

# Elements of the Periodic Table:

H (Hydrogen)

He (Helium)

Li (Lithium)

Be (Beryllium)

B (Boron)

C (Carbon)

N (Nitrogen)

O (Oxygen)

F (Fluorine)

Ne (Neon)

Na (Sodium)

Mg (Magnesium)

Al (Aluminium)

Si (Silicon)

P (Phosphorus)

S (Sulfur)

Cl (Chlorine)

Ar (Argon)

K (Potassium)

Ca (Calcium)

Cr (Chromium)

Mn (Manganese)

Fe (Iron)

Co (Cobalt)  
Ni (Nickel)  
Cu (Copper)  
Zn (Zinc)  
As (Arsenic)  
U (Uranium)  
Se (Selenium)  
Mo (Molybdenum)

Br (Bromine)

Sb (Antimony)

Pd (Palladium)

Bi (Bismuth)

Ag (Silver)

V (Vanadium)

Cd (Cadmium)

Ga (Gallium)

Sn (Tin)

Ge (Germanium)

I (Iodine)

Te (Tellurium)

Kr (Krypton)

Xe (Xenon)

Pt (Platinum)

Au (Gold)

Hg (Mercury)

Rb (Rubidium)

Sr (Strontium)

Cs (Cesium)

Ba (Barium)

Pb (Lead)

W (Tungsten)

Ti (Titanium)

Os (Osmium)

Rn (Radon)

## Scientific Method

### 1. Observe

→ Qualitative: observation not requiring measurement

- color

- state of matter

→ Quantitative: measured value (Number + unit)

### 2. Formulate hypothesis (explain observation)

### 3. Perform experiments to test hypothesis

## Law / Theory

Law = summary of facts / observations

ex: Law of conservation of mass (can be observed)

Laws sometimes come with certain caveats (They only apply in certain conditions)

Theory = the "why" of many experiments. Theories are consistent with many experiments

ex: "Matter is consisted of solid indivisible particles known as atoms"

This is an explanation, Dalton didn't actually observe atom's.

This theory didn't stand the test of time. We can split atoms today.

Theories can be modified as our understanding evolves.

## Measurements

P	n	$\mu$	m	c	d	k	M
$10^{-12}$	$10^{-9}$	$10^{-6}$	$10^{-3}$	$10^{-2}$	$10^{-1}$	$10^3$	$10^6$

### SI Units:

Time: s , Mass: kg , Temperature: K , Length: m  
 $(^{\circ}\text{C} + 273.15)$

Relationship between volume units:  $1 \text{ L} = 1 \text{ dm}^3$   
 $1 \text{ mL} = 1 \text{ cm}^3$

## Density

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad (\text{g/cm}^3)$$

→ density varies with temperature

## Taking Measurements

Accuracy = how close a measurement is to the true value

Precise = Degree of agreement between several measurements (\* Number of decimal places)

ex: Precise within 0.002 g

15.456 g  
15.454 g  
15.455 g

This can still be "inaccurate" (miss-calibrated balance)

## When measuring

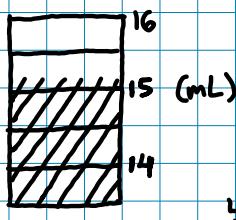
Record

All certain digits

+ 1 uncertain digit

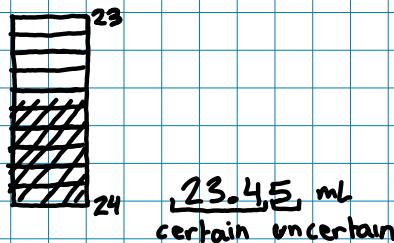
- Digital balance: All digits recorded.  
The uncertain digit is the last visible one on the balance.
- Glassware: Read volume from the bottom of the miniscus

### Graduated Cylinder



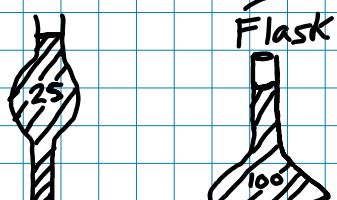
Always record  
1 decimal

### Burret



Always record  
2 decimals

### Volumetric Pipette



If liquid fills up to  
the line:  
"number on pipette"  
+ .00 mL

# Significant Figures

sig figs = certain + 1 uncertain

Rules:

- 1) All non-zero numbers are significant
- 2) Zeroes surrounded by non-zeroes are significant
- 3) Zeroes preceding the 1<sup>st</sup> non-zero is NOT significant
  - These just give us to how small a number is by a factor of 10
  - Does not add precision
  - ex: convert 0.00651 g → 6.51 mg
- 4) Zeroes following the last non-zero (trailing zero):
  - a) A trailing zero after a decimal is significant
  - b) If it is not after a decimal, it is only significant if there is a decimal point following

ex: How many sig figs in 400m?  
1 sig fig  
0 certain numbers  
(sig figs = certain + 1 uncertain)

1000  
View from telescope.  
X located at 400 m.

Express 400 m in 2 sig figs:  $4.0 \times 10^2$  m

Express 400 m in 3 sig figs: 400. m or  $4.00 \times 10^2$  m  
add decimal

## Exact numbers

Not obtained by measurement

→ counted numbers

→ definition (ex:  $1 \times 10^{-6}$  m = 1 μm)

\* Infinite sig figs.

→ When multiplying by exact numbers, assume you are "adding" and use "addition rules".  
(only when this adds precision)

## Multiplication/Division (sig figs)

- The number of sig figs in the product/quotient equals the smallest number of sig figs used in measured quantities.

## Adding / Subtracting (sig figs)

- Round so that there are the same number of decimal places as the measurement with the fewest decimal places.
- When adding, make sure all numbers have decimals

$$\text{ex: } 600 \text{ m} \rightarrow 6. \times 10^2 \text{ m}$$

(1 sig fig)

- When adding, make sure all numbers have the same  $\boxed{10^2}$ .

$$\begin{aligned} \text{ex: } 2.7 \times 10^2 + 5.2 \times 10^2 &\rightarrow 0.27 \times 10^2 + 5.2 \times 10^2 \\ &= 5.47 \times 10^2 \\ &\uparrow 5.5 \times 10^2 \quad (1 \text{ decimal}) \end{aligned}$$

## Converting Units (Unit-Factor)

→ Multiply by one or more unit factors (eliminate units)

- Unit factor :  $\frac{\text{value}}{\text{value}} \frac{\text{unit 1}}{\text{unit 2}} = 1$

ex: what is the mass of 0.3450 L of benzene?  
( $d = 0.8765 \text{ g/cm}^3$ )

$$0.3450 \cancel{L} \cdot \frac{1000 \cancel{mL}}{1 \cancel{L}} \cdot \frac{1 \cancel{cm}^3}{1 \cancel{mL}} \cdot \frac{0.8765 \cancel{g}}{1 \cancel{cm}^3} = 302.3925 \text{ g}$$

$\downarrow \quad \downarrow \quad \downarrow \quad \downarrow$   
4 sig figs    3 sig figs (definition)    4 sig figs    302.4 g

\* Note: you multiply by 1, get the same thing (different unit)

ex: what is  $37.0^\circ\text{C}$  in Kelvin?

$$\begin{array}{r} 37.0^\circ\text{C} \leftarrow 1 \text{ decimal point} \\ + 273.15 \\ \hline 310.15 \text{ K} \end{array}$$

## Averages

- As precise (number of decimals) as the least precise measurement.

# Atom

→ fundamental particle of a given element.  
 → the smallest thing that still is a given element

Particle = tiny entity

Species = thing

Atoms are charged species (they contain charged particles).  
 However the atom as a whole is neutral and has an equal number of  $\oplus$  and  $\ominus$

Subatomic particle:

- Nucleus = central core of atom
- Proton =  $\oplus$  charged particle in nucleus
- Electron =  $\ominus$  charged particle (not nucleus)
- Neutron = neutral particle in nucleus

Atoms are composed of a small  $\oplus$  charged nucleus  
 Large area where  $\ominus$  particles exist (electron cloud)

If nucleus is the size of a marble  
 atom is the size of a baseball field

Subatomic Particle	Location	Relative mass (mass proton = 1)	Relative charge (charge proton = 1+)
Proton	nucleus	1	1+
Neutron	nucleus	1	zero
Electron	large area	1/2000	1-

All atoms of the same element have the same number of protons.

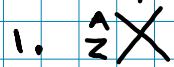
Element: A substance that cannot be decomposed into simpler substances

Compounds: Substance that is composed of 2 or more elements (different) in fixed integer ratios

$\text{SO}_2$  = sulfur dioxide 1:2 ratio S to O

Isotope: atoms with same number of protons  
but different number of neutrons  
↳ different mass.

### Identifying isotopes



2. Element - A

X = symbol for element

A = mass number (sum of protons and neutrons)

Z = atomic number (number of protons)

N = number of neutrons

$$A = Z + N$$

\* Always same number of protons and electrons

### Molecule vs Compound

- Molecule = the smallest particle of a group of 2 or more atoms where electrons are shared between atoms.

→ smallest thing of the compound that still is that given substance

→ molecule doesn't necessarily need different types of elements

- Compound = the entirety of the sample

→ must have different types of elements

\* To be a molecule, electrons have to be shared  
This is only the case between non-metals

- Molecule: shared electrons (covalent)
- Ionic compound: not shared (ionic)

# Chemical Formula

Indicates the chemical composition of a molecule or compound (2 kinds)

## 1. empirical formula

Shows the relative number of each element present, using the integer ratios

## 2. molecular formula

Shows the actual number of each element

ex: benzene  $\rightarrow$  Molecular formula : C<sub>6</sub>H<sub>6</sub>  
Empirical formula : CH

# Ions

Atom that has lost or gained an electron

There exists mono-atomic ions (simple) and poly-atomic ions (group of atoms).

A charge of n+, means there are n more protons  
A charge of n-, means there are n ~~more~~ electrons

Simple ions: 41 Ca<sup>2+</sup>  
protons: 20  
neutrons: 21

electrons: 18 (lost 2)

79 Se<sup>2-</sup>  
protons: 34  
electrons: 36 (gained 2)  
neutron: 45

Polyatomic: CO<sub>3</sub><sup>2-</sup>

$$(1 \times 6 p^+) + (3 \times 8 p^+) = 30 \text{ protons}$$

$$(1 \times 6 n^+) + (3 \times 8 n^o) = 30 \text{ neutrons}$$

electron: 32 (gained 2)

Cation = positive ion (cations = positive/cute)

Anion = negative ion (anions = negative)

Ionic bond = attractive force between cation and anion

Ionic compounds (salt) = held together by ionic bonds; an electrostatic attraction (+/-)

## Periodic Table

Metals: atoms located to the left of the staircase

→ when part of an ionic compound, lose electrons to become (+) cations.

→ conduct electricity and heat

Non-metals: atoms located to the right of the staircase

→ when part of an ionic compound, gain electrons to become (-) anion

Metalloids (semi-metals): Touch the staircase

Rows are called periods ↔ (electron shells)

Columns are called groups ↑ (valence electrons)

Transition metals:

- Columns with top element Sc to Zn
- "Roman numeral" + B

Lanthanides: Top of the 2 bottom rows (14 elements)

Actinides: Bottom of the 2 bottom rows (14 elements)

Main group elements (representative elements):

- "Roman numeral" + A

## Metals

<u>Group</u>	<u>Name</u>	<u>Charge of cation (+)</u>
I A (1)	Alkali metals	1+
II A (2)	Alkaline-earth metals	2+
III A (3)	Aluminium group	3+ (Al) (* Not necessarily for every element in that group)

## Nonmetals

<u>Group</u>	<u>Name</u>	<u>Charge of anion (-)</u>
VA (15)	Nitrogen group	3-
VI A (16)	Oxygen group	2-
VIIA (17)	Halogens	1-
VIIIA (18)	Noble gases	(* won't be found in an ionic compound)

If X is the roman numeral:

- metals form a cation of charge X
- non-metals form an anion of charge 8-X

\* These elements, when not part of an ionic compound are neutral (charge = 0)

# Average Atomic Mass

$Z$
$X$
amu mass

Average atomic mass for an element of that atom (in amu)

amu =  $\frac{1}{12}$  the mass of one atom of carbon -12

The average atomic weight is numerically equal to the average molar mass of an atom.  
 $\hookrightarrow$  (amu = g/mol)

Average  $\rightarrow$  it takes into account the relative distribution of the isotopes of the element.

$\rightarrow$  Based on the % abundance/occurrence

$$\text{Average atomic mass} = \sum \left( \frac{\% \text{ abundance isotope } n}{100} \cdot \text{mass of isotope } n \right)$$

The molar mass (g/mol) is numerically equal to the atomic mass

Mole is a number :  $6.022 \times 10^{23}$  (4 sig figs)

$\rightarrow$  Avogadro's number ( $= N_A$ )

ex: The mass of  $6.022 \times 10^{23}$  (1 mol) atoms of Cl is 35.45 g/mol

The mass of 1 atom of Cl :

- NOT 35.45 amu (average of isotopes)
- Each atom will be one of the isotopes  $^{35}\text{Cl}$  or  $^{37}\text{Cl}$  ( $\approx 35$  amu or  $37$  amu)

ex: How many atoms of P in 16.3 g of P?

$$16.3 \text{ g} / 30.97 \text{ g/mol} / 6.022 \times 10^{23} \text{ atoms/mol}$$

$$16.3 \text{ g} \cdot \frac{1 \text{ mol}}{30.97 \text{ g}} \cdot \frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol}} = 3.1694 \times 10^{23} \text{ atoms} \\ = 3.17 \times 10^{23} \text{ atoms}$$

Sig Figs: 3      4      4

# State of matter

(g) : gas      (s) : solid      (l). liquid      (aq) : aqueous

## Diatomeric Molecules

elements that exist in pairs ( $X_2$ )

I<sub>2</sub>   H<sub>2</sub>   N<sub>2</sub>   Br<sub>2</sub>   O<sub>2</sub>   Cl<sub>2</sub>   F<sub>2</sub>  
I      Have   No      Bright      Or      Clever      Friends

## Allotropes

Different structural forms of an element  
Bonded differently

ex: allotropes of oxygen = O<sub>2</sub>(g), O<sub>3</sub>(g)  
<sup>ozone</sup>

ex: allotropes of carbon = C(s, graphite), C(s, diamond),  
C<sub>60</sub> (buckminsterfullerene)

# Nomenclature

Binary compounds = compounds composed of 2 different elements

Binary ionic compounds = combination of a non-metal and metal  
(gain electron) (lose electron)

## Charge neutrality

\* Only for ionic compounds (not molecules)

→ The total charge of (+) cations equal the total charge of (-) anions  
↳ overall neutral compound

→ The ratio of number of anions to cations is the smallest possible integer ratio

\* 2 exceptions: compounds with Mercury (I) ( $\text{Hg}_2^{2+}$ ) and with Peroxide ( $\text{O}_2^{2-}$ )

$$(\frac{\text{number of cations}}{\text{cations}} \times \text{cation charge}) + (\frac{\text{number of anions}}{\text{anions}} \times \text{anion charge}) = \text{zero}$$

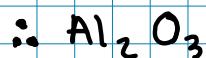
Molecular formula: metal → non-metal

ex: What is the molecular formula of an ionic compound with Al and O?

$$(\frac{\text{number of Al}}{\text{Al}} \times \text{Al charge}) + (\frac{\text{number of O}}{\text{O}} \times \text{O charge}) = \text{zero}$$

$$(\square \times 3^+) + (\square \times 2^-) = \text{zero}$$

$$(2 \times 3^+) + (3 \times 2^-) = \text{zero}$$



## Binary Ionic Compounds (Type I)

→ The cation (metal) must have only one possible number of positive charge when present as part of an ionic compound

In periodic table:

1. Alkali metals (when in compound → 1+)
2. Alkaline-earth metals (when in compound → 2+)
3. Aluminium / Gallium (when in compound → 3+)
4. 3 cases of transition metals
  - Cd<sup>2+</sup>
  - Ag<sup>+</sup>
  - Zn<sup>2+</sup>

Naming compound

- Focus on elements
- Ignore subscripts

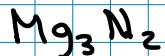
1. Name metal (cation) as is

2. "Root" of non-metal (anion) + "ide"

Root =

- Everything before the vowel in the second syllable
- Everything before "ine" / "ium" in element name

ex:

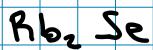


Metal = alkaline-earth  
↳ type I

Magnesium

nitride

nitrogen

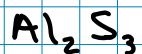


Metal = alkali  
↳ type I

Rubidium

selenide

selenium



Metal = Aluminium  
↳ type I

Aluminium

sulfide

sulfur

## Binary Ionic Compounds (Type 2)

→ The cation (metal) may form more than one numerical value of (+) charge

## In Periodic table :

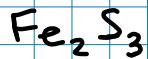
1. Small rectangle, left of staircase  
In, Sn, Sb, Ti, Pb, Bi
  2. All transition metals (other than Cd, Ag, Zn)
  3. Lanthanides, actinides

## Naming compound

- Because the metal can have different charges, we specify its charge
  - Other than that the naming is the same as type 1

1. Name metal (cation)
  2. in parenthesis, put roman numeral indicating the charge of the cation (+)
  3. "Root" non-metal (anion) + "ide"

ex:



Metal = transition

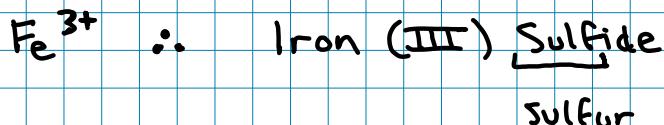
↳ type 2

## Finding charge of Fe

$$\left( \text{charge of Fe} \times \text{number of Fe} \right) + \left( \text{charge of S} \times \text{number of S} \right) = \text{zero}$$

$$(\square \times 2) + (2 \times 3) = \text{zero}$$

$$(3+ x 2) + (2- x 3) = \text{zero}$$



# Ionic compounds: polyatomic ions

## Polyatomic ions

<u>Cation</u>	$\text{Hg}_2^{2+}$	(mercury I)
	$\text{NH}_4^+$	(ammonium) ↳ not ammonia ( $\text{NH}_3$ )
	$\text{OH}^-$	(hydroxide)
	$\text{CN}^-$	(cyanide)
	$\text{NCS}^-$	(thiocyanate)
	$\text{CH}_3\text{COO}^-$	(acetate)
	$\text{MnO}_4^-$	(permanganate)
	$\text{CrO}_4^{2-}$	(chromate) chrome
	$\text{Cr}_2\text{O}_7^{2-}$	(dichromate)
	$\text{O}_2^{2-}$	(peroxide)
	$\text{C}_2\text{O}_4^{2-}$	(oxalate)
	$\text{S}_2\text{O}_3^{2-}$	(thiosulfate)

### I (charge 1-)

$\text{I}^3$

\* Iodate  $\text{IO}_3^-$

Periodate  $\text{IO}_4^-$

Iodite  $\text{IO}_2^-$

Hypiodite  $\text{IO}^-$

### Cl (charge 1-)

$\text{Cl}^3$

\* Chlorate  $\text{ClO}_3^-$

Chlorite  $\text{ClO}_2^-$

Hypo chlorite  $\text{ClO}^-$

Perchlorate  $\text{ClO}_4^-$

Find "ate" form

$\rightarrow +1$  oxygen : per \_\_\_\_\_

$\rightarrow -1$  oxygen : \_\_\_\_\_ ite

$\rightarrow -2$  oxygen : hypo\_\_\_\_\_ ite

\* charge remains

$\rightarrow +1$  hydrogen : hydrogen \_\_\_\_\_

↳ charge +1

\* Find charge in periodic t.

N (charge 3-  $\rightarrow 1^-$ )  
 $\text{tr}_1 = 3$

\* Nitrate  $\text{NO}_3^-$

Nitrite  $\text{NO}_2^-$

S (charge 2-)  
 $\text{f} = 4$

\* Sulfate  $\text{SO}_4^{2-}$

Sulfite  $\text{SO}_3^{2-}$

Hydrogen sulfate  $\text{HSO}_4^-$

P (charge 3-)

$\text{ph} = 4$

\* Phosphate  $\text{PO}_4^{3-}$

Hydrogen phosphate  $\text{HPO}_4^{2-}$

Dihydrogen phosphate  $\text{H}_2\text{PO}_4^-$

C (charge 4-  $\rightarrow 2-$ )

$\text{C}^3$

\* Carbonate  $\text{CO}_3^{2-}$

Hydrogen carbonate  $\text{HCO}_3^-$

Br (charge 1-)

$\text{Br}^3$

\* Bromate  $\text{BrO}_3^-$

Bromite  $\text{BrO}_2^-$

Hypobromite  $\text{BrO}^-$

## Naming compound with polyatomic ions

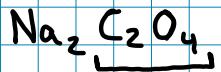
→ Compounds with polyatomic anions (-)

1. Name the metal (cation)

↳ If the metal is of type 2, put a roman numeral to indicate his charge

2. Name the polyatomic anion

ex: Formula to name

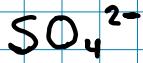
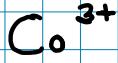


metal = type 1

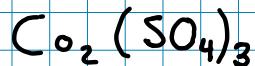
Sodium Oxalate

ex: Name to formula

Cobalt (III) Sulfate



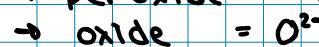
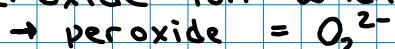
$$\begin{aligned} (\text{number of } \text{Co} \times 3^+) + (\text{number of } \text{SO}_4 \times 2^-) &= 0 \\ (2 \times 3^+) + (3 \times 2^-) &= 0 \end{aligned}$$



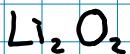
\* Note: parenthesis if more than 1

\* Never simplify ions (peroxide / mercury I always have an even number of O or Hg)

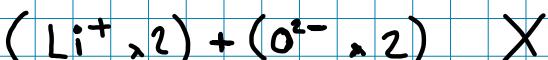
Peroxide ion when naming



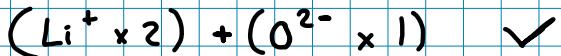
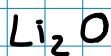
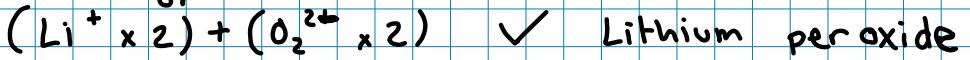
Think of neutrality (what do I need to cancel the charge)



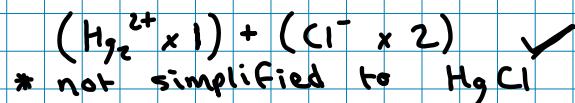
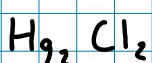
type I metal



or



Lithium oxide



Mercury (I) chloride

\* not simplified to  $\text{HgCl}$

# Metal Hydrate

→ ionic compound

→ solid

→ water in its crystal structure (waters of hydration)

Formula: (ionic compound) • (number) ( $\text{H}_2\text{O}$ )

- For every unit of the compound there are that many  $\text{H}_2\text{O}$  molecules.

Naming:

1. Name ionic compound

2. "Prefix" + "hydrate"

Prefixes (indicating number of  $\text{H}_2\text{O}$ ):

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

ex. name  $\text{Ni}(\text{NO}_3)_2 \cdot 2 \text{ H}_2\text{O}$



type 2 metal

$\text{NO}_3^-$  = nitrate

$$(1 \times \underset{\text{charge}}{\text{Ni}}^{\text{charge}}) + (2 \times 1^-) = 0$$

charge Ni = 2<sup>+</sup>

Nickel (II) nitrate dihydrate

## Binary Covalent compound (type 3)

→ Between 2 non-metals

→ Electrons are shared by a covalent bond

↳ No ions

→ Formula for charge neutrality doesn't apply here as there is no charge

Naming:

Element 1 + Element 2

("prefix" + "element 1") + ("prefix" + "root 2" + "ide")

Note:

→ For the first element, we omit prefix "mono"

→ For "oxide", we drop the "a" or "o" of the prefix

ex:  $\text{N}_2\text{O}_5$  dinitrogen pentoxide

ex. CO carbon monoxide

## Hydrogen Ion (For naming)

→ H can be cation (+) or anion (-)

• H can gain an electron

↳  $\text{H}^-$  (hydride) : anion

↳ Interact with metals (cation)

• H can lose an electron (for naming purposes)

↳  $\text{H}^+$  : cation

↳ Interact with non-metals (anion)

ex: LiH Lithium hydride

Type I

ex  $\text{H}_2\text{S}$  Hydrogen sulfide

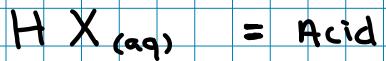
Hydrogen Compounds with special names:

$\text{H}_2\text{O}$  = water

$\text{NH}_3$  = ammonia

$\text{CH}_4$  = methane

# Naming Acids



$\rightarrow X = \text{anion}$

→ Trick: "aq" — queu that you have an acid

Monoprotic acid:  $\text{H}_3\text{X}_{(\text{aq})}$   
→ 1 acidic hydrogen

Polyprotic acid: more than 1 acidic hydrogen

- Diprotic acid:  $H_2X^{(aq)}$

$\Rightarrow 2$  acidic hydrogens

- Triprotic acid:  $H_3X^{(aq)}$   
 $\rightarrow$  3 acidic hydrogens

## Naming:

- Think of hydrogen having a + 1 charge : H<sup>+</sup>

Case 1: Anion X does not contain oxygen

1. ("hydro" + "root X" + "ic") + ("acid")

ex:  $\text{H}\text{I}_{(\text{aq})}$  = hydroiodic acid

ex:  $\text{HCN}_{(\text{aq})}$  = hydrocyanic acid

Case 2: Anion X contains oxygen (oxyacid)

1. If anion is "root" + "ATE"  
→ ("root" + "ic") + ("acid")

2. If anion is "root" + "ife"  
→ ("root" + "ous") + ("acid")

\* The root is "sulfur" and "phosphor"

Ex:  $\text{HClO}_4$

$\text{ClO}_4^-$  = perchlorate  
perchloric acid

ex:  $\text{H}_2\text{SO}_4$

$$\text{SO}_3^{2-} = \text{sulfite}$$

sulfurous acid

ex: H<sub>2</sub>S  
NO oxygen

hydro sulfuric acid

# Molar mass of an ionic compound / molecule

Molar mass = m.m. = molecular weight = M.W.

- Multiply the molar mass of each element by the number of atoms present of that element in the molecule / compound.
- Add all the products

ex. what is the molar mass of  $\text{NaH}_2\text{PO}_4$

Atom	Number of atoms	m.m of atom
Na	1	22.99 g/mol
H	2	1.008 g/mol
P	1	30.97 g/mol
O	4	16.00 g/mol

$$= 119.976 \rightarrow 119.98 \text{ g/mol}$$

(round to 2 decimals)

\*Note (sig figs): when multiplying by a "counted number" use addition rules (round to decimal numbers)

ex: how many carbon atoms in 1.00 g of  $\text{C}_{10}\text{H}_8$ ?

$$\text{m.m} = 128.2 \text{ g/mol}$$

$$\begin{aligned}
 & 1.00 \text{ g } \text{C}_{10}\text{H}_8 \cdot \frac{1 \text{ mol}}{128.2 \text{ g}} \cdot \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \cdot \frac{10 \text{ atoms C}}{1 \text{ molecule } \text{C}_{10}\text{H}_8} \\
 & = 4.6973 \times 10^{22} \text{ atoms} \rightarrow 4.70 \times 10^{22} \text{ atoms}
 \end{aligned}$$

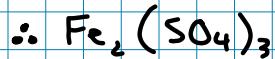
## Percent by mass

→ % by mass of an element in a compound

$$\rightarrow \% \text{ by mass} = \frac{\text{mass of element}}{\text{Total mass of compound}} \times 100$$

ex' Calculate the mass % of sulfur in Iron (III) sulfate

- Calculate as if there were 1 mole of compound  
↳ The mass % doesn't change with the number of atoms/molecules ~~in~~ in the compound.



$$\begin{aligned} 1 \text{ mol of } \text{Fe}_2(\text{SO}_4)_3 &= 2 \text{ mol Fe} \times \frac{55.85 \text{ g}}{\text{mol}} = 11.70 \text{ g Fe} \\ &\quad 3 \text{ mol S} \times \frac{32.07 \text{ g}}{\text{mol}} = 96.21 \text{ g S} \\ &\quad 12 \text{ mol O} \times \frac{16.00 \text{ g}}{\text{mol}} = 192.00 \text{ g O} \\ \hline &\quad 399.91 \text{ g compound} \end{aligned}$$

$$\begin{aligned} \text{mass \%} &= \frac{96.21 \text{ g S}}{399.91 \text{ g compound}} \times 100 \\ &= 24.06 \% \end{aligned}$$

\* Note: use addition laws :: multiply by counted number.

