Chemistry Notes

Joshua Liu

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Contents

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1 Introduction

Google classroom code: 5mxwn3z

2 Scientific Notation

Number written as the production of two numbers: A coefficient and some power of 10 e.g 2.0×10^{-10}

General form of Scientific notation: $M \times 10^x$ where

$M \geq 1$ but < 10

x is an integer and exponent

Determine M by moving the decimal point in the original number to the left or right so that *only one non-zero-digit* remains to the left of the decimal

Exercise: $1400000 = 1.4 \times 10^6$

$0.0000000000667 = 6.67 \times 10^{-11}$

$5.78 \times 10^8 = 578000000$

$85000000 = 8.5 \times 10^7$

$0.0009 = 9 \times 10^{-4}$

$74000 = 7.4 \times 10^4$

$0.0000005 = 5 \times 10^{-6}$

$30000000 = 3 \times 10^7$

$864000 = 8.64 \times 10^5$

2.1 Adding/Subtracting

In order to add or subtract numbers in scientific notation **your exponents must be the same**

Example:

$$\begin{aligned} &6.3 \times 10^4 + 2.1 \times 10^5 \\ &= 0.63 \times 10^5 + 2.1 \times 10^5 \\ &= 2.73 \times 10^5 \end{aligned}$$

2.2 Multiplying

When multiplying scientific notation you multiple the coefficients and **add the exponents together**

Example:

$$4 \times 10^5 \times 2 \times 10^2$$
$$= 4 \times 2 \times 10^{5+2}$$
$$= 8 \times 10^7$$

2.3 Division

When dividing, divide the coefficients and **subtract the exponents**

Example:

$$(4 \times 10^5)/(2 \times 10^7)$$
$$= 4/2 \times 10^{5-7}$$
$$= 2 \times 10^{-2}$$

Exercise:
$2.4 \times 10^{-3} - 1.2 \times 10^{-2} = 2.4 \times 10^{-3} - 12 \times 10^{-3} = -9.6 \times 10^{-3}$
$(7.4 \times 10^{-8})/(1.2 \times 10^{-2}) = 7.4/1.2 \times 10^{-8+2} = 6.2 \times 10^{-6}$
$3.45 \times 10^4 \times 2.3 \times 10^3 = 3.45 * 2.3 \times 10^{4+3} = 7.94 \times 10^7$
$3.6 \times 10^5 + 7.82 \times 10^4 = 3.6 \times 10^5 + 0.782 \times 10^5 = 4.4 \times 10^5$

3 Measurements and Significant digits

3.1 Accurate and Precise

Precision: the closeness of a set of measurement of the same quantities made in the same way (how well repeated measurements of a value agrees with one another)

Accuracy: is determined by the agreement between the measured quantity and the correct value.

A good example between the difference of accuracy and precision

Precision	Accuracy
Reproducibility	Correctness
Check by repeating measurements	Check by using a different method
Poor precision errors from poor techniques	Poor accuracy results from procedural or equipment flaws

3.2 Percentage Error

Percentage Error $= \frac{\text{Accepted Value} - \text{Experimental Value}}{\text{Accepted Value}} \times 100\%$

3.3 Measurement

The number of Significant digits in a value includes all digits that are certain and one that is uncertain

3.4 Reporting Measurements

- Using Significant figures
- Report what is known with certainty
- Add one digit of uncertainty (Estimation)

3.5 Counting Significant Figures

When you report a measured value it is assumed that all the numbers are certain except for the last one, where there is an uncertainty of ± 1.

Example: the nail is 6.3 6cm long. The 6.3 are certain values and the final 6 is uncertain! There are 3 significant figures in the value 6.36cm (2 certain and 1 uncertain). All measured values will have one (and one only) uncertain number (the last one) and all others will be certain. The reader can see that the 6.3 are certain values because they appear on the ruler, but the reader has to estimate the final 6.

3.6 Rounding

In all cases, round like normal except for when the number ends with only 5.

If digit before 5 is Even: you round down.

If digit before 5 is Odd: you round up.

