrostatic bonds of attraction between cations and anions**.
- A cation will always be attracted to an anion and vice-versa – it does not matter which cation donated its electrons to which anions (that doesn't seem to parse well…never mind)

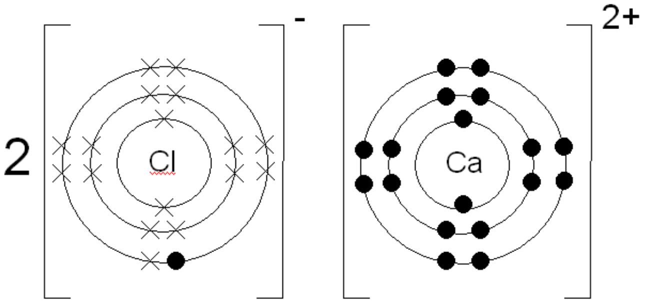
DOT AND CROSS DIAGRAMS

Note that if you were to use a dot and cross diagram to represent ions you will need square brackets around them, with the charge as superscript at the top right corner. Do remember to include a key and use different electron notation (dot/cross) for each element as shown in the diagram for NaCl.

And yes, these were dug up from my S2 EOY notes ☺ Credit goes to Kenneth (3L'09) for making the $CaCl_2$ diagram.



Key
x - Electron of chlorine ion
· - Electron of sodium ion



Key
● Calcium electron
× Chlorine electron

5b

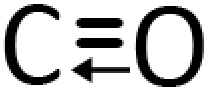
Covalent bonding: dot-and-cross diagrams, single, double and triple bonds, coordinate bonds, polar bonds, macromolecules vs simple molecules

COVALENT BONDING

- One covalent bond → sharing of two electrons (each atom contributes one, think Cl_2)
- Needless to say, triple bonds (N_2) are stronger than double bonds (O_2), which are stronger than single bonds (F_2). Triple bonds are NOT three times stronger than single bonds, but then again we don't need to know that. Yet.

COORDINATE (DATIVE COVALENT) BONDING

- A coordinate bond is defined as a covalent bond where the shared pair of electrons is provided by only **one** of the bonded atoms. In the diagram for CO below, the oxygen atom provides an electron pair to help carbon achieve its stable octet structure
- In coordinate bonding, there is a **donor** (the atom that donates the electron pair, here oxygen) and an **acceptor** (that atom that receives the electron pair, here carbon)
- The acceptor is usually a metal cation/transition metal atom/atom in a molecule…or so they say (e.g. $AlCl_3$ accepts the lone pair in $:NH_3$ and forms a coordinate bond to become NH_3AlCl_3).
- In structural formulae, the arrow is drawn from the donor to the accepter. See below.
- Coordinate bonds are responsible for the phenomenon of complex formation (think tetrahydroxozincate, $Zn(OH)_4^{2-}$ - that's a good four water molecules!).
- Look at the question closely. If you run into problems while drawing dot and cross diagrams, you might want to try coordinate bonding.
- By now you should have realised that I'm alluding to the RI S3 EL EOY 2009 questions. Haha I hope that galvanized you to continue reading and laughing at my bad diction. At least that's thawing the ice between the persona and the reader.



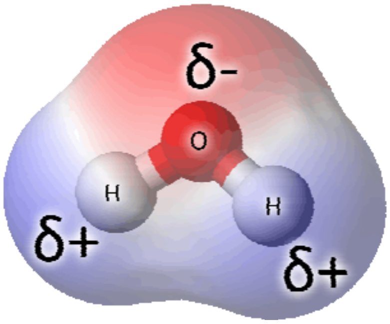
(Don't forget to include the key. I'm just lazy :P)

POLAR (COVALENT) BONDS

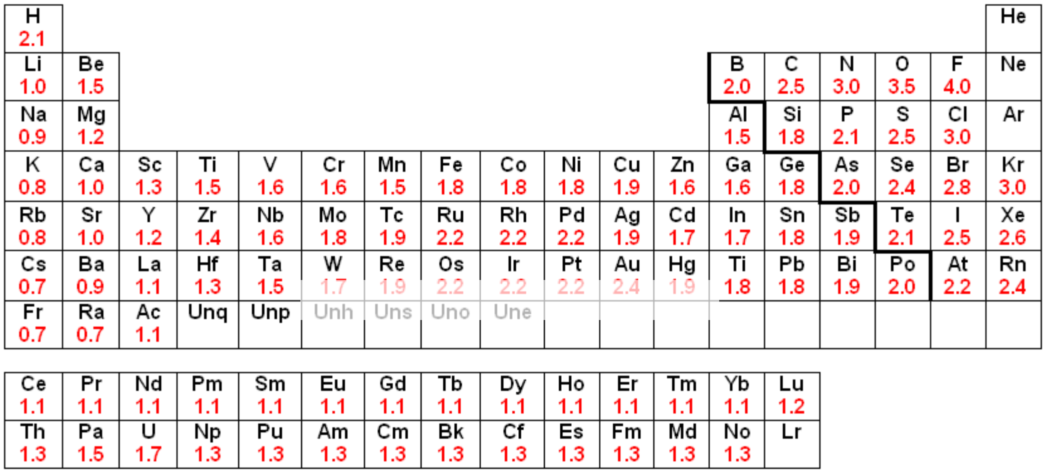
- Different elements have different electronegativity values (the higher the electronegativity, the greater it attracts electrons to itself). See table below.
- Electronegativity increases across periods from left to right and increases asm we go up the group → more electronegative elements are nearer to the upper right of the periodic table.
- Because of possible electronegativity differences, electrons shared in a covalent bond may be pulled towards a more electronegative atom.
- Due to this electron pull, the end with the greater electronegativity acquires a partial negative charge (δ^-) and the other end acquires a partial positive charge (δ^+).
- The covalently bonded molecule is thus polar (with a positive dipole δ^+ and a negative dipole δ^-). The end!