# Determining Empirical Formula

→ Want the relative number of elements

→ Mass doesn't tell you the relative number  
→ Moles tell the relative number

ex. calculate empirical formula of a compound that has 60.29 mg gold, 21.10 mg sodium, 29.37 mg O and 39.24 mg sulfur

$$\frac{g}{\text{mol}} = \frac{\text{mg}}{\text{mmol}}$$

\* Find number of moles of each element-

$$60.29 \text{ mg Au} \times \frac{\text{mmol Au}}{197.0 \text{ mg Au}} = \frac{0.3061 \text{ mmol Au}}{0.3061}$$

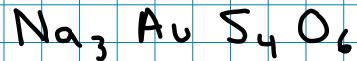
$$21.10 \text{ mg Na} \times \frac{\text{mmol Na}}{22.99 \text{ mg Na}} = \frac{0.9178 \text{ mmol Na}}{0.3061}$$

$$29.37 \text{ mg O} \times \frac{\text{mmol O}}{16.00 \text{ mg O}} = \frac{1.836 \text{ mmol O}}{0.3061}$$

$$39.24 \text{ mg S} \times \frac{\text{mmol S}}{32.07 \text{ mg S}} = \frac{1.224 \text{ mmol S}}{0.3061}$$

\* Divide each amount obtained by the smallest amount of moles to determine ratio

Ratio : 1 Au : 3 Na : 6 O : 4 S



Determine empirical formula from "mass %":

ex: Binary compound is Si and Cl (28.37% Si), determine empirical formula

\* Assume any mass (100 g) ∵ mass doesn't change formula

in 100 g.

$$28.37 \text{ g Si} \times \frac{\text{mol Si}}{28.09 \text{ g Si}} = \frac{1.010 \text{ mol Si}}{1.010} = 1$$

$$71.63 \text{ g Cl} \times \frac{\text{mol Cl}}{35.45 \text{ g Cl}} = \frac{2.021 \text{ mol Cl}}{1.010} = 2$$



## Combustion Analysis (Empirical Formula)

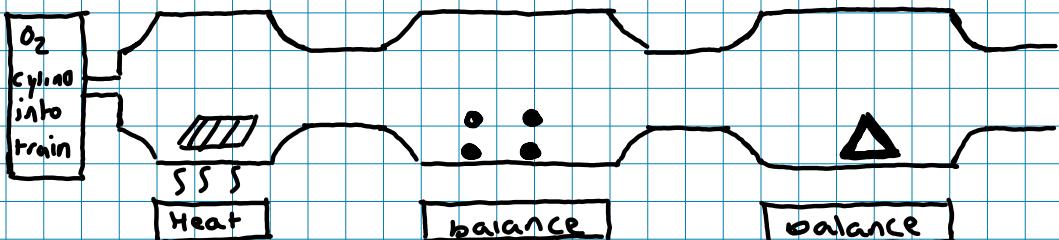
Hydrocarbon = compound containing only C and H

Combustion train:

= sample compound

= desiccant (absorbs  $H_2O$ )

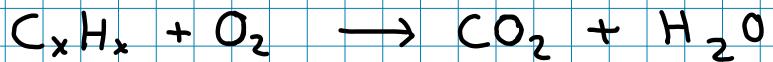
= ( $O_2$  absorbing compound)



- All the C in the initial compound reacts to become  $CO_2$
- All the H in the initial compound reacts to become  $H_2O$

The change in mass in the 2nd and 3rd bulb tell us how much  $CO_2$  and  $H_2O$  was absorbed  
→ for every  $CO_2$ , there were 1 C in sample  
→ For every  $H_2O$ , there were 2 H in sample  
\* If there was O in sample compound, the  $O_2$  and  $H_2O$  unit tell us anything since some of the O came from the cylinder

Combustion of Hydrocarbon



## Combustion problem, Hydrocarbon

- Find number of moles of C and of H
- Divide by smallest number to find integer ratio  
↳ must be less than 0.1 from integer to round to integer → if not, multiply

ex. Hydrocarbon burned to give 477.0 mg CO<sub>2</sub>, 111.60 mg H<sub>2</sub>O

$$\frac{g}{mol} = \frac{mg}{mmol}$$

$$477.0 \text{ mg CO}_2 \times \frac{\text{mmol CO}_2}{44.01 \text{ mg CO}_2} \times \frac{1 \text{ mmol C}}{1 \text{ mmol CO}_2} = \frac{10.84}{10.84} \text{ mmol C}$$

$$111.60 \text{ mg H}_2\text{O} \times \frac{\text{mmol H}_2\text{O}}{18.02 \text{ mg H}_2\text{O}} \times \frac{2 \text{ mmol H}}{1 \text{ mmol H}_2\text{O}} = \frac{12.39}{10.84} \text{ mmol H}$$

1 C : 1.143 H → Trick:  $\frac{1}{\text{decimal}} = \text{multiplier}$  (less than 20)

$$\frac{1}{1.143} = 6.99 \approx 7 \quad C_{(1,7)}H_{(1.143 \times 7)} = C_7H_8$$

\*must be less than 0.1 away from integer to round to integer

## Combustion problem, C + H + X (element missing)

ex: Molecule C, H, S. 1.086 g molecular produces  
2.272 g CO<sub>2</sub> and 0.465 g H<sub>2</sub>O

$$2.272 \text{ g CO}_2 \times \frac{\text{mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.05162 \text{ mol C}$$

$$0.465 \text{ g H}_2\text{O} \times \frac{\text{mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.0516 \text{ mol H}$$

Mass in compound: anything not H or C is S

$$0.05162 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.6200 \text{ g C}$$

$$0.0516 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.0520 \text{ g H}$$

$$\frac{1.086 \text{ g} - 0.6200 \text{ g} - 0.0520 \text{ g}}{= 0.414 \text{ g S}}$$

$$0.414 \text{ g S} \times \frac{\text{mol S}}{32.07 \text{ g S}} = 0.0129 \text{ mol S}$$

Divide all "mol" by the smallest (0.0129 mol)

4 C : 4 H : 1 S

∴ empirical formula: C<sub>4</sub>H<sub>4</sub>S

## Empirical Formula $\rightarrow$ Molecular Formula

$\rightarrow$  Determine empirical formula (% mass / combustion)

\* Molar mass of compound is given

$\rightarrow$  Find molar mass of empirical formula

$\rightarrow$  Multiply each subscript of empirical formula by integer:

$$\frac{\text{m.m. compound}}{\text{m.m. empirical formula}}$$

Ex:  $C_3H_5$  empirical formula  
molar mass of compound = 410.7 g/mol

$$\frac{410.7 \text{ g/mol}}{41.07 \text{ g/mol}} = 10 \quad \therefore C_{(3 \times 10)} H_{(5 \times 10)}$$
$$\therefore C_{30}H_{50}$$

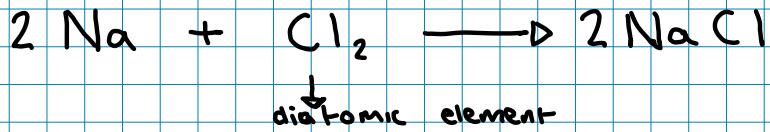
# Chemical Reaction (Stoichiometry)

- Matter is neither created nor destroyed
- What changes are the interactions between atoms
- The number of each element is the same in reactants and products

Reactants (reagents) → Products

- The chemical equation has to be balanced

ex: write the chemical equation for sodium reacting with chlorine to form sodium chloride

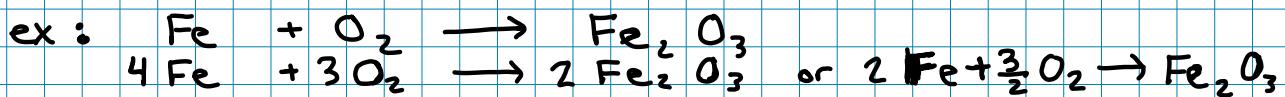


To balance, change only stoichiometric coefficient, not the subscript

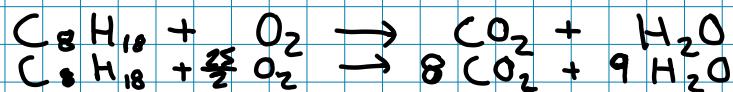
Coefficient = relative number of each reactant that reacts and relative number of products produced (assuming reaction completion)

- 2 molecules of Na react with 1 of Cl<sub>2</sub> to produce 2 molecules of NaCl
- 2 moles of Na react with 1 mole of Cl<sub>2</sub> to produce 2 moles of NaCl

\* This is not true for mass, since mass doesn't tell the relative number of each compound



ex: Combustion of C<sub>8</sub>H<sub>18</sub> (C<sub>x</sub>H<sub>x</sub> + O<sub>2</sub> → CO<sub>2</sub> + H<sub>2</sub>O)

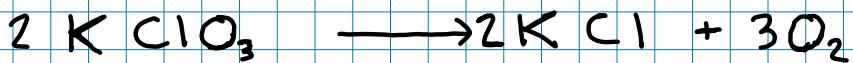


C	8	8
H	18	18
O	25	25

Predict maximum amount of product if a reaction goes to completion  
(for reactions with 1 reactant)

- If we are given masses, convert to moles using molar mass
- Stoichiometry (coefficient) tells us the relation between quantities of products and reactants (molecules, moles)

ex: 3.00g of potassium chlorate decomposes to potassium chloride and oxygen. What is the maximum mass of oxygen gas produced



- Assuming reaction goes to completion, all  $\text{KClO}_3$  is transformed to  $\text{KCl}$  and  $\text{O}_2$

Convert grams to mol:

$$3.00 \text{ g KClO}_3 \times \frac{\text{mol KClO}_3}{122.55 \text{ g KClO}_3} = 0.02448 \text{ mol KClO}_3$$

How many moles of  $\text{O}_2$  produced:

$$0.02448 \text{ mol KClO}_3 \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} = 0.03672 \text{ mol O}_2$$

$$\begin{aligned} 0.03672 \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{\text{mol O}_2} &= 1.175 \text{ g O}_2 \\ &= 1.18 \text{ g O}_2 \end{aligned}$$

A chemical reaction can go 2 ways:

→ Forward directions = reactants (left of arrow)  
products (right of arrow)

→ Reverse direction = products (left of arrow)  
reactants (right of arrow)

# Limiting Reagent

- A reaction with 2 reactants has a limiting reagent  
→ one of the reactants has leftover while the other one completely reacted  
→ The one that completely reacted limits the reaction from continuing (limiting reagent)

Limiting reagent : completely transformed to product

Excess reagent : leftover after reaction completion

If there is a perfect amount of each and neither remains in excess, the reactants react "stoichiometrically"



You make 4 moles of ammonia react with 4 of O<sub>2</sub>  
→ 1 mol of O<sub>2</sub> leftover  
→ Ammonia (NH<sub>3</sub>) is the limiting reagent

To determine which is limiting reagent :

1. Find out how much one reactant would react with how much of the other (use each of the number of moles that are used to react)
  - Whichever reacts with less than the number of moles of the other is limiting
2. Find out how many moles of product each reagent produces if it reacted completely
  - Whichever produces less is limiting

Theoretical yield (rendement) :

→ The maximum amount of a product produced (assuming all the limiting reagent is converted)

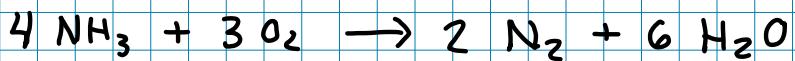
Find out how much excess there is :

- Find out how much of non-limiting reagent reacted
- Subtract the amount of non-limiting from starting (Starting amount) - (reacted amount) = (excess)

% yield :

- sometimes reactions don't go to theoretical completion
- we obtain the "actual yield" (product obtained experimental)

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



ex: 34.0 g ammonia react with 50.0 g O<sub>2</sub>

Limiting reagent:

- How many moles of each

$$34.0 \text{ g NH}_3 \times \frac{\text{mol NH}_3}{17.04 \text{ g NH}_3} = 1.995 \text{ mol NH}_3 = 2.00 \text{ mol NH}_3$$

$$50.0 \text{ g O}_2 \times \frac{\text{mol O}_2}{32.00 \text{ g O}_2} = 1.5625 \text{ mol O}_2 = 1.56 \text{ mol O}_2$$

1.  $\begin{array}{ccc} \text{NH}_3 & & \text{O}_2 \\ \cancel{4 \text{ mol}} & & \cancel{3 \text{ mol}} \\ *2.00 & \xrightarrow{\quad} & 1.50 \\ 2.08 & \leftarrow & *1.56 \end{array}$

NH<sub>3</sub> reacts with less  
∴ NH<sub>3</sub> is limiting

2.  $1.995 \text{ mol NH}_3 \times \frac{6 \text{ mol H}_2\text{O}}{4 \text{ mol NH}_3} = 2.99 \text{ mol H}_2\text{O}$   
 $1.56 \text{ mol O}_2 \times \frac{6 \text{ mol H}_2\text{O}}{3 \text{ mol O}_2} = 3.13 \text{ mol H}_2\text{O}$

NH<sub>3</sub> produces less product  
∴ NH<sub>3</sub> is limiting

Theoretical yield:

- How many moles of product did limiting reagent produce

$$2.99 \text{ mol H}_2\text{O} \times \frac{18.06 \text{ g H}_2\text{O}}{\text{mol H}_2\text{O}} = 54.0 \text{ g H}_2\text{O}$$

Excess of non-limiting reagent:

$$34.0 \text{ g NH}_3 \times \frac{\text{mol NH}_3}{17.04 \text{ g NH}_3} \times \frac{3 \text{ mol O}_2}{4 \text{ mol NH}_3} \times \frac{32.00 \text{ g O}_2}{\text{mol O}_2} = 47.9 \text{ g O}_2 \quad \cancel{\text{X X X X}}$$

$$50.0 \text{ g O}_2 - 47.9 \text{ g O}_2 = 2.1 \text{ g O}_2 \quad \text{(excess)}$$

$$\text{or } 1.56 \text{ mol O}_2 - 1.50 \text{ mol O}_2 = 0.06 \text{ mol O}_2 \quad \text{(excess)}$$

$$0.06 \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{\text{mol O}_2} = 1.92 \text{ g} = 2 \text{ g O}_2$$

% Yield:

During an experiment, we collect 50.0 g water (actual yield)

→ Theoretical yield = 54.0 g

$$\% \text{ yield} = \frac{50.0}{54.0} \times 100 = 92.6 \%$$

# Solutions

**Solubility:** amount of substance that dissolves in a given solvent

**Solution:** homogeneous mixture

**Solvent:** dissolving medium

**Solute:** substance dissolved

## Concentration:

$$\text{* Molarity (M)} = \frac{\text{moles of solute}}{\text{L of solution (not solvent)}}$$

$$\% \text{ mass for a solution} = \frac{\text{mass of solute}}{(\text{mass solute}) + (\text{mass solvent})} \times 100$$

↳ must be same units

**Concentrated X solution:** a solution that has the maximum amount of a solute dissolved in it

ex: conc. HCl solution is 12.1 M HCl. What is the mass % HCl? (density HCl solution = 1.19 g/cm³)

$$\left( \frac{12.1 \text{ mol HCl}}{\text{L solution}} \times \frac{36.46 \text{ g HCl}}{\text{mol HCl}} \times \frac{\text{cm}^3}{1.19 \text{ g solution}} \times \frac{\text{mL}}{\text{cm}^3} \times \frac{1}{1000 \text{ mL}} \right) \times 100$$

= 37.1 %

ex: How many g of solid potassium chloride is required to prepare 100.00 mL of 0.250 M solution?

$$100.00 \text{ mL solution} \times \frac{\text{L solution}}{1000 \text{ mL}} \times \frac{0.250 \text{ mol KCl}}{\text{L solution}} \times \frac{74.55 \text{ g KCl}}{\text{mol KCl}} =$$

= 1.86 g KCl

## Dilution

- Add solvent to stock solution (more conc.)
- Get a more dilute solution (less conc.)

\* If one only adds solvent, the moles of solute stay the same

$$C_1 V_1 = C_2 V_2$$

concentration  $\times$  volume = concentration  $\times$  volume  
(stock soln.) (dilute soln)

$$C \times V = \frac{\text{mol solute}}{\text{L solution}}, \text{ L solution} = \text{mol solute} \text{ (remains the same)}$$

ex: What initial volume of a 0.200 M NaOH solution be used to prepare a 0.0100 M solution with final volume 100.00 mL? And how much solvent added?

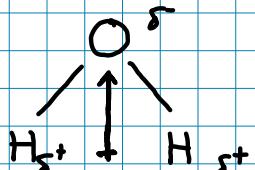
$$C_1 \times V_1 = C_2 \cdot V_2$$
$$(0.200 \text{ M}) \times V_1 = (0.0100 \text{ M})(100.00 \text{ mL})$$
$$\frac{V_1}{V_1} = 5.00 \text{ mL}$$

∴ add 95.00 mL

## Water Solvent

- water is a very good solvent because it's a polar molecule
  - ↳ electrons are polarized (pulled to one part of molecule)

\* Oxygen is one of the best atoms at pulling electrons towards itself



$\delta^-$  = partial negative charge  
 $\delta^+$  = partial positive charge

Dipole moment arrow:



Ionic compound dissolved in water :

→ ionize : dissociate into cation and anion

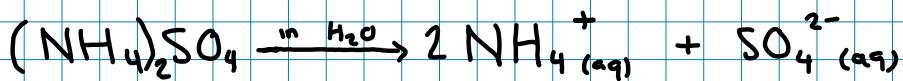
Covalent compound dissolved in water :

→ stay intact

ex: potassium nitrate



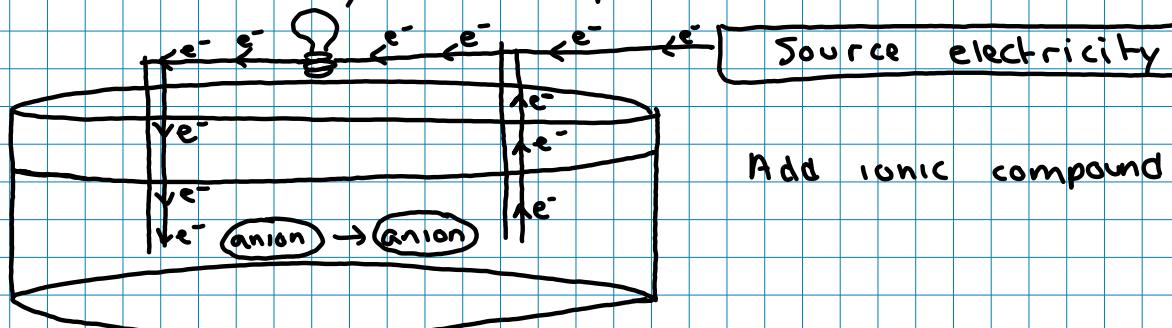
ex ammonium sulfate



Water solvent :

$\delta^-$  (O) surrounds cation (+)  
 $\delta^+$  (H) surrounds anion (-)

Proof that ionic compounds split up into anion / cations in water : Lightbulb experiment



The lightbulb lights up, meaning something (-) completed the circuit (anions)

Electrolyte : Solution or solute that permits the conduction of electricity

→ weak electrolyte : slightly conducts (dim light)

→ strong electrolyte : strong conduction (bright light)

Non-electrolyte : solution or solute that does not conduct electricity

→ non-ionic

→ ex: sugar solution

## 2 kinds of strong electrolytes (to learn) \* Split

### ① Strong Acids

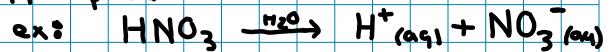
- HCl (aq)
- H<sub>2</sub>SO<sub>4</sub> (aq)
- HBr (aq)
- HI (aq)
- HClO<sub>4</sub> (aq)
- HNO<sub>3</sub> (aq)
- \* HClO<sub>3</sub> ⇒ borderline strong

### ② Soluble ionic compounds

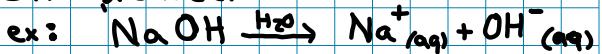
## Acids and Bases (reactants)

### Arrhenius definition

Acid = H<sup>+</sup> producer



Base = OH<sup>-</sup> producer

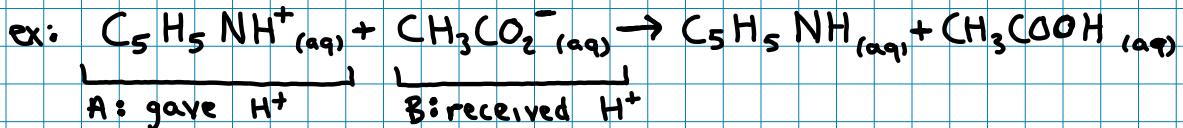


### Bronsted-Lowry

Acid = H<sup>+</sup> donor

Base = H<sup>+</sup> acceptor

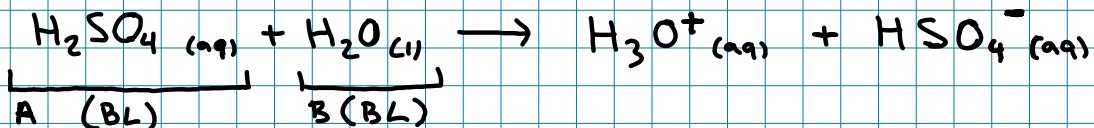
\* B-L Acid always coupled with B-L base



Strong Acid = acid HX (aq) that splits completely into H<sup>+</sup> (aq) and X<sup>-</sup> (aq)

Hydronium (H<sub>3</sub>O<sup>+</sup>) → H<sup>+</sup> (aq) is in water but doesn't exist as H<sup>+</sup> in water. It exists as H<sub>3</sub>O<sup>+</sup>. (H<sup>+</sup> short way of writing)

Dissociation of a strong acid:



\* Give 1 H<sup>+</sup> at a time

\* HSO<sub>4</sub><sup>-</sup> is not one of the strong acids so we don't write that it dissociates into SO<sub>4</sub><sup>2-</sup> and H<sup>+</sup>

## Weak electrolytes

\* Don't split

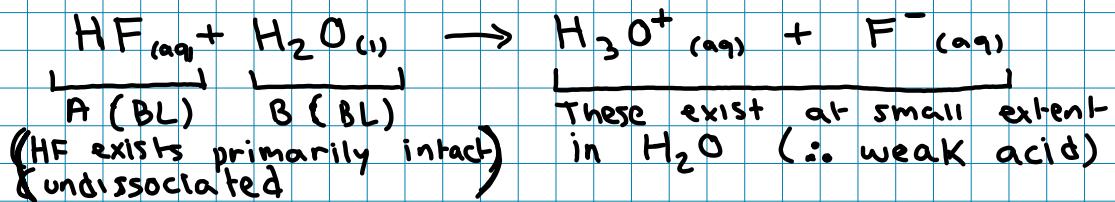
① weak acids

② weak base

③ slightly soluble ionic compounds

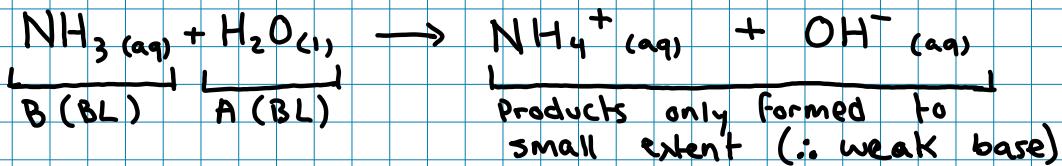
weak acid =  $\text{HX}_{(\text{aq})}$  that splits only to a small extent into  $\text{H}^+$  and  $\text{X}^-$

→ All acids other than the 6 are weak (unless told otherwise)



weak base =  $\text{B}_{(\text{aq})}$  (base) that splits into  $\text{BH}^+_{(\text{aq})}$  and  $\text{OH}^-_{(\text{aq})}$ , to a small extent.

→  $\text{NH}_3_{(\text{aq})}$  (ammonia) is the only substance to recognize as weak base.



\* Strong ionic compounds (soluble) dissolve and split into cations and anions

\* Weak ionic compounds (slightly soluble) exist overwhelmingly in solid form in water and split minimally into cations and anions.

## Solubility Rules

### Soluble ionic compounds (strong)

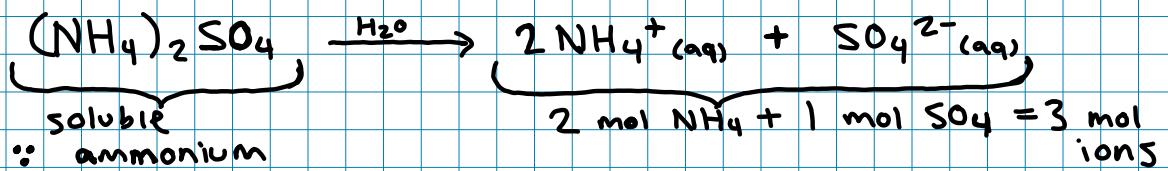
- Ionic compounds with nitrate  $\Rightarrow \text{NO}_3^-$
  - Ionic compounds with alkali metals or ammonium  $\Rightarrow \text{X}^+$ ,  $\text{NH}_4^+$
  - Ionic compounds with chloride, bromide or iodide  $\Rightarrow \text{Cl}^-, \text{Br}^-, \text{I}^-$
- \* There H A P ens to be 3 exceptions  
 $\text{Hg}_2\text{X}_2$ ,  $\text{AgX}$ ,  $\text{PbX}_2$
- Ionic compounds with sulfate  $\Rightarrow \text{SO}_4^{2-}$
- \* There H A P ens to be 5. why? B 'Cause  
 $\text{Hg}_2\text{X}$ ,  $\text{Ag}_2\text{X}$ ,  $\text{PbX}$ ,  $\text{BaX}$ ,  $\text{CaX}$
- Ionic compounds with perchlorate and chlorate  $\Rightarrow \text{ClO}_4^-$ ,  $\text{ClO}_3^-$

\* Note similarities with strong acids

### Slightly soluble ionic compounds (weak)

- Ionic compounds with most other anions such as hydroxide, sulfide, carbonate, chromate phosphate
- \* Exception: if these anions are present with an alkali metal or ammonium

ex: calculate the concentration of all ions of the following ionic compound  $\Rightarrow$  2.00 mol of ammonium sulfate in 1.00 L of solution

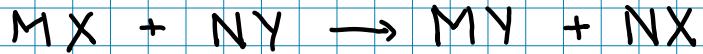


$$\frac{2.00 \text{ mol } (\text{NH}_4)_2\text{SO}_4}{1.00 \text{ L solution}} \times \frac{3 \text{ mol ions}}{\text{mol } (\text{NH}_4)_2\text{SO}_4} = \frac{6.00 \text{ mol ions}}{\text{L solution}}$$

## Precipitation (reaction type 1)

- 2 solutions, containing soluble ionic compounds are mixed
- After reaction, a solid is formed (precipitated or crashed out)

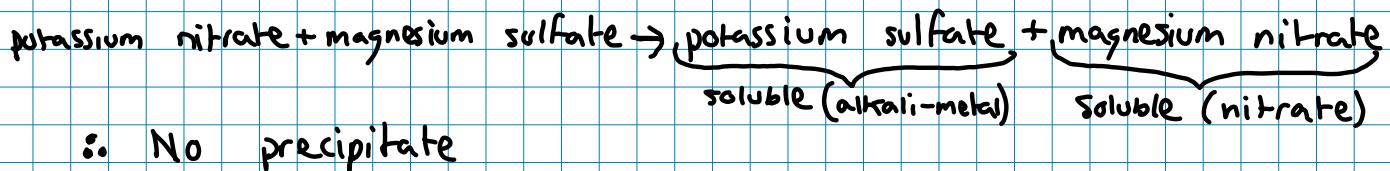
Precipitation = exchange reaction (double displacement or metathesis)



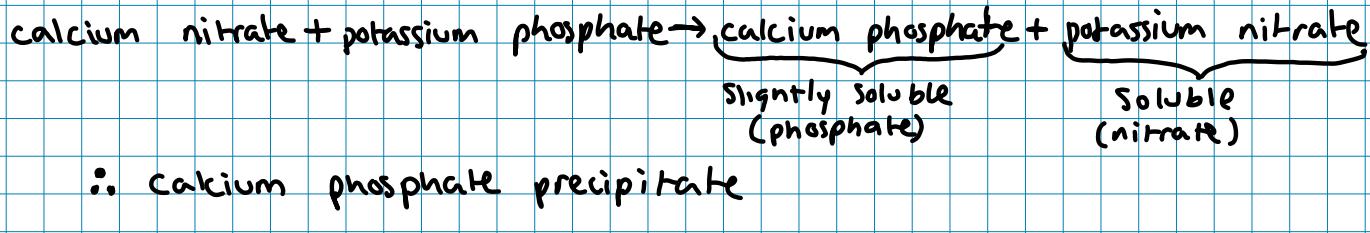
- cation 1 with anion 2
- cation 2 with anion 1

\* If one of the products contains a slightly soluble ionic compound ∴ that product precipitates

ex: mix aqueous solutions: potassium nitrate and magnesium sulfate



ex: mix aqueous solutions: calcium nitrate  
potassium phosphate



\* See on next page different ways of writing the equation

### 3 ways of writing equations of reactions in aqueous soln

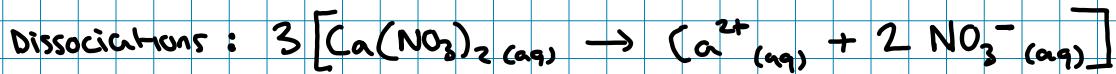
#### ① Overall balanced equation

- Write compounds (soluble or not) undissociated
- Molecular formula
- Coefficient to balance equation
- Soluble ionic compound (aq)
- Slightly soluble ionic compounds (s)

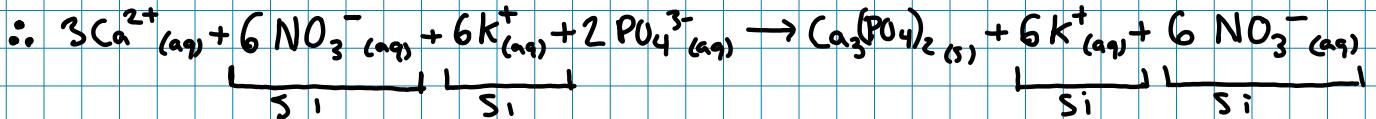
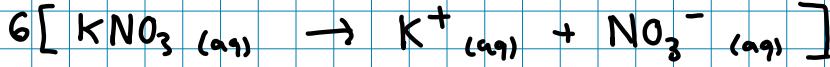


#### ② Complete ionic equation

- Write ions present
- Strong acids ( $\text{HX}$ ):  $\text{H}^+$ ,  $\text{X}^-$
- Soluble ionic compound: cation<sup>+</sup>, anion<sup>-</sup>
- Weak electrolytes (precipitates): intact form (also gas/solid)



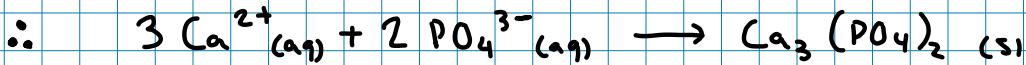
$\text{Ca}_3(\text{PO}_4)_2(s) \rightarrow$  Intact



Si = spectator ions  $\Rightarrow$  nothing happened to these ions

#### ③ Net ionic equation

- Remove spectator ions from the equation
- ↳ spectator ions: ions present on both sides



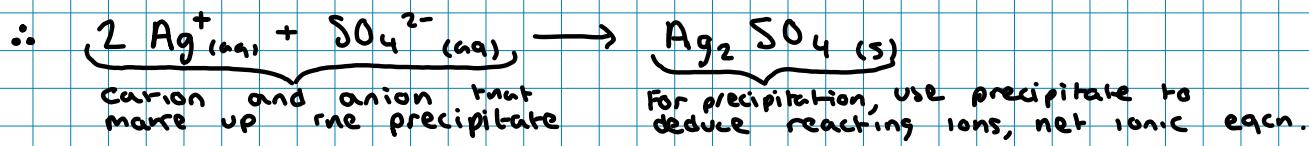
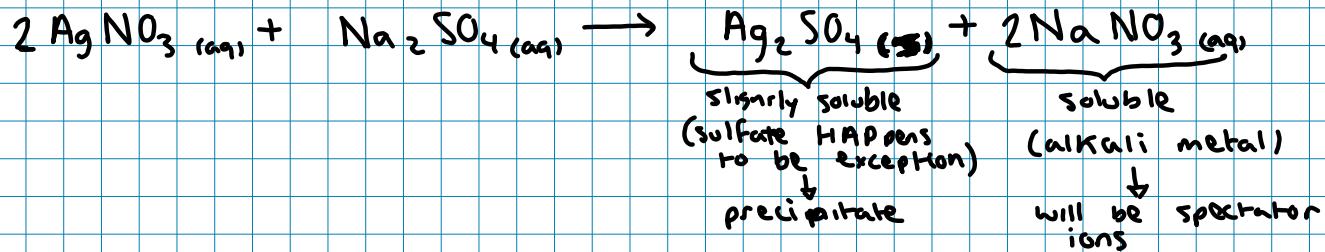
All that actually happened is  $\text{Ca}^+$  and  $\text{PO}_4^{3-}$  reacted.  
The other ions didn't do anything.

