CHEMISTRY 11 : FINAL EXAM NOTES

chapter one : activities of science

- **observations:** what your sight detects
- **inferences:** what you interpret from observations
- **classifications:** process of placing similar things into categories
- **hypothesis:** temporary explanation for an observed regularity
  (**theory**: a hypothesis in science)
- **model:** representation intending to convey info

MATTER

- **matter:** anything with mass & volume; matter is conserved
- **inertia:** resistance to change an object's motion (ex. to stop, move, change direction)
- **mass:** what gives matter its weight & inertia
- exists in physical states: solid, liquid, gas, **plasma** (like gas but composed of charged particles like electrons instead of uncharged ones)

ENERGY

- **energy**; anything that isn't matter & can cause a change in matter; energy is conserved
  - **potential energy:** a matter's capacity for changes due to its position/configuration
  - **kinetic energy:** energy of motion

MEASUREMENTS IN SCIENCE

- quantity has both magnitude & unit! (ex. 60 kg)
  * only like quantities (same units) can be subtracted or added*
- **SI**: international measurement system based on powers of 10

  ↳ SI base units...

  length – metre (m)
  mass – kilogram (kg)
  time – second (s)
  temperature – Kelvin (K)
  electric current – ampere (A)
  amount of substance – mole (mol)

  ↳ their prefixes...

  10^9 giga (G)
  10^6 mega (M)
  10^3 kilo (k)
  10^2 hecto (h)
  10^1 deca (da)
  10^{-1} deci (d)
  10^{-2} centi (c)
  10^{-3} mili (m)
  10^{-6} micro (μ)
  10^{-9} nano (n)

  ↳ SI derived quantities...

  area – square metre (m^2)
  volume – cubic metre (m^3)
  speed – metre per second (m/s)
  wave number – 1 per meter (m^{-1})
  density – kilogram per cubic metre (kg/m^3)
  force – Newton (N from $kg{\cdot}m/s^2$)
  pressure – Pascal (Pa from N/m^2)
  energy – joule (J from N/m)
- KNOW THESE: 1 ml = 1 cm^3 / 1000 L = 1 m^3 / 1 tonne = 1000 kg

THE CONVERSION METHOD

STEPS:

1. Multiply by the conversion factor.
   ↳ place the desired unit on top
   ↳ put "1" in front of the larger unit

ex.700 cm = ____ hm?

$7000 \text{ cm} \times \frac{1 \text{ hm}}{10000 \text{ cm}} = 0.7 \text{ hm}$

ex. 0.0000034 mm^3 = ___ μm^3?

2. Calculate (cancel units). $0.0000034\ mm^3 \times (\frac{1000\ \mu m}{1\ mm})^3 = 3400\ \mu m^3$

DENSITY

ex. What mass of mercury (density 13.6g/ml) will occupy 25000 µl?

$25000\ \mu l \times \frac{1\ ml}{1000\ \mu l} = 25\ ml$

density = mass/volume

13.6 g/ml = mass/25 ml ∴ mass = 340g

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chapter two : finding out about matter

Physical properties can be determined without changing the substance.
ex. colour, smell, conductivity, density, freezing point

↳ intensive ones are used to identify the substance (ex. freezing point, smell)

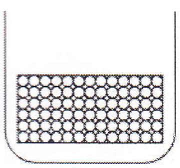
↳ extensive ones don't tell us anything (ex. mass – 5 grams..of what?)

Chemical properties is the ability of a substance to undergo chemical reactions

Physical changes could be reversible; no new substances formed.
Chemical changes are harder to reverse; form new substance with different properties.

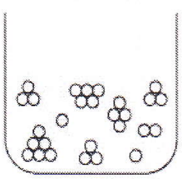
↳ indicators: heat, light, sound produced, colour change, bubbles, cloudiness, precipitate forms, reactants are used up

PHYSICAL STATES



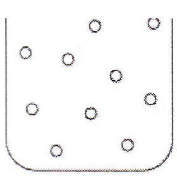
SOLID
mass = definite
volume = definite
shape = definite

Δ temperature = no effect
Δ pressure = no effect
(vibrational energy)



LIQUID
mass = definite
volume = definite
shape = indefinite

Δ temperature = ↑temp, ↑volume
Δ pressure = no effect
(vibrational, rotational, translational energy)



GAS
mass = definite
volume = indefinite
shape = indefinite

Δ temperature = ↑temp, ↑volume
Δ pressure = ↑pressure, ↓ volume
(mainly translational energy)

CLASSIFICATION OF MATTER

Homogeneous substances appear to be the same throughout; could be a...

↳ **pure substance:** element or compound

↳ **homogeneous solution:** components evenly distributed & miscible

Heterogeneous substances are mixtures with observable segregation of its components, could be...

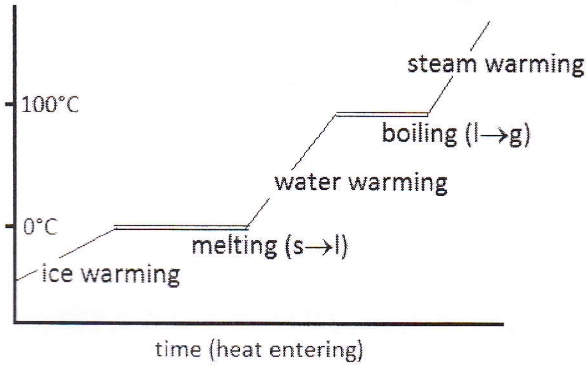
↳ **mechanical:** ex. trail mix

↳ **suspension:** ex. lemonade with pulp

MIXTURES : SEPARATION METHODS

- **hand separation –** hand, sieve, magnet
- **filtration** – ex. filter out sand from salt water
- **evaporation** – using differing evaporation points (ex. boil water to dry out water, leaving salt behind)
- **distillation –** 2 liquids with different boiling points; collect & condense the gas of the one with lower boiling point
- **chromatography** – separates a solution with 2^+ solutes; use of their different solubilities
- **solvent extraction** – adding a solvent that better dissolves the solute than the current solvent
- **separatory funnel –** components separate b/c different densities
- **centrifuge –** spins to separate substances of different densities