Electronegativity of oxygen is 3.5
Electronegativity of hydrogen is 2.1
Electrons are pulled closer towards oxygen → dipole moments (areas of partial charge) arise, molecule is consequently polar covalent.

Note that molecules like CO_2 are non-polar covalent because the dipole moments cancel out. Well I don't think we need to know that.



Modified from http://upload.wikimedia.org/wikipedia/commons/1/15/Water-elpot-transparent-3D-balls.png



H 2.1																	He
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Ti 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn 2.4
Fr 0.7	Ra 0.7	Ac 1.1	Unq	Unp	Unh	Uns	Uno	Une									

Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.1	Sm 1.1	Eu 1.1	Gd 1.1	Tb 1.1	Dy 1.1	Ho 1.1	Er 1.1	Tm 1.1	Yb 1.1	Lu 1.2
Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr

This is a table showing the Pauling electronegativity values. There are other ways to quantify electronegativity, but then Pauling's scale has the most widespread use. I'm not sure if the values are accurate; this revolution is for display purposes only. ☺
Taken from http://webassign.net/graphics/electronegativities.gif

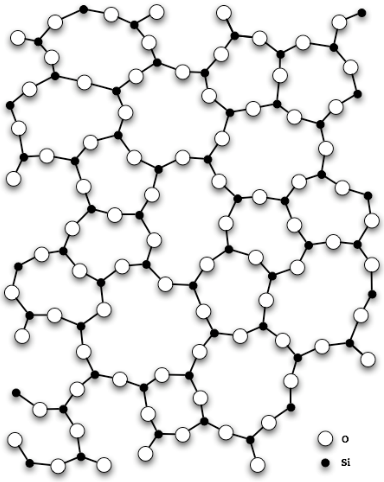
NOTE (continued from previous page): There are thresholds which are attempts to define covalent and ionic bonds by electronegativity differences.

$0 < \Delta EN < 0.5$	Non-polar covalent
$0.5 < \Delta EN < 2.0$	Polar covalent
$2.0 < \Delta EN < 4.0$	Ionic

Stuff like H_2, Cl_2, N_2 and all the other diatomic molecules are effectively non-polar covalent since the electronegativity difference is zero – though ID-ID interactions are possible (see 5c)

MACROMOLECULES

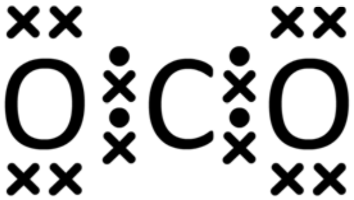
- Macromolecules are NOT IONIC LATTICES. Those are different.
- Macromolecules are large molecules where many atoms are joined together by covalent bonds. Well, relatively many compared to simple covalent molecules like CO_2.
- Because of their extensively bonded structure, macromolecules have **high melting and boiling points** (**high thermal energy** is needed to **overcome** the **strong covalent bonds** between atoms in the macromolecular structure → high m.p., b.p.)
- Examples of macromolecules include diamond, graphite (more on these two allotropes of carbon later) and SiO_2.



Snakes and ladders, anyone?
Taken from http://upload.wikimedia.org/wikipedia/commons/1/19/SiO2_-_Glas_-_2D.png

DOT AND CROSS DIAGRAMS

Yes, I put them here because I kind of forgot about them.
HELLO THERE SECONDARY TWO CHEMISTRY SYLLABUS. Ah, the good old days…



(Don't forget to include the key. I'm just lazy :P)

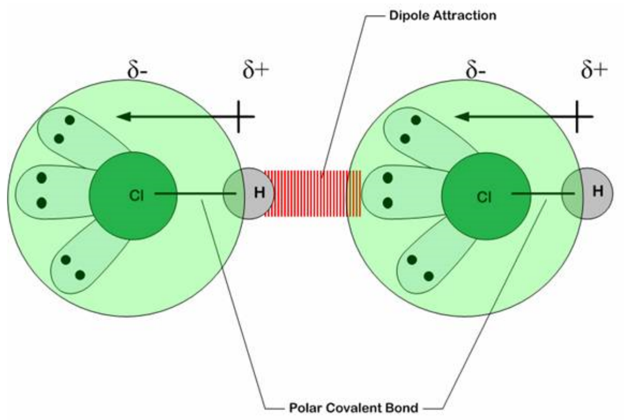
5c | Weak intermolecular forces

There are two types of intermolecular forces which we will be dealing with:

- Van der Waals forces (100x weaker than a covalent bond)
- Hydrogen bonding

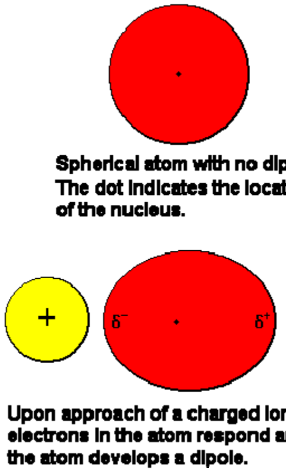
Van der Waals FORCES (haha that was weird)
There are three subtypes. Oh wow.

