

# 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

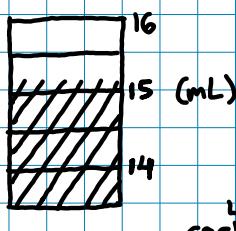
Record

All certain digits

+ 1 uncertain digit

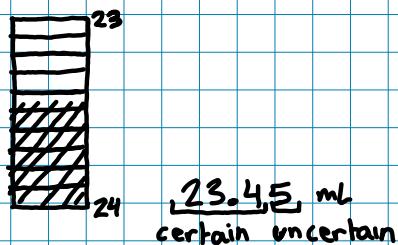
- Digital balance: All digits recorded.  
The uncertain digit is the last visible one on the balance.
- Glassware: Read volume from the bottom of the miniscus

### Graduated Cylinder



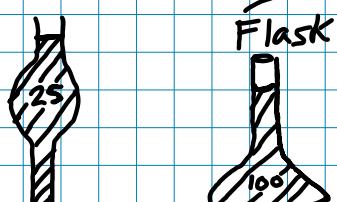
Always record  
1 decimal

### 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  
X  
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:  $\underline{400.}$  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 :  $\frac{\text{value}_1}{\text{value}_2} \frac{\text{unit}_1}{\text{unit}_2} = 1$

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.

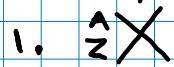
Element: A substance that cannot be decomposed into simpler substances

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 :  $C_6H_6$   
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:  ${}^{41}Ca^{2+}$  protons: 20  
neutrons: 21

electrons: 18 (lost 2)

${}^{79}Se^{2-}$  protons: 34  
electrons: 36 (gained 2)  
neutron: 45

Polyatomic:  $CO_3^{2-}$

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

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

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(g)</sub>, O<sub>3(g)</sub>  
<sup>ozone</sup>

ex: allotropes of carbon = C<sub>{3, graphite}</sub>, C<sub>(s, diamond)</sub>, 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 ( $\text{O}_2^{2-}$ )

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

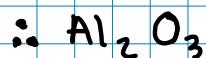
Molecular formula: metal → non-metal

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

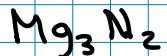
1. Name metal (cation) as is

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

Root =

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

ex:

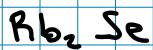


Metal = alkaline-earth  
↳ type I

Magnesium

nitride

nitrogen

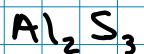


Metal = alkali  
↳ type I

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:

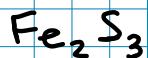
1. Small rectangle, left of staircase  
In, Sn, Sb, Ti, Pb, Bi
2. All transition metals, (other than Cd, Ag, Zn)
3. Lanthanides, actinides

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

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

$$(\boxed{\phantom{0}} \times 2) + (2^- \times 3) = \text{zero}$$

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

$\text{Fe}^{3+}$  ∴ Iron (III) Sulfide  
Sulfur

# 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

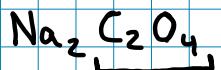
→ Compounds with polyatomic anions (-)

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

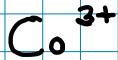


metal = type 1

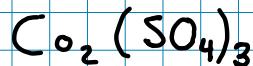
Sodium Oxalate

ex: Name to formula

Cobalt (III) Sulfate



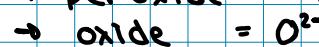
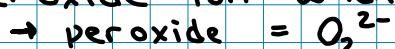
$$\begin{aligned} (\text{number of } \text{Co} \times 3^+) + (\text{number of } \text{SO}_4 \times 2^-) &= 0 \\ (2 \times 3^+) + (3 \times 2^-) &= 0 \end{aligned}$$



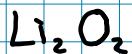
\* Note: parenthesis if more than 1

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

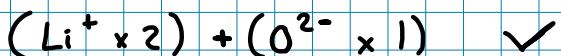
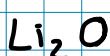
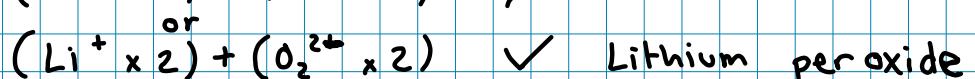
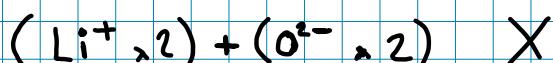
Peroxide ion when naming



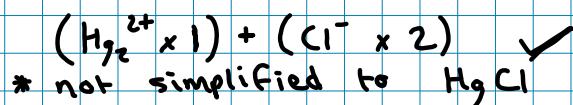
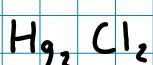
Think of neutrality (what do I need to cancel the charge)



type I metal



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) (H<sub>2</sub>O)

- For every unit of the compound there are that many H<sub>2</sub>O molecules.