HEATING CURVES



- PURE SUBSTANCE: temperature stays constant for duration of the physical state change process
- MIXTURES: not usualyl

ELEMENTS & COMPOUNDS

- **elements:** pure substances that can't be broken down into simpler substances
- **compounds:** chemically combined elements
  - ↳ **law of multiple proportions** (ex. carbon & oxygen can make CO_2, CO, CO_3)
  - ↳ **law of definite composition** (ex. water's H & O ratio is always 2-1)
- atoms, molecules, ions (cations, anions, polyatomic ions)
- metals – conduct heat & electricity, malleable, high melting points, positive ions
- non-metals – doesn't conduct, brittle, negative ions
- **binary compounds**: compounds consisting of only 2 elements
- **diatomic molecules:** H_2 O_2 F_2 Br_2 I_2 N_2 Cl_2
- **electrolysis**: passing through electric current to cause chemical change

NAMING COMPOUNDS & WRITING FORMULAS

- IONIC: - ide endings
  - multivalent cases
  - reduce subscript ratio to lowest possible
  - brackets around polyatomic if more than 1
- COVALENT: - mono, di, tri, tetra, penta, hexa, hepta, octa, nona, deca
  - no prefix if there's only 1 of the first element
  - usually non-metals with non-metals

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c h a p t e r   t h r e e : n u m b e r s   l a r g e & s m a l l

Precision is the uncertainty in measurement.

- precision ↑, uncertainty ↓
- the more sig. figs a measuring device gives, the ↑ precision

- **absolute uncertainty**: ex. a ruler calibrated to 0.1cm has uncertainty of ±0.01cm
- **relative uncertainty**:

  ex. 10.55cm with 0.01cm uncertainty – $\frac{0.01\ cm}{10.55\ cm}$ = 0.0009479

  ex. 0.016cm with 0.001cm uncertainty – $\frac{0.001\ cm}{0.016\ cm}$ = 0.0625

  ∴ relative uncertainty of 2nd example is greater, even though this instrument measures more precisely

Accuracy is how close one comes to the correct value.

SIGNIFICANT DIGITS

= all the certain digits + one uncertain

- don't count placement zeros (ex. <u>0.00</u>78 has 2 sig. figs)
- if zero is a sig. fig, show in scientific notation (ex. 3.00 x 10^2 instead of 300)
- any zeros between sig. figs are significant (ex. 23000001)

  * 200. → 3 sig. figs 200 → 1 sig. fit*
- when doing calculations, your answer cannot get more precise...

  ↳ <u>ADD/SUBTRACT:</u> use the least # of decimal places

  ex. 3.<u>462</u> + 7.00924 = 10.<u>471</u> (3 decimal places)

  ex. 564 – 7.1 = 557 (no decimal places)

  ↳ <u>MULTIPLY/DIVIDE:</u> use the least # of sig. figs

  ex. 3.4617 x 10^7 ÷ 0.00<u>80</u> x 10^2 = <u>4.3</u> x 10^7 (2 sig. figs)
- defined numbers & counting numbers don't affect sig. figs (ex. 12 in a dozen, 2 H atoms in water, 100 cm in a metre)
- rounding fives – round up if it gives the next even number

VOCABULARY

- **extensive property:** dependent on the amount measured (ex. mass)
- **intensive property:** not dependent on the amount measured (ex. water's boiling point is same regardless of how much water we have)
- **specific volume:** volume per gram (cm^3/g)

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chapter four : the mole

Atomic mass number in periodic table represents:

- mass of one atom (carbon atom weights 12.0 amu)
- mass of one **mole (6.02 x 10^{23})** of atoms (carbon mole weights 12.0g)

MOLAR MASS OF COMPOUNDS

ex. $K_2Cr_2O_7$

K 2(39.1) +

Cr 2(52.0) +

O 7(16.0) = 294.2 g/mol

ex. neon gas

N 2(20.2) = 40.4 g/mol

molar mass always has 1 decimal

CONVERSION FACTORS

- remember: $\frac{6.02 \text{ x } 10 \text{ p23}}{\text{mol}}$, $\frac{\text{(molar mass)}}{\text{mol}}$, $\frac{22.4 \text{ L}}{\text{mol}}$ gas @ STP, $\frac{\text{x atoms}}{\text{molecule}}$

ex. How many oxygen atoms are in 75.0L of SO_3 (g) at STP?

$$75.0\text{L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} \times \frac{6.02 \text{ x } 10 \text{ p23 molecules}}{1 \text{ mol}} \times \frac{3 \text{ oxygen atoms}}{1 \text{ SO3 molecule}} = 6.05 \times 10^{24} \text{ oxygen atoms}$$

ex. What is the volume occupied by 50.0g of NH_3 (g) at STP?

N 14.0 + $\quad$ $50.0g \times \frac{1\ mol}{17.0\ g} \times \frac{22.4\ L}{1\ mol} = 65.9L$ @ STP

H 3(1.0) = 17.0 g/mol

DENSITY

ex. What is the volume occupied by 3.00 mol of ethanol, CH_3CH_2OH (l), with density 0.790 g/ml?

C 2(12.0) + $\quad$ $3.00\ mol \times \frac{46.0\ g}{mol} \times \frac{ml}{0.790\ g} = 175\ mL$

H 6(1.0) +

O 16.0 = 46.0 g/mol

ex. Al_2O_3 (s) has a density of 3.97 g/mL. How many atoms of Al are in 100 ml?