- (Permanent) Dipole-dipole attraction
- Dipole - induced dipole attraction (D-ID)
- Induced dipole – induced dipole attraction (ID-ID, haha KYX)



http://www.chemprofessor.com/imf_files/image005.jpg

Dipole-dipole attraction



http://www.chem.purdue.edu/gchelp/liquids/ioninddp.gif

Dipole - induced dipole attraction

I know that the yellow circle is an ion, not a dipole…but then you can pretend that there's a (permanent) dipole causing the induced dipole (red ellipse)

More on (permanent) dipoles
Permanent dipoles are…well…permanent. The partial charges always exist because the dipole moments don't cancel out.

More on ID-ID interactions
Electrons orbit atoms and are moving. At any instance, the distribution of electrons in a molecule may not be uniform, even in a non-polar covalent molecule like Br_2, and an instantaneous dipole exists. This instantaneous dipole may induce a dipole in another nearby molecule and an ID-ID interaction occurs.

Even though the average dipole of every molecule over a period of time is zero, the resultant forces between molecules at any instant are non-zero.

NOTE (continued from previous page): VdW forces increase with molecule size (because larger molecules have a greater surface area for the forces to act). Because of this, I_2 has a higher melting and boiling point than Br_2 (iodine atoms are larger than bromine atoms → an I_2 molecule is larger than a Br_2 molecule → stronger Van der Waals intermolecular (specifically ID-ID) forces at work → greater thermal energy needed to overcome these forces → higher m.p., b.p.)

I don't think they're going to test that anyway. Don't hold me responsible if you get penalized for writing something erroneous after reading my notes. No one told you to read them ☺

HYDROGEN BONDING

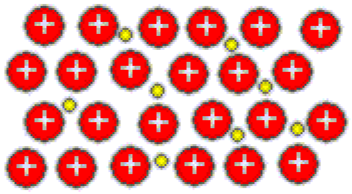
This occurs when the below two conditions are met:

- There is a highly positively polarized hydrogen atom
- There is a highly electronegative atom with a lone pair (so that it provides a δ^- dipole moment to attract the δ^+ dipole moment at the hydrogen atom of another molecule)

Because of this, hydrogen bonding usually occurs only when molecules have N-H, F-H or O-H groups. Recall that N, F and O have insanely high electronegativities.

Molecules capable of hydrogen bonding (e.g. water, alcohols [with their –OH functional groups]) have higher melting and boiling points (more thermal energy needed to overcome intermolecular forces [specifically hydrogen bonding] → higher m.p., b.p.)

5d Metallic bond and structure



Taken from http://www.revisioncentre.co.uk/gcse/chemistry/metallic_bonding.gif
I couldn't find a better picture that caught my fancy.

METALLIC BONDING 101

- Metallic bonds are electrostatic forces of attraction (?) between positively charged metal ions in a sea of freely moving delocalized negatively charged valence electrons
- As a result, metals generally have high melting points (high amounts of thermal energy needed to overcome strong metallic bonds between metal cations and freely moving delocalized electrons → higher m.p.)

5e/f/g How structure and bonding affect melting and boiling points/solubility in water and in non-polar solvents/electrical conductivity

I am really adverse to the idea of typing out the entire table in the worksheet. Try saying that in the EOYs. Ah well.

	Ionic	Simple Molecular	Macromolecular	Metallic
Melting/ Boiling Points	**High m.p. and b.p.** Ions are held together by strong ionic bonds (electrostatic forces of attraction). High amounts of heat energy are needed to break these bonds	**Low m.p. and b.p.** Molecules are held together by weak intermolecular forces of attraction which require little heat energy to break	**High m.p. and b.p.** Atoms in macromolecular structures are held together by an extensive network of strong covalent bonds that require high amounts of heat energy to break	**Generally high m.p. and b.p. (with the exception of Hg, group I/II metals)** Strong metallic bonding (electrostatic forces of attraction) between positive metal ions and sea of freely moving delocalized negatively charged electrons requires high amounts of heat energy to break
Electrical conductivity	**Conducts only when molten/aqueous** In aqueous/molten states, ions are no longer in fixed positions in the ionic lattice and are thus mobile and can then act as charge carriers to carry current and conduct electricity	**Does not conduct in any state (WTEO polar covalent molecules such as HCl)** Made up of neutral molecules, thus there are no mobile charge carriers to carry current and conduct electricity	**Does not conduct in any state** Made up of neutral atoms fixed in the macromolecular structure, thus there are no mobile charge carriers to carry current and conduct electricity	**Conducts in solid/liquid states** Metallic lattice consists of positive metal ions and sea of freely moving delocalized negatively charged electrons. These mobile electrons can act as charge carriers to carry current and conduct electricity

	Ionic	Simple Molecular	Macromolecular	Metallic
Solubility in solvents	**Most are soluble in water but are insoluble in organic solvents** Water molecules are polar and can thus interact with the ions and weaken the ionic bonds (electrostatic forces of attraction) that hold them in place in the ionic lattice. The ions then separate and the ionic compound dissolves.	**Most are soluble in organic solvents but are insoluble in water** Like dissolves like. Ah well.	**Not soluble in any type of solvent.** Lalalala, lalalala, Elmo's World…	**Not soluble in any type of solvent WTEO Group I/II metals which REACT WITH water** This is a chemical change as new compounds are formed (metal hydroxide + H_2) – it is NOT the physical change of salvation as seen in ionic compounds dissolving in water and other cases.