\* Note: spectator ion moles are the same before and after the reaction  $\therefore$  nothing happens to the spectator ions.

## Cumulative Problem

ex. Will precipitate form when 0.200 M  $\text{AgNO}_3$  (100.00 mL) is mixed with 0.200 M  $\text{Na}_2\text{SO}_4$  (100.00 mL)? What is the theoretical yield of the precipitate?

\* Find precipitate and ions on both sides of the reaction



$$\therefore \text{Spectator ions} = 2 \text{NO}_3^- \text{(aq)} \text{ and } 2 \text{Na}^+ \text{(aq)}$$

\* Find limiting reagent to find out how much precipitate is formed (limiting ion reagent)

$$100.00 \text{ mL} \times \frac{0.200 \text{ mol AgNO}_3}{\text{L soln}} \cdot \frac{1 \text{ mol}}{1000 \text{ mL}} \cdot \frac{1 \text{ mol Ag}^+}{\text{mol AgNO}_3} = 0.0200 \text{ mol Ag}^+$$

$$100.00 \text{ mL} \times \frac{0.200 \text{ mol Na}_2\text{SO}_4}{\text{L soln}} \cdot \frac{1 \text{ mol}}{1000 \text{ mL}} \cdot \frac{1 \text{ mol SO}_4^{2-}}{\text{mol Na}_2\text{SO}_4} = 0.0200 \text{ mol SO}_4^{2-}$$

How much product?

$$\left. \begin{aligned} 0.0200 \text{ mol Ag}^+ &\cdot \frac{1 \text{ mol Ag}_2\text{SO}_4}{2 \text{ mol Ag}^+} = 0.0100 \text{ mol Ag}_2\text{SO}_4 \\ 0.0200 \text{ mol SO}_4^{2-} &\cdot \frac{1 \text{ mol Ag}_2\text{SO}_4}{1 \text{ mol SO}_4^{2-}} = 0.0200 \text{ mol Ag}_2\text{SO}_4 \end{aligned} \right\} \therefore \text{Ag}^+ \text{ is limiting reagent}$$

Theoretical yield

$$0.0100 \text{ mol Ag}_2\text{SO}_4 \cdot \frac{311.87 \text{ g Ag}_2\text{SO}_4}{\text{mol Ag}_2\text{SO}_4} = 3.12 \text{ g Ag}_2\text{SO}_4$$

What is the concentration of each ion remaining in solution once precipitation is complete?

\* Which ions are still present after reaction:

- ① Spectator ions (didn't participate)
- ② Non-limiting reagent (leftover)

\* Volume of final solution = volume of both initial solutions

~~Initial amount~~ ~~each ion~~

~~0.0200 mol~~

① Concentration Spectator Ions (moles are the same. all that happens is they get diluted)

$$C_1 \cdot V_1 = C_2 \cdot V_2$$

$$NO_3^- : \left( \frac{0.200 \text{ mol } AgNO_3}{L} \cdot \frac{1 \text{ mol } NO_3^-}{\text{mol } AgNO_3} \right) (100.00 \text{ mL}) = C_2 \cdot (200.00 \text{ mL})$$

$$\boxed{C_2 = 0.100 \text{ M } NO_3^-}$$

$$Na^+ : \left( \frac{0.200 \text{ mol } Na_2SO_4}{L} \cdot \frac{2 \text{ mol } Na^+}{\text{mol } Na_2SO_4} \right) (100.00 \text{ mL}) = C_2 \cdot (200.00 \text{ mL})$$

$$\boxed{C_2 = 0.200 \text{ M } Na^+}$$

② Non spectator, non limiting reagent

$$SO_4^{2-} : \begin{aligned} \text{Initial amount} &= 0.0200 \text{ mol } SO_4^{2-} \\ \text{Reacted amount} &= \underbrace{0.0100 \text{ mol } Ag_2SO_4}_{\text{moles formed}} \cdot \frac{1 \text{ mol } SO_4^{2-}}{\text{mol } Ag_2SO_4} \\ &= 0.0100 \text{ mol } SO_4^{2-} \end{aligned}$$

$$\text{Initial} - \text{Reacted} = \text{Final} = 0.0100 \text{ mol } SO_4^{2-} \text{ excess}$$

$$\frac{0.0100 \text{ mol } SO_4^{2-}}{200.0 \text{ mL}} \cdot \frac{1000 \text{ mL}}{L} = \boxed{0.0500 \text{ M } SO_4^{2-}}$$

\* There is no more  $Ag^+$  remaining

$\therefore Ag^+$  is limiting reagent  $\therefore$  all of it reacts

$\therefore Ag^+$  isn't spectator ion  $\therefore$  everything that reacts becomes precipitated compound

ex: 0.887 g sample of ~~XXXXXX~~ mixture (not only NaCl, other substance) of NaCl and KCl is dissolved in water. The solution is then treated with an excess of  $\text{Ag}_2\text{NO}_3$  to give a theoretical yield of 1.913 g of AgCl. Calculate % mass of each compound in mixture

- \*  $\text{AgNO}_3$  is not limiting reagent (excess)
- \* Both NaCl and KCl will react with  $\text{AgNO}_3$
- \* NaCl and KCl don't react (exchange anion  $\rightarrow$  get same thing)
- \* Can't write:  $\text{NaCl} + \text{KCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{NaNO}_3 + \text{KNO}_3$   
 ↳ can't assume that every one mol of NaCl that reacts, one mol KCl reacts  
 ↳ both reactions are separate



- \* Mixture  $\rightarrow$  exactly 2 different compounds react with the same reagent

All NaCl and KCl transformed to product (limiting)  
 Theoretical yield AgCl = 1.913 g

$$A = X \text{ g NaCl} \cdot \frac{\text{mol NaCl}}{58.44 \text{ g NaCl}} \cdot \frac{1 \text{ mol AgCl}}{1 \text{ mol NaCl}} \cdot \frac{143.45 \text{ g AgCl}}{\text{mol AgCl}} = \text{g AgCl from NaCl} \quad \Rightarrow A$$

$$B = Y \text{ g KCl} \cdot \frac{\text{mol KCl}}{74.55 \text{ g KCl}} \cdot \frac{1 \text{ mol AgCl}}{1 \text{ mol KCl}} \cdot \frac{143.45 \text{ g AgCl}}{\text{mol AgCl}} = \text{g AgCl from KCl} \quad \Rightarrow B$$

$$A + B = 1.913 \text{ g}$$

$$\boxed{① X \left( \frac{143.45}{58.44} \right) \text{ g } + Y \left( \frac{143.45}{74.55} \right) \text{ g } = 1.913 \text{ g AgCl}}$$

$$\boxed{② X \text{ g NaCl} + Y \text{ g KCl} = 0.887 \text{ g}}$$

$$\text{Solve: } X = 0.388 \text{ g NaCl}, Y = 0.499 \text{ g KCl}$$

Divide by 0.887 g (mass sample), multiply by 100

$$\therefore 43.7\% \text{ NaCl by mass and } 56.3\% \text{ KCl by mass}$$

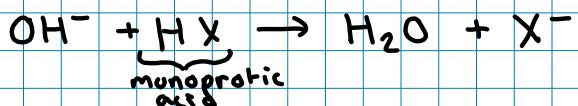
# Acid/Base reaction (reaction type 2)

\* Hydroxide ( $\text{OH}^-$ ):

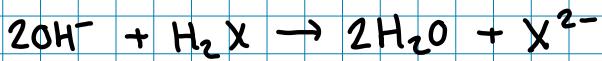
→  $\text{OH}^-$  = strongest base

→ Reacts completely with all acids (weak or strong)

→  $\text{OH}^-$  will deprotonate (remove  $\text{H}^+$ ) completely all acids

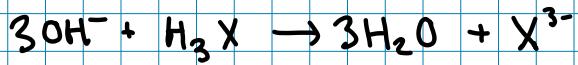


(Remove 1  $\text{H}^+$  from monoprotic)



(Remove 2  $\text{H}^+$  from diprotic acid)

Just enough  $\text{OH}^-$  to remove the 2  $\text{H}^+$  → second equivalence



(Remove 3  $\text{H}^+$  from triprotic)

Just enough  $\text{OH}^-$  to remove the 3  $\text{H}^+$  → third equivalence

\* Hydronium ( $\text{H}_3\text{O}^+$ ):

→  $\text{H}_3\text{O}^+$  = strong acid

→ React completely with all bases (weak or strong)

→  $\text{H}_3\text{O}^+$  will protonate (give  $\text{H}^+$ ) all bases completely



## Titration

- one reactant added to another little by little (mixing)
- exactly enough for stoichiometric reaction (no limiting reagent)

Analyte: reactant 1 placed in Erlenmeyer flask

Titrant: reactant 2 placed in buret being added to the analyte

Equivalence point (stoichiometric) = when stoichiometric reaction  
↳ Neutralization point for acid/base reactions

Indicator: compound change colour when equivalence point

End point: volume of titrant added to react stoichiometrically with the analyte

## Acid / base titrations

Indicator = phenolphthalein

↳ equivalence point when strong acid = analyte  
strong base = titrant

Color of phenolphthalein in analyte:

→ excess strong acid = colorless

→ excess strong base = hot pink (fuchsia)

\* At equivalence point = lightest possible permanent shade of pink  
(swirling soln; pink might be on and off)

There exist other indicators

Writing formula equation for acid/base reaction  
 → Think of it like exchange reaction

Complete ionic equations

→ Strong acid and base split into anion and cation

- Strong acids = HCl, HNO<sub>3</sub>, HI, HBr, H<sub>2</sub>SO<sub>4</sub>, HClO<sub>4</sub>

- Strong bases = alkali metal + hydroxide

alkaline-earth metal + hydroxide (moderately soluble - up until a certain molarity)

→ Weak acid and base written intact

- Only dissociate to a small extent

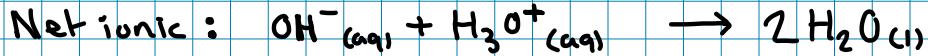
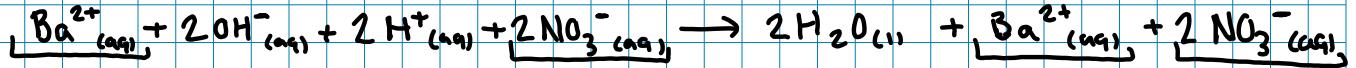
Net ionic equation (most important)

- Shows who H<sup>+</sup> donor and who H<sup>+</sup> receiver are

→ alkaline-earth and alkali metals, and cation of strong acid = spectator ions

\* For all reactions between strong acids and strong bases,  
 Net ionic equation = H<sub>3</sub>O<sup>+</sup><sub>(aq)</sub> + OH<sup>-</sup><sub>(aq)</sub> → 2 H<sub>2</sub>O<sub>(l)</sub>

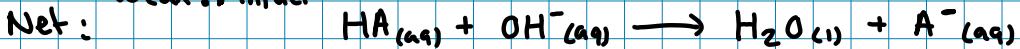
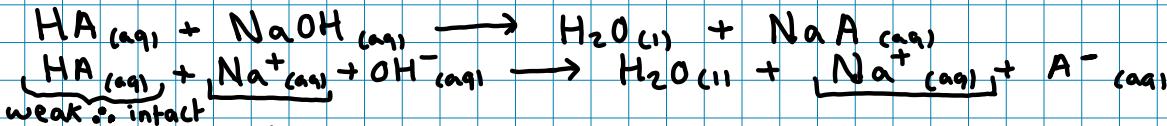
ex: (only strong A/B): 0.0100M soln barium hydroxide (20.15mL) is added by titration to soln nitric acid (5.00mL) to reach equivalence point. What is concentration of nitric acid?



$$0.02015 \text{ L} \times \frac{0.0100 \text{ mol Ba(OH)}_2}{\text{L soln}} \times \frac{2 \text{ mol HNO}_3}{1 \text{ mol Ba(OH)}_2} = 4.030 \times 10^{-4} \text{ mol HNO}_3 \text{ at equivalence}$$

$$M = \frac{4.030 \times 10^{-4} \text{ mol}}{0.00500 \text{ L}} = 0.0806 \text{ M HNO}_3$$

ex: (with weak A/B): Finding molar mass from equivalence  
 2.20g of monoprotic weak acid in water. It takes 25.0mL of 0.500M NaOH to reach equivalence.



$$25.0 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.500 \text{ M NaOH}}{\text{L}} \times \frac{1 \text{ mol OH}^-}{1 \text{ mol NaOH}} \times \frac{1 \text{ mol HA}}{1 \text{ mol OH}^-} = 0.0125 \text{ mol weak acid}$$

$$\frac{2.20 \text{ g HA}}{0.0125 \text{ mol HA}} = 176 \text{ g/mol} \quad \therefore \text{titration gives us molar mass}$$

\*Reminder: empirical formula + molar mass → molecular formula

# Redox reaction (reaction type 3)

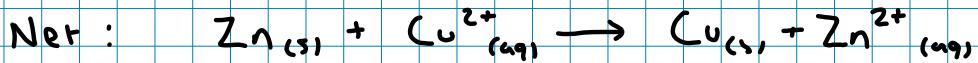
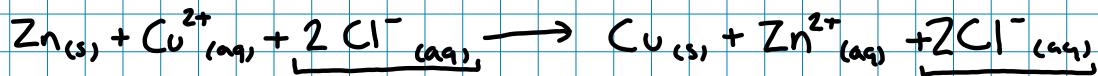
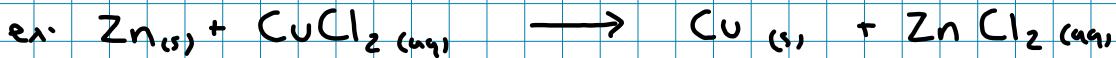
Oxydation  
Involves  
Loss (of electrons)

Reduction  
Involves  
Gain (of electrons)

- \* If there's an oxidation (loss) there must be a reduction (gain)

Oxidized species : reagent that lost electrons  
↳ Reducing Agent (cause reduction)

Reduced species : reagent that gained electrons  
↳ Oxidizing Agent (cause oxidation)



Species oxidized:  $Zn(s)$  (lost electrons)

Species reduced:  $Cu^{2+}(aq)$  (gained electrons)

Oxidizing agent:  $Cu^{2+}(aq)$  (caused the loss - oxidation)

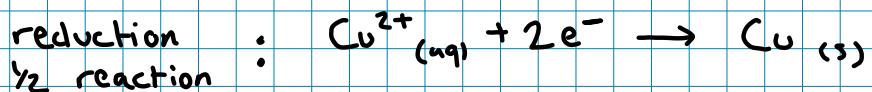
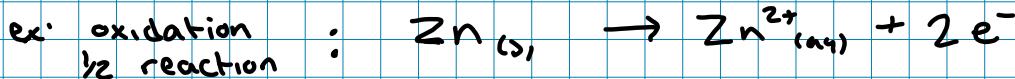
Reducing agent:  $Zn(s)$  (caused the gain - reduction)

- \* Always write reactants (what is on the left that has been modified - oxidized or reduced)

Redox reaction = reaction with an oxidation and a reduction

Half reactions: we can write a half reaction to show just what has been oxidized and another to show just what has been reduced

- \* Make reactants and products have the same charge

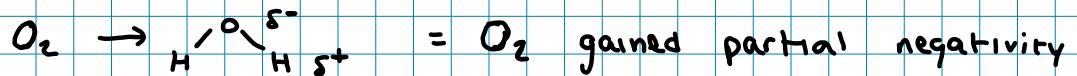
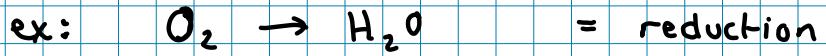


- \* Electrons in product = oxidation (loss)

Electrons in reagent = reduction (gain)

## Oxidation number (state)

- In a compound, some atoms can be partially negative and others partially positive
- If one gains partial negativity, there is a reduction, even if the charge as a whole of the compound stays the same



Oxidation number = number to keep track of who is partially negative or positive and determine which species has been oxidized or reduced

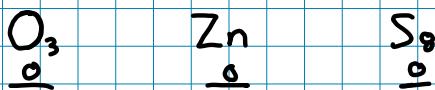
\* All elements are assigned oxidation number

\* Oxidation number =  $\pm$  "number"

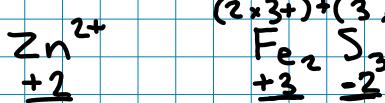
+ charge = "number"  $\pm$

Rules  $\sum (\text{oxidation num} \times \text{subscript}) = \text{charge}$

1. An atom of a neutral element on its own or a molecule with only one type of element  
 $\Rightarrow$  oxidation num. = 0



2. Monoatomic ions (in ionic compound or on their own)  
 $\Rightarrow$  oxidation num. = charge



3. F  $\Rightarrow$  oxidation num. = -1 (in compound)

\* exception: F and  $F_2$  alone  $\Rightarrow$  oxidation num. = 0

4. H  $\Rightarrow$  oxidation num = +1 (in compound)

\* exception: in  $H_2$   $\Rightarrow$  oxidation num. = 0

\* H in a compound with metal (hydride)  $\Rightarrow$  oxidation num = -1

5. O  $\Rightarrow$  oxidation num. = -2

\* exception: in  $O_2$  or  $O_3$   $\Rightarrow$  oxidation num = 0

\* in molecule that contains only O and F (fluorine is stronger)

\* in peroxide ion ( $O_2^{2-}$ ) or  $H_2O_2 \Rightarrow$  oxidation num. = -1

6. Cl  $\Rightarrow$  oxidation num. = -1 (except with F or O)

Br  $\Rightarrow$  oxidation num. = -1 (except with F, O, Cl)

I  $\Rightarrow$  oxidation num. = -1 (except with F, O, Cl, Br)

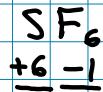
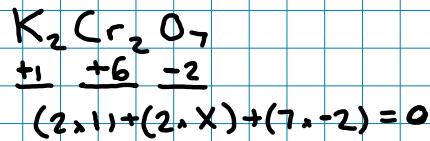
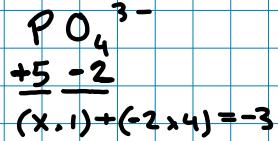
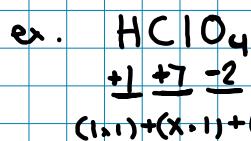
$(Cl_2 \Rightarrow 0)$

$(Br_2 \Rightarrow 0)$

$(I_2 \Rightarrow 0)$

## Oxidation number problems:

\* When writing oxidation num for an element, do not multiply by the subscript

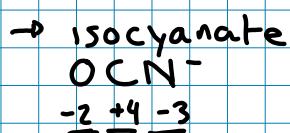
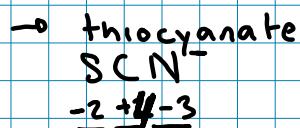
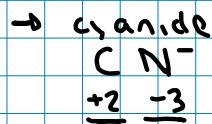


\* The oxidation num. of F is -1 (not -6), since we do not multiply by subscript

### Tricks:

- Group IVA element oxidation num. [-4, 4]
- Group VA element oxidation num. [-3, 5]
- Group VIA element oxidation num. [-2, 6]
- Group VIIA element oxidation num. [-1, 7]

### 3 ions that don't follow rules:



If the oxidation num. of an element increases while going from reactant to product, the reactant has been oxidized (lose electron ∵ partial charge increases)

If the oxidation num. of an element decreases, the reactant has been reduced (gain electron ∵ partial charge decreases)

## Balancing Redox Reaction

- Ensure overall same charge
- Use half-reaction method (ion-electron method)

### Half-reaction method in ACIDIC soln:

→ Given unbalanced net ionic equation

1. Assign oxidation num. to all atoms of the unbalanced eqn.
2. Find reactant oxidized (lose electron) and reduced (gain) and divide into oxidation  $\frac{1}{2}$  reaction and reduction  $\frac{1}{2}$  reaction

→ Element with increase in oxidation num. = atom

↳ Species oxidized = the entire compound (reactant)

3. Write both half reactions

4. For each half reaction:
  - Balance all elements other than O and H
    - ① add subscript
    - ② Add substance present in original equation
  - O: Balance number of O by adding  $H_2O$
  - H: Balance number of H by adding  $H^+$
  - $e^-$ : Balance charge by adding electrons ( $e^-$ )

5. Multiply both half-reactions so that the number of  $e^-$  is the same in both

→ oxidation  $\frac{1}{2}$  : electrons in products (lose)

→ reduction  $\frac{1}{2}$  : electrons in reactants (gain)

6. Add both  $\frac{1}{2}$  reactions

→ any species present on both sides cancel out

### Half-reaction method in BASIC soln:

→ Balance as if it were an acidic soln and then

1. Add  $OH^-$  for every  $H^+$  to both sides of equation

2.  $OH^-$  and  $H^+$  on the same side become  $H_2O$

→ cancel out  $H_2O$  on both sides

\* Because it's in basic soln, we have  $OH^-$  (not  $H^+$ )

\* We can add  $H_2O$ ,  $H^+$ ,  $OH^-$

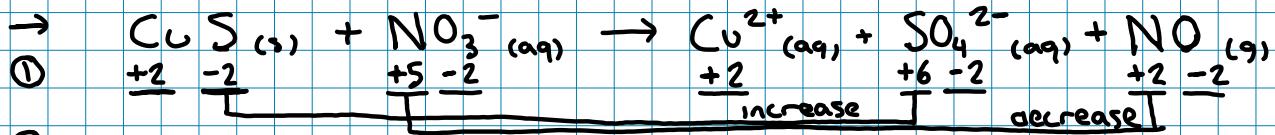
$\downarrow$   
(acidic)       $\downarrow$   
(basic)

Note: Aqueous soln are neutral. The reason there's a charge is because this is the net ionic equation and spectator ions aren't shown

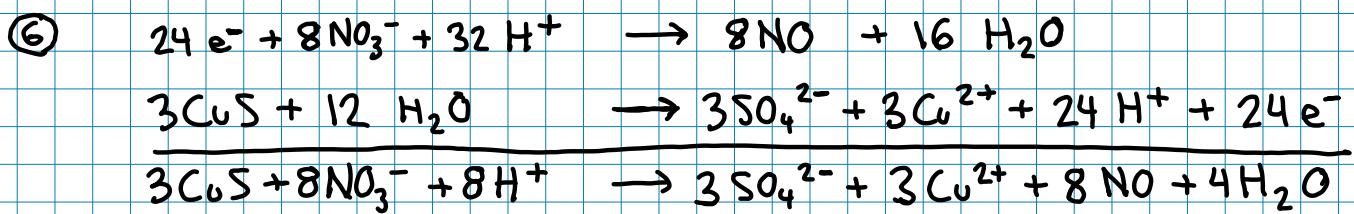
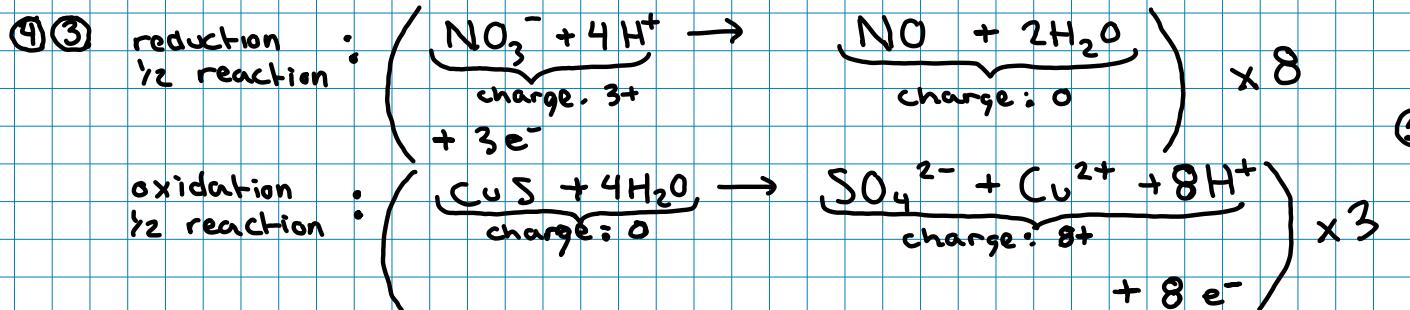
\* Double check atoms and charge

\*  $e^-$  in  $\frac{1}{2}$  reactions, not in balanced redox reaction

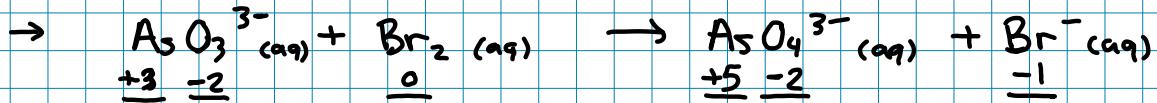
ex: redox in acidic soln.



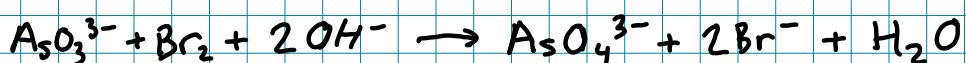
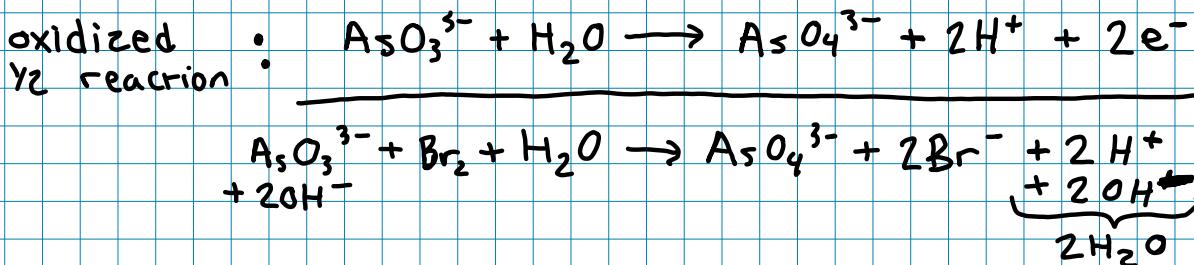
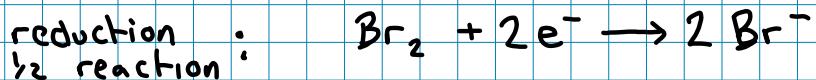
\textcircled{2} N decrease  
 $\hookrightarrow \text{NO}_3^-$  : reduced  
 S increase  
 $\hookrightarrow \text{CuS}$  : oxidized  
 \* Not just element



ex: redox in basic soln.