Al 2(27.0) + $\quad$ $100\ mL \times \frac{3.97\ g}{ml} \times \frac{mol}{102.0\ g} \times \frac{6.02 \times 10\ p23\ molecules}{mol} \times \frac{2\ Al\ atoms}{molecule} = 4.69 \times 10^{24}$ atoms

O 3(16.0) = 102.0 g/mol

ex. What's the density of uranium hexafluoride gas at STP?

U 238.0 + $\quad$ 352.0 g/mol @ STP equals $\frac{352.0\ g}{22.4\ L}$ @ STP

F 6(19.0) = 352.0 g/mol $\quad$ $\frac{352.0\ g}{22.4\ L} = \frac{352.0\ g}{22400\ mL} = 0.0157$ g/mL

FINDING EMPIRICAL FORMULAS

Empirical formula is the smallest whole-number ratio of atoms in compound. We need to have more information to find the **molecular formula** which isn't always in lowest ratio (ex. $C_6H_{12}O_6$).

STEPS: 1. Use percent composition to find out mass of element in 100g.

* skip this if the specific mass of each element is already given*

2. Determine # of moles in that.
3. Determine ratio between moles (by dividing by the smallest #)
4. Create formula using the same ratio on atoms.

ex. Find the empirical & molecular formulas for a compound with 85.7% carbon & 14.3% hydrogen, with molecular weight 28u.

→ *in 100g of this compound, there would be*

$85.7\ g\ C \times \frac{mol}{12.0\ g} = 7.14$ mol $\quad$ & $\quad$ $14.3\ g\ H \times \frac{mol}{1.0\ g} \times 14.3$ mol

$7.14 \div 7.14 = 1$ mol $\quad$ $14.3 \div 7.14 = 2$ mol

→ *the molar ratio between C & H is 1-2 ∴ empirical formula is CH_2*

→ *now check the molecular weight using this empirical formula*

C 12.0 + $\quad$ 14.0u x 2 = 28u

H 2(1.0) = 14.0u $\quad$ ∴ multiply empirical formula by 2 to obtain molecular formula of C_2H_4

ex. In 35.73g of a compound, there's 9.26g nitrogen & 26.47g oxygen.

$9.26\ g\ N \times \frac{mol}{14.0\ g} = 0.66$ mol $\quad$ $26.47\ g \times \frac{mol}{16.0\ g} \times 1.654$ mol

$0.66 \div 0.66 = 1$ mol $\quad$ $1.654 \div 0.66 = 2.5$ mol

→ *multiply ratio 1-2.5 by two to get 2-5 ∴ empirical formula is N_2O_5*

ex. In a hydrate compound there's 43.7g magnesium sulphate & 56.3g water. Find empirical formula.

$MgSO_4$ is 120.4 g/mol $\quad$ H_2O is 18.0 g/mol

$43.7g\ MgSO_4 \times \frac{mol}{120.4\ g} = 0.363$ mol $\quad$ $56.3g\ H_2O \times \frac{mol}{18.0\ g} = 3.13$ mol

$0.363 \div 0.363 = 1$ mol $\quad$ $3.13 \div 0.363 = 8.62$ mol

→ *multiply ratio 1-8.6 by five to get 5-43 ∴ empirical formula is* $5MgSO_4 \cdot 43H_2O$

MOLARITY & CONCENTRATION

- unit: mol/L or M

ex. How many grams of KNO_3 should be used to prepare 2.00L of 0.500M solution?

KNO_3 is 101.1 g/mol

$2.00\text{L} \times \frac{0.500\text{ mol}}{\text{L}} \times \frac{101.1\text{ g}}{\text{mol}} = 101.1\text{ g}$

ex. What's the molarity of H_2SO_4, having density 1.839 g/mL?

H_2SO_4 is 98.1 g/mol

$\frac{1.839\text{ g}}{\text{mL}} \times \frac{\text{mol}}{98.1\text{ g}} = \frac{0.0187\text{ mol}}{\text{mL}} = \frac{0.0187\text{ mol}}{0.001\text{ L}} = 18.7\text{ M}$

Since moles of substance before dilution = moles of substance after dilution, $\mathbf{M_1V_1 = M_2V_2}$

ex. To what volume should 25mL of 15M nitric acid be diluted to prepare a 3.0M solution?

$M_1V_1 = M_2V_2 \rightarrow (15M)(0.025L) = (3.0M)\,V_2 \rightarrow V_2 = 0.12\text{ L}$

VOCABULARY

- **solution:** homogeneous mixture of 2+ substances
- **solvent:** component of solution of greater quantity
- **solute:** component of solution of lesser quantity
- **Avogadro's number:** 6.02×10^{23}
- **Avogadro's hypothesis:** equal volume of any gas at same temperature & pressure contain equal number of particles/moles
- **STP (standard temp. pressure):** 22.4 L/mol
- **RTP (room temp. pressure):** 24.0 L/mol

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c h a p t e r   f i v e   :   c h e m i c a l   r e a c t i o n s

- **law of conservation of mass:** total mass of reactants = total mass of products

TYPES OF REACTIONS

- **synthesis:** $A + B \rightarrow AB$
- **decomposition:** $AB \rightarrow A + B$
- **single replacement:** $X + AB \rightarrow B + AX$ where X is a non-metal with higher **activity**
  $Y + AB \rightarrow A + YB$ where Y is a metal with higher activity
- **double replacement:** $AY\,(aq) + XB\,(aq) \rightarrow AB\,(s) + XY\,(aq)$
  * can be water-forming **neutralization:** acid + base → salt + water
- **combustion:** any reaction with O_2 as a reactant; creates lots of heat
  * combustion of hydrocarbons (CH_4, C_2H_6, etc) → $CO_2 + H_2O$
  * combustion of metals / non-metals → metal oxide / non-metal oxide