5h Graphite vs diamond

	M.p.	Density	Appearence	Hardness	Electrical conductivity
Diamond	3700°C	3.5 g/cm^3	Colourless transparent crystals	Hardest natural substance – structure consists of many carbon atoms in strong covalent bonds in tetrahedral units → rigid, strong, hard	Does not conduct electricity – each of the carbon atom's four valence electrons is involved in covalent bonding with other carbon atoms → no delocalized electrons to move through the structure and carry current to conduct electricity
Graphite	3300 °C	2.2 g/cm^3	Black shiny powder	Soft – used as lubricant – consists of parallel layers of carbon atoms covalently bonded within layers. Weak intermolecular (Van der Waals) forces between layers allow them to slide against one another → soft, slippery	Conducts electricity – each carbon atom bonded to three others, each carbon atom has one valence electron not involved in bonding, which then delocalizes and can move freely along carbon atom layers to carry current and conduct electricity

6 Stoichiometry and Mole Concept

6a Writing formulae and balancing equations

THE GOOD OLD DAYS REVISITED

Thanks go to Kenneth (3L'09) again for the great work from last year's S2 EOY notes.

1. Atoms cannot disappear or appear on any side of the equation. So, you cannot have 3 hydrogen atoms and 4 carbon atoms on one side of the equation, and 3 hydrogen atoms and 5 carbon atoms on the other side. For example:
   $H_2 + O_2 \rightarrow H_2O$ (wrong)
   $2H_2 + O_2 \rightarrow 2H_2O$ (correct)
2. For chemical formulas of salts, the number of individual ions in them is equal to the charge of the other ion. For example, for aluminium sulphate (VI):
   Aluminium Sulphate (VI)
   $Al^{\textcircled{3+}}$ $SO_4^{\textcircled{2-}}$
   $Al_{\textcircled{2}}(SO_4)_{\textcircled{3}}$
3. The name of the positive ion in a compound is the same as that of the atom
4. Negative ions with at least 1 oxide ion end in -ate, except for hydroxide, which ends in -ide. All other negative ions end in -ide.
5. State symbols should be written in subscript after every compound in a reaction. (s) should be written for a solid, (l) for a liquid, (g) for a gas, and (aq) for an aqueous solution. For example:
   $NaCl(aq) + AgNO_3\,(aq) \rightarrow AgCl(s) + NaNO_3\,(s)$
6. Acids are considered to be aqueous

CHEMICAL EQUATIONS

1. Write word equation (eg. Hydrogen + Oxygen → Water)
2. Convert chemical name to chemical formula (eg. $H_2 + O_2 \rightarrow H_2O$)
3. Check if compounds are correct (eg. Hydrogen = H_2, Oxygen = O_2, Water = H_2O)
4. Balance equations (eg. $2H_2 + O_2 \rightarrow 2H_2O$)
5. Write in state symbols [(s), (l), (g), or (aq)] (eg. $2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O(l)$)
6. Check if equation is balanced (eg. LHS: 4 hydrogen, 2 oxygen from $2H_2$ and O_2, RHS: 4 hydrogen, 2 oxygen from $2H_2O$)

6b | Calculate % composition by mass

HERE WE GO AGAIN

1. Find relevant A_r of the elements in the compound
2. Take total relative mass of element/compound in question and divide it by M_r/relative formula mass of the entire compound
3. Multiply by 100% to obtain a percentage

Example
Find the percentage composition of oxygen in sodium thiosulphate, $Na_2S_2O_3\cdot 5H_2O$
A_r of Na = 23.1
A_r of S = 32.1
A_r of O = 16.0
A_r of H = 1.0
Percentage composition

$$= \frac{(3\times16.0)+(5\times16.0)}{(2\times23.1)+(2\times32.1)+(3\times16.0)+(5\times2\times1.0)+(5\times16.0)}\times100\%$$

=51.5% (3sf)

That was really painful.

6c | Relative atomic mass (A_r) versus relative molecular mass (M_r)

Um… is there anything to say?
A_r of N is 14.0
M_r of N_2 is 28.0
THE END!

6d | Rate of diffusion of a gas is dependent on M_r

GRAHAM'S LAW

$$rate\ of\ diffusion\ of\ gas \propto \frac{1}{\sqrt{M_r\ of\ gas}}$$

$$\frac{rate\ of\ diffusion\ of\ gas\ A}{rate\ of\ diffusion\ of\ gas\ B} = \frac{\sqrt{M_r\ of\ gas\ B}}{\sqrt{M_r\ of\ gas\ A}}$$

In general, the greater the M_r of a gas, the slower it diffuses.
Remember: the average M_r of air is 29.0 (3sf). Unorthodox, but potentially useful.

6e How many is one mole?

AVOGADRO CONSTANT

One mole of particles contains $6.02214179 \times 10^{23}$ particles.
Just use the 6.02×10^{23} value, okay?

6f Calculate empirical formula and molecular formula

Empirical formula → Simplest whole number ratio of atoms/ions present in a compound
Molecular formula → Actual ratio of atoms/ions present in a compound

Example

Find the empirical and molecular formula of ascorbic acid (that's Vitamin C) if it contains 4.55% hydrogen, 40.91% carbon and 54.54% oxygen by mass and its M_r is 176.