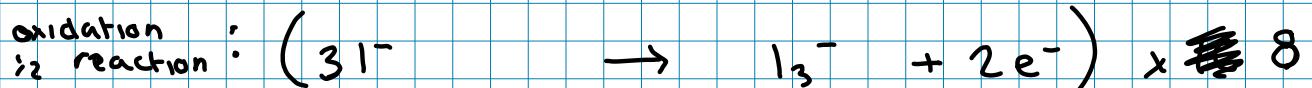
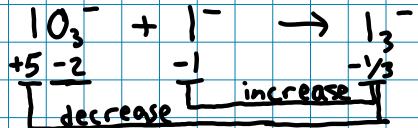
As increase  
 $\hookrightarrow \text{AsO}_3^{3-}$  : oxidized  
 Br decrease  
 $\hookrightarrow \text{Br}_2$  : reduced



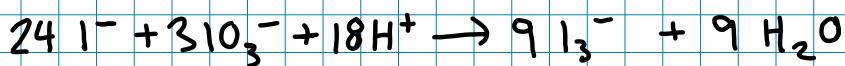
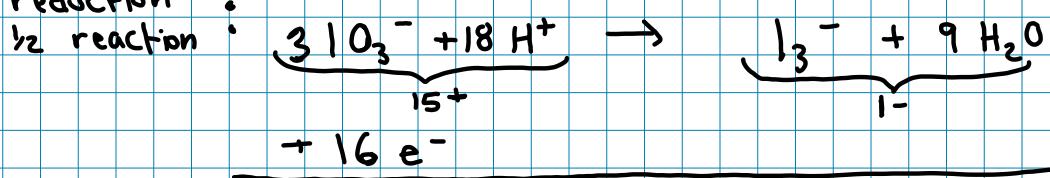
## Hard Redox Problem

- Fractional oxidation number
- Reactant or product present in both  $\frac{1}{2}$  reactions

ex: balance in acidic



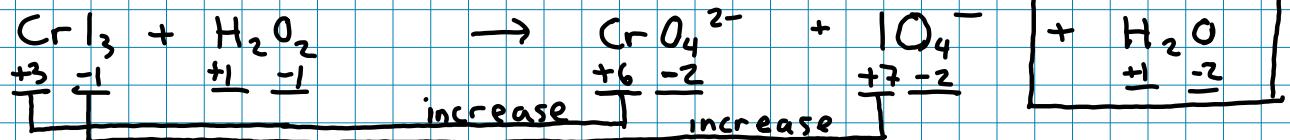
reduction:



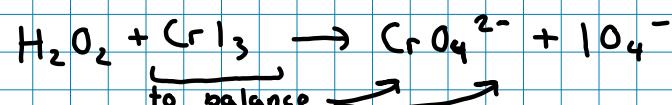
## Hard Redox Problem

- There can't be both a reduction and an oxidation in a  $\frac{1}{2}$  reaction
- If you can't find anything,  $\text{H}_2\text{O}$  can be an unwritten reactant or product.

ex: balance in acidic

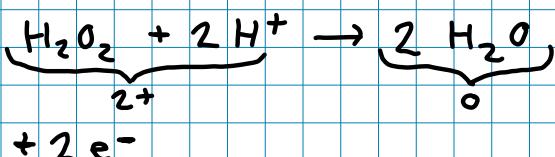


Reduction  $\frac{1}{2}$  reaction  
(BAD)



\* Now I have a reduction and oxidation in  $\frac{1}{2}$   
\* Restart and add  $\text{H}_2\text{O}$  product

Reduction  $\frac{1}{2}$  reaction  
(Good)



# GASES

→ " Ideally ", gases are gas particles moving freely as an individual atom or molecule in space ( container ).

4 variables :

- Pressure (P)

- SI unit : Pa ( Pascal )

$$\text{Pa} = \frac{\text{kg}}{\text{m}^2 \cdot \text{s}^2} \Rightarrow P = \frac{F}{A} = \frac{\text{mass} \cdot \text{acceleration}}{\text{area}}$$

$$= \frac{\text{kg} \cdot \text{m/s}^2}{\text{m}^2}$$

- Volume (V)

- L

- Moles (n)

- Temperature (T)

- K ( Kelvin = Celsius + 273 )

Units for pressure

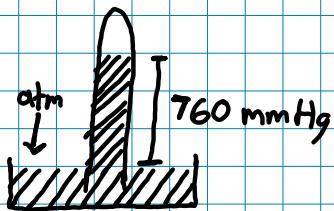
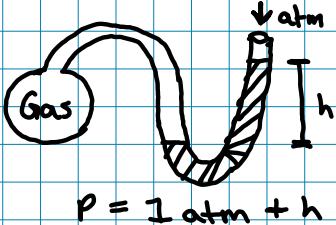
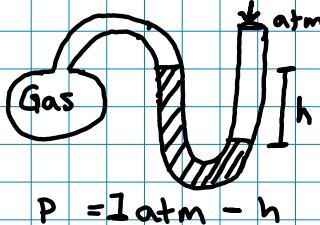
→ 1 atm = 1 standard atmosphere

= 101.3 kPa = 101300 Pa

= 14.7 psi ( pounds per square inch - 1b/in<sup>2</sup> )

= 760 torr = 760 mm Hg ( exact number )

Barometer and manometer



## Gas Laws

$$\frac{P_1 V_1}{n_1 R T_1} = \frac{P_2 V_2}{n_2 R T_2}$$

### Boyle's Law

→ constant moles and temp.

- Pressure and volume are inversely proportional

$$P_1 V_1 = P_2 V_2 \Rightarrow P \propto \frac{1}{V}$$

### Charles Law

→ constant pressure and moles

- Volume and temperature are directly proportional

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow V \propto T$$

Absolute zero : 0 Kelvins

→ Molecular motion comes to halt (no motion)

→ Zero volume

### Avogadro's Law ( $V \propto n$ )

→ constant pressure and temp.

- Volume and moles are directly proportional

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \Rightarrow V \propto n$$

## IDEAL GAS LAW

$$PV = nRT$$

$$R = 8.314 \frac{\text{J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}}{\text{mol}}$$

$$\text{or } R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}$$

(given on exam)

\* Ideal gas law isn't always true

Only under ideal gas conditions

- Above 200 K (high temp)

- Below 10 atm (low pressure)

The more one respects these conditions, the better IGL works

\* Assume IGL holds true unless told otherwise.

How to find molar mass of a gas, given its density?

$$PV = nRT$$

$$PV = \left(\frac{m}{M_m}\right)RT$$

$$M_m = \frac{(m)}{V} R \cdot T$$

$$\boxed{M_m = \frac{d \cdot R \cdot T}{P}}$$

$$m = \text{mass (g)}$$

$$M_m = \text{molar mass (g/mol)}$$

$$n = \frac{m}{M_m} (\text{mol})$$

$$d = \text{density} = \frac{m}{V} (\text{g/L})$$

Standard Temperature and ~~Pressure~~ (STP)

$$\rightarrow 0^\circ\text{C} (273 \text{ K})$$

$$\rightarrow 1 \text{ atm}$$

Standard molar volume = volume of 1 mol of gas at STP

$$= 22.4 \text{ L/mol}$$

$$\frac{V}{n} = \frac{RT}{P} = \frac{(0.08206)(273 \text{ K})}{1 \text{ atm}}$$

↳ Always the same regardless the identity of the gas.

## Gaseous Mixture

Partial pressure = pressure that a gaseous component of a mixture would exert if that gaseous component were present alone in the container.

Dalton's Law of partial pressure:

$$P_A + P_B + P_C + \dots + P_n = P_{TOT}$$

∴ The pressure of a given gas is unaffected by the presence of other gases in the mixture

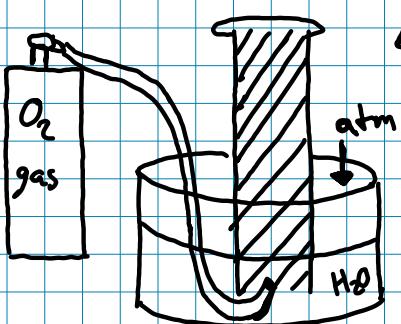
If one has a gas in a container at a given pressure, and one adds a second gas

→ Pressure of gas 1 doesn't change

→ Total pressure of the mixture increases

Gas collected over water

- When gas is collected over water, there is
  - pressure of gas collected
  - pressure of water vapour
  - \*  $P_{TOT} = \text{sum of both partial pressures.}$

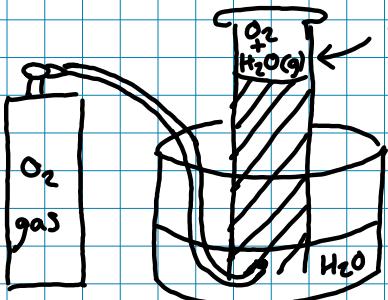


upside down grad cylinder filled with  $H_2O(l)$

### Closed Gas Cylinder

- Atm is strong enough to push down  $H_2O$
- $H_2O$  in cylinder doesn't go down

↓ open O₂ gas cylinder



$O_2 + H_2O(g)$   
\* Total pressure of gaseous mixture is the same as atmospheric pressure

### Open Gas cylinder

- Air bubbles (gas) go up
- Causes water to get displaced by gas
- Close valve

\* Problems where gases are collected over water and then dried  $\Rightarrow$  water vapour removed from mixture

ex: He(g) collected over water at 25°C and 1.00 atm  
What volume of gas must be collected to obtain  
0.586 g He?  
The vapour pressure of H<sub>2</sub>O is 23.8 torr

$$P_{\text{tot}} = P_{\text{He}} + P_{\text{vap}} = 1.00 \text{ atm}$$

$$0.586 \text{ g He} \cdot \frac{\text{mol He}}{4.003 \text{ g He}} = 0.1464 \text{ mol He}$$

$$P_{\text{tot}} = \left( \underbrace{\frac{P_{\text{He}} \cdot R \cdot T}{V}}_{P_{\text{He}}} \right) + \left( 23.8 \text{ torr} \cdot \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 1.00 \text{ atm}$$

$$1.00 = \frac{(0.1464)(0.08206)(298)}{V} + \frac{23.8}{760}$$

$$\boxed{V = 3.7 \text{ L}}$$

## Mole fraction / mole percentage

→ 2 gas mixture (gas A + gas B)

Mole fraction (of gas A) :

$$\underline{X}_A = \frac{n_A}{n_{TOT}} \quad \text{or} \quad \begin{matrix} (\text{for mixture of} \\ 2 \text{ gases } A+B) \end{matrix} \quad \underline{X}_A = \frac{n_A}{n_A+n_B}$$

Mole percentage (of gas A) : \* For gaseous mixtures  
% volume = mole %.

$$\underline{X}_A \times 100$$

Mole fraction and partial pressure

$$\boxed{P_A = \underline{X}_A \times P_{TOT}}$$

Proof:

$$\frac{P_A}{P_{TOT}} = \frac{\frac{n_A \cdot R \cdot T}{n_{TOT} \cdot R \cdot T}}{1} = \frac{n_A}{n_{TOT}} = \underline{X}_A$$

$$\therefore P_A = \underline{X}_A \cdot P_{TOT}$$

ex. Atmospheric pressure surface of Mars  $\cdot 5.92 \times 10^{-3}$  atm  
Martian atmosphere 95.3% CO<sub>2</sub> and 2.7% N<sub>2</sub>.  
Find mole fraction and partial pressure of N<sub>2</sub>.

$$\frac{\% \text{ volume}}{100} = \underline{X}_{N_2} = \frac{2.7}{100} = 0.027$$

$$P_{N_2} = \underline{X}_{N_2} \cdot P_{TOT} = 0.027 \cdot (5.92 \times 10^{-3} \text{ atm}) = 1.6 \times 10^{-4}$$

## Kinetic Molecular Theory (KMT)

→ explains  $T$ , ideal gas laws

Law: Observed

Theory: consistent explanation of law

1. Gas particle volume is negligible  
∴ separated by large distance
2. Gas particles are continuously in motion and colliding
  - collisions between particles are elastic
  - collisions with walls creates pressure
3. Average kinetic energy (energy of motion) of gas particles is proportional to the temperature.  
→  $T \uparrow$ : move faster  
 $T \downarrow$ : move slower
4. Ideal gas particles exert no force (attract / repel) on other gas particles

example of explanation:

Avogadro's law:  $n \propto V$

- constant average kinetic energy at constant temp.
- to maintain constant pressure (collisions with walls) as one increases the number of moles of gas, one must increase the volume of its container

## Kinetic Energy (K.E.)

$$\left\{ \frac{1}{2} N_A \cdot m \bar{v}^2 = RT \right.$$

Base equation

Forces

more force → more mass → more particles → more collision  
 ↳ more velocity → more frequent collisions

∴ Pressure  $\propto$  to Force  $\propto$  collisions

$$K.E. = \frac{1}{2} mv^2 \quad (\text{of a single particle})$$

$$\text{Unit : Joule (J)} = \text{kg m}^2/\text{s}^2$$

\* For a mol of particles

$$\boxed{(K.E.)_{avg} = \frac{3}{2} (R \cdot T)} \quad [J/mol]$$

$$R = 8.314 \frac{J}{K \cdot mol} \quad \text{or} \quad \frac{\text{Kg m}^2}{\text{K} \cdot \text{mol}}$$

→  $(K.E.)_{avg}$  depends solely on the temperature,  
 Not the identity of the gas

Root-mean-square velocity (urms) (extremely close to  $v_{avg}$ )

$$\boxed{\sqrt{\bar{v}^2} = urms = \sqrt{\frac{R \cdot T \cdot 3}{m \cdot m}}}$$

$$R = 8.314 \frac{\text{Kg} \cdot \frac{\text{m}^2}{\text{s}^2}}{\text{K} \cdot \text{mol}}$$

$T \Rightarrow$  in kelvin

$m \cdot m \Rightarrow$  in Kg/mol

$$\rightarrow urms \propto \sqrt{\text{temp}}$$

$$\rightarrow urms \propto \frac{1}{\sqrt{m \cdot m}}$$

\* NOT g/mol → Kg/mol

↑ T, speed ↑  
 ↑ mass, speed ↓  
 Lighter = faster / heavier = slower

Ex: compare urms and  $(K.E.)_{avg}$  of H<sub>2</sub> and SF<sub>6</sub>

$$urms (\text{of H}_2) = \sqrt{\frac{3RT}{mm}} = \sqrt{\frac{3 \cdot 8.314 \cdot 273}{0.00202 \text{ Kg/mol}}} = 1840 \text{ m/s}$$

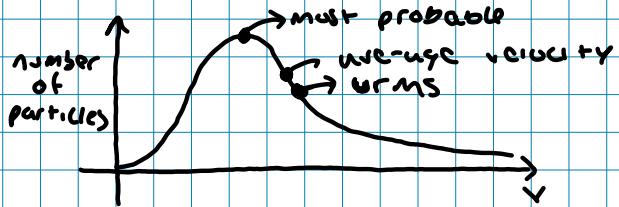
$$urms (\text{of SF}_6) = \sqrt{\frac{3 \cdot 8.314 \cdot 273}{146 \text{ g/mol} \cdot \text{Kg/1000g}}} = 216 \text{ m/s}$$

$(K.E.)_{avg}$  is the same for both. It depend solely on temperature

\* Root-mean-square velocity is an average; some particles are going faster and some are going slower.

## Root-mean-square velocity vs. mean velocity

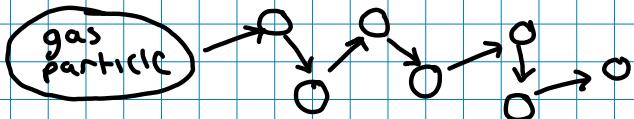
- $v_{rms}$  is slightly higher velocity than mean velocity
- But, it is very close and gives us a good sense of average velocity.



Gas particles travel extremely quickly ( $>500 \text{ m/s}$ ), however it can still take up to 10s for a particle to get from one side to another of a room.

Mean free path - distance a particle travels before colliding with another particle

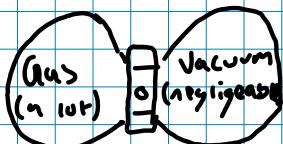
In a room filled with air, the mean free path is very small. The gas will travel in a zig-zag from the back of the front of the room as it hits and collides with different gas particles.



## Effusion

Rate of effusion = rate at which a gas travels through a pin-hole into a vacuum

⇒ open stopcock and effusion occurs.



Rate of effusion  $\propto$   $U_{rms}$

## Graham's Law

$$\frac{\text{Rate of effusion (gas A)}}{\text{Rate of effusion (gas B)}} = \sqrt{\frac{m.m.(B)}{m.m.(A)}}$$

\* Assuming one is comparing same moles of gas.

$$\frac{\text{Rate effusion (A)}}{\text{Rate effusion (B)}} = \frac{\text{Time effusion (B)}}{\text{Time effusion (A)}} = \sqrt{\frac{m.m.(B)}{m.m.(A)}}$$

→ Faster takes less time

→ Lighter particles effuse faster (time ↓, rate ↑)

\* units must be the same for numerator/denominator

ex: Sample of argon gas effuses in 147 s. How much time for the same number of moles of  $SO_2$ ?

$$\frac{X \text{ s } (SO_2)}{147 \text{ s } (Ar)} = \sqrt{\frac{m.m.(SO_2)}{m.m.(Ar)}}$$

$$X = 186 \text{ s}$$

check.  $SO_2$  heavier:  
takes more time

$$\frac{\text{Rate effusion} \cdot \text{Time}}{(Ar)} \cdot \frac{\text{mol}}{(Ar)} - \frac{\text{mol}}{(SO_2)} = \frac{\text{Rate effusion} \cdot \text{Time}}{(SO_2)} \cdot \frac{\text{mol}}{(SO_2)} - \frac{\text{mol}}{(SO_2)}$$

$$\frac{\text{mol/s}}{\text{s}} = \frac{\text{mol}}{\text{s}}$$

## Real Gases

→ Ideal gases follow ideal gas law:  $PV/nRT = 1$   
• Only true at low pressure and high temp

\* KMT assumes no interparticle forces (attract/repel) and that particles have no volume.

Gas particles don't follow KMT at high pressure:

- High pressure → greater concentration → greater likelihood of particles approaching each other and be ATTRACTED / REPELLED

Gas particles don't follow KMT at low temperature:

- High EK at high temp.  
→ move fast, not likely to be affected by attractive or repulsive forces of other particles
- Low EK at low temp.  
→ move slow, more likely to be attracted/repelled  
ex: beautiful rose: run by → don't pay attention  
pass by slowly → attracted

\* USE ideal gas law unless told otherwise

At high pressure and low temp. use:

Van der Waals equation

$$\boxed{(P + a(\frac{n}{V})^2)(V - nb) = nRT}$$

a = Van der Waals "a" constant

b = Van der Waals "b" constant

a → in atm  $\cdot \frac{L^2}{mol^2}$

b → in  $\frac{L}{mol}$

"b" typically increases as the size of gas particle increases  
"a" related to strength also increases with the size

\* As "a" and "b" constant values increase, we say that the gas deviates more from the  $PV = nRT$  value

## Theory behind Van der Waals equation :

1. - Particles can't move in the entirety of the container.
  - Only in volume not being occupied by other gas particle(s) (free space)
  - $V - nb$   
volume container" - "volume filled with other particles"
  - $nb$  = volume correction factor
  - bigger the volume of a particle = bigger  $b$
2. - Ideal gas  $\Rightarrow$  no attraction
  - If there are attractions, the attraction results in the path of a particle towards a wall to be diverted  $\Rightarrow$  less collisions  
↳ pressure with attraction < ideal gas pressure
  - $P$  = measured pressure
  - $a(n/V)^2$  = correction factor
    - concentration ( $n/V$ )  $\uparrow$   
 $\therefore$  chance of attraction  $\uparrow$   
 $\therefore$  collision with wall  $\downarrow$   
 $\therefore$  pressure  $\downarrow$
  - use  $(n/V)^2 \Rightarrow$  square because 2 particles required

ex: 8.00 kg N<sub>2</sub>(g) fills 100. L flask at 300°C  
Calculate pressure.

$$\text{For N}_2 : a = 1.390 \text{ atm } \frac{\text{L}^2}{\text{mol}^2}, b = 0.03913 \text{ L/mol}$$

$$(P + a(n/V)^2)(V - nb) = nRT \quad 8.00 \text{ kg} \cdot \frac{1000 \text{ g}}{\text{kg}} \cdot \frac{\text{mol}}{28.02 \text{ g/mol}} = 286$$

~~$$(P + 1.390 \left( \frac{286}{100} \right)^2)(100 - 286 \cdot 0.03913) = 286 \cdot 0.08206 \cdot 573$$~~

$$(P + 1.390 \left( \frac{286}{100} \right)^2)(100 - 286 \cdot 0.03913) = 286 \cdot 0.08206 \cdot 573$$

$$\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

$$P = 140 \text{ atm}$$

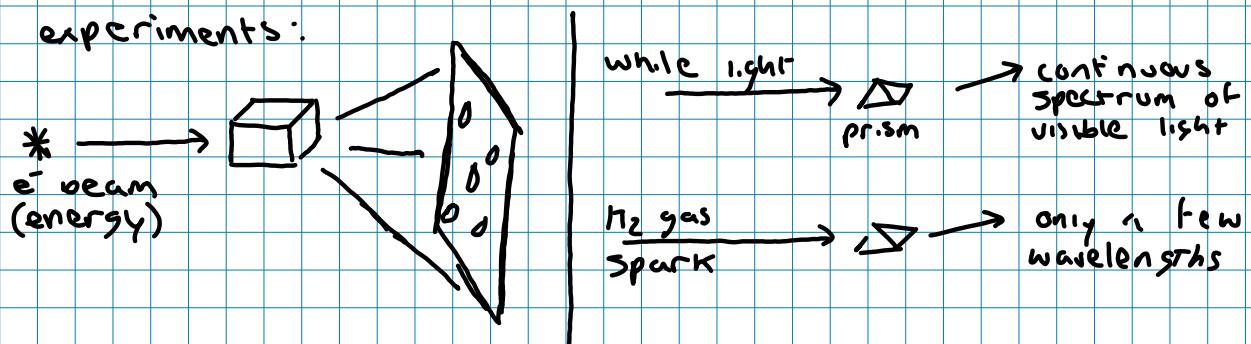
\* Note! use  $R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$  value

# Quantum Numbers

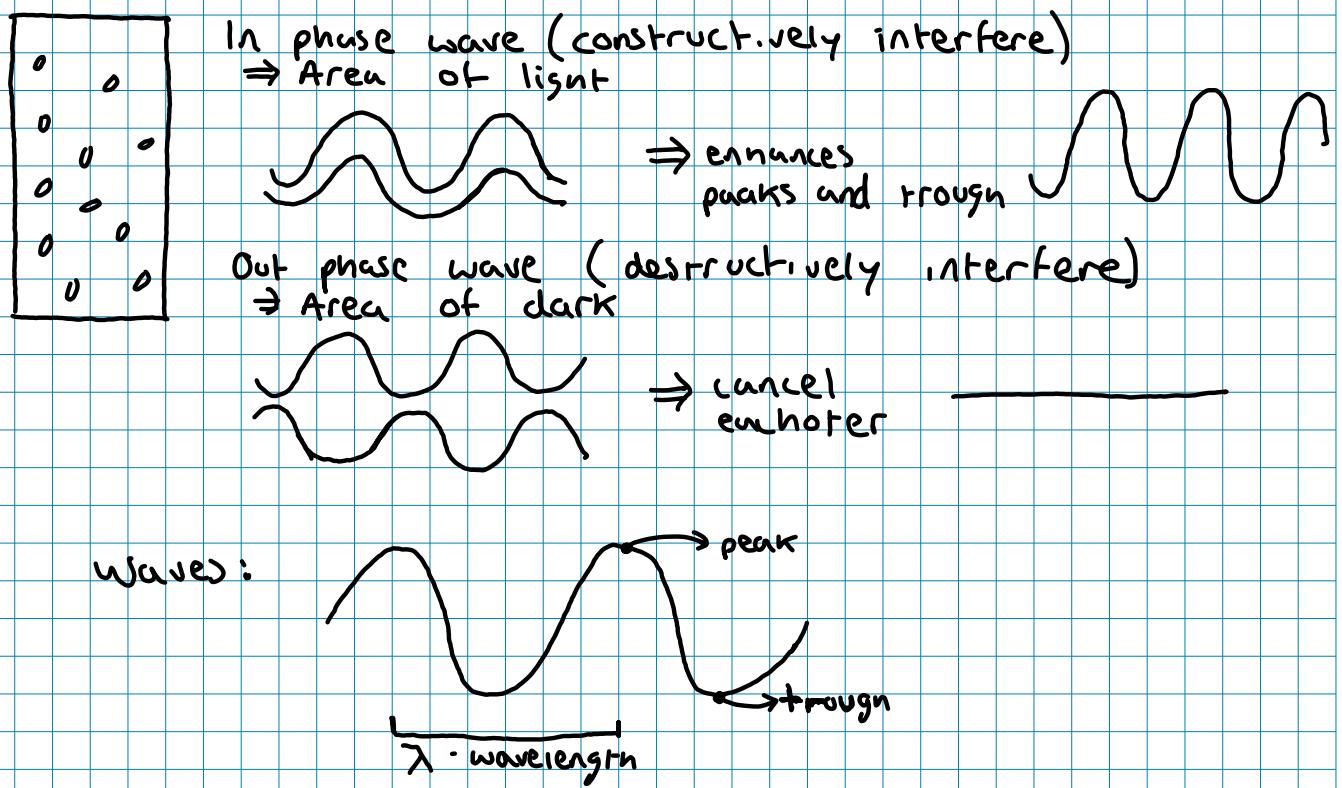
Electrons are the most important subatomic particles in chemical reactivity

One thinks of an electron as a particle and as a wave

experiments:

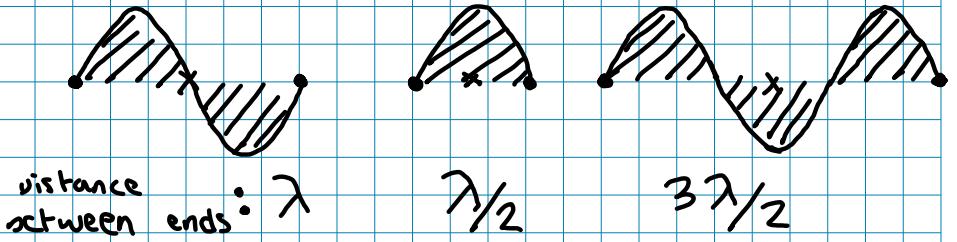


- So only certain wavelengths associated with  $H_2(g)$
- Diffraction pattern when electron beam hits crystal



It is theorized that electrons are standing waves (waves with fixed ends). Fixed ends because the nucleus keeps the electrons close.

X - nucleus  
● - electron  
/// = wave



distance between ends:  $\lambda$

$\lambda/2$

$3\lambda/2$

There are only certain allowed integer values to explain how these electrons live in the electron cloud around the nucleus.

Orbital = (wave function)<sup>2</sup>  
= Area in which electrons have a high probability of existing

Quantum numbers are used to describe orbitals

Quantum numbers:

$\rightarrow n$   
 $\rightarrow l$   
 $\rightarrow m_l$   
 $\rightarrow m_s$

} orbitals  
} electrons per orbital

{ \* More stable = lower energy }

Quantum numbers  $\rightarrow$  quantization of energy  
(quantization = only certain allowed values)

1.  $n$  = (shell number) = (principal quantum number)

→ size of orbital (larger number = larger size)

\* Allowed values : positive integer (1, 2, 3, ...)

ex: which is more stable :  $n=1$  or  $n=2$ ?

$n=1$  is smaller orbital than  $n=2$

∴ electron in  $n=1$  is closer to nucleus

∴ more stable ∴ lower energy

2.  $\ell$  = (subshell number) = (angular momentum QN)

→ shape of orbital

\* Allowed values : all integers between 0 to  $n-1$

ex: values of  $\ell$  for  $n=3$ ?