Naming:

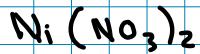
1. Name ionic compound

2. "Prefix" + "hydrate"

Prefixes (indicating number of H<sub>2</sub>O):

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

ex. name Ni(NO<sub>3</sub>)<sub>2</sub> • 2 H<sub>2</sub>O



type 2 metal

NO<sub>3</sub><sup>-</sup> = nitrate

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

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

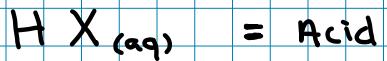
Hydrogen Compounds with special names:

$\text{H}_2\text{O}$  = 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 :  $H^+$

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  $\text{C}_{10}\text{H}_8$ ?

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

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

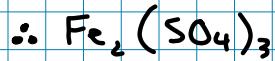
## Percent by mass

→ % 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-

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

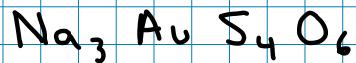
$$21.10 \text{ mg Na} \times \frac{\text{mmol Na}}{22.99 \text{ mg Na}} = \frac{0.9178 \text{ mmol Na}}{0.3061}$$

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

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

\* 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)

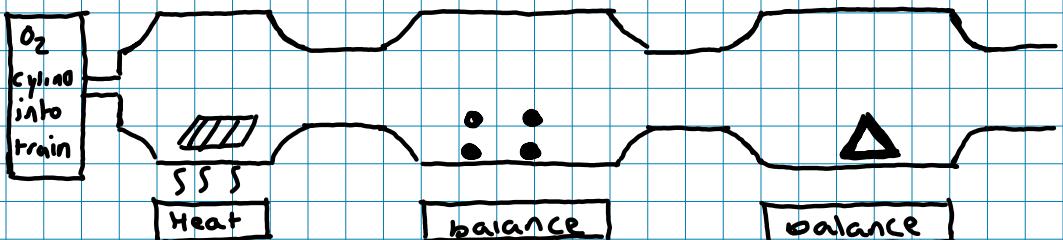
Hydrocarbon = compound containing only C and H

Combustion train:

= sample compound

= desiccant (absorbs  $H_2O$ )

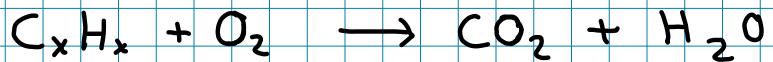
= ( $O_2$  absorbing compound)



- 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

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

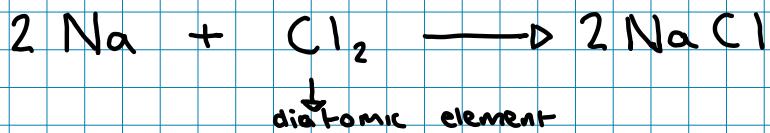
# Chemical Reaction (Stoichiometry)

- 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

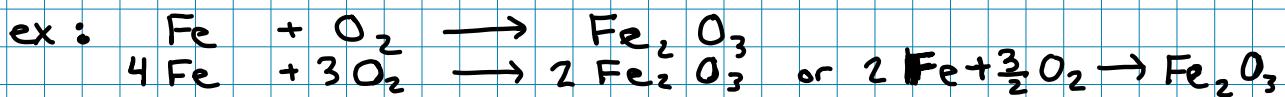


To balance, change only stoichiometric coefficient, not the subscript

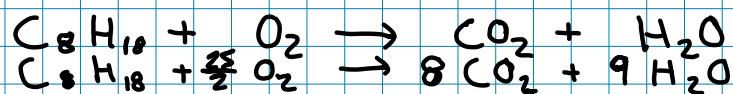
Coefficient = relative number of each reactant that reacts and relative number of products produced (assuming reaction completion)

- 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)

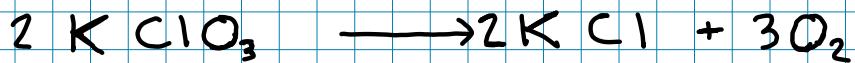


C	8	8
H	18	18
O	25	25

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)

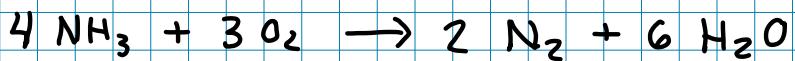
Find out how much excess there is :

- 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:

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

$$50.0 \text{ g O}_2 - 47.9 \text{ g O}_2 = 2.1 \text{ g O}_2 \quad \begin{matrix} \text{(start)} \\ \text{(react)} \end{matrix} \quad \begin{matrix} \text{(excess)} \end{matrix}$$

$$\text{or } 1.56 \text{ mol O}_2 - 1.50 \text{ mol O}_2 = 0.06 \text{ mol O}_2 \quad \begin{matrix} \text{(start)} \\ \text{(react)} \end{matrix} \quad \begin{matrix} \text{(excess)} \end{matrix}$$

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

% Yield:

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

→ Theoretical yield = 54.0 g

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

# Solutions

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

**Solution:** homogeneous mixture

**Solvent:** dissolving medium

**Solute:** substance dissolved

## Concentration:

\* **Molarity (M) = moles of solute / L of solution (not solvent)**

% 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/cm<sup>3</sup>)

$$\left( \frac{12.1 \text{ mol HCl}}{L \text{ solution}} \times \frac{36.46 \text{ g HCl}}{\text{mol HCl}} \times \frac{\cancel{cm^3}}{1.19 \text{ g solution}} \times \frac{\cancel{mL}}{\cancel{cm^3}} \times \frac{1}{1000 \text{ mL}} \right) \times 100 \\ = 37.1\%$$

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

$$\frac{100.00 \text{ mL solution}}{1000 \text{ mL}} \times \frac{0.250 \text{ mol KCl}}{L \text{ solution}} \times \frac{74.55 \text{ g KCl}}{\text{mol KCl}} = \\ = 1.86 \text{ g KCl}$$

## 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)

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

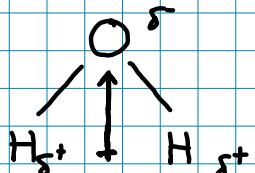
ex: What initial volume of a 0.200 M NaOH solution be used to prepare a 0.0100 M solution with final volume 100.00 mL? And how much solvent added?

$$\begin{aligned} C_1 \times V_1 &= C_2 \cdot V_2 \\ (0.200 \text{ M}) \times V_1 &= (0.0100 \text{ M})(100.00 \text{ mL}) \\ V_1 &= 5.00 \text{ mL} \\ &\therefore \text{add } 95.00 \text{ mL} \end{aligned}$$

## 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

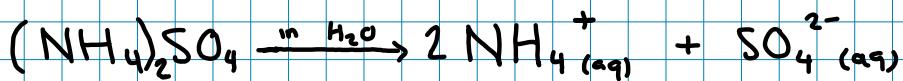
Covalent compound dissolved in water :

→ stay intact

ex: potassium nitrate



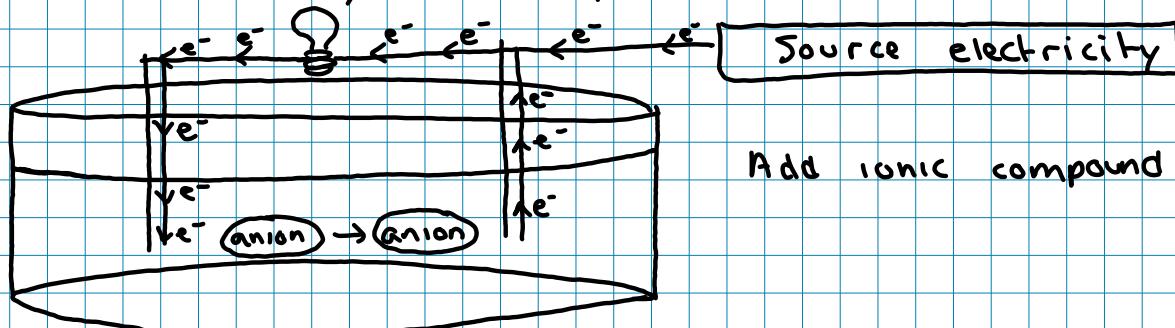
ex ammonium sulfate



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)

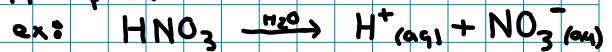
\* HClO<sub>3</sub>  $\Rightarrow$  borderline strong

### ② Soluble ionic compounds

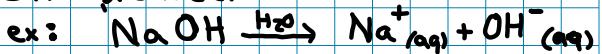
## Acids and Bases (reactants)

### Arrhenius definition

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



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

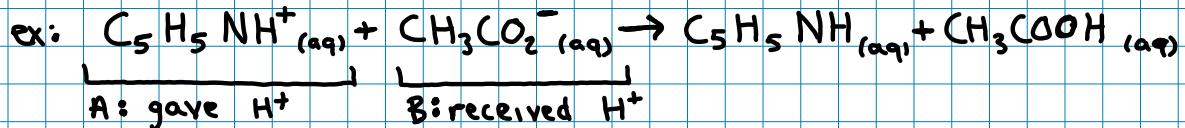


### Bronsted-Lowry

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

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

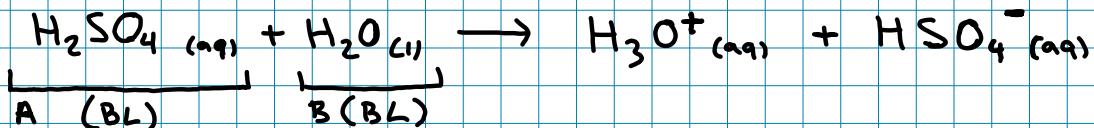
\* B-L Acid always coupled with B-L base