	C	H	O
Mass (assuming 100g)	40.91g	4.55g	54.54g
Amount	40.91/12.0 = 3.41mol (3sf)	4.55/1.0 = 4.55mol	54.54/16.0 = 3.40875mol
Divide by smallest amount	3.41/3.40875 = 1.00 (3sf)	4.55/3.40875 = 1.33 (3sf)	3.40875/3.40875 = 1
Simplest ratio	3	4	3

Therefore, empirical formula of ascorbic acid is $C_3H_4O_{3\#}$
Let molecular formula of ascorbic acid be $C_{3n}H_{4n}O_{3n}$
M_r of $C_3H_4O_3 = (3 \times 12.0) + (4 \times 1.0) + (3 \times 16.0) = 88$
$n = 176/88 = 2$
Therefore, molecular formula of ascorbic acid is $C_6H_8O_{6\#}$

6g Molar volume at (r.t.p.) [of gases] and Avogadro's Law

Molar volume of gases at r.t.p. [room temperature (25°C) and pressure (1 atm)] is **$24.0dm^3$.**

Avogadro's Law states that equal volumes of all gases under the same conditions of temperature and pressure contain the same number of molecules.

The ideal gas law pV = nRT (p: Pressure; V: Volume; n: Amount (in moles); R: Ideal gas constant 8.314472 $JK^{-1}mol^{-1}$; T: Temperature (in Kelvin)) is fun to exploit (e.g. it can tell you that doubling pressure halves volume since RHS remains constant but p doubles) but shouldn't be used if you don't know how to.

6h Molar mass and calculations involving masses; limiting reactant

If you're bad at this, the only way is to PRACTISE. I can't help you here…
Some tips:

1. **Good presentation.** Provide notes at the side (e.g. Amount of H_2SO_4 (aq) = 25/1000 x 0.001 = 0.000025mol) which explicitly state what you are doing so that the marker doesn't get confused. It also helps you when you check anyway.
2. **When in doubt, check units.** Don't get a "g/mol" value when you're trying to find something like the amount of HCl (aq) in a solution, because that means something has gone HORRIBLY WRONG.
3. **Write the balanced equation out.** This is so that you don't make careless mistakes (TWO units of HCl react with one unit of $Ca(OH)_2$…be careful.)

Have problems with presentation? Check the worksheets or use Chemistry Insights.

6i % purity and % yield

Percentage purity of substance in sample = $\frac{Mass\ of\ pure\ substance\ in\ sample}{Mass\ of\ given\ sample} \times 100\%$

Percentage yield of a product of a reaction = $\frac{Actual\ Yield}{Expected\ Yield} \times 100\%$

6j Concentration of a solution

There are two measures of concentration: mole concentration (mol/dm^3, a.k.a. molarity, M); and mass concentration (g/dm^3).

Molarity of a solution = $\frac{Amount\ of\ solute\ (in\ mol)}{Volume\ of\ solution\ (in\ dm^3)}$ mol/dm^3 or M

Mass concentration of a solution $\frac{Mass\ of\ solute\ (in\ g)}{Volume\ of\ solution\ (in\ dm^3)}$ g/dm^3

We end off with the epic (commonsensical) relationship:

$$\underset{(g/dm^3)}{\text{Mass Concentration}} = \underset{(mol/dm^3)}{\text{Molarity}} \times \underset{(g/mol)}{\text{Molar Mass}}$$

6k Titration techniques and calculation

Um….what?

(not very complete) TITRATION STEPS

1. Wash everything with tap water then rinse with distilled water
2. Place metal clip in place
3. Rinse burette with small amount of solution to be used (ensure that water is flushed out)
4. Fill burette with solution to be used during filter funnel (it is not necessary to fill to the 0.00cm^3 mark)
5. Flash out any air bubble by releasing and closing the metal clip
6. Leave the burette aside for air bubbles to settle (if they settle later during experimentation you are going to end up using less solution than you think you did)
7. Fix pipette-filler to pipette
8. Rinse pipette by drawing small amount of solution to be filled and discard (ensure that water is removed)
9. Fill to 25.0cm^3 line, ensuring lower meniscus just touches the thin line at eye level.
10. Transfer to conical flask
11. CONICAL FLASK MUST NOT BE WASHED WITH SOLUTION – ONLY DISTILLED WATER WILL DO.

RANDOM ADDITIONAL THINGS IN THE CHECKLIST THAT MIGHT BE IMPORTANT

- Remove filter funnel from burette top when not in use
- Do not push more than 1cm of the pipette into the pipette filler
- Ensure that there are no air bubbles in pipette before transfer
- Add only 2 or 3 drops of indicator (if the titration requires one)
- Endpoint colour for methyl orange is orange, endpoint colour for phenolphthalein is colourless
- When reaching endpoint, flush down drops of titrant at the side with distilled water
- RECORD BURETTE READING TO 2DP and pipette reading to 1dp (then again it's usually 25.0dm^3 so don't worry too much)
- Report calculation results to 3sf (we always do anyway haha)

7 Acids and Bases

7a,b Chemistry of acids and bases (haha shortcut!)

Reactants	Products	Example
Acid + Base	Salt + H_2O (l)	$NaOH$ (aq) + HCl (aq) $\rightarrow$ $NaCl$ (aq) + H_2O (l)
Acid + Carbonate	Salt + H_2O (l) + CO_2 (g)	H_2SO_4 (aq) + $MgCO_3$ (s) $\rightarrow$ $MgSO_4$ (aq) + H_2O (l) + CO_2 (g)
Acid + Reactive Metal	Salt + H_2 (g)	H_2SO_4 (aq) + Mg (s) $\rightarrow$ $MgSO_4$ (aq) + H_2 (g)
Acid + Sulfite salt	Salt + H_2O (l) + SO_2 (g)	H_2SO_4 (aq) + Na_2SO_3 $\rightarrow$ Na_2SO_4 (aq) + H_2O (l) + SO_2 (g)
Base + Ammonium salt	Salt + H_2O (l) + NH_3 (g)	$NaOH$ (aq) + NH_4Cl (aq) $\rightarrow$ $NaCl$ (aq) + H_2O (l) + NH_3 (g)
Alkali + Aqueous Salt (some)	Salt + Hydroxide ppt (s)	$2NaOH$ (aq) + $Ca(NO_3)_2$ $\rightarrow$ $2NaNO_3$ (aq) + $Ca(OH)_2$ (s)

7c Arrhenius model

Arrhenius acids ionize/dissociate to yield H^+ in aqueous solution, while Arrhenius bases ionize/dissociate to yield OH^- in aqueous solution. The other acid-base theories are the Brønsted-Lowry and Lewis ones, but that's not relevant now.