ENERGY CHANGES IN REACTIONS

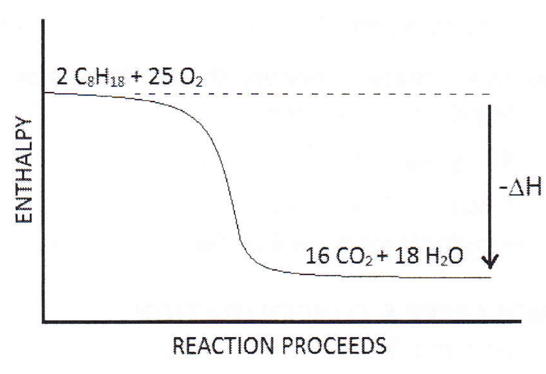
- all substances have **enthalpy H** (chemical potential energy)

<u>**Exothermic**</u>

- converts enthalpy into heat, which is released into surroundings (feels hot)
  reactant → product + kJ (heat)
- products have lower energy
- ΔH = negative
- energy to break bonds < energy released when bonds form

ex. $H + Cl \rightarrow HCl + \underline{432kJ}$

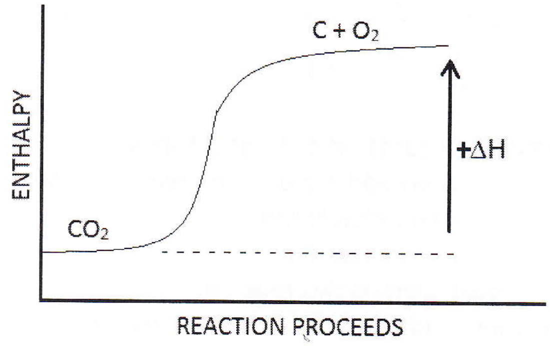
or $H + Cl \rightarrow HC$ $\underline{\Delta H = -432kJ}$



ex. $H_2O_{(g)} \rightarrow H_2O_{(l)}$

Endothermic

- absorbs heat from surroundings & converted into enthalpy (feels cold)
  reactant + kJ (heat) $\rightarrow$ product
- products have higher energy
- ΔH = positive
- energy to break bonds > energy released when bonds form



ex. $CO_2 + \underline{433kJ} \rightarrow C + O_2$

or $CO_2 \rightarrow C + O_2$ $\underline{\Delta H = 433kJ}$

ex. $H_2O_{(s)} \rightarrow H_2O_{(l)}$

ENTHALPY & MOLES

Given $N_2 + 3H_2 \rightarrow 2NH_3 + 46.2$ kJ , it means that

- 46.2 kJ is released for every 1 mol N_2 $\therefore \frac{46.2 \text{ kJ}}{1 \text{ mol N2}}$
- 46.2 kJ is released for every 3 mol H_2 $\therefore \frac{46.2 \text{ kJ}}{3 \text{ mol H2}}$
- 46.2 kJ is released for every 2 mol NH_3 $\therefore \frac{46.2 \text{ kJ}}{2 \text{ mol NH3}}$

ex, How much heat is released during the formation of 14.6 moles of NH_3?

$14.6 \text{ mol} \times \frac{46.2 \text{ kJ}}{2 \text{ mol NH3}} = 337 \text{ kJ}$

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c h a p t e r s i x : s t o i c h i o m e t r y

In a balanced equation, the coefficients = ratio

ex. $2Mg + O_2 \rightarrow 2MgO$ $\therefore$ ratio of Mg-O_2-MgO is 2-1-2 *it's the ratio of molecules/moles, not mass!*

do all calculations using conversion method (refer back to Ch1 notes)

STOICHIOMETRY WITH MASS & VOLUME OF GAS

ex. In $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$, what mass of C_3H_8 is required to produce 100.0g of H_2O?

C_3H_8 is 44.0 g/mol & H_2O is 18.0 g/mol

$100.0\text{g} \times \frac{\text{mol}}{18.0 \text{ g}} \times \frac{1 \text{ mol C3H8}}{4 \text{ mol H2O}} \times \frac{44.0 \text{ g}}{\text{mol}} = 61.1\text{g}$

ex. In that same reaction, how many molecules of CO_2 are produced if 1.25×10^{-6}g of C_3H_8 is burned?

$1.25 \times 10^{-6}\text{g} \times \frac{\text{mol}}{44.0 \text{ g}} \times \frac{3 \text{ mol CO2}}{1 \text{ mol C3H8}} \times \frac{6.02 \text{ x } 10 \text{ p23}}{\text{mol}} = 5.13 \times 10^{16}$ molecules

ex. In $2C_6H_6 + 15O_2 \rightarrow 6H_2O + 12CO_2$, find the mass of C_6H_6 needed to produce 2.66L of CO_2 at STP.

C_6H_6 is 78.0 g/mol

$2.66\text{L} \times \frac{\text{mol}}{22.4 \text{ L}} \times \frac{2 \text{ mol C6H6}}{12 \text{ mol C02}} \times \frac{78.0 \text{ g}}{\text{mol}} = 1.54 \text{ g}$

STOICHIOMETRY WITH MOLARITY

When a volume is mentioned, don't assume 22.4L unless it's STP & a gas!

ex. In $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$, what volume of HCl (0.0010M) is neutralized by 0.750g $CaCO_3$?

$CaCO_3$ is 100.1 g/mol

$0.750\text{g} \times \frac{\text{mol}}{100.1 \text{ g}} \times \frac{2 \text{ mol HCl}}{1 \text{ mol CaCO3}} \times \frac{\text{L}}{0{,}0010 \text{ mol}} = 15.0 \text{ L}$

ex. Above equation: what volume of $CO_{2\,(g)}$ at STP is produced if 1.25L of 0.0055M $HCl_{(aq)}$ reacts with $CaCO_{3\,(s)}$?