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>)  $\rightarrow$  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

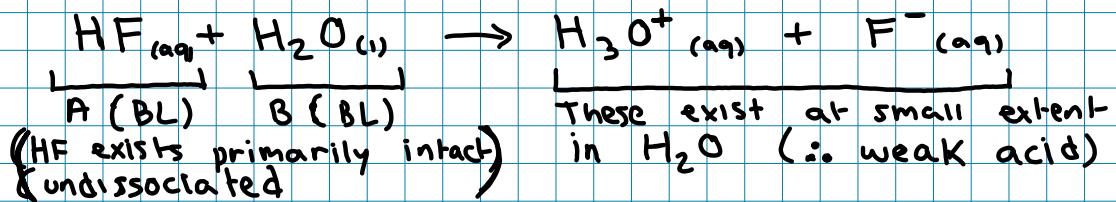
① weak acids

② weak base

③ slightly soluble ionic compounds

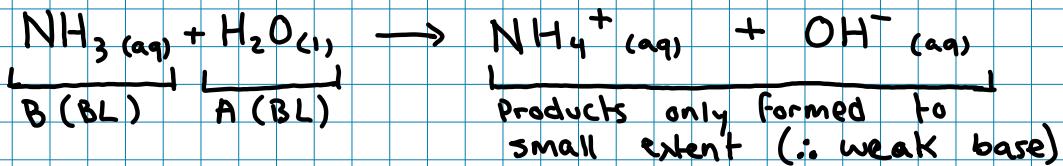
weak acid =  $\text{HX}_{(\text{aq})}$  that splits only to a small extent into  $\text{H}^+$  and  $\text{X}^-$

→ 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)

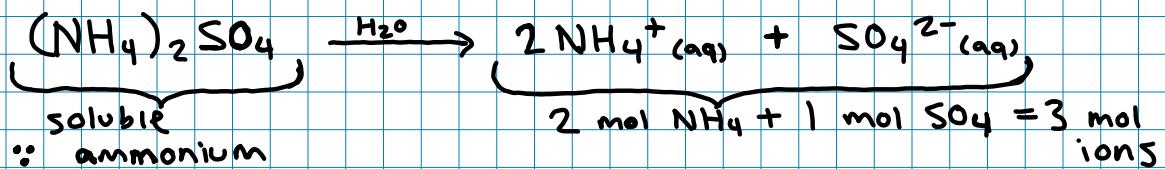
- Ionic compounds with nitrate  $\Rightarrow \text{NO}_3^-$
  - Ionic compounds with alkali metals or ammonium  $\Rightarrow \text{X}^+$ ,  $\text{NH}_4^+$
  - Ionic compounds with chloride, bromide or iodide  $\Rightarrow \text{Cl}^-, \text{Br}^-, \text{I}^-$
- \* There H A P ens to be 3 exceptions  
 $\text{Hg}_2\text{X}_2$ ,  $\text{AgX}$ ,  $\text{PbX}_2$
- 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

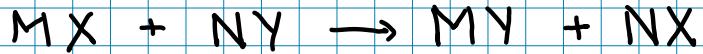


$$\frac{2.00 \text{ mol } (\text{NH}_4)_2\text{SO}_4}{1.00 \text{ L solution}} \times \frac{3 \text{ mol ions}}{\text{mol } (\text{NH}_4)_2\text{SO}_4} = \frac{6.00 \text{ mol ions}}{\text{L solution}}$$

## Precipitation (reaction type 1)

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

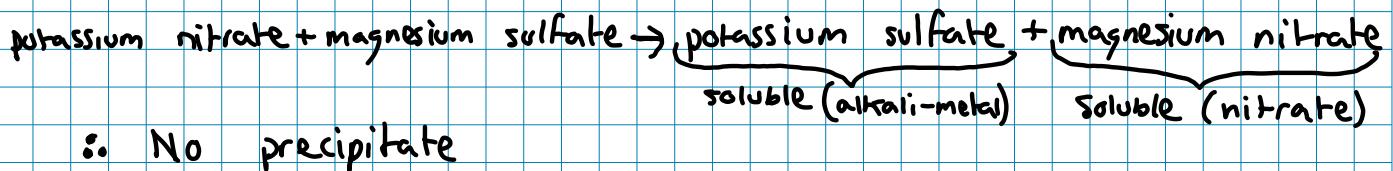
Precipitation = exchange reaction (double displacement or metathesis)