7d Difference between strength and basicity; difference between strength and concentration

Strength	Degree of dissociation of an acid or alkali (e.g. virtually complete dissociation → strong acid/alkali; partial dissociation → weak acid/alkali) If you see $HA \rightleftharpoons H^+$ (aq) + A^- (aq), then HA is a weak acid! If you see $HA \rightarrow H^+$ (aq) + A^- (aq), then HA is a strong acid!
Basicity	Number of H^+/OH^- that a molecule of the acid/alkali can yield. HCl (aq) and $NaOH$ (aq) are monobasic; H_2SO_4 and $Ba(OH)_2$ are dibasic; H_3PO_4 is tribasic etc. n-basic acids can form n salts (e.g. tribasic H_3PO_4 can form MH_2PO_4, M_2HPO_4 or M_3PO_4) If you're wondering, solubility of Group II hydroxides increase down the group, so $Ba(OH)_2$ is reasonably soluble.
Concentration	See 6j. Haha.

7e Types of oxides (acidic, basic, neutral, amphoteric)

	Oxides			
	Basic	**Amphoteric**	**Neutral**	**Acidic**
Range	Most metal oxides	Zinc, aluminium and lead oxides. Oh, and water.	Some other oxides.	Even more oxides.
Examples	Na_2O, CaO, CuO	ZnO, Al_2O_3, PbO, H_2O	CO, NO	CO_2, SO_2, NO_2

Hydroxides behave like oxides so the above still applies for the hydroxides.

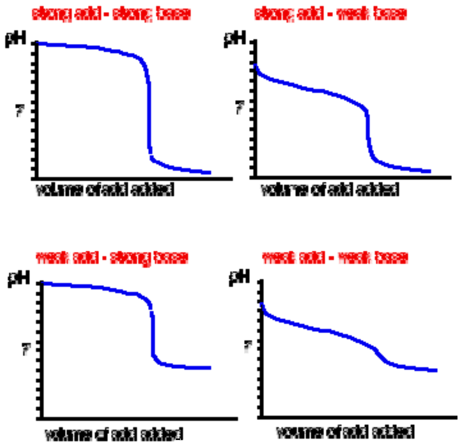
7f Calculate pH

$pH = -\log[H^+] \rightarrow [H^+] = -\log^{pH}$

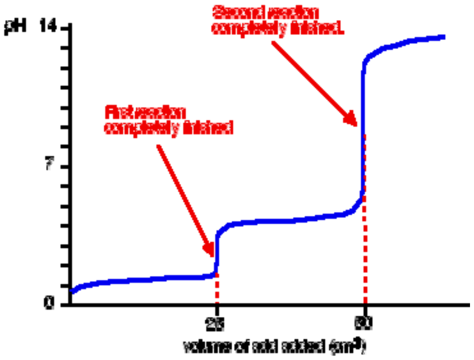
$$pH = -\log(percentage\,dissociation \times molarity)$$

Water has pH 7 → $[H^+] = [OH^-] = 10^{-7} mol/dm^3$

7g Interpreting acid-base titration curves



Titration graph for the four possible titrations (assuming acid added to alkali)



Titration graph for a monobasic alkali added to a dibasic acid

Taken from http://www.chemguide.co.uk/physical/acidbaseeqia/phcurves.html
The website from which I took this picture has a good discussion of the pH curves.

Endpoint → point at which indicator changes colour
Equivalence point → the point at which you reach the correct stoichiometrical ratio (i.e. 1:1 for HCl and NaOH – neutralization success!)
They aren't the same thing.

NOTE (continued from previous page):

- pH of strong acid is usually 2
- pH of weak acid is usually 5
- pH of strong base is usually 11
- pH of weak base is usually 9
- When selecting indicators, the pH at which the indicator changes colour should lie inside the range of sharp pH rise

7h Managing acidic air pollutants and acid rain

Acid rain can arise from **wet deposition** (acidic oxide gases dissolve in precipitation such as rain or snow which then falls) and **dry deposition** (the acidic oxide gases just get onto surfaces and might dissolve in running water to worsen acid rain). I don't think that's important though. Then again I don't know what IS.

Normal rain is naturally weakly acidic (because acidic gases do exist in the atmosphere) but acid rain has a pH of 4. Wow. Chemicals responsible for acid rain are SO_2 and NO_x (oxides of nitrogen).

	Source	Effect	Ways to minimize effect
SO_2	Combustion of sulphur-containing fossil fuels	• Damage lungs; causes bronchitis • Damages metal bridges and stone buildings • Lowers pH of water bodies and kills marine life • Combine with water and oxygen to form strong acids: o $2SO_2\ (g) + 2H_2O\ (g) + O_2\ (g) \rightarrow 2H_2SO_4\ (aq)$ o $4NO_2\ (g) + 2H_2O\ (g) + O_2\ (g) \rightarrow 4HNO_3\ (aq)$	• Remove S from fossil fuels before burning • Remove SO2 from waste gases (flue gas desulphurization) o $CaCO_3\ (s) + SO_2\ (g) \rightarrow CaSO_3\ (?) + CO_2\ (g)$
NO_x	Produced in car engines as N_2 and O_2 from air in the engine combine under high temperature to form NO (which then reacts further with O_2 to give NO_2)		• Fit cars with catalytic converters o $2NO\ (g) + 2CO\ (g) \rightarrow N_2\ (g) + 2CO_2\ (g)$

7i Writing ionic equations

IONIC EQUATIONS 101

I'm just extracting the steps from some worksheet, because it works. Glorious!