$1.25L \times \frac{0.0055 \text{ mol}}{L} \times \frac{1 \text{ mol CO2}}{2 \text{ mol HCl}} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 0.077 \text{ L}$

ex. $H_3PO_4 + 2KOH \rightarrow K_2HPO_4 + 2H_2O$: 19.8mL of H_3PO_4 reacts with 25.0mL of 0.500M KOH, what's its molarity?

$0.025L \times \frac{0.500 \text{ mol}}{L} \times \frac{1 \text{ mol H3PO4}}{2 \text{ mol KOH}} = 6.25 \times 10^{-3}$ mol of H_3PO_4 in that 19.8 mL

$\frac{6.25 \times 10{-}3 \text{ mol}}{0.0198 \text{ L}} = 0.32 \text{ M}$

EXCESS & LIMITING REACTANTS

- we usually add more of one reactant (the cheaper/more accessible one) to ensure that the other reactant will be completely used up.
- **limiting reactant** limits the amount of product we can obtain ∴ we base calculations for theoretical & actual yield on this reactant
- must find out which is limiting when reactants aren't present in exact ratios given in their chem. equation
- **percent yield** = $\frac{actual}{theoretical} \times 100$

ex. If $2H_2 + O_2 \rightarrow 2H_2O$ and 20g of H_2 is burnt with 100g of O_2, which reactant is excess & by how many grams?

moles of H_2 in that? $20g \times \frac{\text{mol}}{2.0 \text{ g}} = 10 \text{ mol } H_2$

moles of O_2 in that? $100g \times \frac{\text{mol}}{32.0 \text{ g}} = 3.12 \text{ mol } O_2$

*Think: for 3.12 mol of O_2, how many moles of H_2 do we really need?

$3.12 \text{ mol } O_2 \times \frac{2 \text{ mol H2}}{1 \text{ mol O2}} = 6.24 \text{ mol } H_2$ needed ∴ H_2 is in excess

$10 \text{ mol } H_2 - 6.24 \text{ mol } H_2 = 3.76 \text{ mol extra}$

$3.76 \text{ mol} \times \frac{2.0 \text{ g}}{\text{mol}} = 7.52 \text{ g}$

ex. If 0.250g of $Ba(OH)_2$ is mixed with 15.0mL of 0.125M HBr, what mass of $BaBr_2$ can be formed if the percent yield is 83%? [Equation: $Ba(OH)_2 + 2HBr \rightarrow BaBr_2 + 2H_2O$]

moles of $Ba(OH)_2$? $0.250g \times \frac{\text{mol}}{171.3 \text{ g}} = 1.46 \times 10^{-3} \text{ mol } Ba(OH)_2$

moles of HBr? $0.015L \times \frac{0.125 \text{ mol}}{L} = 1.88 \times 10^{-3} \text{ mol HBr}$

$1.88 \times 10^{-3} \text{ mol HBr} \times \frac{1 \text{ mol Ba(OH)2}}{2 \text{ mol HBr}} = 0.94 \times 10^{-3} \text{ mol } Ba(OH)_2$ needed ∴ $Ba(OH)_2$ is in excess

*Base all calculations on the LIMITING one from now..

$BaBr_2$ is 297.1 g/mol

$1.88 \times 10^{-3} \text{ mol HBr} \times \frac{1 \text{ mol BaBr2}}{2 \text{ mol HBr}} \times \frac{297.1 \text{ g}}{\text{mol}} \times 83\% = 0.232 \text{ g}$

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chapter eight : composition of the atom

ATOM HISTORY

- early Greeks (Aristotle) - matter's properties: moist, dry, hot, cold (water, earth, fire, air)
- medieval chemists - 'corpuscles'
- *Dalton* - atom is smallest particle of matter
  - atoms of same element have same mass & same chemical behaviour
  - atoms of diff. elements can combine in fixed ratios to produce specific compounds
  - each compound is unique & has certain atoms arranged in particular way

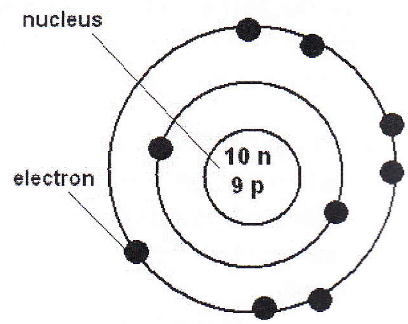
  * Law of Definite Proportions – a compound always contains exactly the same proportion of elements by mass

  * Law of Multiple Proportions – same elements can form different compounds when combined in different whole number ratios

  *Law of Conservation of Mass – mass of a closed system stays the same
- *Thompson* - discovered charged electrons & protons

- **plum pudding model** (bulk is protons, with electrons embedded)
- *Rutherford* - alpha particle gold paper experiment → discovered nucleus
  - equal # of electrons & protons
  - **planetary model** (tiny dense positive nucleus surround by electrons; mostly empty)
- *Chadwick* - found neutrons
- *Bohr* - electrons are restricted to a certain path ('orbital'); remains fixed distance from nucleus
  - electrons can only emit/absorb **quantized** energy when they move orbitals
- *Schrodinger* - quantum mechanics

THE ATOM
- protons: 1.0 amu / 1+
- neutrons: 1.0 amu / no charge
- electrons: 0 amu / 1-
- **ions**: charged particles due to gaining or losing electrons
- **isotopes**: same atomic number but different atomic masses (= different number of neutrons)
  ex. A sample of carbon is found to have 97.00% C-12, 1.00% C-13, and 2.00% C-14. Find the molar mass.
  (0.97 x 12.0 g/mol) + (0.01 x 13.0 g/mol) + (0.02 x 14.0 g/mol) = 12.05 g/mol