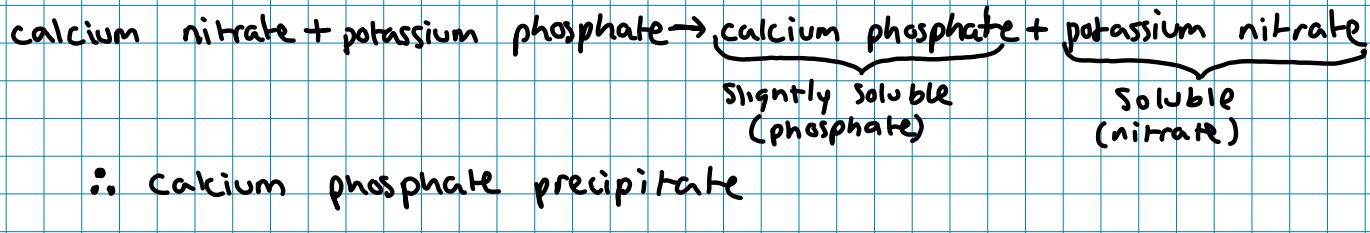
- cation 1 with anion 2
- cation 2 with anion 1

\* 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

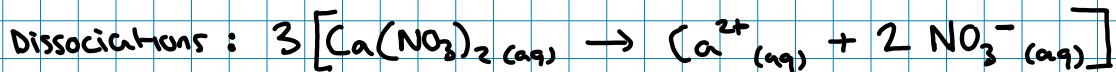
#### ① Overall balanced equation

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

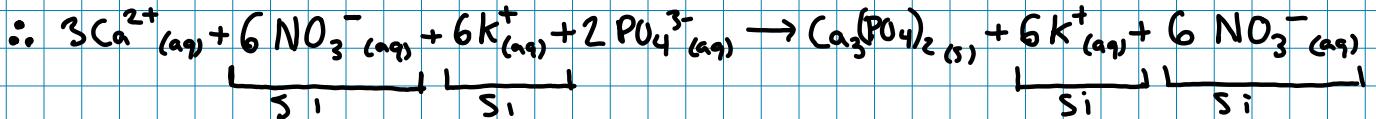


#### ② Complete ionic equation

- 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)



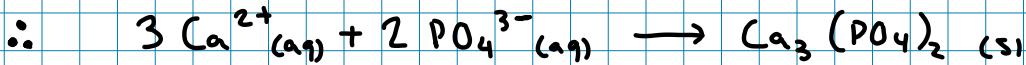
$\text{Ca}_3(\text{PO}_4)_2(s) \rightarrow$  Intact



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



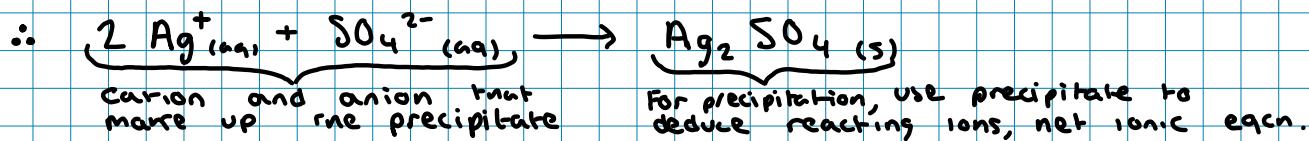
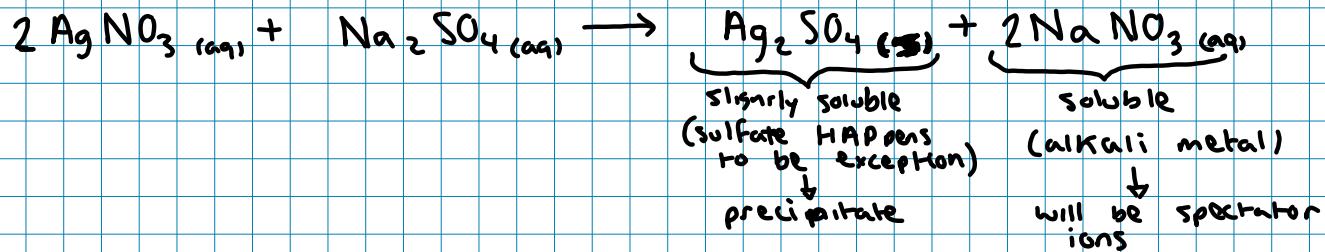
All that actually happened is  $\text{Ca}^+$  and  $\text{PO}_4^{3-}$  reacted.  
The other ions didn't do anything.

\* 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?