1. Write balanced chemical equation
2. Put in state symbols
3. Write out free ions in aqueous solution
4. Cancel away spectator ions
5. Write down uncancelled ions

TIPS

- Ensure that number of atoms and charge are balanced
- Water-soluble substances are in the aqueous state
- Substances in solid/liquid/gas states do not form free ions
- Check that there are no repeated ions on both sides
- All acid-alkali neutralization reactions can be simplified to the reaction: H^+ (aq) + OH^- (aq) → H_2O (l)
- The ionic equation of a precipitation reaction usually involves the ions constituting the ppt: Ba^{2+} (aq) + SO_4^{2-} (aq) → $BaSO_4$ (s)

8 Redox Reactions

8a Definition 1: Oxidation is loss of electrons, reduction is gain of electrons

Consider the equation below.

$5Fe^{2+}$ (aq) $8H^+$ (aq) + MnO_4^- (aq) → $5Fe^{3+}$ (aq) + $4H_2O$ (l) + Mn^{2+} (aq)

Iron in Fe^{2+} loses an electron to form Fe^{3+} and is thus oxidized.

Manganese in MnO_4^- gains 5 electrons to form Mn^{2+} and is thus reduced.

8b Definition 2: Oxidation is increase in oxidation state/number; reduction is decrease in oxidation state/number

Consider the equation $(NH_4)_2Cr_2O_7$ (s) → Cr_2O_3 (s) + N_2 (g) + $4H_2O$ (g).

Chromium's oxidation state decreases from +6 in $(NH_4)_2Cr_2O_7$ to +3 in Cr_2O_3 and is thus reduced.

Nitrogen's oxidation state increases from -3 in $(NH_4)_2Cr_2O_7$ to 0 in N_2 and is thus oxidized.

8c Writing half-equations

In case of emergency, here are a few steps that can save your life.
No, it's not DRABC.

1. Write half-equation
2. Balance element oxidized/reduced
3. Balance O: add H_2O
4. Balance H: add H^+
5. Balance charge: add e^-
6. Balance e^- transfer: multiply half-equations
7. Add together half-equations
8. Eliminate common terms

To put this into perspective, try it with the reduction half-equation of manganese (MnO_4^-).

STEP 1: WRITE HALF-EQUATION

We put the manganate (VII) ion on the LHS and the manganate (II) ion Mn^{2+} on the RHS.

MnO_4^- (aq)	→	Mn^{2+} (aq)

STEP 2: BALANCE ELEMENT OXIDISED/REDUCED

There's already one Mn so don't worry about that for now.

MnO_4^- (aq)	→	Mn^{2+} (aq)

STEP 3: Balance O: add H_2O

There are four O atoms on the LHS, so we have to add four H_2O on the RHS.

MnO_4^- (aq)	→	Mn^{2+} (aq) **+ $4H_2O$ (l)**

STEP 4: Balance H: add H^+

There are eight H atoms on the RHS, so we have to add eight H^+ on the LHS.

MnO_4^- (aq) **+ $8H^+$ (aq)**	→	Mn^{2+} (aq) + $4H_2O$ (l)

STEP 5: Balance charge: add e^-

The charge on the LHS is (-1) +8 = +7, while the charge on the RHS is +2. Hence, we can add five electrons to the LHS to bring the charge on the LHS down to +2.

MnO_4^- (aq) + $8H^+$ (aq) **+ $5e^-$**	→	Mn^{2+} (aq) + $4H_2O$ (l)

STEP 6: Balance e^- transfer: multiply half-equations

For illustration purposes, we will use the oxidation of iron (II) ions to iron (III) ions.

Oxidation half-equation: Fe^{2+} (aq) → Fe^{3+} (aq) + **e^-**
Reduction half-equation: MnO_4^- (aq) + $8H^+$ (aq) + **$5e^-$** (aq) → Mn^{2+} (aq) + $4H_2O$ (l)

As you can see, the oxidation half-equation spits out one e^-, while the reduction half-equation requires $5e^-$. We thus multiply the oxidation half-equation by 5 so that both sides involve 5 electrons.

Therefore,
Oxidation half-equation: $5Fe^{2+}$ (aq) → $5Fe^{3+}$ (aq) + **$5e^-$**
Reduction half-equation: MnO_4^- (aq) + $8H^+$ (aq) + **$5e^-$** (aq) → Mn^{2+} (aq) + $4H_2O$ (l)

STEP 7,8: ADD TOGETHER HALF-EQUATIONS, ELIMINATE COMMON TERMS

Adding everything together and cancelling out the electrons [since (-5) + 5 = 0] yields the following redox equation.
$5Fe^{2+}$ (aq) + MnO_4^- (aq) + $8H^+$ (aq) → $5Fe^{3+}$ (aq) + Mn^{2+} (aq) + $4H_2O$ (l)

8d Rules for assigning oxidation state/number to elements

THE EIGHT COMMANDMENTS OF ASSIGNING OXIDATION STATES

After a while everything becomes instinctual. After a while.