$n=3$        $\ell$  from 0 to  $3-1$

0 to 2      ∴  $\ell = 0, 1, 2$

\*  $\ell = 0, 1, 2, 3$       Sets Poor Daniel Free

Value $\ell$	Subshell name	Shape ( $x$ = nucleus)
0	S	spherical
1	P	dumbbell 2 lobbed
2	d	4 lobbed
3	f	8 lobbed

Note:  $\ell=4$  : g and  $\ell=5$  : h (no known atoms)

Describing orbitals :  $n$  (subshell name)

ex:  $n=3, \ell=1 \rightarrow 3p$  orbital

$n=2, \ell=2 \rightarrow 2d$  orbital

## Energy in orbitals:

- For polyelectronic atoms (all atoms but H), if first and second orbital have the same principal quantum number ( $n$ )
  - The orbital with the smaller  $l$  value has lower energy
- Penetration theory: orbital with lower subshell number (assuming same shell number) penetrates the area near the nucleus more effectively
- Closer to nucleus ∴ more stable ∴ lower energy

Ex: For polyelectronic atom, what is lower energy?

$3s$  or  $3p$  orbital?

Both have same shell number, but  $3s$  has lower subshell number ∴  $3s$

3.  $m_l = \text{(magnetic QN)} = \text{(final QN to describe orbital)}$

→ orientation of orbital in 3D

\* Allowed values: all integers between  $-l$  and  $l$   
(including zero)

Ex: If  $l=3$

$$m_l = -3, -2, -1, 0, +1, +2, +3$$

⊗ A combination of these three QN ( $n, l, m_l$ ) uniquely describes an orbital.

Ex: how many possible orbitals for  $n=3$ ?

$$n = 3$$

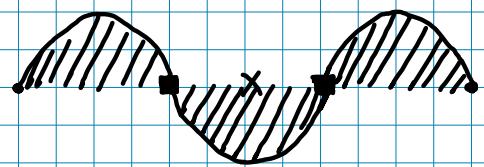
$$l = 0, 1, 2$$

$$m_l = 0 \quad -1, 0, +1 \quad -2, -1, 0, +1, +2$$

↑      ↑      ↑  
one  $3s$     three  $3p$     five  $3d$

→ 9 orbitals

## Sketching orbitals



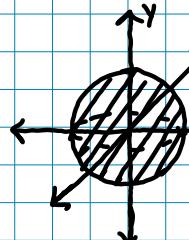
$\circ$  = fixed end  
 $x$  = nucleus  
 $\backslash\backslash\backslash$  = area where electrons exist  
 ■ = node

node of wave = zero amplitude region (there is no wave). zero probability of finding electrons  
 → node, nodal surface or nodal plane

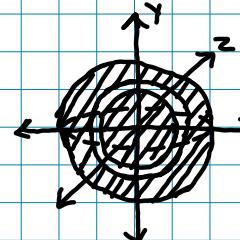
number of nodes =  $n-1$  (for s orbital)

S: Draw 1s and 2s orbital: (spherical)

$$1s(n=1) \\ (l=0) \\ (m_l=0) \\ \text{node} = n-1 = 0$$



$$2s(n=2) \\ (l=0) \\ (m_l=0) \\ \text{node} = n-1 = 1$$

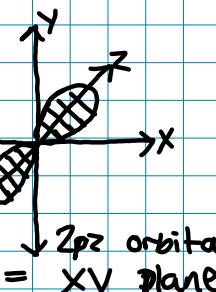
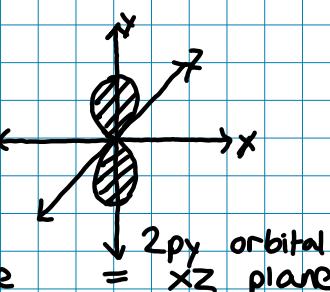
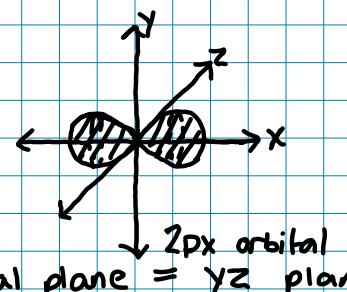


node is unshaded razor thin area

\* 3 coordinate axis for each orbital (origin (0,0) is nucleus)

P: Draw 2p orbitals: (dumbbell)

$$3 \left\{ \begin{array}{l} (n=2, l=1, m_l=-1) \\ (n=2, l=1, m_l=0) \\ (n=2, l=1, m_l=+1) \end{array} \right. \\ \text{node} = n-1 = 1 \\ (\text{nodal plane splits in two})$$



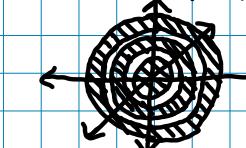
- The 3 orbitals are mutually orthogonal ( $\perp$ )
- 3 orbitals for P ( $l=1$ ):  $p_x, p_y, p_z$

d: 3d orbitals:

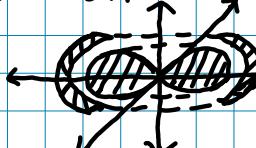
$$\begin{array}{l} n=3 \\ l=2 \\ m_l=-2, -1, 0, +1, +2 \end{array}$$

5 orbitals:  $d_{xy}, d_{xz}, d_{yz}, d_{x^2-y^2}, d_{z^2}$

ex: 3s orbital (2 nodes)



3px orbital (2 nodes)



- between dumbbell  
 - razor thin layer

2 electrons per orbital

→ each electron differs from the other on how they interact with a magnetic field

$$4. m_s = (\text{electron spin QN}) = (\text{final QN})$$

→ describes electro

\* Allowed values :  $m_s = +\frac{1}{2}$  (spin up  $e^-$ )  
 $m_s = -\frac{1}{2}$  (spin down  $e^-$ )

electron =  $\uparrow$   $\boxed{\uparrow \downarrow}$  is orbital

(\*) every  $e^-$  can be described by 4 QN ( $n, l, m_l, m_s$ )  
→ Pauli exclusion principle

ex: What is max number of  $e^-$  for ( $n=2, m_s = -\frac{1}{2}$ )?

$$\begin{array}{l} n=2 \\ l=0 \\ m_l=0 \end{array} \quad \begin{array}{c} 1 \\ -1, 0, +1 \end{array} \quad \begin{array}{c} 4 \\ \text{orbital} \end{array} \times \begin{array}{c} 1e^- \\ (\text{per orbital}) \end{array} = 4 e^- \text{ max.}$$

ex: What is max number of  $e^-$  for ( $n=4, l=2$ )?

$$\begin{array}{l} n=4 \\ l=2 \\ m_l=-2, -1, 0, +1, +2 \end{array} \quad \begin{array}{c} 5 \\ \text{orbital} \end{array} \times \begin{array}{c} 2 e^- \\ (\text{per orbital}) \end{array} = 10 e^- \text{ max.}$$

ex: What is max number of  $e^-$  for ( $n=3, m_l=-3$ )?

$$\begin{array}{l} n=3 \\ l=0 \\ m_l=0 \end{array} \quad \begin{array}{c} 1 \\ -1, 0, +1 \end{array} \quad \begin{array}{c} 2 \\ -2, -1, 0, +1, +2 \end{array}$$

∴ zero  $e^-$   
b/c  $m_l \neq -3$

## Quantum Numbers and Periodic Table

→ Modern atomic theory shows how electrons are theorized to exist in orbitals

Aufbau Principle = "building up" = states that for each proton in an atom, one places an electron in the lowest energy available orbital  
→ lower energy = more stable

Energy:

- Increase in energy as "n" increase
- Increase in energy as "l" increase
- Lower " $n+l$ " = lower energy.

Atomic orbital box / line diagram:

- Arrow designates electron (first pointing up)

Alternative notation:

- Shell number, subshell letter, superscript x  $\Rightarrow n^l^x$   
 $\Rightarrow x$  = number of electrons in subshell

ex: Hydrogen



Helium



Ground State Electron Configuration (GSEC):

- Electrons for an atom go to the most stable (low energy) possible orbital

\* If electrons are not written in the most stable possible orbital, it is ~~%~~ an excited state  
→ energy must be applied to reach excited state

ex: excited Helium :  $\frac{\uparrow}{1s} \quad \frac{\uparrow}{2s}$

\* When writing GSEC for elements, one can replace filled orbitals by the previous noble gas

ex: replace  $1s^2$  by [He]

Note: relation between shell number and period (row) where element is found.

Hunds Rule: When given the choice to place an electron in 2 degenerate orbitals, one chooses to place the electron in an empty orbital rather than in an orbital with another electron.

- Degenerate = equal in energy  
= 2 orbitals that have same "n" and "l"
  - Paired  $e^-$  =  $e^-$  found in orbital with other  $e^-$
  - unpaired  $e^-$  =  $e^-$  found in orbital alone
- \* Electrons prefer to be unpaired to avoid  $e^-/e^-$  repulsion

Valence electrons = electrons found in the highest shell number containing electrons

- Valence shell = shell number containing val.  $e^-$
  - core electrons = non-valence  $e^-$  of an atom
- \* Valence electrons are found on shell number corresponding to row number

Main group elements, valence  $e^-$  correspond to Roman numeral.

- Let  $n$  = row number in periodic table

- Valence  $e^-$  are:

- gr IA  $\Rightarrow$  1 val  $e^- \Rightarrow nS^1$
- gr IIA  $\Rightarrow$  2 val  $e^- \Rightarrow nS^2$
- gr IIIA  $\Rightarrow$  3 val  $e^- \Rightarrow nS^2 nP^1$
- gr IVA  $\Rightarrow$  4 val  $e^- \Rightarrow nS^2 nP^2$
- gr VA  $\Rightarrow$  5 val  $e^- \Rightarrow nS^2 nP^3$
- gr VIA  $\Rightarrow$  6 val  $e^- \Rightarrow nS^2 nP^4$
- gr VIIA  $\Rightarrow$  7 val  $e^- \Rightarrow nS^2 nP^5$
- gr VIIIA  $\Rightarrow$  8 val  $e^- \Rightarrow nS^2 nP^6$

Transition metal: last electrons fill d orbital

Lanthanides : last electrons fill 4f orbital  
Actinides : last electrons fill 5f orbital

## Rule for expected order of filling $e^-$ in orbitals

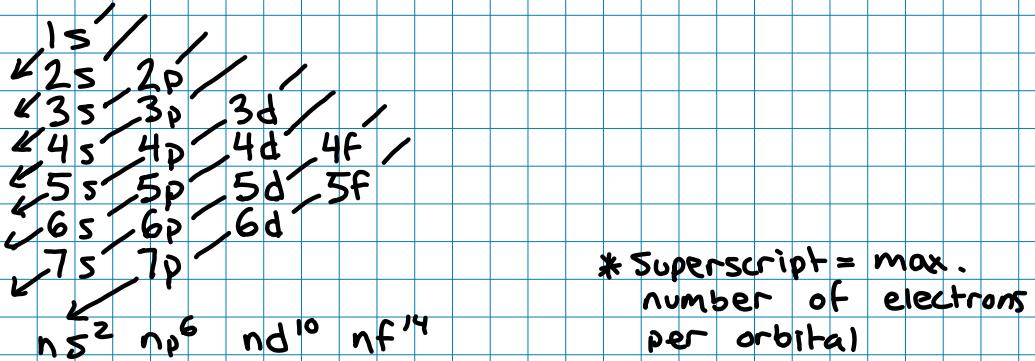
→ Assigned in order of  $n+l$  value

→ For 2 orbitals with equal value, assign to lower  $n$

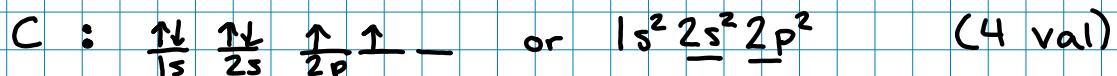
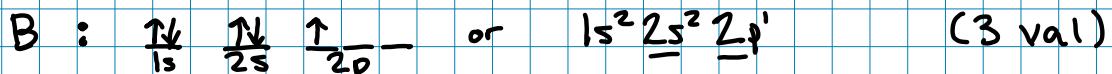
$$\text{ex: } \begin{array}{ccccccccc} 3s & 3p & 3d & & 4s & 4p & 4d \\ n+l = & 3 & 4 & 5 & & 4 & 5 & 6 \end{array} \quad \therefore 3s, 3p, 4s, 3d, 4p$$

(lower "n")

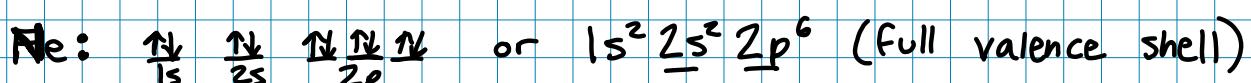
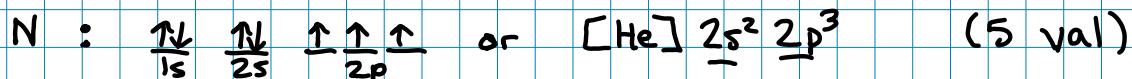
Trick:



GSEC for main group elements: (Row 2)

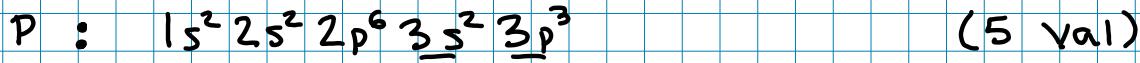


$\frac{\uparrow\downarrow}{1s} \quad \frac{\uparrow\downarrow}{2s} \quad \frac{\uparrow\downarrow}{2p} -$  Wrong b/c paired  $e^-$  (Hund's rule)  
This is excited state (repulsion)



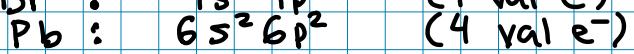
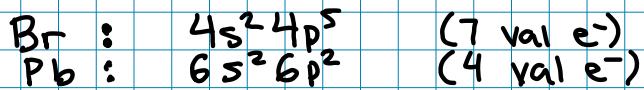
\* Note: Valence shell underlined  
↳  $n=2$  shell (corresponds to row 2)

GSEC for main group elements: (Row 3)

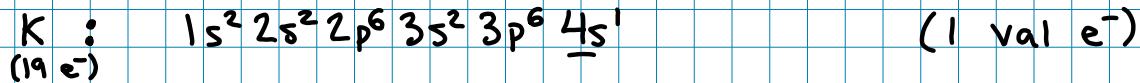


→ same number of val.  $e^-$  as N  
↳ Both in group VIA

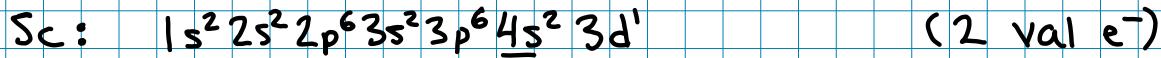
Valence  $e^-$  for main gr. elements: (underlined)



GSEC for elements: (Row 4) - main and transition



→ What happened to 3d orbital  
↳ follow  $n+l$



→ Final  $e^-$  went into "d" orbital

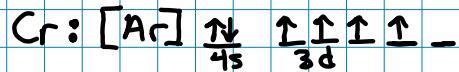
Final  $e^-$  go into...

3d orbital	⇒ Sc - Zn	T metal
4p orbital	⇒ Ga - Kr	
5s orbital	⇒ Rb - Sr	
4d orbital	⇒ Y - Cd	T metal
5p orbital	⇒ In - Xe	
6s orbital	⇒ Cs - Ba	
4f orbital	⇒ Lanthanides	
5f orbital	⇒ actinides	
5d orbital	⇒ La - Hg	T metal

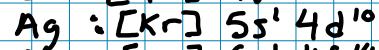
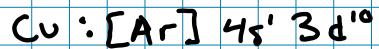
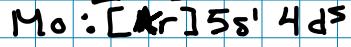
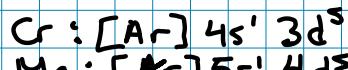
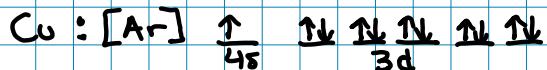
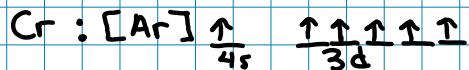
## Exceptions to expected order of filling

→ Group VII B : Cr, Mo  
 → Group I B : Cu, Ag, Au

Expected GSEC for:



Experimentally:



Theorized: Cr and Mo exception →

To have all electrons unpaired (4s 3d or 5s 4d)  
 To avoid unfavourable repulsion.

Cu, Ag and Au →

Having all 5 d orbitals filled provides  
 a certain stability

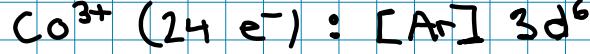
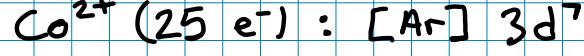
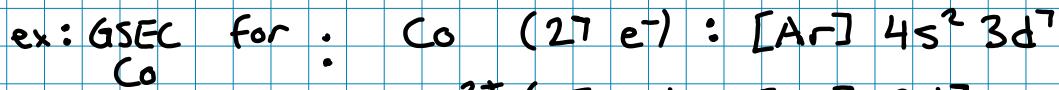
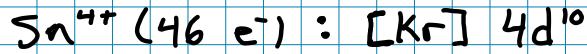
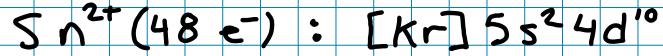
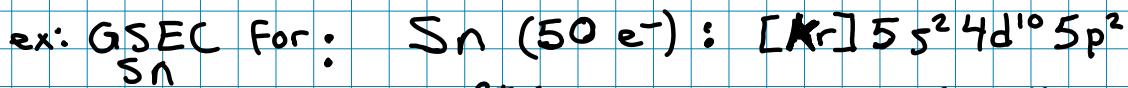
## Electron configuration for cations (+)

\* Remove e<sup>-</sup> from furthest valence shell

- Let n = highest shell number containing e<sup>-</sup>

• Main group elements: remove first from np orbitals (if present), then ns

• Transition metals: remove from ns, and only once emptied, remove from (n-1) d



# Periodic Trends

- Properties of atoms related to their position in the periodic table
- ex: ionization energy, electron affinity, electronegativity, atomic radius, density, etc.

## 1. Ionization energy

= energy required to remove a single electron from nucleus, which is attractive (+ and -), of a gaseous atom



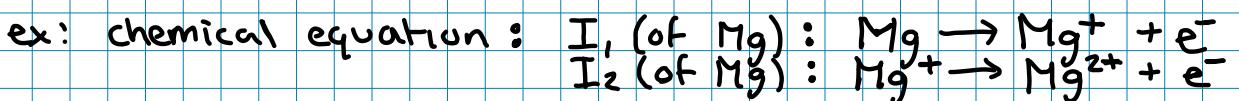
It requires energy to pull electron away from attracting positive nucleus

$I_1$  = first ionization energy

energy required to remove first  $e^-$

$I_2$  = second ionization energy

energy to remove second  $e^-$



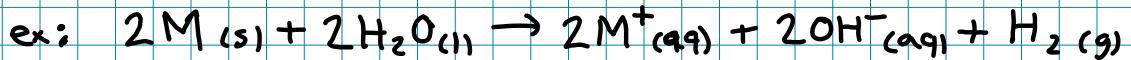
Up → Down trend

- First ionization energy decreases going down a group

Theory: Electrons are removed from furthest valence shell. The higher you are, the ~~higher / up / made~~ closer the valence shell is to the attractive nucleus → the harder it is to pull an  $e^-$

\* Same logic, an element will be a better reducing agent when located lower in periodic table (more reactive)

Easier to remove  $e^-$  ∴ requires less energy



Which will be better oxidized (=reducing agent)

$M = K$  or  $M = Na$ ?

K b/c valence shell larger than Na ∴ val  $e^-$  further from nucleus ∴ better at giving  $e^-$   
∴ more reactive

Left  $\rightarrow$  Right

~~Biggest to Left~~ trend (general)

$I_1$  increases going left to right across period

Theory: number of protons increase going left to right

$\therefore$  electrons held more tightly in a given shell  
 $\therefore$  harder to remove  $\therefore$  requires more energy (more protons = more effective nuclear charge  $\therefore$  greater ionization energy)

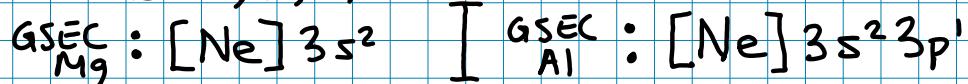
ex:  $I_1$  for row 3:

Na	Mg	Al	Si	P	S	Cl	Ar
495	735	580,	780	1060	1005,	1225	1527

exception

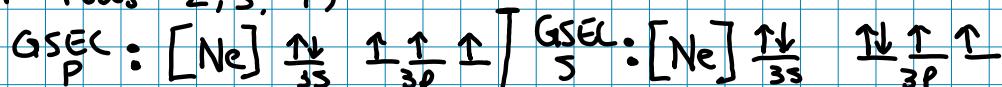
### Exceptions:

- Group IIIA element has smaller  $I_1$  than IIA (for rows 2, 3, 4)



The first  $e^-$  being removed from Al is from a  $p$  orbital, which is less stable than an  $s$  orbital  $\therefore$  easier to remove

- Group VIA element has smaller  $I_1$  than ~~VIIA~~, VIIA (for rows 2, 3, 4)



It's easier to remove an  $e^-$  from sulfur b/c one is relieving an unfavorable  $e^-/e^-$  repulsion

- \* Exception disappears once we reach row 5 b/c  $e^-/e^-$  repulsion is not severe in a very large  $5p$  orbital.

Trend in ionization energy in a given atom  
ex: Mg  $\rightarrow$  GSEC:  $1s^2 2s^2 2p^6 3s^2$

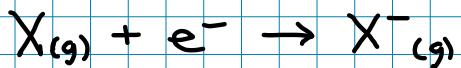
$$I_1 = 735, I_2 = 1440, I_3 = 7730 \text{ kJ/mol}$$

General increase in ionization energy b/c it becomes harder to rip off  $e^-$  when proton/electron ratio increases

$\rightarrow$  There is a dramatic jump for  $I_3$  b/c the 3rd  $e^-$  is removed from a  $n=2$  shell, which is closer to nucleus than  $n=3$  shell (valence) from which first 2  $e^-$  were removed.

## 2. Electron Affinity

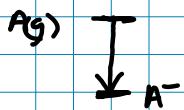
= energy change associated with a gaseous atom or ion accepting a single electron



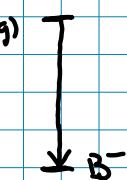
The more negative the value of electron affinity (greater number), the greater the electron affinity  
 ∴ the more exothermic the electron affinity  
 ∴ the more favourable

The lower the energy, the more stable  
 ∴ greater electron affinity means a greater drop in energy  
 ∴ more favourable b/c it makes an atom more stable

ex:  $A_{(g)}$  e<sup>-</sup> affinity = -50 kJ



$B_{(g)}$  e<sup>-</sup> affinity = -100 kJ



$B^-$  is more stable (low energy) relative to  $B_{(g)}$  than  
 $A^-$  is relative to  $A_{(g)}$   
 ∴  $B_{(g)}$  accepts more favourably.

Left → Right trend

- electron affinity becomes more favourable (greater num.)

Theory: greater number of protons as one moves to the right, so better ability to pick up e<sup>-</sup>

Up → Down trend

- electron affinity becomes less favourable

Theory: as one moves down valence shell is further from nucleus ∴ harder to pick up e<sup>-</sup>.

Exceptions:

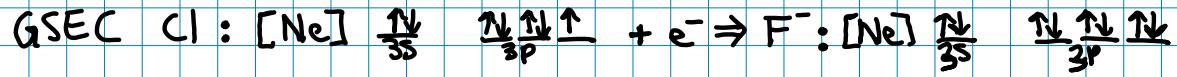
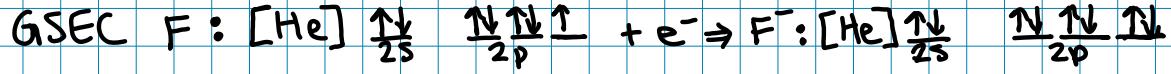
- Group VIII A (noble gases) are an exception to increase in electron affinity

Noble gases have unfavourable electron affinity b/c they have a full valence shell.

ex: Valence e<sup>-</sup> Ar:  $\frac{N}{3s} \frac{1\uparrow 1\downarrow 1\uparrow}{3p} + e^- = \frac{N}{3s} \frac{1\uparrow 1\downarrow 1\uparrow}{3p} \frac{1}{4s}$

- exception to up  $\rightarrow$  down (less favourable) trend
  - Cl more favourable than F
  - S more favourable than O
  - P more favourable than N

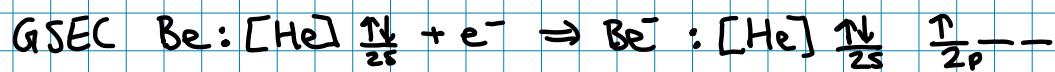
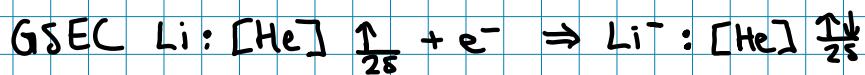
N	O	F
P	S	Cl



F has a less favourable  $e^-$  affinity than Cl b/c it accepts the extra  $e^-$  in a  $2p$  orbital (smallest  $p$  orbital)  
 $\therefore e^-/e^-$  repulsion is very strong and unfavourable.

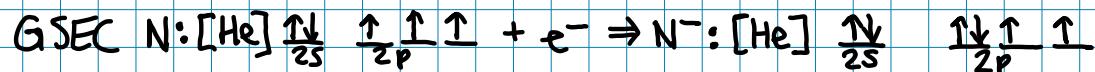
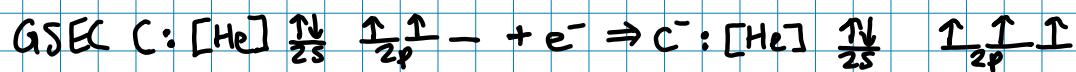
ex: $e^-$ affinity of:	F	- 327.8 KJ/mol
	Cl	- 348.7 KJ/mol
	Br	- 324.5 KJ/mol
	I	- 295.2 KJ/mol

- exception to left  $\rightarrow$  right (more favourable) trend
  - Group IA has more favourable  $e^-$  affinity than IIA (for rows 2, 3, 4)



Li is more favourable  $e^-$  affinity than Be  
 b/c  $2p$  orbital is less stable than  $2s$ , and  
 Be accepts extra  $e^-$  on  $2p$ , while Li  
 accepts the extra  $e^-$  on  $2s$  orbital

- Group IVA has more favourable  $e^-$  affinity than VA (for rows 2, 3, 4)



C is more favourable  $e^-$  affinity than N  
 b/c C accepts extra  $e^-$  into an empty orbital  
 but N accepts extra  $e^-$  paired with another  $e^-$   
 creating an unfavourable  $e^-/e^-$  repulsion

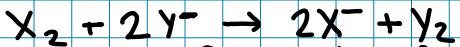
### 3. Halogen ~~as~~ oxidizing agent

Up  $\rightarrow$  Down trend

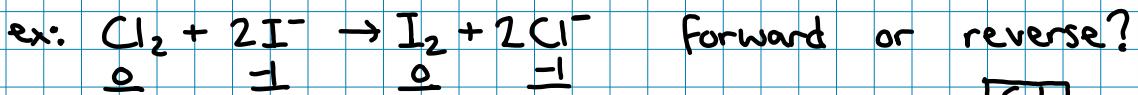
- Going down in the halogen group, halogens become less good oxidizing agents (or less easily reduced)

Theory: valence shell further from attractive force of nucleus as one goes down so less capable of picking up  $e^-$

Note: In reaction between halogen ( $X_2$ ) and halide ( $Y^-$ ),



- will go forward if halogen reactant ( $X_2$ ) is a better oxidizing agent than halogen product ( $Y_2$ )
- will go reverse if opposite



Forward b/c  $\text{Cl}_2$  is a better oxidizing agent (higher in table) b/c valence shell is closer to attractive force of nucleus

Cl
Br
I

### 4. Electronegativity

= ability of an atom to draw  $e^-$  toward itself within a compound (partial charge)

Left  $\rightarrow$  Right trend

- Atom becomes more electronegative

Theory: more protons  $\therefore$  better able to pull  $e^-$

Up  $\rightarrow$  Down trend

- Atom becomes less electronegative

Theory: as one moves down, valence shell is further away from nucleus  $\therefore$  harder to pull  $e^-$

Exceptions:

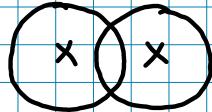
- Group VIII A (noble gases) are exception to left  $\rightarrow$  right increase

Noble gases are not very electro negative due to their full valence shell.

## 5. Atomic Radius

=  $\frac{1}{2}$  the distance between the nuclei of 2 identical atoms

Units:  $\text{\AA} = 10^{-10} \text{ m}$  (Angstroms)



Right  $\rightarrow$  Left trend

- Atomic radius decreases

Theory: more protons  $\therefore$  pulls electrons more tightly  
 $\therefore$  smaller radius

Up  $\rightarrow$  Down trend

- Atomic radius of atoms and ionic radius of identically charged ions increases

Theory: add shells moving down a group  $\therefore$  size of atom increases.

## Ion sizes

$\rightarrow$  As radius of an ion gets larger, the size gets larger

- Cation (+) always smaller than its parent (atom) (neutral)
- Anion (-) always larger than its parent atom
- $\rightarrow$  with higher proton/electron ratio, electrons held more tightly (smaller size)

## Isoelectronic Ions (same number of $e^-$ )

$\rightarrow$  The isoelectronic ion with more protons pulls  $e^-$  more tightly  $\therefore$  smaller radius.