It is a good practice to write your numbers in scientific notation to better show your significant digits

3.7 Adding/Subtracting

When adding and Subtracting the answer should have the **same number of decimal places** as the one with the **least number of decimal places**.

e.g $12.734 - 3.0 = 9.734- > 9.7$
$13.64 + 0.075 + 67 = 80.715- > 81$
$267.8 - 9.36 = 258.44- > 258.4$

3.8 Multiplying/Dividing

When multiplying the answer should have the same number of significant digits as the significant figure with the **least** significant digits.

e.g $61 \times 0.00745 = 0.45445 = 4.5 \times 10^{-1}$
$608.3 \times 3.45 = 2098.635 = 2.10 \times 10^{3}$
$4.8 \div 392 = 0.012245 = 1.2 \times 10^{-2}$

4 Lab Safety

1. Read and follow all directions exactly as they are written. If in doubt, ask your teacher.

2. Never mix chemicals (or perform tests) without your teacher's permission.

3. Keep your work area clean and keep all materials (clothing, hair, papers, ect.) away from a flame or heat source.

4. Never run or push someone else in the lab. This rule applies at all times.

5. Always wear safety goggles whenever you are working with chemicals or other substances that might get into your eyes.

6. Immediately notify your teacher if any chemicals gets on your skin or clothing to find out what to do to clean it off.

7. Immediately notify your teacher if you get cut or have another injury when performing an experiment.

8. Never reach across a flame.

9. Never look directly into a test tube when mixing or heating chemicals.

10. Always point a test tube away from you and others when heating over a flame or other heat source.

11. Never smell a chemical directly from the container. Wave your hand over the opening of the container and ”waft” the fumes toward your nose.

12. Never taste a chemical unless you are instructed by your teacher to do so.

13. Never use broken or chipped glassware.

14. Keep lids on bottles and containers when not in use.

15. Wash your hands before and after each experiment.

16. Always clean up your work area and equipment after an experiment us completed. Equipment must be returned to its proper place.

5 Atom and Periodic Table

Name	Relative Mass	Electric Charge	Location
Proton	1	+	Nucleus
Neutron	1	0	Nucleus
Electron	1/10000	-	Outside the Nucleus

5.1 Atomic Math

Atomic Number: Number of protons/electrons
Atomic Mass: # of protons + # of neutrons

5.2 Representing Elements

Standard Notation: ${}^{22.990}_{11}\mathrm{Na}$

5.3 Isotopes

Isotopes: atoms of the same element with different numbers of neutrons

5.4 Bhor-Rutherford Diagram

Put number of protons and neutrons in the middle

Draw every electrons

e.g Sodium

5.5 Lewis Diagram

Lewis is simpler than Bhor-Rutherford

Put symbol of the element in the centre

Add valence shell around the symbol

e.g Sodium

$\cdot\overset{\cdot\cdot}{\underset{\cdot}{\mathrm{Na}}}\colon$

6 Isotopes

Isotopes-Atoms that have the *same* number of protons but *different* number of neutrons are called isotopes.

Isotopes have different mass number.

6.1 Average Atomic Mass

Most elements occur naturally as mixtures of isotopes.

The mass number on the periodic table are the weighted average of the most abundant isotopes mass numbers

The atomic mass of an element is a weighted average mass of the atoms in a naturally occurring sample of the element

6.2 Radioisotopes

Some isotopes are stable, others break apart easily.

Difference in stability is due to the composition of the nuclei

Unstable isotopes emit nuclear radiation are known radioisotopes.

When a radioactive isotope breaks apart we get **Radioactive Decay**

Three types

- Alpha particles
- Beta particles
- Gamma rays

Alpha Decay: loses two protons and two neutrons this particle is called Alpha particle. ${}^{4}_{2}\mathrm{He}$

${}^{226}_{88}\mathrm{Ra} \longrightarrow {}^{222}_{86}\mathrm{Rn} + {}^{4}_{2}\mathrm{He}$

Beta Decay: loses one electron this particle is called Beta particles. ${}^{0}_{-1}\beta$

${}^{14}_{6}\mathrm{C} \longrightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}\beta$

Gamma Decay: loses no mass or charge but loses energy this is called Gamma particles. ${}^{0}_{0}\gamma$

${}^{60}_{27}\mathrm{Co} \longrightarrow {}^{60}_{27}\mathrm{Co} + {}^{0}_{0}\gamma$

In terms of size, Alpha particles are the largest, Beta particles are the second, and Gamma particles are the smallest.

Alpha particles can't pass through paper

Beta particles can't pass through aluminum

Gamma particles can partially pass through lead

6.3 Calculating Average Atomic Mass

Basic math, I don't see a reason to write about this

7 The Modern Periodic Table

Groups: vertical columns (1-18)
Periodic: horizontal rows (1-7)
Periodicity: The similarities of the elements in the same group is explained by the arrangement of valence electrons
Atomic Radius: is the distance from the centre of the nucleus of an atom to the outermost electron.