1. Elements that are uncombined with other elements is assigned an oxidation state of 0.
2. In simple (monoatomic) ions, the oxidation number is simply the charge on the ion.
3. In complex (polyatomic) ions, the sum of the oxidation numbers of all the atoms in the complex (polyatomic) ion gives the charge of the complex ion.
4. The oxidation number of hydrogen in all compounds is +1, except in metal hydrides where is oxidation number is -1.
5. The oxidation number of oxygen in all compounds is -2, except in peroxides where its oxidation number is -1.
6. The oxidation number of Group I elements in their compounds is +1. For Group II elements, their oxidation number in their compounds is +2. For aluminum, its oxidation state in its compounds is +3.
7. There are many oxidation numbers for Group VII elements in their compounds but the usual one is -1.
8. In neutral molecules with no charge, the sum of the oxidation numbers of all the atoms in the molecule is 0.

TIPS

- Start with putting the oxidation states of H and O as well as any Group I or II atoms. This is because stuff like those in Group V (e.g. nitrogen) have many oxidation states. For example, the oxidation state of N in HNO_3 is $0 - (+1) - [3 \times (-2)] = +5$. It's easier to use what you know to find what you don't.
- Remember to check charge at the end.
- CHARGE IS NOT THE SAME AS OXIDATION STATE. They may coincide, though.

8e$^-$ (LOL) Redox reagents and colour changes ($K_2Cr_2O_7$, $KMnO_4$, KI)

Overkill!

State of sample	Oxidising Agent Test	Reducing Agent Test
Insoluble Solid	**Add Conc. HCl (aq) + Heat** Pale green gas (Cl_2) evolved that bleaches moist blue litmus paper Oxidation: $2Cl^-$ (aq) → Cl_2 (g) + $2e^-$	**Add Conc. HNO_3 (aq) + Heat** Brown gas (NO_2) evolved Reduction: NO_3^- (aq) → NO_2 (g) + O
Solution	**Add Acidified KI (aq) /+ Starch (aq)** Brown solution obtained/ Dark blue solution obtained Oxidation: $2I^-$ (aq) → I_2 (g) + $2e^-$ **Add Acidified $FeSO_4$ (aq)** Pale green solution becomes yellow Oxidation: Fe^{2+} (aq) → Fe^{3+} (aq) + e^-	**Add Acidified $K_2Cr_2O_7$** Orange solution becomes green Reduction: $Cr_2O_7^{2-}$ (aq) + $14H^+$ (aq) + $6e^-$ → $2Cr^{3+}$ (aq) + $7H_2O$ (l) **Add Acidified $KMnO_4$ (aq)** Purple solution becomes colourless Reduction: MnO_4^- (aq) + $8H^+$ (aq) + $5e^-$ → Mn^{2+} (aq) + $4H_2O$ (l) **Add Acidified $FeCl_3$ (aq)** Yellow solution becomes pale green Oxidation: Fe^{3+} (aq) + e^- → Fe^{2+} (aq)
Gas	**Bubble into Acidified KI (aq)** Brown solution obtained Oxidation: $2I^-$ (aq) → I_2 (g) + $2e^-$ **Bubble into Acidified $FeSO_4$ (aq)** Pale green solution becomes yellow Oxidation: Fe^{2+} (aq) → Fe^{3+} (aq) + e^-	**Bubble into Acidified $K_2Cr_2O_7$** Orange solution becomes green Reduction: $Cr_2O_7^{2-}$ (aq) + $14H^+$ (aq) + $6e^-$ → $2Cr^{3+}$ (aq) + $7H_2O$ (l) **Bubble into Acidified $KMnO_4$ (aq)** Purple solution becomes colourless Reduction: MnO_4^- (aq) + $8H^+$ (aq) + $5e^-$ → Mn^{2+} (aq) + $4H_2O$ (l) **Bubble into Acidified $FeCl_3$ (aq)** Yellow solution becomes pale green Oxidation: Fe^{3+} (aq) + e^- → Fe^{2+} (aq)
-	**Reducing agents for test** $Na_2S_2O_3$, SO_2, CO, H_2, Mg, Zn	**Oxidising agents for test** H_2O_2, MnO_2, Cl_2

EXTRA | Solubility rules

For some evil reason, this wasn't in the scope. Woopsies.

Soluble	Insoluble	
Chlorides	Carbonates	
Sulphates	Hydroxides	
Nitrates	Oxides	
Group I salts	$PbCl_2$	$PbSO_4$
Ammonium salts	AgCl	$CaSO_4$
		$BaSO_4$

EXTRA | An Appeal for Donations

If these notes have benefitted you/made your life a lot easier, you might want to consider buying some of my Red Cross International Bazzar Tickets.

The Singapore Red Cross has organised an International Bazaar held on Sunday, 25th of October, 2009 and the Royal Thai Embassy (This is the Sunday after your Bio/MT papers ☺). Funds from ticket sales will be used to further humanitarian work locally and overseas (SRC's aid for Typhoon Ketsana and the Padang earthquakes, for example).

Each ticket costs $10, but you can always team up with others to buy a ticket (5 x $2 = $10, yay!) Tickets can be used as "cash" to buy items from booths at the Bazaar or participate in games there. Or you can just donate money so that I have to sell less of these, haha…

It's okay if you don't want to donate…either way, send me an email at jeremyteoh [at] gmail [dot] com if you liked the notes, along with any comments/constructive criticism so that I can improve on my next set of notes (COUGH BIO COUGH)

Happy mugging!
Jeremy (3K'09)