BOHR DIAGRAM
- shows # of protons & neutrons + electrons in circular paths around nucleus
- orbitals' electrons: 2, 8, 8, 18, 18

RUTHERFORD MODEL
- tiny dense positively charged nucleus
- negatively charged electrons in free motion around nucleus
- electrostatic attraction between electrons & nucleus (like gravitational attraction in solar system)

BOHR MODEL
- electrons orbit nucleus in orbits of specific size & energy (quantum)
- energy of orbit is related to its size (lowest energy is found in smallest orbit)
- radiation is absorbed or emitted when an electron moves from one orbit to another
- shell structure of atom (electrons in specific orbits in energy levels)

QUANTUM MECHANICAL MODEL
- electrons exist in highest probability orbitals
- principal quantum number represents energy level (there are 7)
- subshells within each level (s, p, d, f) with varying number of orbitals in each subshell

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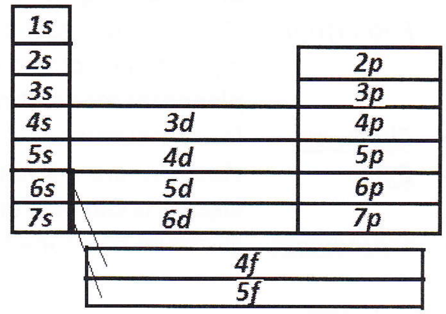
chapter ten : electrons in an atom

- **ground state:** lowest energy state of an atom
- **ionization energy:** energy needed to remove an electron from atom
- **principal quantum number (n):** represents the energy level of an electron
- **isoelectronic**: same electron configuration
- electrons can only exist in specific energy states
- electrons absorb energy = move from an orbit to another (↑orbit's distance from nucleus, ↑energy)
- energy added to atom is absorbed & re-emitted as **photons** → line spectrum when pass through prism
- **energy level:** specific amount of energy electrons can possess

↳ line spectrum is caused by energy level differences

- **quantum** of energy: difference between two different energy levels
- electrons occupy regions of space called **orbital** (hold 2 electrons)



ELECTRON CONFIGURATION

- **sub shell**: a set of orbitals of the same type (ex. there are three 'p' orbitals in 2p subshell)
  - **'s'**- spherical, 1 orbital/2 electrons
  - **'p'** - balloon shaped, 3 orbitals/6 electrons
  - **'d'** - 5 orbitals/10 electrons
  - **'f'** - 7 orbitals/14 electrons

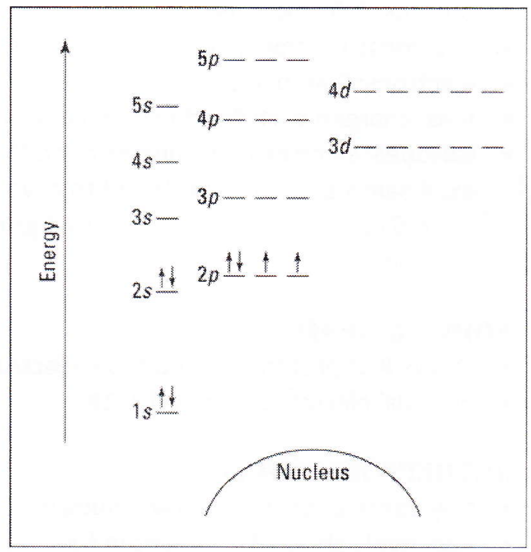
* add electrons to lowest energy orbitals first (ex. 4s > 3d > 4p)

* electrons must be placed singly in sub shells until can be paired

- **core electrons** - same as nearest noble gas
- **valence electrons** - involved in chemical reactions; when counting, exclude core & filled 'd','f' electrons

ex. Ga: [Ar] $4s^2\ 3d^{10}\ 4p^1 \rightarrow$ 3 valence (ignore the [Ar] & $3d^{10}$)

ex. Xe: [Kr] $5s^2\ 4d^{10}\ 5p^6 \rightarrow$ 0 valence (this is a noble gas configuration)



EXCEPTION

- when one electron short of a filled/half-filled 'd', shift an electron from the highest energy 's' shell

ex. Cr: [Ar] $4s^2\ 3d^4$ *..$3d^4$ is one short of a half-filled 'd' shell*

$\rightarrow$ [Ar] $4s^1\ 3d^5$ *..now both are half-filled shells*

ex. Cu: [Ar] $4s^2\ 3d^9$ *..$3d^9$ is one short of a filled 'd' shell*

$\rightarrow$ [Ar] $4s^1\ 3d^{10}$ *..now one is half-filled while the other is filled*

ELECTRON CONFIGURATION: ANIONS

- isoelectronic to nearest noble gas

ex. P: [Ne] $3s^2\ 3p^3 + 3e^- \rightarrow$ [Ne] $3s^2\ 3p^6$ *..same as Argon*

ELECTRON CONFIGURATION: CATIONS

- remove electrons in outermost shell (largest n value); remove p-electrons > s-electrons > d-electrons

ex. Sn: [Kr] $5s^2\ 4d^{10}\ 5p^2 \rightarrow$ [Kr] $5s^2\ 4d^{10} + 2e^-$ *..electrons removed from 5p*

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c h a p t e r   e l e v e n : t h e   p e r i o d i c   t a b l e

DEVELOPMENT

- 1st periodic table organized by atomic mass (Dmitri Meneleev 1869)
- gaps = elements not yet discovered; table could be used to predict properties of missing elements
- **periodic law**: elements' properties recur periodically when elements are arranged in increasing atomic #

DIVISIONS WITHIN THE TABLE

- metals: good conductors, malleable, ductile, high density, high boiling/melting point, solids at room temp.
- non-metals: poor conductors, no luster, low densities, low boiling/melting points
- metalloids: properties of both
- **alkali metals**: group 1 ~~(H)~~, react vigorously with water/oxygen, low density & melting point, conductive