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

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

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

$\because Ag^+$  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 Ag<sub>2</sub>NO<sub>3</sub> to give a theoretical yield of 1.913 g of AgCl. Calculate % mass of each compound in mixture

- \* AgNO<sub>3</sub> is not limiting reagent (excess)
- \* Both NaCl and KCl will react with AgNO<sub>3</sub>
- \* NaCl and KCl don't react (exchange anion → get same thing)
- \* Can't write: NaCl + KCl + AgNO<sub>3</sub> → AgCl + NaNO<sub>3</sub> + KNO<sub>3</sub>  
 ↳ can't assume that every one mol of NaCl that reacts, one mol KCl reacts  
 ↳ both reactions are separate



- \* Mixture → ~~consist~~ 2 different compounds react with the same reagent

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

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

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

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

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

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

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

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

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

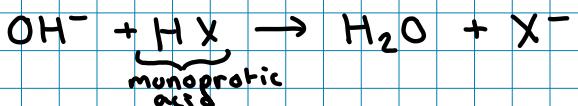
# Acid/Base reaction (reaction type 2)

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

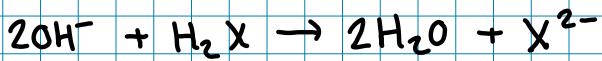
→  $\text{OH}^-$  = strongest base

→ Reacts completely with all acids (weak or strong)

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

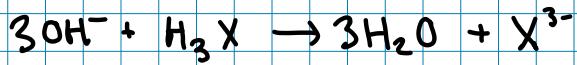


(Remove 1  $\text{H}^+$  from monoprotic)



(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

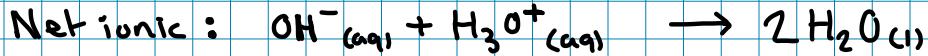
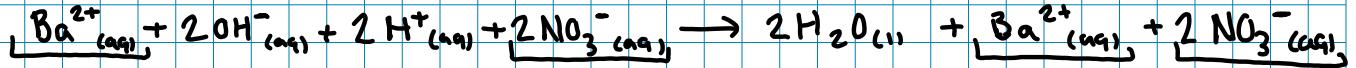
Net ionic equation (most important)

- 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>

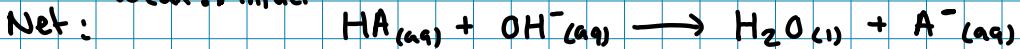
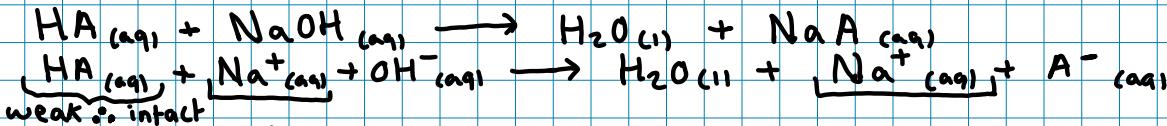
ex: (only strong A/B): 0.0100M soln barium hydroxide (20.15mL) is added by titration to soln nitric acid (5.00mL) to reach equivalence point. What is concentration of nitric acid?



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

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

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.



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

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

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

# Redox reaction (reaction type 3)

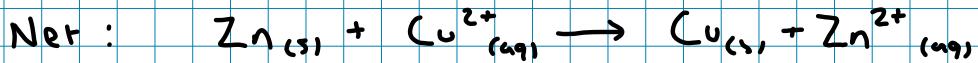
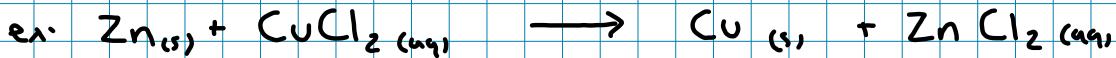
Oxydation  
Involves  
Loss (of electrons)

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)

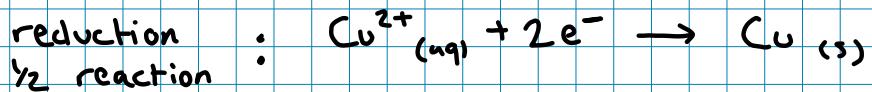
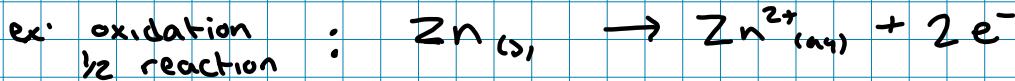
Reducing agent:  $Zn(s)$  (caused the gain - reduction)

- \* 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

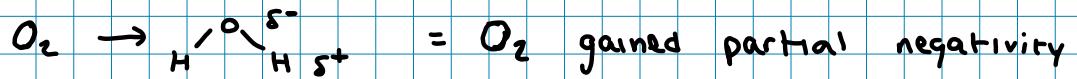
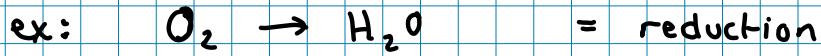


- \* Electrons in product = oxidation (loss)

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

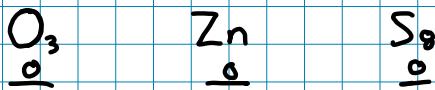
\* All elements are assigned oxidation number

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

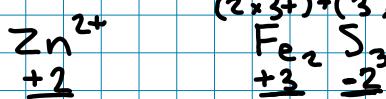
+ charge = "number"  $\pm$

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)