The greater the number of energy levels the greater is the distance of the outermost electron to the centre of it atom's nucleus.

The size of the atomic radius cannot be measured exactly because it does not have a sharply defined boundary. The atomic radius is measured by the

distance of the centre of two nuclei of two atoms beside each other divided by 2.
Trends in atomic radii: atomic radii decreases as you move up in a group, and decreases as you move across a period.

Force of attraction: The attraction of the electrons to the nucleus is what keeps the electrons with the nucleus.
There are two factors that affect the force of attraction

The size of the positive charge determined by the number of protons.
The distance between the outermost electron and the nucleus
The balance exists between the attraction of the electron to the nucleus and the repulsion between the electrons themselves.

The valence electrons receive a positive charge from the nucleus as the inner electrons weaken the attraction for the valence electrons this is called **Shielding Effect**

7.1 Effective Nuclear Charge

A number assigned to elements to describe the amount of shielding to valence electrons

$$\text{ENC} = \text{\# of Protons} - \text{\# of Inner Electrons}$$

The greater the ENC the stronger the attraction to valence electrons
The greater the ENC the smaller atomic radius

7.2 Atomic Radius of Ions

Sodium: Na

$$ENC = 11 - 10 = 1$$

Sodium ion: Na^+

$$ENC = 11 - 2 = 9$$

Cations will have smaller ionic radius that neutral atom.
Anions will have a larger ionic radius than the neutral atom.

7.3 Ionization Energy

The amount of energy to remove the valence electron to form an ion.

Refer to first graph

Moving up the group, it takes more energy.
Moving up the period, it takes more energy.

7.4 Successive Ionization Energy

After the first Ionization Energy (IE) is removed, the successive IE increases as it becomes more difficult to move the next electrons since the pull from the nucleus becomes stronger.

There's a huge jump in IE between removing the last valence electron and removing the first electron in the new ring.

7.5 Electron Affinity

Electron Affinity is the amount of energy released when an electron is captured by an atom to form a negative ion (anion).

$$\mathrm{Cl} + \mathrm{Electron} \longrightarrow \mathrm{Cl}^{-} + \mathrm{Energy}$$

Periodic Trend: Electron Affinity increases as you move up the periods.
Group Trend: Electron Affinity increases as you move to the right of the group.

Electron Affinity and Electronegativity follow the same general Trend.

Halogens are the peaks while noble gases are the troughs.

7.6 Electronegativity

Electronegativity is the tendency an electron would be attracted to a atom when combining with another element.

The scale is does not have a value like km/h but is an arbitrary value between 0 and 4.

In general, **metals** have a **low** EN while **nonmetals** have a **high** EN.

Fun fact: Fluorine has the highest EN at 4.

Refer to the trends of Electron Affinity trend for trend of Electronegativity.

7.7 Reactivity

When looking at **nonmetals**, they lose electrons and the reactivity trend follows the trend of *Electron Affinity*

When looking at metals, they gain electrons and the reactivity trend follows the trend of *Ionization Energy*

7.8 Summary

Name	Definition	Group Trend going down	Periodic Trend going left to right
Atomic Radius	Distance from centre of nucleus to outermost electron (pm)	Increases	Decreases
First IE	Energy to remove outmost electron	Decreases	Increases
Electron Affinity	Energy released when gaining electron	Decreases	Increases excluding group 18
Electronegativity	Tendency to gain electrons	Decreases	Increases excluding group 18
Reactivity Metal	Degree to which metal react	Increases	Decreases
Reactivity Nonmetals	Degree to which nonmetal react	Decreases	Increases excluding group 18

8 Chemical Bonding

All atoms are trying to achieve a **stable octet**

The proton in one nucleus are attracted to the electron of another nucleus based of their electronegativity
Ionic bonding: Forms ionic compounds through atoms giving and taking electrons
Covalent bonding: Forms molecules through atoms by sharing electrons
Metallic bonding: Creates a an excess of electrons as both are positive ions (Mobile Electrons). Allows it to conduct electricity.

Metallic Characteristics:

- High melting point, ductile, malleable, shiny
- Hard substance
- Good conductor as solid and liquid

8.1 Ionic Bonding

Electrons are **transferred** between valence shells of atoms.
Ionic compounds are made of ions
They are called salts or crystals

Always forms between metals and non-metals
Ionic compounds formed a difference in electronegativity of 1.7 or greater

Polyatomic ions always form ionic compounds
Properties

- Hard solids are 22°C.
- High melting point
- Non-conductive as solids
- Good conductors as liquids or when dissolved in water (aq)

8.2 Covalent Bonding

Electrons are **shared** between non-metal atoms.
Electronegativity difference less than 1.7

Covalent bonding forms (two or more) polyatomic ions
Properties

- Low melting point, low boiling point
- softer solid compared to ionic compounds
- Non conductors, period.