- **alkaline earth metals**: group 2, less reactive than first group
- **halogens**: group 17, form salts with first & second group
- **noble gases**: group 18, generally non-reactive, gas at room temp.
- **transition metals**: groups 3-12, shiny, good conductors
- **lanthanides**: top row of inner trans. metals
- **actinides**: bottom row of inner trans. metals

TRENDS IN THE TABLE

- *<u>**atomic radius** increases down a group</u>*
  More electrons in higher energy levels (larger orbitals).
  More electrons, more **inter-electron repulsion**.
  Even though **nuclear charge** increases, more shielding electrons are there to block the attractive pull.
- *<u>atomic radius decreases across a period</u>*
  More electrons, more protons. Greater nuclear charge pulls electrons closer.
- *<u>cation is smaller than its neutral atom</u>*
  Lose electrons, less electron repulsion, net positive charge pulls electrons closer
- *<u>anion is greater than its neutral atom</u>*
  Gain electrons, more inter-electron repulsion
- *<u>smaller the atom, higher the **ionization energy**</u>*
  Greater atoms, less attractive force on outer electrons.
  Smaller atoms, more attractive force on outer electrons.
- *<u>smaller the atom, higher the **electron affinity** (energy released when electron is added)</u>*
  When electron is added close to nucleus, greater attraction = greater energy.
  exclude noble gases from trend
- *<u>smaller the atom, higher the **electronegativity** (ability to attract electrons)</u>*
  Smaller atoms, closer to nucleus, greater attractive pull.
  * exclude noble gases from trend*
- *<u>metallic properties increases down a group + left a period</u>*
- *<u>reactivity increases down a group of metals</u>*
  More energy levels, more shielding electrons. Attractive pull weakens, valence electrons easier to lose.
- *<u>reactivity decreases down a group of non-metals</u>*
  More energy levels, more shielding electrons. Weakened attractive pull is less likely to attract electrons to fill valence shell.

Always write the electron configuration first !
ex. Which has a higher IE, aluminum or magnesium?
Mg: [Ne] $3s^2$ Al: [Ne] $3s^2$ $3p^1$ $\rightarrow$ Harder to remove a paired electron $\therefore$ Mg has higher IE

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chapter twelve : chemical bonding

UNEQUAL SHARING OF ELECTRONS

- atoms share electrons unequally (in covalent bonds!) b/c electronegativity difference
  - 0.4 or less – covalent; shared equally
  - 0.5-1.6 – **polar covalent**; **dipole** (bond or molecule with δ+ and δ-)
  - 1.7 or more – ionic; electrostatic attraction between + & - ions
- non-polar bonds always = non-polar molecules
- polar bonds arranged symmetrically around central atom = non-polar molecule
- polar bonds arranged asymmetrically around central atom = polar molecule

CHEMICAL BONDS

- potential energy decreases as atoms get closer
- energy released when bonds form, energy absorbed when bonds break
- **bond energy**: energy required to break/form a mole of bond
  ex. $H(g) + Br(g) \rightarrow HBr(g) + 364$ kJ
  *bond strength (greatest to least): triple → double → single bond

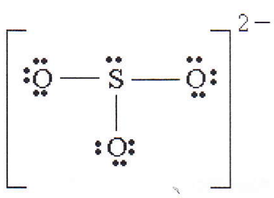
Ionic Compounds & Bonds

- shape: **crystal lattice** arrangement
- crystalline solids at room temp.
- hard / conduct electricity / high melting & boiling points

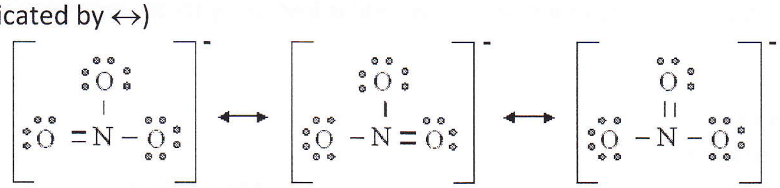
Covalent Compounds & Bonds

- soft / don't conduct electricity / low melting & boiling points
- **octet rule:** bonded non-metals have 8 electrons in outermost shell *(see exceptions on pg 343)*
  ↳ exceptions: H, B, Be fewer than 8; P, S more than 8
- steps to drawing Lewis structures:
  1. Figure out total valence electrons.
  2. Draw that many dots around elements to fulfill octet rule. If there aren't enough…
     → multiply number of atoms by 8 ex. CO_2: 3 atoms X 8 = 24
     → subtract number of total valence electrons 24-16 [C has 4, O each has 6] = 8
     → that's the number of shared electrons :Ö=C=Ö:

  *dative covalent bond: when a pair of electrons are donated by one atom b/c it has spare & the other atom is lacking
  *polyatomic ions: add/subtract from total valence electrons
  ex. SO_3^{-2}: S has 6, O each has 6, plus 2 extra electrons = 26 total valence electrons



- **resonance:** various structures for a molecule that can't be adequately represented by a single structure (indicated by ↔)
  ex.