$(Cl_2 \Rightarrow 0)$

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

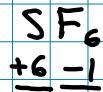
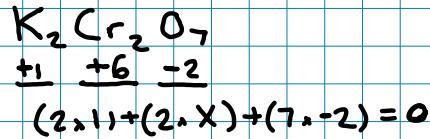
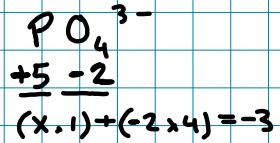
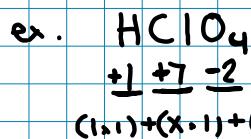
$(Br_2 \Rightarrow 0)$

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

$(I_2 \Rightarrow 0)$

## Oxidation number problems:

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

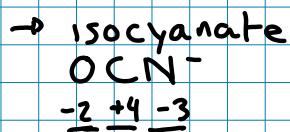
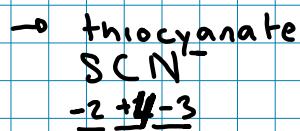
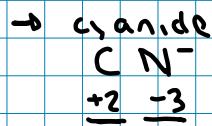


\* The oxidation num. of F is -1 (not -6), since we do not multiply by subscript

### 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  $H_2O$ ,  $H^+$ ,  $OH^-$

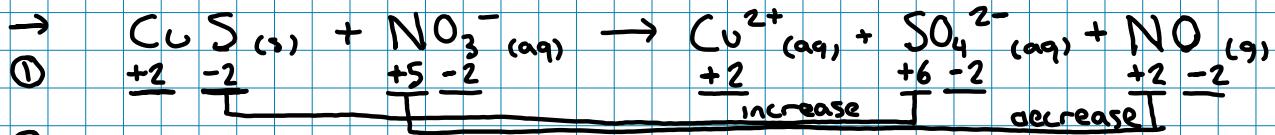
$\downarrow$        $\downarrow$   
(acidic)    (basic)

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

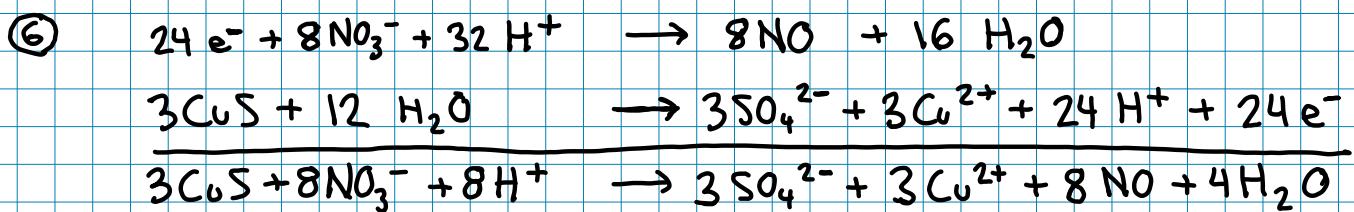
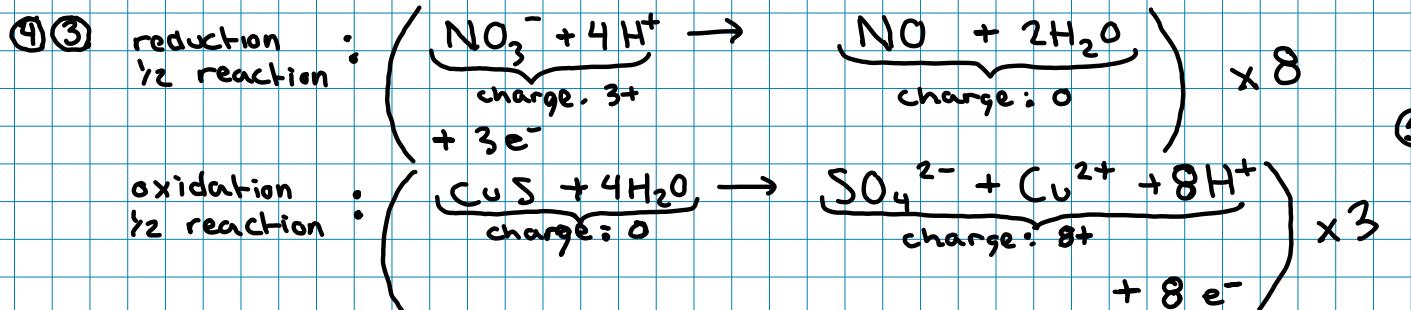
\* Double check atoms and charge

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

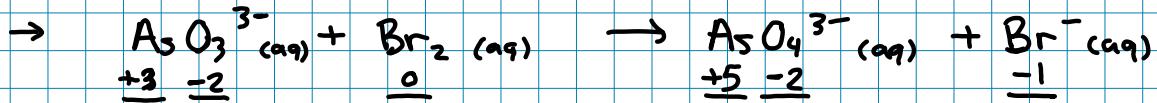
ex: redox in acidic soln.