8.3 Drawing ionic compounds

Lewis Dot Diagram

Represents the valence electrons as dots
e.g: NaCl

$$[\mathrm{Na}]^+[:\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}:]^-$$

e.g: AlS

$$2[\mathrm{Al}]^{3+}\ 3[\mathrm{:\ddot{\underset{..}{S}}:}]^{2-}$$

Practice

LiF	$[\ \mathrm{Li}\ ]^{+}[\mathrm{:\ddot{\underset{..}{F}}:}]^{-}$
MgO	$[\ \mathrm{Mg}\ ]^{2+}[\mathrm{:\ddot{\underset{..}{O}}:}]^{2-}$
$CaCl_2$	$[\ \mathrm{Cl}\ ]^{2+}2[\mathrm{:\ddot{\underset{..}{Cl}}:}]^{-}$
K_2S	$2[\ \mathrm{K}\ ]^{+}[\mathrm{:\ddot{\underset{..}{S}}:}]^{2-}$

8.4 Drawing Molecules

I don't do these but they might help

Steps

1. Count total valence electrons involved
2. Connect the central atom (usually the first in the formula) to the other with single bonds
3. Complete valence shells of outer atoms
4. Add any extra electrons to central atom

Double Bond are atoms that share 2 pairs (4 total) electrons

$\mathrm{:\ddot{O}=\ddot{O}:}$

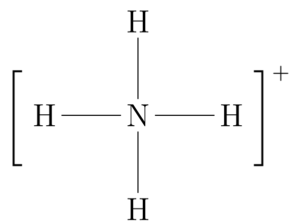
Triple Bond are atoms that share 3 pairs (6 total) electrons

$\mathrm{:N\equiv N:}$

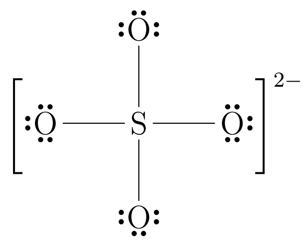
8.5 Drawing Polyatomic Ions

Count all valence electrons needed for covalent bonding
Add or subtract other electrons based on the charge

Ammonium: $NH_4{}^{+}$



Sulfate: $SO_4{}^{2-}$



8.6 Types of Covalent Bonds

Non-Polar Bonds

Electrons shared evenly in the bond
Electronegativity difference less than 0.4

Polar Bonds

Electrons shared unevenly in the bond
EN greater than 0.4 and less than 1.7

When an atom in a polar bond has an electron longer than another atom it's represented at $\delta-$ while the other atom is represented as $\delta+$

8.7 VSEPR Theory

VSEPR: Valence Shell Electron Pair Repulsion Theory

The theory where electron pairs do their best to spread themselves out so that their repulsion is minimal

Bonding Pairs: Forms bonds **Lone Pair**: Do not form bonds

Lone pairs repel just a bit more than bonding pairs

8.8 Shapes of Molecules

Linear
Bent
Trigonal Pyramid
Tetrahedral

8.9 Intermolecular Attraction

Attraction between molecules

van der Waals Forces: Weak attraction forces between non-polar molecules
Hydrogen Bonding: Strong attraction between special polar molecules
Van der Waals

Non-polar can exist in solid and liquid phases as van der Waal forces keep them together

Exists in diatomics and monoatomics.

Periodicity: increases with molecular mass and closer distance between moleules.

9 Intermolecular Forces

Forces can be strong or weak.

Force of attraction dictates how easy it is for gases to turn into liquid and liquid turn into solid

9.1 Forces of attraction

There are two types of forces in liquids and gases

Intermolecular - between molecules

Intramolecular - inside molecule

Intramolecular forces are what makes chemical bonds

Intramolecular forces are stronger than intermolecular forces (30-400x)

Intermolecular forces are known as van der Waals forces

9.2 Types of Intermolecular forces

There are three types of intermolecular forces,

London Dispersion force

Dipole-Dipole force

Hydrogen bond Note that Dispersion is weaker than Dipole which is weaker than Hydrogen bonds

A true covalent bond is 400kcal

9.3 London Dispersion Forces

London Dispersion Forces (Induced Dipole Forces)

Dispersion forces are where molecules have temporary dipoles as a result of two molecules being close to each other.