Ion	# $e^-$	# protons	
$\text{Sn}^{2-}$	18	16	largest
$\text{Cl}^-$	18	17	
$\text{K}^+$	18	19	
$\text{Ca}^{2+}$	18	20	smallest

These ions have "noble gas configuration" or "complete octet"

$[\text{Ar}] 1s^2 2s^2 2p^6 3s^2 3p^6$  (surrounded by 8 val.  $e^-$ )

- Metals don't typically lose  $e^-$  beyond noble gas config.
- Non-metals don't typically gain  $e^-$  beyond noble gas config.
- \* soon try to have  $e^-$  config. of noble gas

## 6. Density (Up $\rightarrow$ Down):

- mass ~~increases~~ increases more rapidly than atomic size going down  
 $\therefore$  density increases. ex: Br more dense than Cl

# Bonding

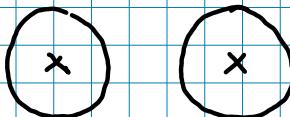
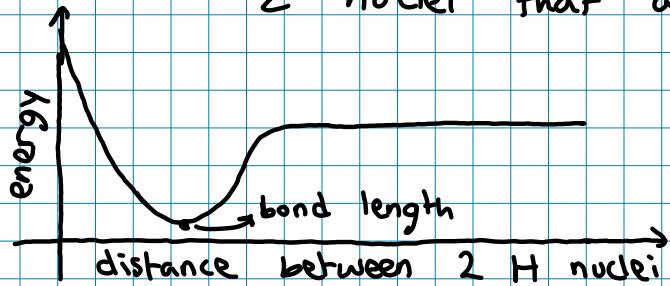
→ Forces that hold atoms in close proximity in compounds or polyatomic ions.

Ionic bond = electrostatic attraction between oppositely charged ions

- Write cation next to anion :  $\text{Na}^+ \text{Cl}^-$
- ✗ Do NOT represent ionic bond by a line (line = covalent bond) : ~~Na-Cl~~

Covalent bond = a bond where electrons are shared between 2 nuclei  
(holds non-metals together in molecules)

Bond length = most stable (low energy) distance between 2 nuclei that are adjacent in compound.



- energy rises when we approach zero distance b/c nuclei repel each other
- energy rises to the right b/c takes energy to break bond
- ✗ The electrons of one H atom are attracted to the proton of a second H atom ∴ ideal (stable / low energy) distance between 2 nuclei.

## Polarity of bonds

- Pauling numbers (Fig 8.3) describe relative electronegativity  
→ Greater Pauling number = more electronegative the atom

Note: • F is the most electronegative atom, followed by O.

- Non-metals are more electronegative than metals

Bond Ionic in character:

"Difference between Pauling numbers for 2 atoms"  $\gg 2$

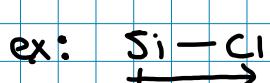
Covalent bond:

"Difference between Pauling numbers for 2 atoms"  $< 2$

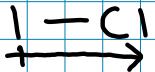
- \* Even in covalent bonds between non-identical atoms, there is a difference in electronegativities
- ∴ unequal share of electrons
- ∴ polar covalent bond

The greater the difference in electronegativities (Pauling num.), the greater the polarity

More electronegative atom:  $\delta^-$



Less electronegative atom:  $\delta^+$



ex: which bond is more polar: C-N or C-O?  
(Pauling numbers: C = 2.5, N = 3.0, O = 3.5)

$$\text{C}-\text{N}: 3.0 - 2.5 = 0.5$$

$$\text{C}-\text{O}: 3.5 - 2.5 = 1.0 \quad \therefore \text{C}-\text{O} \text{ is more polar (greater difference)}$$

Not all covalent bonds are polar (if difference = 0)

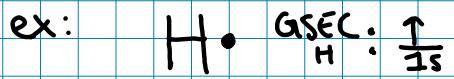
→ Bond between 2 identical atoms

→ Rare 2 atoms with same Paulin number such as P-H (Pauling num.: P = 2.1, H = 2.1)

\* C-H is thought as non-polar

## Lewis dot symbol

→ Nucleus and core electrons = symbol for atom  
→ Electrons on valence shell = dot •



\* Note: Hydrogen only needs 2 e<sup>-</sup> to complete its valence shell



- bonding pairs (shared) : pair of e<sup>-</sup> shared between 2 atoms.  
Represent by line: —

- lone pairs (nonbonding) : pair of e<sup>-</sup> that are attached to only one atom

Duet rule: H follows duet rule to always be surrounded by 2 e<sup>-</sup> in molecules or ions

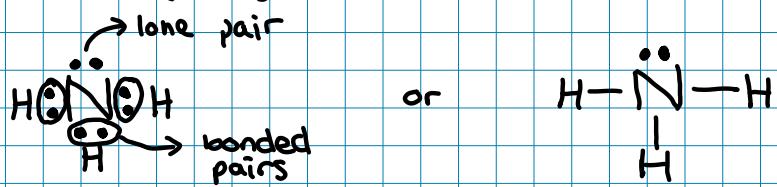
Octet rule: Atoms are surrounded by 8 val. e<sup>-</sup> in molecules

→ Second period non-metals, carbon, nitrogen oxygen, fluorine follow octet

→ Terminal atoms follow octet (atoms that are attached to only 1 other atom)

\* Atoms tend to combine to complete their valence shell

ex: Ammonia (NH<sub>3</sub>)



## Localized Electron theory (L.E): Lewis Structures

- For molecules like HF or NH<sub>3</sub>, drawing Lewis structures is pretty easy
- For many simple ions or molecules, one must follow certain steps to be able to draw a structure

### ① Choose the central atom

- often atom with lowest electronegativity
- never H
- atom that is attached to 2 or more other atoms  
note: if only 2 atoms, none are central
- \* Some molecules have several central atoms

### ② Add up total number of electrons

- If ion has negative charge: add val e<sup>-</sup>
- If ion has positive charge: remove val e<sup>-</sup>

### ③ Place one bonding pair (represented by a line) between each bonded atoms

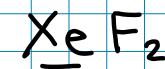
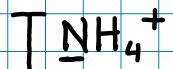
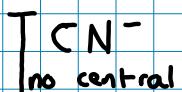
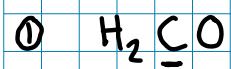
### ④ Use remaining val e<sup>-</sup> as lone pairs around each terminal (non-central) atom

- If possible, have 8 val e<sup>-</sup> around each terminal atom (other than H).

### ⑤ 3 possibilities:

- A)
  - o All atoms (terminal and central) are surrounded by 8 e<sup>-</sup>
  - o All val e<sup>-</sup> are used
- B)
  - o All val e<sup>-</sup> are used up
  - o But not all atoms obey octet rule
  - More lone pair from one atom to bonded pair
  - \* Exceptions:
    - B, Be and certain central atoms exceed octet rule
    - Atoms row 3 and heavier can exceed octet b/c they have space in d orbital
- C)
  - o Octet/Duet rule is satisfied
  - o But not all val e<sup>-</sup> are used
  - Place remaining e<sup>-</sup> on central atom (lone pairs)

ex: formaldehyde, cyanide, ammonium, xenon difluoride  
 underline central atom



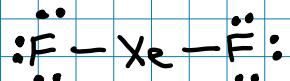
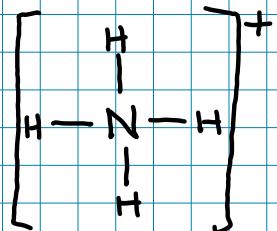
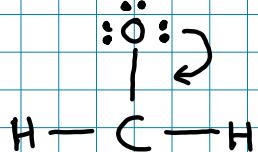
$$\begin{aligned} \textcircled{2} \quad & (\frac{2}{\text{H}} \times 1e^-) + (\frac{1}{\text{C}} \times 4e^-) + (\frac{1}{\text{O}} \times 6e^-) \\ & = 12 \text{ val. } e^- \end{aligned}$$

$$\begin{aligned} & (\frac{1}{\text{C}} \times 4) + (\frac{1}{\text{N}} \times 5) + \underline{\text{1 } e^-} \\ & = 10 \text{ val. } e^- \end{aligned}$$

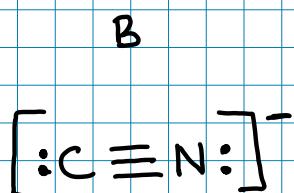
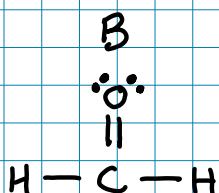
$$\begin{aligned} & (\frac{1}{\text{N}} \times 5) + (\frac{4}{\text{H}} \times 1) - \underline{\text{1 } e^-} \\ & = 8 \text{ val. } e^- \end{aligned}$$

$$\begin{aligned} & (\frac{1}{\text{Xe}} \times 8) + (\frac{2}{\text{F}} \times 7) \\ & = 22 \text{ val. } e^- \end{aligned}$$

③  
④



⑤



Notes:

- curved arrow

pair  $e^-$   
 start                  end

- $\text{Xe}$  can use a d orbital to be surrounded by more than 8 val.  $e^-$

$$\text{Val. } e^- : \frac{1\downarrow}{4s} \frac{1\uparrow}{4p} \frac{1\downarrow}{4p} \frac{1\uparrow}{4p} \frac{1\downarrow}{4p}$$

Single bond = covalent bond with  
 Double bond = covalent bond with  
 Triple bond = covalent bond with

$\frac{2}{4} e^-$   
 $\frac{4}{4} e^-$   
 $\frac{6}{4} e^-$

=  
 =  
 ≡

\* Never exceed triple bond

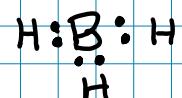
For larger molecules:

→ When atom X is neutral (zero formal charge)  
 → surrounded by Y bonded and Z lone pairs

X (obey octet)	Y (bonded)	Z (lone)	
C	4	0	-C-      =C=        C
N	3	1	-N-      :N≡      -N=
O (other VI A)	2	2	:O:      :O=
F (other halogen)	1	3	:F: —

Exceptions to octet

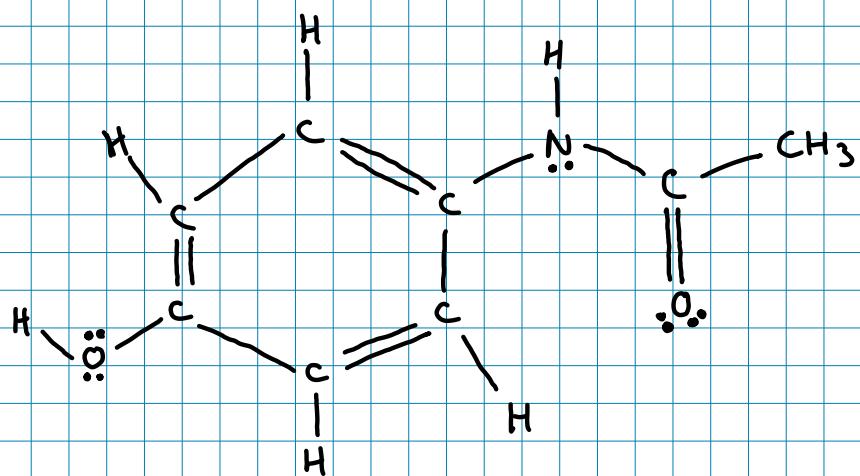
- Be neutral (in compounds) surrounded by 4 val. e<sup>-</sup>
- B neutral (in compounds) surrounded by 6 val e<sup>-</sup>



Note: just because these molecules are neutral, they are still very reactive (space for e<sup>-</sup>)

complete skeletal structure:

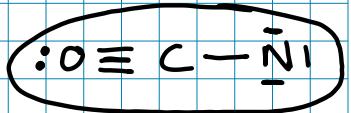
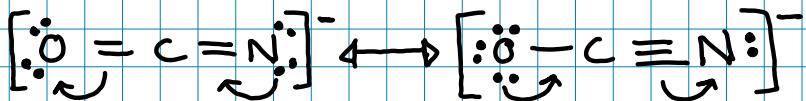
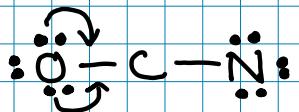
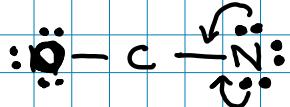
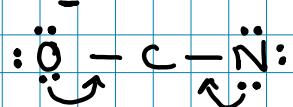
- Add lone pairs and bonded pairs
- Add no new single bonds



## Resonance

- sometimes more than one Lewis structure is valid
  - occurs when one has a choice of where to put a multiple bond

ex:  $\text{OCN}^-$  (16 val. e<sup>-</sup>)



Bad F.C.

\* Arrow shows how one goes from one valid Lewis structure to another

→ show how electrons spread to some extent (delocalize)

Resonance = concept explaining how the actual structure of a molecule/ion is a hybrid of all valid Lewis structures

Double headed arrow: placed between valid structures

Hybrid = composite = supercomposition

Each valid Lewis structure (if more than one) is a resonance structure/contributing structure.  
→ each of these contribute to the resonance hybrid

Resonance hybrid = composite of all resonance structures.

\* The hybrid is the actual structure

Note: single bond ( $A-B$ ) weaker and longer than double bond ( $A=B$ ) weaker and longer than triple bond ( $A\equiv B$ )

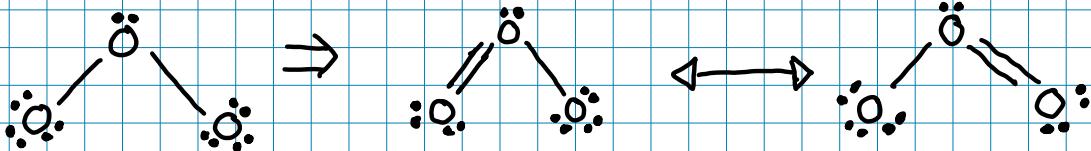
ex: Length and strength of bonds in  $\text{OCN}^-$

- C-O bond length and strength is between that of a C=O double bond and C≡O triple bond
- C-N bond length and strength is between that of a C=N double bond and C≡N triple bond

\* Electrons are said to be delocalized (curved arrow), which explains how bond is intermediate or hybrid

For valid resonance structures, electrons are located on or between different atoms ON PAPER.

Ex: Draw  $O_3$  resonance structures and describe bond different on paper



$\therefore$  O-O bond length and strength between that of single and double bond.

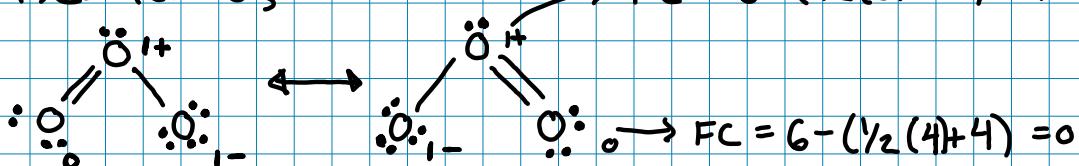
## Formal Charge (F.C.)

→ comparing number of val e<sup>-</sup> of an atom in molecule vs as neutral atom on its own

$$F.C. = \frac{\text{number of val } e^- \text{ surrounding atom when neutral}}{2} - \left( \frac{1}{2} \left( \begin{array}{l} \text{number of shared } e^- \\ \text{around given atom} \end{array} \right) + \left( \begin{array}{l} \text{number of unshared } e^- \\ \text{around atom} \end{array} \right) \right)$$

- Notes:
- Formal charge of all atoms should add up to overall charge of ion or compound.
  - Formal charge shown near atom

ex. F.C. for  $O_3$



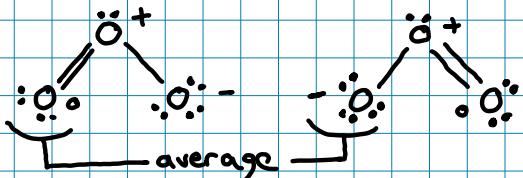
\* Equivalent resonance structures:

- every atom X in one resonance structure has a certain F.C. and is surrounded by some bond combo

→ There exist atom X in 2nd resonance structure with same F.C. and bond combo

If one has equivalent resonance structures, the resonance hybrid (actual charge) is an average of all equivalent structures.

ex: F.C. of left oxygen in  $O_3$



$$F.C. = \frac{O\text{ F.C.} + 1^-\text{ F.C.}}{2} = \frac{1}{2}^-$$

hybrid  
(left O)

If structures aren't equivalent, there are:

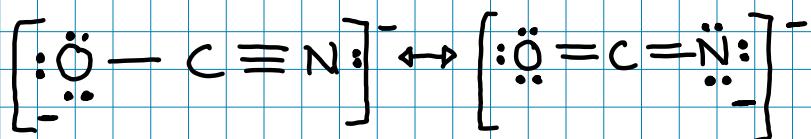
- Major resonance structure: contributing more to hybrid, preferred
- Minor resonance structure: contributes to hybrid but less than the "major"

A structure contributes more if:

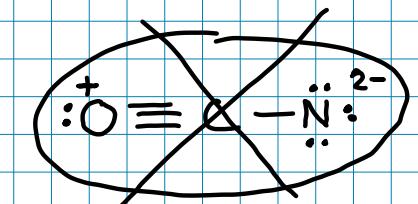
1. Resonance structure minimizes F.C. (close to zero)
2. Negative value for formal charge is on more electronegative atom

Note: If F.C. is not minimized and negative formal charge is on the less electronegative atom (pos. FC on most electro.)  $\therefore$  not valid structure.

ex:  $OCN^-$



Major b/c more electronegative oxygen holds negative F.C. better

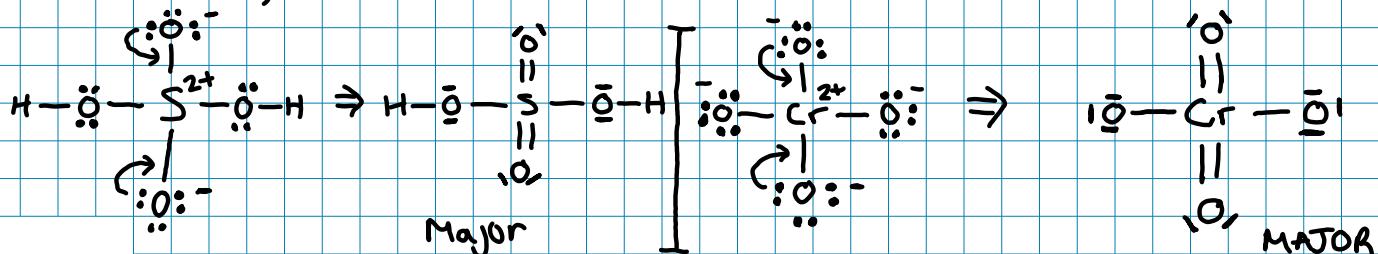


Bad b/c oxygen is more electronegative

Charge N: between O and  $1^-$   
but closer to O

Charge O: between  $1^-$  and 0  
but closer to  $1^-$

Minimizing F.C. (note: S and Cr can exceed octet)



# VSEPR Theory / Orbital Hybridization

(Localized electron model)

VSEPR = Valence Shell Electron Pair Repulsion

VSEPR theory and molecular geometry/bond angles:

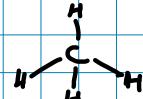
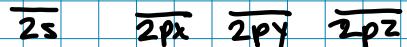
- Minimizing the electron pair repulsion of  $e^-$  pairs dictates bond angles and geometry

ex:  $\text{BeH}_2 \Rightarrow \text{H}-\text{Be}-\text{H}$  electron pairs want to be as far away from each other

Hybrid orbitals describe where pairs of electrons exist around an atom when part of a molecule.

Remember: Atomic Orbital = area in which  $e^-$  have high probability of existing in an ATOM (NOT in molecule)

ex: valence shell C :



However, we can't say that in a molecule like  $\text{CH}_4$  the 8 electrons surrounding C are in  $2s, 2px, 2py, 2pz$ .

It doesn't make sense b/c  $2s$  orbital is lower energy than  $2p$ . It doesn't make sense that one C-H bond is stronger than another.

Instead, there is another theory to explain how  $e^-$  exist around an atom in a molecule.

## Hybrid Orbitals

- pairs of  $e^-$  that surround an atom in a molecule exist in hybrid orbitals
- hybrid orbitals of the same name  $\Rightarrow$  equal energy

- Valence s, p and d orbitals combine to create hybrid orbitals
- The number of orbitals that combine is equal to the number of hybrid orbitals produced

\* Each hybrid orbital can hold 2  $e^-$

Region of electron density : - lone pair  
 (effective  $e^-$  pair) - single bond  
 - multiple bond (1 region)

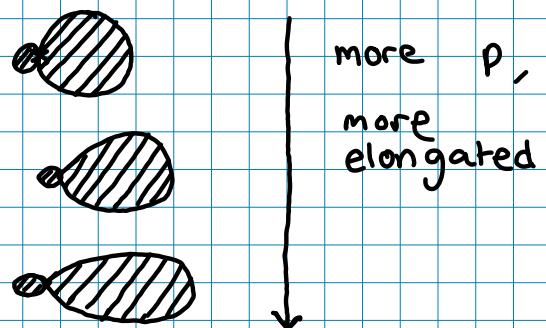
Hybrid orbitals are given names based upon the number of regions of  $e^-$  density surrounding a given central atom.

Hybridization of central atom (name of hybrid orbitals)	Number of regions of $e^-$ density	example: A = central atom X = region $e^-$ density
SP	1+1	2
$SP^2$	1+2	3
$SP^3$	1+3	4
$SP^3d$	1+3+1	5
$SP^3d^2$	1+3+2	6

From theory, Hydrogen, (H), is surrounded by an atomic orbital ('1s') rather than a hybrid orbital when part of a molecule.

Orbital drawing:

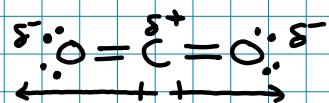
- SP (hybrid of s/p) ∴
- $SP^2$  (more p like) ∴
- $SP^3$  (even more p like) ∴



## Polarity of molecules

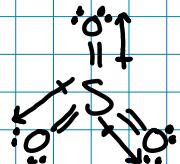
- A bond is polar based on electronegativity (difference in Pauling numbers) of 2 atoms in a bond
    - Molecule is polar if e<sup>-</sup> density is shifted more to one side.
    - Molecule is non-polar if dipole moment arrow vectors cancel each other out
      - \* If at the end of all hybrid orbitals surrounding the central atom, there's the same atom group

ex:  $\text{CO}_2$



non-polar b/c vectors cancel out.

ex.:  $\text{SO}_3$



 non-polar b/c dipole moment  
arrow vectors cancel out

## SP orbitals

$\rightarrow 1 \text{ valence } s + 1 \text{ valence } p \Rightarrow 2 \text{ sp hybrid orbitals}$

Geometry : Linear

ex:	Lewis	Bond pairs	Lone pairs	Arrangement e <sup>-</sup> pairs	molecular geometry	3D sketch	Polar
BeCl <sub>2</sub>		2	0	linear	linear		No

Note: Cl +  
||  
Be 2-

||  
Cl +

Bad b/c Cl more electronegative

Cl-Be-Cl  $\angle = 180^\circ$   
(bond angle)

look at e<sup>-</sup> pairs      look at atoms      no lone pairs

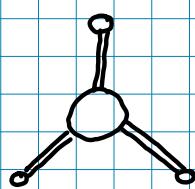
## SP<sup>2</sup> orbitals

$\rightarrow 1 \text{ valence } s + 2 \text{ valence } p \Rightarrow 3 \text{ sp}^2 \text{ hybrid orbitals}$

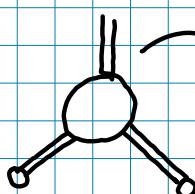
Geometry : Trigonal Planar

ex:	Lewis	Bond pairs	Lone pairs	Arrangement e <sup>-</sup> pairs	Molecular geometry	3D sketch	Polar
BF <sub>3</sub>		3	0	trigonal planar	trigonal planar		No
SnCl <sub>2</sub>		2	1		V-shaped bent angular		Yes

BF<sub>3</sub>



SnCl<sub>2</sub>



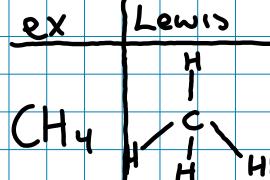
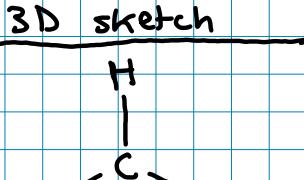
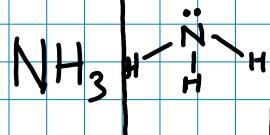
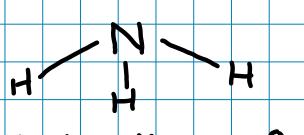
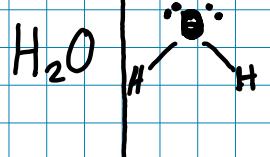
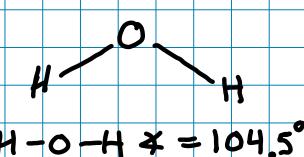
lone pair counts  
as a hybrid orbital

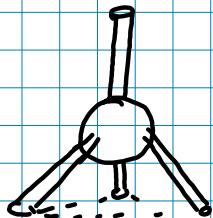
## $sp^3$ orbitals

$\rightarrow 1 \text{ valence } s + 3 \text{ valence } p \Rightarrow 4 \text{ } sp^3 \text{ hybrid orbitals}$

Geometry: Tetrahedral

Note: When a central atom in a molecule is surrounded by lone pairs and bonded pairs, each lone pair causes a bond compression of  $2.5^\circ$ .  
 \* Lone pairs repel more

ex	Lewis	Bond pairs	Lone pairs	Arrangement e <sup>-</sup> pairs	molecular geometry	3D sketch	Polar
$CH_4$		4	0	Tetrahedral	Tetrahedral	 $H-C-H \approx 109.5^\circ$	No
$NH_3$		3	1		Trigonal pyramidal	 $H-N-H \approx 107^\circ$	Yes
$H_2O$		2	2		V-shaped bent angular	 $H-O-H \approx 104.5^\circ$	Yes



## $sp^3d$ orbitals

→ 1 valence s + 3 valence p + 1 valence d = 5 hybridized  $sp^3d$  orbitals

Geometry: Trigonal Bipyramidal

Note: • equatorial : e<sup>-</sup> pairs in trigonal bipyramidal that are  $120^\circ$  apart.  
("equator" - think circle around earth)

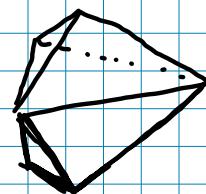
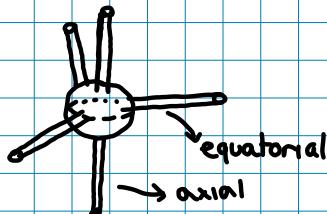
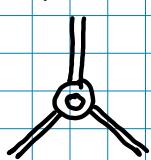
• axial : e<sup>-</sup> pairs in trigonal bipyramidal that are  $180^\circ$  apart  
("axe" - think  $180^\circ$  axis)

\* Lone pairs go into an equatorial place to minimize e<sup>-</sup>/e<sup>-</sup> repulsion (VSEPR)  
→ lone pairs repel more

- in axial :  $90^\circ$  relative to 3 e<sup>-</sup> pairs
- in equatorial :  $90^\circ$  relative to 2 e<sup>-</sup> pairs

ex:	Lewis	Bond Pairs	Lone Pairs	Arrangement e <sup>-</sup> pairs	molecular geometry	3D sketch	Polar
PCl <sub>5</sub>		5	0	Trigonal bipyramidal	Trigonal bipyramidal		No
SF <sub>4</sub>		4	1	See-saw of lone pairs		$\angle S-F-S \approx 120^\circ, 90^\circ$ $F-S-F \approx 120^\circ, 90^\circ$	Yes
BrF <sub>3</sub>		3	2	T-shaped		$F-Br-F \approx 90^\circ$	Yes
XeF <sub>2</sub>		2	3	Linear		$F-Xe-F = 180^\circ$	No

Top view



2 pyramids superposed

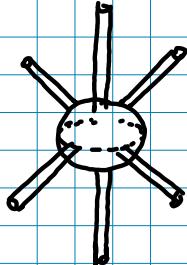
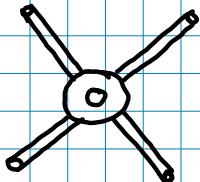
## $sp^3d^2$ orbitals

→ 1 valence s + 3 valence p + 2 valence d = 6  $sp^3d^2$  hybrid orbitals

Geometry: octahedral

ex:	Lewis	Bond Pairs	Lone Pairs	Arrangement $e^-$ pairs	molecular geometry	3D sketch	Polar...
$SF_6$		6	0	Octahedral	Octahedral		No
$IF_5$		5	1		Square pyramid		Yes
$XeF_4$		4	2		Square planar		Nb

Top View



Number of e <sup>-</sup> pair (central atom)	Bond Pairs	Lone Pairs	Arrangement of e <sup>-</sup> pairs	Molecular Geometry	$\angle$
2	2	0	Linear	Linear	180°
3	3	0	Trigonal planar	Trigonal planar	120°
	2	1		V-shaped	
4	4	0	Tetrahedral	Tetrahedral	
	3	1		Trigonal pyramid	
	2	2		V - shaped	109.5°
5	5	0		Trigonal bipyramidal	
	4	1	Trigonal Bipyramidal	see-saw	120°
	3	2		T-shaped	and 90°
	2	3		Linear	
6	6	0		Octahedral	
	5	1	Octahedral	Square pyramid	90°
	4	2		Square planar	

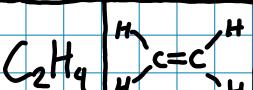
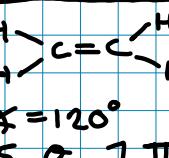
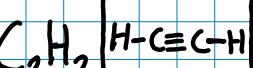
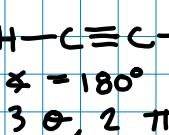
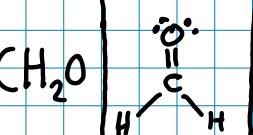
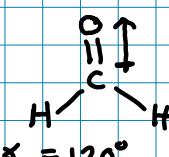
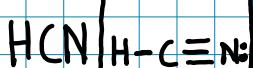
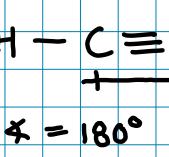
\*Note: VSEPR theory is a brilliant theory but it is not infallible.  
 For example, experimental data shows that bond angles in  $\text{PH}_3$  are closer to 94° than 107°.