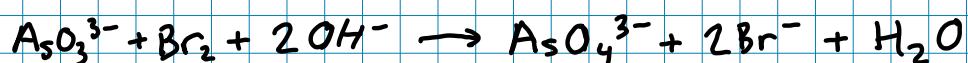
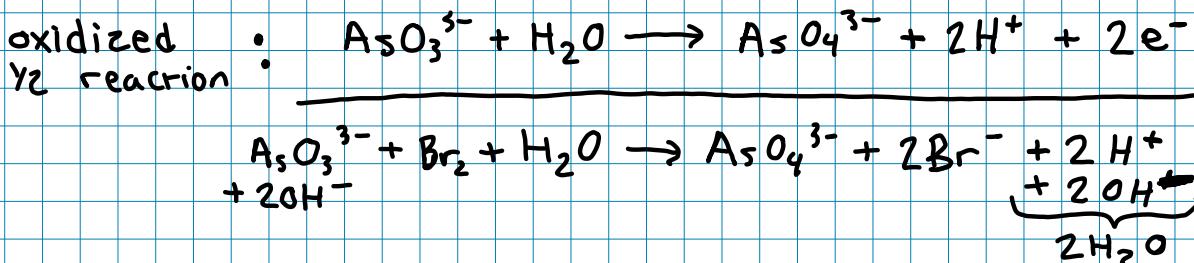
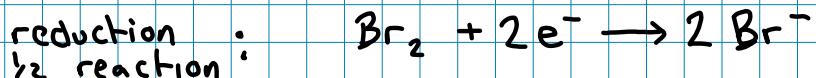
② N decrease  
 ↳  $\text{NO}_3^-$  : reduced  
 S increase  
 ↳ CuS : oxidized  
 \* Not just element



ex: redox in basic soln.



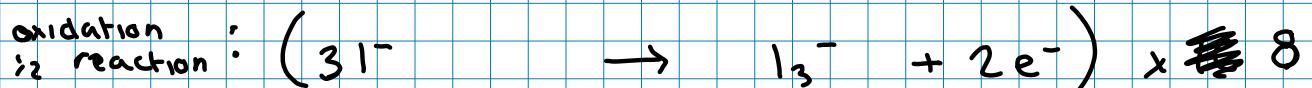
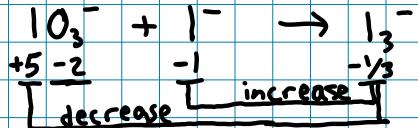
As increase  
 ↳  $\text{AsO}_3^{3-}$  : oxidized  
 Br decrease  
 ↳  $\text{Br}_2$  : reduced



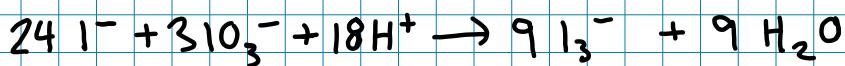
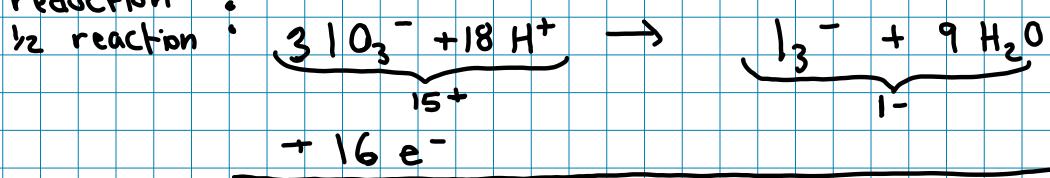
## Hard Redox Problem

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

ex: balance in acidic



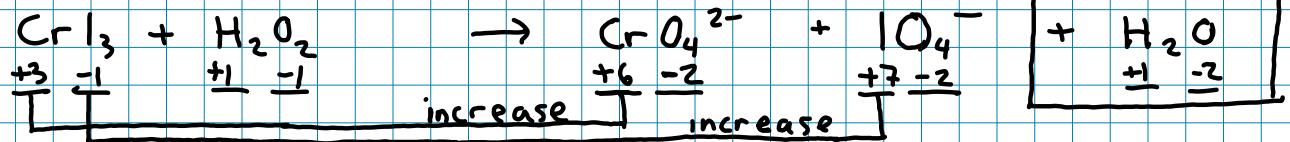
reduction:



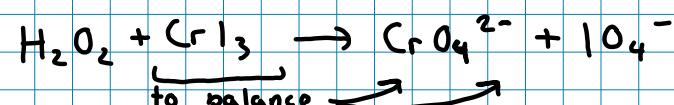
## Hard Redox Problem

- 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)