MOLECULES & GEOMETRIC SHAPES

- **valence shell electron pair repulsion (VSEPR) theory**: theory describing the equal distribution of electron pairs (bonding & non-bonding) to be as far apart as possible to minimize repulsions
  - 1 bond = linear
  - 2 bonds = linear/linear triatomic
  - 2 bonds + 2 lone pairs = angular/bent
  - 3 bonds = trigonal planar *see pg 352*
  - 3 bonds + 1 lone pair = trigonal pyramidal
  - 4 bonds = tetrahedral
  - 5 bonds = trigonal bipyramidal
  - 6 bonds = octahedral
- for complex molecules: look only at a central atom

INTERMOLECULAR ATTRACTION FORCES

- **van der Waals forces:** name for intermolecular forces (ex. London, dipole-dipole)
- **dipole-dipole:** attraction caused by polar molecules' partially charged ends (can be intramolecular too)
  ∴ more energy to separate → higher boiling points
- **hydrogen bond**: attraction between positive H atom in one molecule & an extremely electronegative atom of another similar molecule (ex. H_2O)
  ↳ H is small & has no unbonded electrons ∴ hydrogen bond is stronger than other dipole-dipole forces
- **London/dispersion**: weak, short-lived attraction between non-polar molecules b/c of momentary dipoles (random movements of electrons in atom = more electrons on one side of atom for brief time)
  ↳ the force increases as # of electrons/size increases
- **electrostatic:** attraction force between anions & cations

VOCABULARY

- **chemical bond**: result of mutual attraction of two atomic nuclei for electrons
- **crystal lattice**: pattern created by + & - ions in crystal structure
- **electronegativity:** ability to attract electrons

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c h a p t e r   s i x t e e n   :   s o l u t i o n s

VOCABULARY

- **mixture:** combination of different matters that still retain own properties
  ↳ **homogeneous:** components evenly distributed (ex. salt water)
  ↳ **heterogeneous** components evidently segregated (ex. muddy water
- **solution:** homogeneous mixture of 2 or more substances
- **salvation:** interaction between **solvent** (substance of greater quantity) & **solute** (substance of lesser quantity)
- **soluble/insoluble & miscible/immiscible:** ability/inability to dissolve or mix in solution
- **solubility (g/ml):** amount of solute able to dissolve in solvent at specific temp. (↑temp., ↑solubility)
- **saturated solution:** max. quantity of solute has been dissolved at given temp.
- **concentrated/diluted:** relatively large/small amount of solute dissolved
- **precipitate:** insoluble solid product from chemical reaction (causes cloudiness, settle to bottom, etc)
- **filtrate:** solid precipitate captured on filter paper
- **electrolytic sol'n:** conducts electricity b/c contains ions when dissolving ionic solids
  ↳ water needed to **dissociate** (break down ionic compound into ions)

LIKE DISSOLVES LIKE

- polar dissolves polar solute (*organic acids & alcohols = polar)
  ↳ dipoles of solute attract dipoles of solvent
- non-polar dissolves non-polar solutes
- organic dissolves organic
- ionic dissolve in water (which is polar) only!

EQUATIONS FOR SOLUTION CHEMISTRY

- "[ ]" = concentration of whatever substance inside the bracket
- use solubility chart to determine whether a solid ppt will form

Molecular equation show all as neutral:

ex. $2NaOH_{(aq)} + Pb(NO_3)_{2\,(aq)} \rightarrow Pb(OH)_{2\,(s)} + 2NaNO_{3\,(aq)}$

Total ionic equation show all, except the precipitate, as dissociated:

ex. $2Na^{1+}{}_{(aq)} + 2OH^{1-}{}_{(aq)} + Pb^{2+}{}_{(aq)} + 2NO_3{}^{1-}{}_{(aq)} \rightarrow Pb(OH)_{2\,(s)} + 2Na^{1+}{}_{(aq)} + 2NO_3{}^{1-}{}_{(aq)}$

Net ionic equation exclude spectator ions:

ex. $2OH^{1-}_{(aq)} + Pb^{2+}_{(aq)} \rightarrow Pb(OH)_{2\,(s)}$

*if no precipitate is formed, there's no net ionic equ'n

CONCENTRATION, DILUTION, TITRATION

- **$M_1V_1 = M_2V_2$**
- be aware of new volume!
- use mole ratios for concentrations for individual ions (ex. 2 mol Cl^- for every mol of $CaCl_2$)

ex. How many molecules of solute are in 0.050 mL of 18M H_2SO_4?

$$5.0 \times 10^{-5}\text{ L} \times \frac{18\text{ mol}}{\text{L}} \times \frac{6.02 \times 10\text{ p23 molecules}}{\text{mol}} = 5.4 \times 10^{20}\text{ molecules}$$

ex. What volume of 2.50M HCl is needed to neutralize 32.5 mL of 0.412M $Ca(OH)_2$?

balanced equation: $2HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O$

$$0.0325\text{L} \times \frac{0.412\text{ mol}}{\text{L}} \times \frac{2\text{ mol HCl}}{1\text{ mol Ca(OH)2}} \times \frac{\text{L}}{2.50\text{ mol}} = 0.0107\text{ L}$$

ex. What concentration results when 150 mL of 0.36M $MgSO_4$ are added to 750 mL water?

new volume: 0.15L + 0.75L = 0.90 L

$$0.15\text{ L} \times \frac{0.36\text{ mol}}{\text{L}} = 0.054\text{ mol} \therefore \text{ we have } \frac{0.054\text{ mol}}{0.90\text{ L}} = 0.060\text{M}$$

ex. What final $[Cl^-]$ results when 162 mL of 1.50M $CaCl_2$ + 278 mL of 4.00M $AlCl_3$?

new volume: 0.162 L + 0.278 L = 0.440 L

mol of Cl^- in the 0.162L: $0.162\text{L} \times \frac{1.50\text{ mol}}{\text{L}} \times \frac{2\text{ mol Cl}-}{1\text{ mol of CaCl2}} = 0.486\text{ mol}$

mol of Cl^- in the 0.278L: $0.278\text{L} \times \frac{4.00\text{ mol}}{\text{L}} \times \frac{3\text{ mol Cl}-}{1\text{ mol of AlCl3}} = 3.336\text{ mol}$

$$\frac{0.486\text{ mol Cl} + 3.336\text{ mol Cl}}{0.440\text{ L}} = 8.69\text{ M}$$