However, we are predicting according to theory.

## Multiple bonds (VSEPR)

Around 1 central C

ex:	Lewis	Bond Pairs	Lone Pairs	Regions of density	Arrangement/Geometry	3D sketch	Polar
$C_2H_4$		4	0	$\therefore 3$ $SP^2$	Trigonal planar		No
$C_2H_2$		4	0	$\therefore 2$ $SP$	Linear		No
$CH_2O$		4	0	$\therefore 3$ $SP^2$	Trigonal planar		Yes
$HCN$		4	0	$\therefore 2$ $SP$	Linear		Yes

\* Note:

Every bond has  
1  $\sigma$  bond

In multiple bond, every  
bond that isn't  
 $\sigma$ , is a  $\pi$  bond

# Bonding in Multiple Bonds

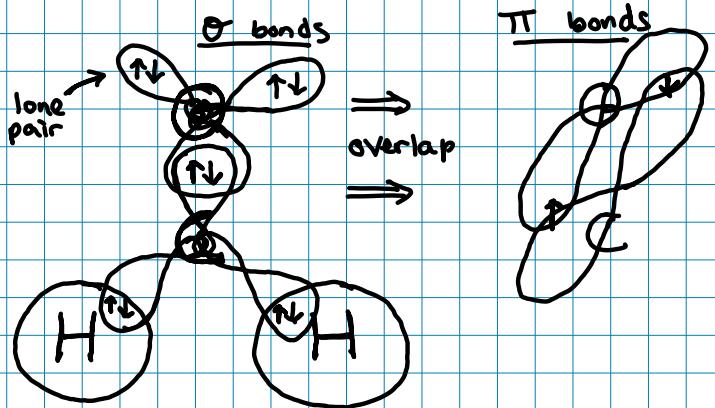
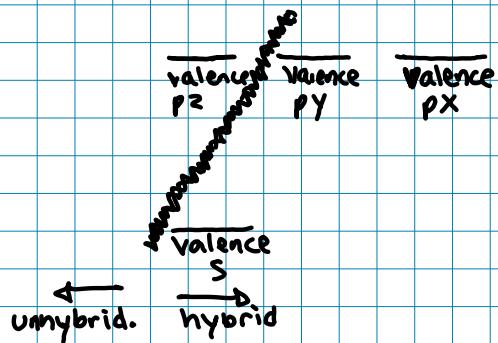
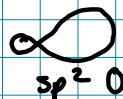
→ For predicting bond angles, assume multiple bonds repel the same as single bonds

- Sigma bond ( $\sigma$ ): The overlap of 2 hybrid orbitals (or  $1s$  for H)
  - bond located between 2 nuclei
  - The  $\sigma$  bond holds only  $2 e^-$   
(So other  $e^-$  in double or triple bond go in  $\pi$  bond)
- Pi bond ( $\pi$ ): The overlap of 2 unhybridized orbitals
  - can hold only  $2 e^-$
  - second bond of double bond
  - second/third bond of triple bond

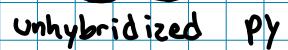
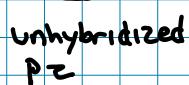
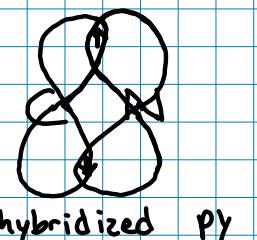
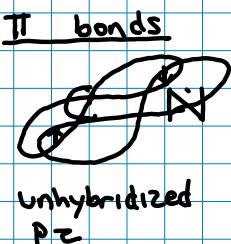
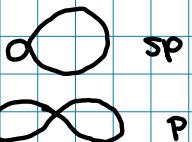
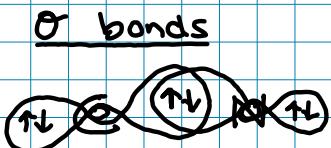
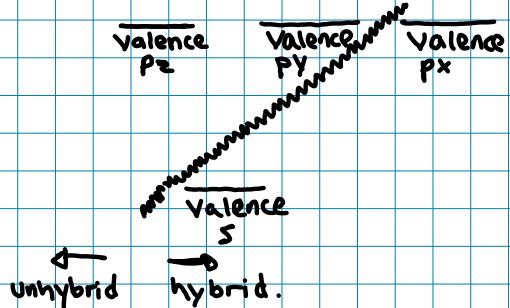
Summary:

single bond	$\Rightarrow$	1 $\sigma$ bond	,	0 $\pi$ bonds
double bond	$\Rightarrow$	1 $\sigma$ bond	,	1 $\pi$ bond
triple bond	$\Rightarrow$	1 $\sigma$ bond	,	2 $\pi$ bonds

ex:  $\text{CH}_2\text{O}$



ex:  $\text{CN}^-$

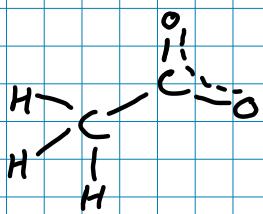
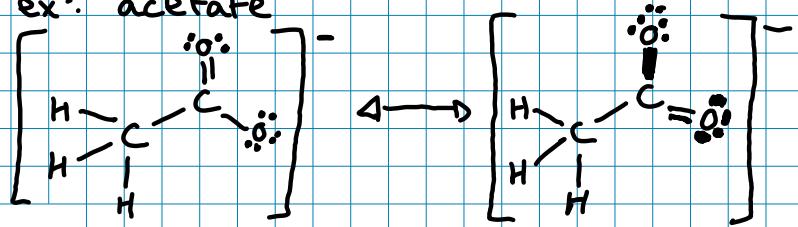


When  $\pi$  bond appears in different places on different valid resonance structures, there is delocalized  $\pi$  bonding.  $\pi$  bond only exists in double bonds or triple bonds.

Delocalize according to dotted curve: , / - \ ,

Delocalized  $\pi$  bonding  
Resonance hybrid

ex: acetate



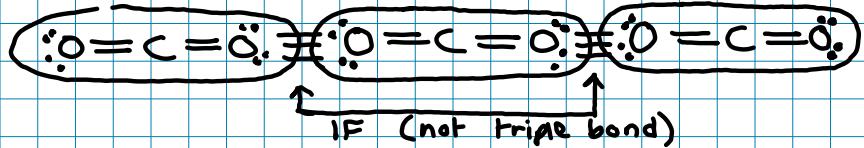
# Interparticle Forces (IF)

Condensed states of matter: solid and liquid

Interparticle forces: hold particles together in condensed states of matter

Note: Bond  $\neq$  IFs

ex: IFs between  $\text{CO}_2$  particles



Liquid vs solid: solid particles have less freedom of motion than liquid particles relative to other particles. Solid particles stay in fixed positions relative to other particles.

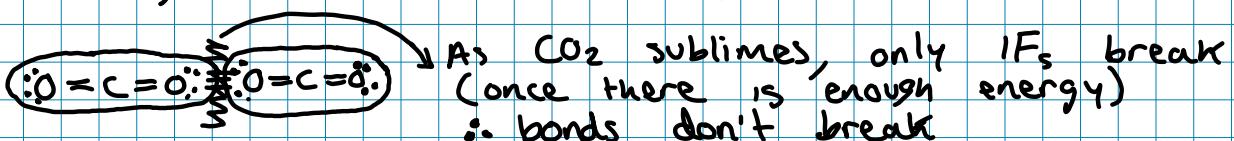
- solid particles: vibrational motion
- liquid particles: translational motion can tumble past other liquid particles

Phase transition

\* At 1 atm, we say "normal X point"

↑	Solid $\rightarrow$ Liquid (melt)	melting point
T	Solid $\rightarrow$ Gas (sublime)	sublimation point
	Liquid $\rightarrow$ Gas (evaporation)	boiling point
↓	Gas $\rightarrow$ Liquid (condensation)	boiling point
T	Liquid $\rightarrow$ Solid (freeze)	freezing point (=melting)
	Gas $\rightarrow$ Solid (deposition)	sublimation point

ex: during sublimation of  $\text{CO}_2$ , do bonds break?



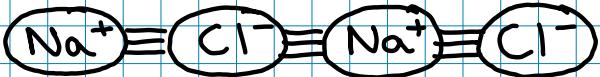
\* Can use IFs to predict relative boiling/melting points

## ① Interionic interaction (ionic bond)

- strongest interparticle force ( $\therefore$  demands more energy to break).

- High melting and boiling points

IF = attraction of cation (+) and anion (-)



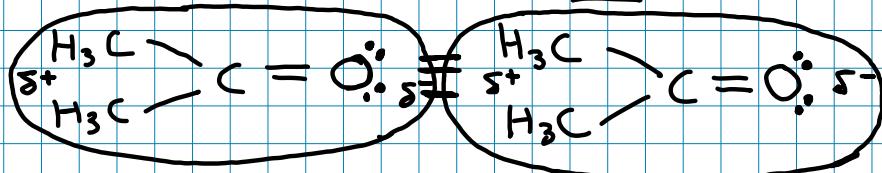
\* In this case, break IF  $\therefore$  break ionic bond

Following IFs useful for compounds with covalent bonds

## ② Dipole / Dipole attraction

- Polar molecules are predicted to have higher boiling points as similarly sized nonpolar molecules (think elements of a same row have the same size and it doesn't affect size).

IF = interaction between  $\delta^+$  side of a molecule with  $\delta^-$  side of a second molecule of the same kind. Must be polar.



ex: What has higher boiling point?



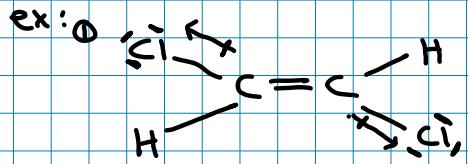
polar



non-polar (only C-C)

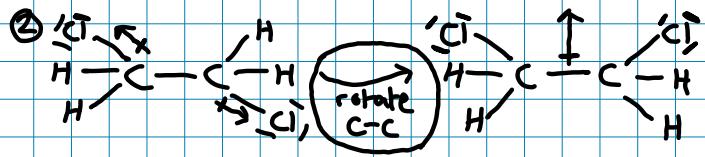
Similar size: 4 second row elements

$\therefore$  ① has a higher boiling point.



non-polar (vectors cancel)

CAN'T rotate double bond (break pi bond)



polar  $\therefore$  Higher boiling point

CAN rotate (single bond)

\* Stronger IF = higher BP (only if similar size)

### ③ H-bonding (strongest dipole-dipole)

- HO, HF, HN covalent bonds
- Causes strong dipole/dipole IF

• Normal boiling points of H-containing compounds:

$\boxed{\text{NH}_3 (-33)}$

$\text{PH}_3 (-88)$

$\text{AsH}_3 (-63)$

$\text{SbH}_3 (-17)$

$\boxed{\text{H}_2\text{O} (100)}$

$\text{H}_2\text{S} (-60)$

$\text{H}_2\text{Se} (-42)$

$\text{H}_2\text{Te} (-2)$

$\boxed{\text{HF} (19.5)}$

$\text{HCl} (-85)$

$\text{HBr} (-66)$

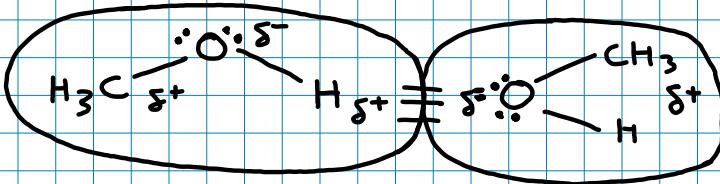
$\text{HI} (-35)$

General Pattern: As the non-hydrogen atom size increases, going down a column, the BP also increases.

\* Exception: O-H, N-H, F-H bonds (Highest BP)

IF = H-bonding occurs between 2 molecules when H is covalently bonded to O, F or N (small highly electronegative atoms).

It is the strongest dipole/dipole interaction ( $\Rightarrow$  highest BP)



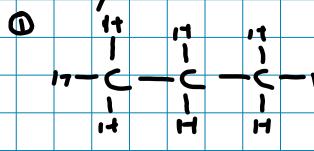
\* H becomes only a proton with no e-  $\Rightarrow$  extremely  $\oplus$

powerful attraction  
 $\because$  e- density pulled away from H attached to highly electronegative O, N, F  $\therefore$  exposed  $\oplus$  charge on H highly attracted to lone-pair.

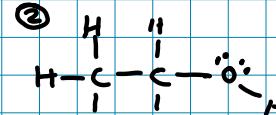
Note:

- H bonding stronger between similarly sized molecules with OH bonds rather than NH or NH<sub>2</sub> (O more electronegative than N  $\therefore$  more positive H)
- The more "OH groups" (or NH or NH<sub>2</sub>), the more opportunities for H-bonding  $\therefore$  stronger IF

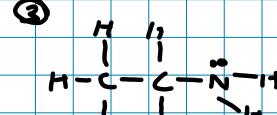
ex: Highest to lowest BP



non-polar (C-H)



H-bonding (OH)



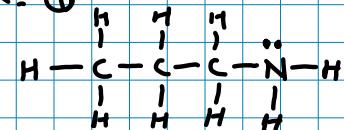
H-bonding (NH)

| similar size

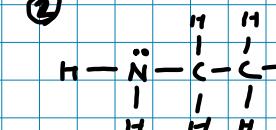
② > ③ > ①

OH stronger than NH

ex: ①



②



② > ①

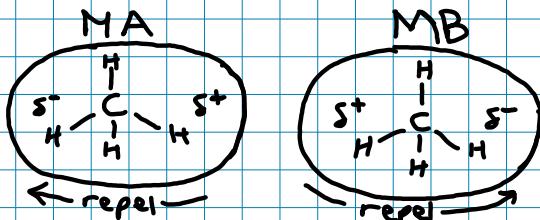
b/c more H-N covalent bonds  $\therefore$  more opportunities for H-bonding

## ④ London Dispersion Forces (LD)

London Dispersion Forces (LDF)  
- (Van der Waals, Induced, Instantaneous, Temporary)

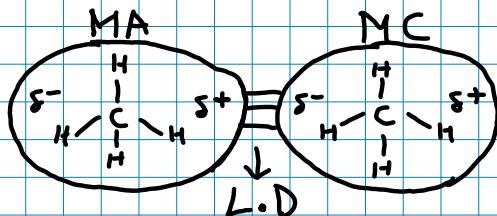
- Explains how noble gases and non-polar molecules can exist in Condensed state of matter (but, weaker forces so lower BP  $\therefore$  commonly gas)

IF = 2 non-polar molecules approach each other  $\Rightarrow$  temporary dipole  
Sample non-polar molecules



Molecules approaching  
 $\therefore$  electrons A repel  
electrons B  
 $\therefore$  Temporary / induced dipole

The induced dipole in MA can now be attracted by a MC and create an induced dipole:



\* All molecules have e<sup>-</sup> ∴ all molecules have L.D forces

Polar molecules of similar size as non-polar molecules have the same LD. However they also have dipole/dipole.

More shells with  $e^-$ 's than an atom has ~~>>>~~  
 $\therefore$  greater likelihood of  $e^-/e^-$  repulsion (causing LD)  
 $\Rightarrow$  Larger atoms within same column, has  $e^-$  held less tightly  $\therefore$  more polarizable' (possibility of creating LD)  
 $\Rightarrow$  Presence of larger atom (column  $\downarrow$ )  $\therefore$  stronger LD.

ex:  $|\bar{B}_r - \bar{B}_{r'}|$

$$\text{--} \bar{\in} \text{---} \bar{\in} \text{---}$$

Cl  
Br  
I

$\therefore \text{Br}_2$  has higher BP b/c larger  $e^-$  shells  
 $\therefore$  more

ex.: ① 

② 

similar size ( same LD forces)  
but ② has dipole/d.pole b/c polar  
∴ ② is higher BP

\* Reminder: larger  $VdW$  "a" constant  $\Rightarrow$  stronger forces

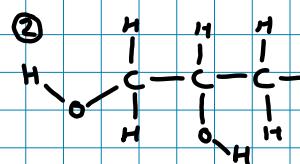
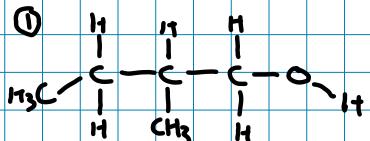
IFs can be used to predict properties:

- Viscosity = resistance of a liquid to flow  
(ex: honey is more viscous than water)
- Surface tension = resistance of a liquid to increase surface area
- Volatility = tendency of a liquid / solid to escape to become gas
  - \* If liquid A ~~is~~ is more viscous than liquid B, A has higher vapour pressure  $\Rightarrow$  b/c more tendency to become gas

A substance  $\otimes$  has stronger IFs than  $\circlearrowleft$

- More viscous :  $\otimes$  (strong interaction  $\therefore$  resist flow)
- Greater surface tension :  $\otimes$
- More volatile :  $\circlearrowleft$  (weaker IF  $\therefore$  easier to escape)  
 $\Rightarrow$  Higher vapour pressure :  $\circlearrowleft$

ex: Which is more viscous



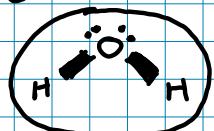
Similar size

②  $\Rightarrow$  more OH bonds  $\therefore$  more opportunities H-bonding  $\therefore$  stronger IF  $\therefore$  More viscous

ex: Which has highest to lowest VdW "a" constant?

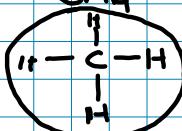
\* Reminder: greater "a" constant = greater attraction between particles

H<sub>2</sub>O



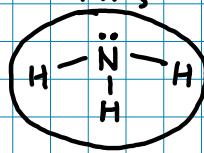
most powerful  
H-bonding (OH)

CH<sub>4</sub>



non-polar  
(only LD)

NH<sub>3</sub>



H-bonding  
(NH)

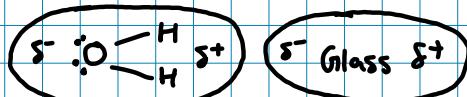
$\therefore \text{H}_2\text{O} > \text{NH}_3 > \text{CH}_4$  in terms of "a" constant

- Capillary Action = spontaneous rising of liquid in narrow tube

Cohesive force: IF between molecules of a liquid

Adhesive force: IF between molecules of a liquid and their container.

Theory: Water meniscus in glass tube is concave  
 $\Rightarrow$  considerable attraction between polar  $H_2O$  and polar glass (adhesive) causing  $H_2O$  to move up the tube



Theory: Mercury's meniscus in glass tube is convex  
 $\Rightarrow$  non-polar Hg is not attracted to  $H_2O$

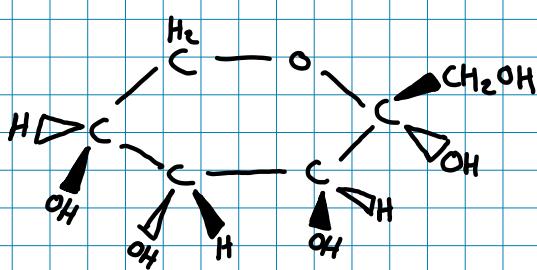


### Non-Volatile Substances

- $\rightarrow$  substances with very strong IF
- ionic compounds (salts)
- sugars
- very large molecules

Note: - IF you leave these substances out for 5 years, they will still be here and not have become gas  
\* Negligible vapour pressure

ex: Fructose



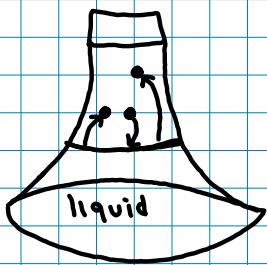
$\blacktriangleleft$  = bond coming out of page

$\blacktriangleright$  = bond pointing away from page

\* Many O-H bonds

\* Many L.D (many e- shells)  
 $\therefore$  prevents fructose from being volatile

## Vapour Pressures

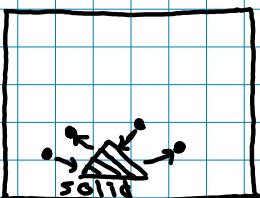


↑ = liquid particle evaporating  
↓ = gas particle condensing

- A liquid particle would most likely be at the surface to escape to become a gas b/c at the surface, there are no IF above the liquid (which is what is keeping it condensed)

- Gas particles may hit the liquid surface and be recaptured by IFs  $\Rightarrow$  re-condensation

When the "rate of vaporization" = "rate of condensation" a constant equilibrium vapour pressure is exerted above the liquid.



For solids, when rate of deposition = rate of sublimation

$\Rightarrow$  vapour pressure observed above solid

\* As temperature increases, the vapour pressure above a solid or liquid increases exponentially.

Boiling point: Temp at which "P<sub>vap</sub> over liquid" = "external pressure"

Sublimation point: Temp. at which "P<sub>vap</sub> over solid" = "external pressure"

Melting point: Temp at which "P<sub>vap</sub> of liquid" = "P<sub>vap</sub> of solid"  
(Note: of the same substance)

$\Rightarrow$  For "normal" point, consider external pressure = 1 atm

Note: At higher altitude  $\Rightarrow$  external pressure is lower  
 $\therefore$  demands less energy for P<sub>vap</sub> = external pressure  
 $\therefore$  BP is lower

Note: At temp < BP, gas bubble inside liquid gets crushed

Note: Melting point = temp at which the solid and liquid forms of a substance can coexist indefinitely.  
 $\rightarrow$  If P<sub>vap</sub> solid > P<sub>vap</sub> liquid, eventually all solid converted to liquid  
 $\rightarrow$  If P<sub>vap</sub> solid < P<sub>vap</sub> liquid, eventually all liquid converted to solid

# Heating/Cooling Curves

- Heating curve = temp vs time for a substance being heated at a constant rate.
- Cooling curve = temp vs time for a substance being cooled at a constant rate.



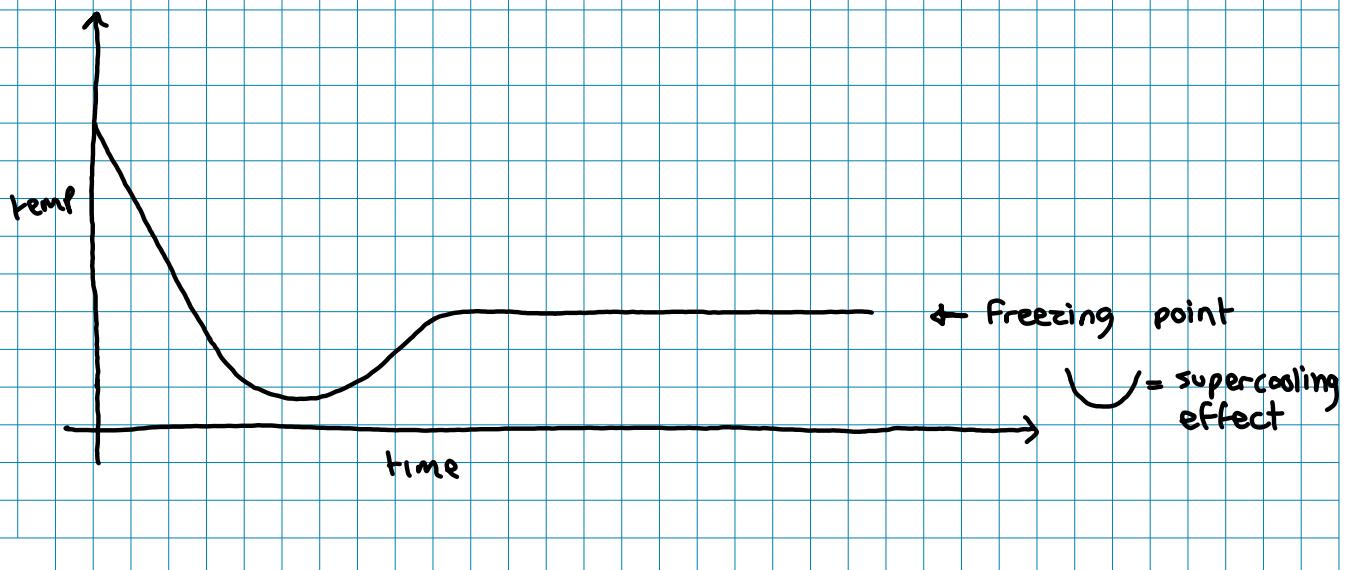
- -10 to 0  $\Rightarrow$  heat going towards raising temp. of ice
- At 0  $\Rightarrow$  energy goes towards breaking enough IFs to allow tumbling of molecules past one another (flow).
- 0 to 100  $\Rightarrow$  heat going towards raising temp. of water
- 100  $\Rightarrow$  remaining IFs are broken (liquid becomes gas)

Note: Flat line at 100°C longer than at 0°C b/c all IFs are broken at 100°C but only some at 0°C.

$\Delta H_{\text{fus}}$  (heat of fusion) = energy to melt a solid  
 $\Delta H_{\text{vap}}$  (heat of vaporization) = energy to vaporize a liquid.

- \* If heated too rapidly, a liquid may achieve temp  $>$  BP  
 $\Rightarrow$  superheated liquid
- \* If cooled rapidly, a liquid may achieve temp  $<$  MP  
 $\Rightarrow$  super-cooled liquid

Ex: cooling curve for tert-butyl alcohol



## Phase Diagram

- phase changes with only temp. and pressure

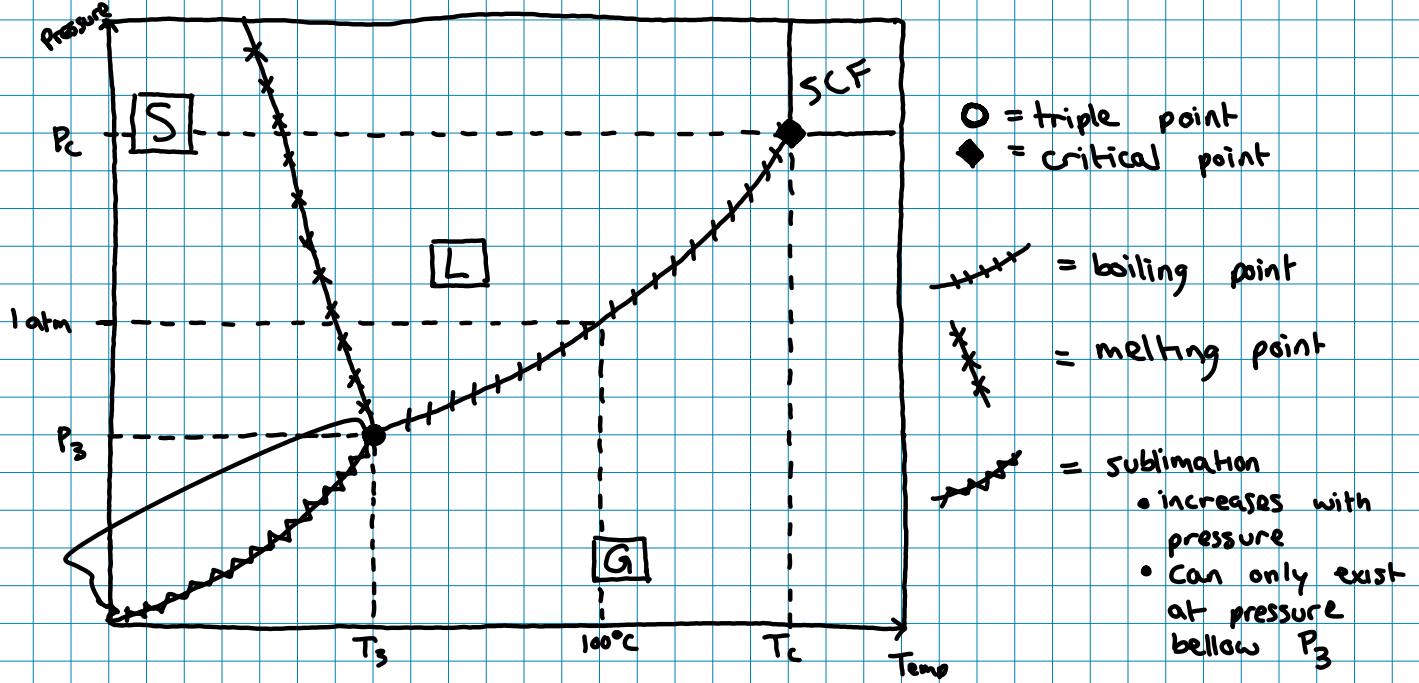
- Solid predominates at low temp, high pressure
- Liquid predominates at ~~high~~ avg. temp. and pressure
- Gas predominates at high temp, low pressure

Solid lines between regions = where 2 phases can coexist

Triple point = at  $(T_3, P_3)$  all 3 phases can coexist (S, L, G)  
 \* can also be where 2 crystal solids and liquid coexist

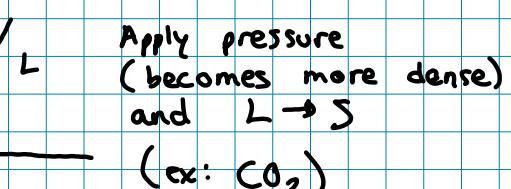
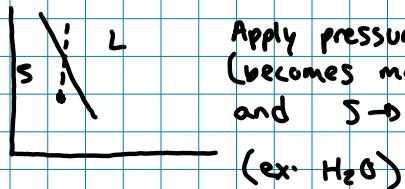
Critical point = Above  $T_c$ , no matter how much pressure is applied, one cannot convert the gas to liquid. pt  $(T_c, P_c)$   
 \* If one applies greater pressure than  $P_c$  and greater temp than  $T_c$  ∴ the gas is instead converted to supercritical fluid  
 ⇒ State of matter that is neither liquid or gas but has properties between the 2 states

ex: Water phase diagram



Note: Solid-liquid boundary (MP line):

- Tilt up-right (positive slope) ∵ solid denser than liquid
- Tilt up-left (negative slope) ∵ liquid denser than solid



# Properties of Solutions

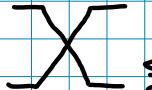
Colligative property = property that depends on the relative number of solvent to solute particles

Reminder: • molarity =  $\frac{\text{mol (solute)}}{\text{L (solution)}}$  / mass % =  $\frac{\text{mass (solute)}}{\text{mass (solution)}} \times 100$

2 new concentration units are used to describe the 4 colligative properties of solutions:

- Vapour pressure
- Freezing point depression
- Boiling point elevation
- Osmotic pressure

Molality ( $m$ ) =  $\frac{\text{moles (solute)}}{\text{kg (solvent)}}$

Mole fraction ( $X$ ):  =  $\frac{n_{\text{sol'n component}}}{n_{\text{total}}}$

- Sol'n component = solute or solvent
- $n_{\text{total}} = n_{\text{solute}} + n_{\text{solvent}}$

Converting molarity  $\rightarrow$  molality

ex: A 4.98 M sol'n of  $\text{H}_2\text{SO}_4$  has a density of 1.286 g/mL  
(mm  $\text{H}_2\text{SO}_4 = 98.09 \text{ g/mol}$ )

$$\boxed{m = \frac{\text{mol H}_2\text{SO}_4}{\text{kg H}_2\text{O}}}$$

\* Assume any volume (1 L)

$$\textcircled{1} \quad 1 \text{ L soln} \cdot \frac{4.98 \text{ mol H}_2\text{SO}_4}{\text{L soln}} = \textcircled{4.98 \text{ mol H}_2\text{SO}_4}$$

$$\textcircled{2} \quad 1 \text{ L soln} \cdot \frac{1.286 \text{ g soln}}{0.001 \text{ L soln}} = \textcircled{1286 \text{ g soln}}$$

$$1286 \text{ g} = 4.98 \text{ mol H}_2\text{SO}_4 \cdot \frac{98.09 \text{ g}}{\text{mol}} + X \text{ g solvent}$$

$$X = 789 \text{ g}$$

$$X = 0.789 \text{ kg solvent}$$

$$m = \frac{4.98 \text{ mol H}_2\text{SO}_4}{0.789 \text{ kg solvent}}$$

$$= \boxed{6.24 \text{ m}}$$

## Solution Vapour Pressure

Type 1 (non-volatile solute in solvent)

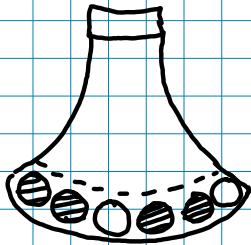
non-volatile solute = sugars or salts

Ideal solution = solution where solute-solvent interactions are the same as solvent-solvent or solute-solute

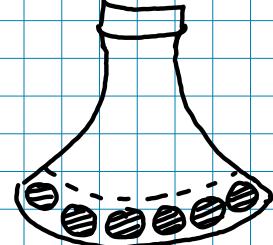
Imagine sol'n: 2/3 solvent, 1/3 solute

$$\chi_{\text{solute}} = 0.33, \quad \chi_{\text{solvent}} = 0.67$$

Solution described above:



Pure solvent:



$\circ$  = x moles of non-volatile solute

$\bullet$  = x moles of solvent

----- = surface

For this sol'n, there is 2/3 the number of solvent molecules at the surface relative to the pure solvent.

$\therefore P_{\text{vap}}$  for the ideal solution is 2/3 of the  $P_{\text{vap}}$  of the pure solvent

$\therefore$  liquid molecules must be at the surface to escape to gas

Raoult's Law (non volatile solutes)

$$P_{\text{vap}(\text{sol'n})} = \chi_{\text{solvent}} \cdot P^{\circ}$$

$\bullet P_{\text{vap}}$  = vapour pressure above soln

$\bullet \chi_{\text{solvent}}$  = mole fraction of solvent

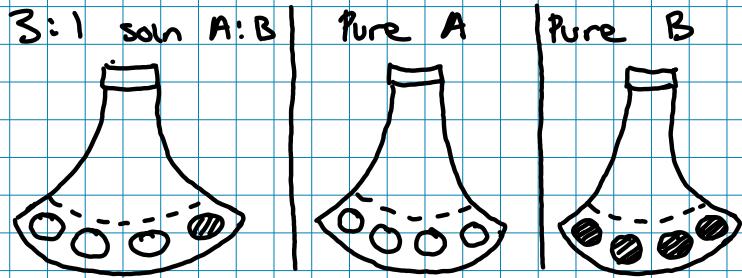
$\bullet P^{\circ}$  = vapour pressure of pure (halo) solvent

\* Units of pressure must be the same.

## Type 2 ( 2 volatile solution components)

2 miscible liquids = 2 liquids soluble in one another

Imagine soln with 3:1 molar mixture A to B



$\textcircled{O}$  =  $x$  moles of A

$\textcircled{\textcircled{B}}$  =  $x$  moles of B

Here, in the 3:1 soln

- $P_A = \text{vapour pressure due to A particles} = \frac{3}{4} P_A^\circ$

- $P_B = \text{vapour pressure due to B particles} = \frac{1}{4} P_B^\circ$

$$\therefore P_{\text{TOT}} = P_{\text{vap(soln)}} = \frac{3}{4} P_A^\circ + \frac{1}{4} P_B^\circ$$

### Modified Raoult's Law (Volatile solvents)

$$P_{\text{vap(soln)}} = \left( X_{\text{solvent}_A} \cdot P_A^\circ \right) + \left( X_{\text{solvent}_B} \cdot P_B^\circ \right)$$

$\underbrace{\phantom{X_{\text{solvent}} \cdot P_A^\circ}}_{P_A}$        $\underbrace{\phantom{X_{\text{solvent}} \cdot P_B^\circ}}_{P_B}$

Note  $X_{\text{solvent}_A}$  = mole fraction of A in liquid phase

Note: 
$$X_A = \frac{P_A}{P_{\text{TOT}}} = \text{mol fraction of A in vapour phase}$$

- $P_{\text{TOT}}$  = total pressure of gas over soln

- $P_A = X_{\text{solvent}_A} \cdot P_A^\circ$

- $P_B = X_{\text{solvent}_B} \cdot P_B^\circ$

\*  $X_A + X_B = 1$

ex:  $H_2O$  vapour pressure at  $80^\circ C$  is 0.4672 atm. What mass of sucrose ( $C_{12}H_{22}O_{11}$ ) must be dissolved in 180.2g of  $H_2O$  to lower  $P_{\text{vap}}(\text{soln})$  to 0.4640 atm.

Sucrose (non volatile) $\rightarrow$ Raoult's Law	$180.2 \text{ g } H_2O \cdot \frac{\text{mol}}{18 \text{ g}} = 10.00 \text{ mol } H_2O$
$P_{\text{vap}} = X_{\text{solvent}} \cdot P^{\circ}$	
$X_{\text{solvent}} = \frac{0.4640}{0.4672} = 0.9932$	
$X_{\text{solvent}} + X_{\text{solute}} = 1$	
$X_{\text{solute}} = 1 - 0.9932 = 0.0068$	
$0.0068 = \frac{X \text{ mol sucrose}}{10.00 \text{ mol } H_2O + X}$	
$0.068 + 0.0068 X = X$	
$X = 0.068 \text{ mol sucrose}$	

$$0.068 \text{ mol} \cdot \frac{342.30 \text{ g}}{\text{mol}} = \boxed{23 \text{ g sucrose}}$$

ex: 0.300 mol benzene mixed with 0.800 mol toluene ( $P_{\text{benzene}}^{\circ} = 1.34 \text{ atm}$ ,  $P_{\text{toluene}}^{\circ} = 0.534 \text{ atm}$ ). What is mole fraction of toluene in vapour?

Both have  $P_{\text{vap}}$   $\therefore$  both volatile  $\rightarrow$  Modified Raoult's Law

$$\begin{aligned} P_{\text{vap}}(\text{soln}) &= (X_{\text{benzene}} \cdot P_{\text{benzene}}^{\circ}) + (X_{\text{toluene}} \cdot P_{\text{toluene}}^{\circ}) \\ &= \left( \frac{0.300 \text{ mol}}{0.300 + 0.800} \cdot 1.34 \text{ atm} \right) + \left( \frac{0.800}{0.300 + 0.800} \cdot 0.534 \text{ atm} \right) \\ &= (0.365 \text{ atm}) + (0.388 \text{ atm}) \end{aligned}$$

$$P_{\text{vap}}(\text{soln}) = 0.753 \text{ atm}$$

$$X'_{\text{toluene}} = \frac{P_{\text{toluene}}}{P_{\text{tot}}} = \frac{0.388}{0.753} = \boxed{0.515}$$

↑

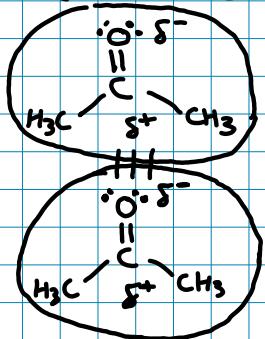
Makes sense:  
 - pure toluene has lower  $P^{\circ}$  than benzene  
 $\therefore X'$  vapour phase less than  $X$  in liquid  
 (lower vapour pressure = less liquid becomes vapour)

## Non-Ideal Solutions

→ to be ideal solute-solvent interactions = solvent-solvent or solute-solute

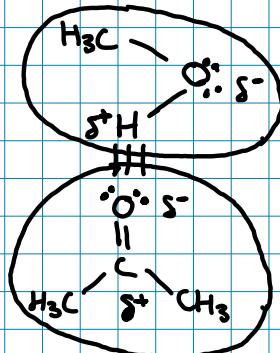
Imagine miscible mixture of acetone and methanol:

Acetone-Acetone interact:



- LD
- dipole-dipole

Acetone-Methanol interact:

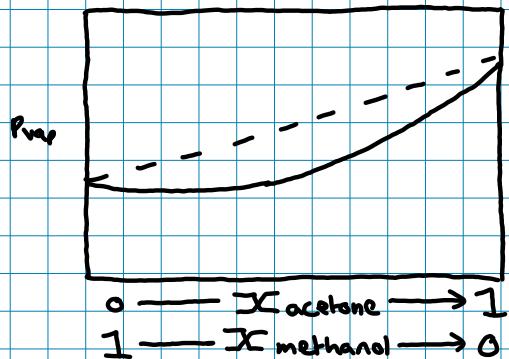


- LD
- H-bonding (strongest)

∴ Stronger interaction

If solute attracts solvent more than solvent-solvent, vapour pressure is less than expected b/c more attraction = less tendency

Negative Deviation from Raoult's Law: solute-solvent interaction stronger than solvent-solvent  
⇒ When  $P_{\text{app}}(\text{soln})$  is less than expected from Raoult's

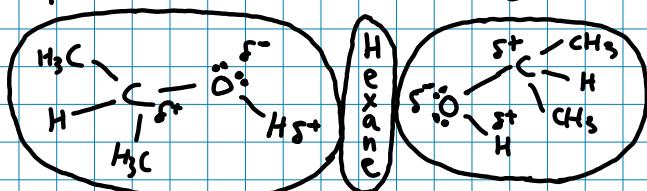


----- = predicted from Raoult's  
— = experimentally observed

Note:  $\chi_{\text{methanol}} = 0 \rightarrow$  higher pressure b/c less methanol ⇒ less H-bonding  
∴ easier to escape

Positive Deviation: solute-solvent interaction weaker than solvent-solvent  
⇒ When  $P_{\text{app}}(\text{soln})$  is more than expected from Raoult's

This can happen when non-polar liquid (ex: hexane) is mixed with polar solvent (ex:  $(\text{CH}_3)_2\text{COH} \cdot \text{H}$ )



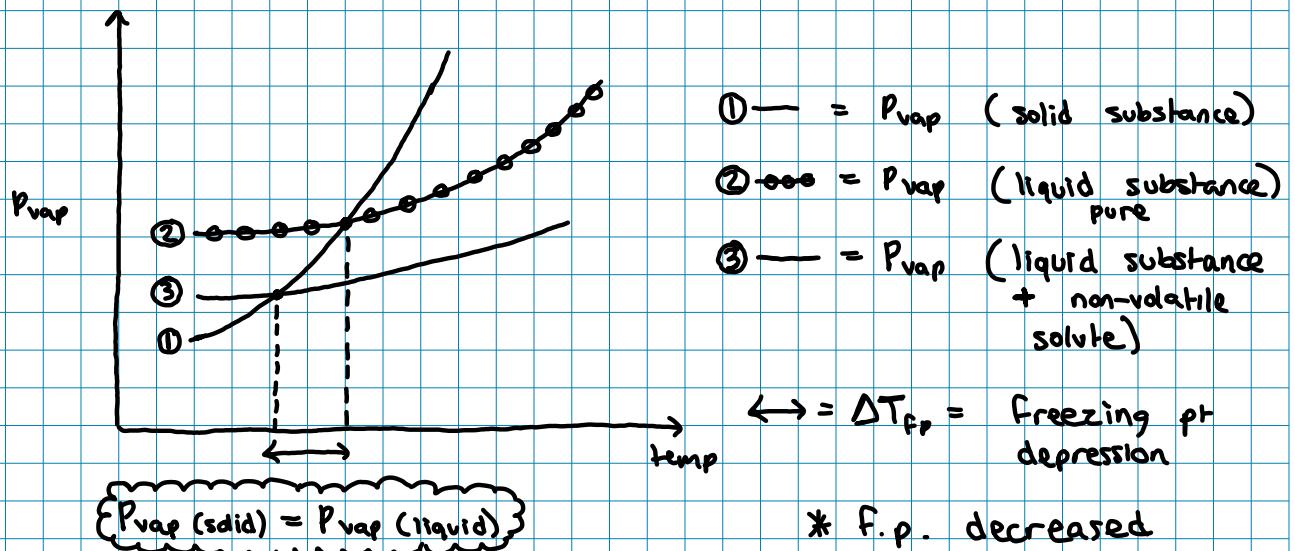
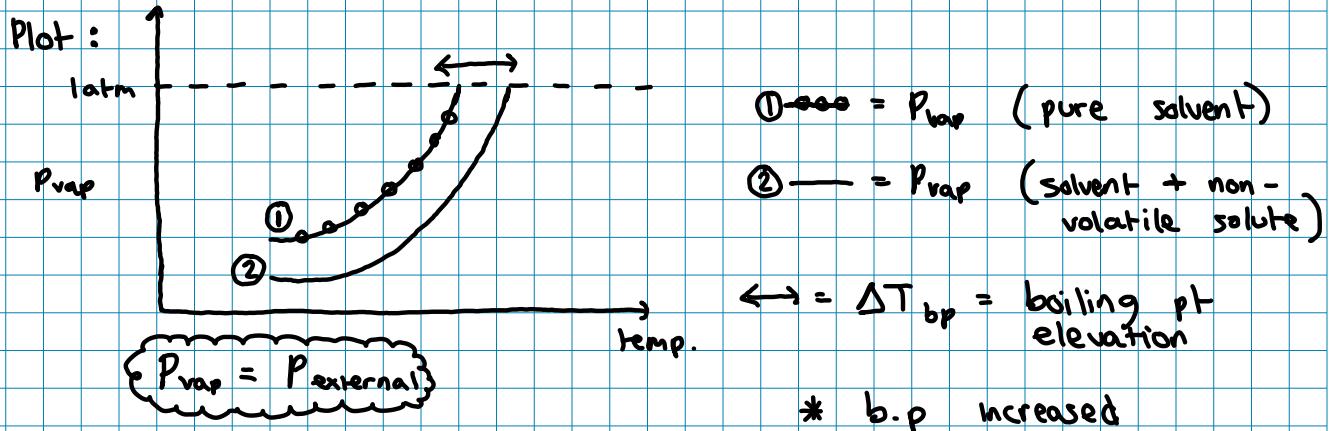
2 isopropane molecules have strong H-bonding. Presence of non-polar disrupts powerful interaction. Weaker interaction ∴ more particles can escape  
∴ higher  $P_{\text{app}}$  than expected

## Freezing-Point Depression and Boiling-Point Elevation

\* Occurs when non-volatile solute added to liquid

Reminder:

- Increase temp  $\Rightarrow P_{\text{vap}}$  increases exponentially
- Boiling Point  $\Rightarrow P_{\text{vap}}(\text{liquid}) = \text{external pressure}$
- Melting Point  $\Rightarrow P_{\text{vap}}(\text{solid}) = P_{\text{vap}}(\text{liquid})$
- In presence of a non-volatile solute,  $P_{\text{vap}}$  of a liquid decreases (Raoult's)



$\Delta T_{bp}$  = boiling point elevation =  $\frac{\text{temp at which liquid boils with non-volatile solute}}{\text{temp at which pure liquid boils}}$

$\Delta T_{fp}$  = freezing point depression =  $\frac{\text{temp at which solvent freezes with non-volatile solute}}{\text{temp at which pure solvent freezes}}$

\* There is a greater  $\Delta T$  (bp or fp) if there are more solute particles

$$\boxed{\Delta T_{bp} = K_b \cdot m}$$

$$\boxed{\Delta T_{fp} = K_f \cdot m}$$

$$m = \text{molarity} = \frac{\text{mol solute}}{\text{kg solvent}}$$

$K_b$  = molal bp elevation constant

$K_f$  = molal f.p. depression constant

\*  $K_f$  and  $K_b$  are specific to the solvent

units:  $^{\circ}\text{C}/\text{m}$  or  $\text{K}/\text{m}$  ( $\frac{\text{change of } 1^{\circ}\text{C}}{\text{change of } 1\text{K}}$ )

~~XXXXXX~~

## Determining Molar Mass

ex: A 39.8 g sample of non-volatile compound is dissolved in 200.0 g of  $\text{H}_2\text{O}$  and the boiling point is raised by  $0.30^{\circ}\text{C}$ . What is the molar mass of compound?

$$K_b(\text{H}_2\text{O}) = 0.512 \text{ } ^{\circ}\text{C}/\text{mol} = 0.512 \text{ K}/\text{mol}$$

} Change of  $1^{\circ}\text{C}$  is same as  $1\text{K}$

$$\Delta T_{bp} = 0.30^{\circ}\text{C}$$

$$\Delta T_{bp} = K_b \cdot m$$

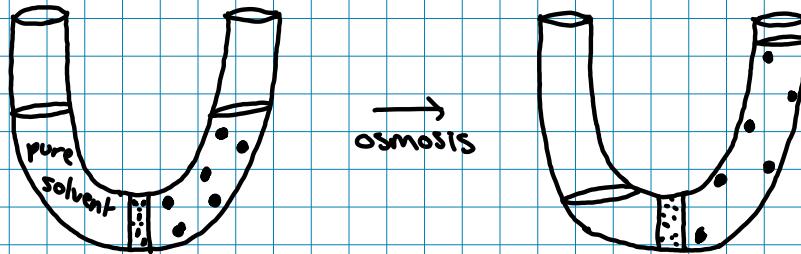
$$0.30^{\circ}\text{C} = 0.512 \cdot m$$

$$\boxed{m = 0.585 \frac{\text{mol}}{\text{kg}}}$$

$$\frac{39.8 \text{ g solute}}{200.0 \text{ g solvent}} \cdot \frac{1000 \text{ g solvent}}{\text{kg solvent}} \cdot \frac{\text{kg solvent}}{0.585 \text{ mol solute}} = \boxed{340 \text{ g/mol}}$$

## Osmotic Pressure

Osmosis = net movement of a solvent through a semi-permeable membrane towards the side of the membrane with the greater concentration



Osmosis occurs naturally to attempt to equalize the concentration. Pure solvent moves to dilute the solution.

The pressure due to osmosis = Osmotic Pressure ( $\Pi$ )  
 $\Rightarrow$  pressure required to stop the solvent from flowing through the membrane

$$\boxed{\Pi = C R T}$$

$$P = \frac{n}{V} R T$$

$$R = 0.08206 \left( \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \right)$$

$C$  = concentration of soln pre-osmosis (mol/L)

$T$  = temp. (K)

- \* A useful benefit about the 4 colligative properties is that they can be used to find the molar mass.  
 $\Rightarrow$  combined with elemental analysis / combustion (find empirical formula), one can find the molecular formula

ex: A dilute aqueous sol'n of non-volatile compound contains 1.19 g compound/L of soln and has  $\Pi = 0.0288 \text{ atm}$  at 37°C. Calculate molar mass

$$\Pi = C R T$$

$$0.0288 \text{ atm} = C \cdot (0.08206) (310)$$

$$\boxed{C = 1.13 \times 10^{-3} \text{ mol/L}}$$

$$\frac{1.19 \text{ g}}{\text{L}} \cdot \frac{\text{L}}{1.13 \times 10^{-3} \text{ mol}} = 1.05 \times 10^3 \text{ g/mol}$$

Osmotic pressure is useful b/c even small concentrations produce high osmotic pressures

If we were to calculate the  $\Delta T_{fp}$  of the same substance as on the previous page:

Assume  $0.001 \text{ mol/L} \approx 0.001 \text{ m}$

$$\Delta T_{fp} = K_f \cdot m$$

$$= (1.86)(0.001) = 2 \times 10^{-3} \text{ }^{\circ}\text{C}$$

- Under most lab conditions, one cannot be confident about calculating this change.
- However, using osmotic pressure, one can find better values

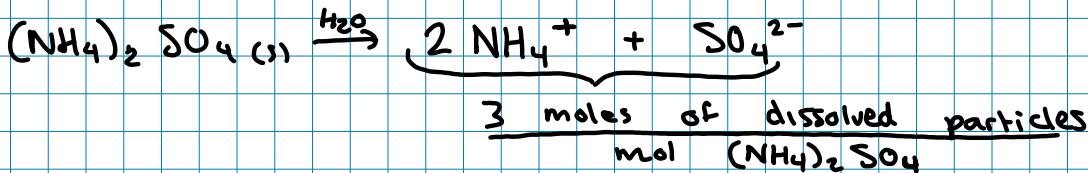
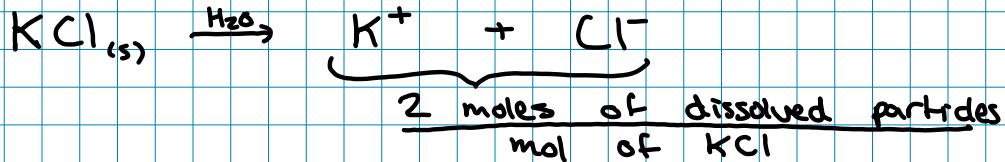
### Reverse Osmosis

→ If one applies pressure greater than osmotic pressure to U-shaped tube it is possible to reverse osmosis.

→ Solvent moves to the side where there is lower concentr.  
ex: desalination occurs by reverse osmosis  
⇒ Does NOT occur naturally

## Ionic Compounds as Solutes

- Colligative properties depend on the number of dissolved particles



We must multiply the molarity or molality of ionic compounds by the van't Hoff factor ( $i$ )

$$i = 2 \text{ for KCl}$$

$$i = 3 \text{ for } (\text{NH}_4)_2\text{SO}_4$$

With the corrective factor:

$$\bullet \Pi = i \cdot CRT$$

$$\bullet \Delta T_{bp} = i \cdot K_b \cdot m$$

$$\bullet \Delta T_{fp} = i \cdot K_f \cdot m$$

Experimentally,  $i$  values may be lower than expected

$$i = \frac{\Pi}{CRT}, \quad i = \frac{\Delta T}{K \cdot m}$$

- Consider  $i = 1$  for non dissociating compound (covalent bonds)
- If strong acid or soluble ionic compound, use expected " $i$ "

Ex: 2.00g KCl dissolved in 10.0g H<sub>2</sub>O at 30°C. What is vapour pressure of soln? (P<sub>vp</sub> of pure H<sub>2</sub>O = 31.82 torr)

KCl (non-volatile and ionic) → Raoult's Law

$$P_{vp \text{ (soln)}} = \sum_{\text{solvent}} \cdot P_{H_2O}^\circ$$

$$2.00 \text{ g KCl} \cdot \frac{\text{mol KCl}}{74.55 \text{ g}} \cdot \frac{2 \text{ mol particles}}{\text{mol KCl}} = 0.0537 \quad / \quad 10.0 \text{ g H}_2\text{O} \cdot \frac{\text{mol H}_2\text{O}}{18.02 \text{ g}} = 0.555$$

$$\sum_{\text{solvent}} = \frac{0.555}{0.555 + 0.0537} = 0.911$$

$$\therefore P_{vp} = (0.911) (31.82 \text{ torr}) = 29.0 \text{ torr}$$

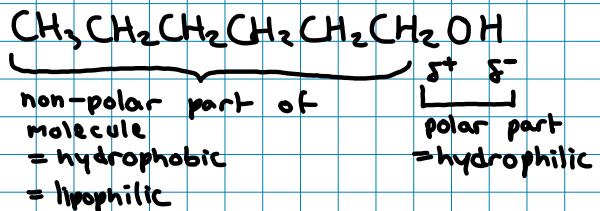
# Solubility of Compounds (covalent)

"Like dissolves like"

- Polar compounds dissolve other polar compounds best
- Non-polar compounds dissolve non-polar compounds best

trick: → popular (polar) so attractive hang out together  
→ unpopular (non-polar) left out, hang out together

ex.: Between 1-hexanol ( $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$ ) and methanol ( $\text{CH}_3\text{OH}$ ), which is more ~~soluble~~ / ~~hydrophilic~~ / ~~hydrophobic~~ soluble in  $\text{H}_2\text{O}$  (polar) and which is in  $\text{CCl}_4$  (non-polar)



1-hexanol has a larger hydrophobic area ∴ it is less soluble in  $\text{H}_2\text{O}$ .  
However it is more soluble in  $\text{CCl}_4$ .

## Gas dissolved in solvent

Henry's Law: concentration of gas dissolved in soln is proportional to external partial pressure of that gas.

$$\frac{P_1}{C_1} = \frac{P_2}{C_2} \quad / P \propto C \quad / C = K_H \cdot P$$

-  $P$  = partial pressure of gas above soln

-  $C$  = concentration of gas dissolved in soln

-  $K_H$  = Henry's Law constant is specific to gas at a temp.

ex:  $N_2(g)$  at 0.78 atm is dissolved in water ( $K_H$  of  $N_2$  is  $8.42 \times 10^{-7}$  M/torr). Calculate molarity.

\* have same units of pressure

$$\begin{aligned} C &= K_H \cdot P \\ &= (0.82 \text{ M/torr}) (0.78 \text{ atm} \cdot \frac{760 \text{ torr}}{\text{atm}}) \times 10^{-7} \\ &= 5.0 \times 10^{-4} \text{ M} \end{aligned}$$

ex: Fish need  $O_2$  in water at  $1.25 \times 10^{-4}$  M to survive. What partial pressure of  $O_2$  above water is needed to obtain this concentration? (solubility of  $O_2$  at 1 atm is  $2.21 \times 10^{-3}$  M)

$$\frac{P_1}{C_1} = \frac{P_2}{C_2}$$

$$\frac{1 \text{ atm}}{2.21 \times 10^{-3} \text{ M}} = \frac{P_2}{1.25 \times 10^{-4} \text{ M}} \Rightarrow P_2 = 5.66 \times 10^{-2} \text{ atm}$$

ex: To celebrate finishing the course, you pop a bottle of champagne!!

Henry's Law is relevant b/c the drink is stored under high pressure  $C$