Dispersion forces are the main forces of non-polar compounds

Using the same idea that you can usually see a pattern in randomness, molecules moving around randomly in the air can be bunched up in one moment and dispersed in another.

As the get closer, they get a temporary dipole.

The larger the molecule the easier δ charges are formed

Molecules that have an even electron distribution are

- Single Atoms
- Molecules of the same element
- Hydrocarbons
- Symmetrical molecules

Trend: As molecular weight increases, so does the boiling point

9.4 Dipole-Dipole Forces

Molecules with permanent dipoles are polar.

Remember polar molecules always have one area in which it is positive and one area in which is negative

Dipole-Dipole forces are stronger than London dispersion forces (2-5 kcal)

9.5 Hydrogen Bond Forces

Hydrogen bond is a special form of Dipole-Dipole force but not a true chemical bond

Strongest intermolecular force

Only exists when

- There is a hydrogen atom in the molecule
- Is with N, O, or F in the molecule (These have the highest EN)

The hydrogen atom is always the positive dipole

N, O, or F is always the negative dipole.

Hydrogen bonding is where you have two said molecules, and the hydrogen of one end bonds with one of the three element of the other end.

Some Hydrogen bonds are stronger than others.

$$\mathrm{F} > \mathrm{O} >> \mathrm{N}$$

HF has a higher boiling point than NH_3

10 Post-Test Notes

In terms of IE, anion(-) ¡ neutral ¡ cation(+)

Review on the melting points of ionic, polar and non-polar molecules (Which has a higher melting point compared to which)

Properties of ionic, polar, non-polar molecules.

11 Chemical Equations

11.1 Word Equations

Is identifies what are the reactants and products of a chemical reaction.

$\mathrm{Sodium} + \mathrm{Chloride} \longrightarrow \mathrm{Sodium\,Chloride}$

11.2 Skeleton Equations

Lists the chemical formula of each reactant on the left and products on the right.

$Na_{(s)} + Cl_{2(g)} \longrightarrow NaCl_{(s)}$

11.3 Law of Conservation of Mass

In a chemical reaction, the mass of the products is always equal to the mass of the reactants

e.g.
Copper(II) Nitrate + Potassium Hydroxide $\longrightarrow$ Potassium Nitrate + Copper(II) Hydroxide

$Cu(NO_3)_{2(aq)} + KOH_{(aq)} \longrightarrow Cu(OH)_{2(s)} + KNO_{3(aq)}$

Atom/Polyatomic Ion	Left Side	Right Side
Cu	1	1
NO_3	2	1
K	1	1
OH	1	2

$Cu(NO_3)_{2(aq)} + 2\,KOH_{(aq)} \longrightarrow Cu(OH)_{2(s)} + 2\,KNO_{3(aq)}$

12 Chemical Reactions

12.1 Synthesis Reactions

Take two or more elements to form a new substance (AKA combination, formation reactions)
$A + B \longrightarrow AB$

There are three types simple synthesis reactions

- Metal/Non-metal reacts with oxygen to make an **oxide**
- Metal and non-metal combine to form a **binary ionic compound**
- Two non-metals for form a **binary molecule**

Examples

- $2\,Fe_{(s)} + 3\,O_{2(g)} \longrightarrow 2\,Fe_2O_{3(s)}$
- $2\,H_{2(g)} + O_{2(g)} \longrightarrow 2\,H_2O_{(g)}$
- $2\,K_{(s)} + Cl_{2(g)} \longrightarrow 2\,KCl_{(s)}$
- $Mg_{(s)} + F_{2(s)} \longrightarrow MgF_{2(s)}$

- $2\,N_{2(g)} + O_{2(g)} \longrightarrow 2\,N_2O_{(g)}$

There are still reactions where there's one or more compounds in the reactant

- When a non-metallic oxide reacts with water, the product is an **acid**
- A metallic oxide reacts with water the product is a metallic hydroxide, this is a **base**

Examples

- $SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(aq)}$
- $CO_{2(g)} + H_2O_{(l)} \longrightarrow H_2CO_{3(aq)}$
- $CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)}$
- $Na_2O_{(s)} + H_2O_{(l)} \longrightarrow 2\,NaOH_{(aq)}$

There are some examples there synthesis is ambiguous.
$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$
$2\,C_{(s)} + O_{2(g)} \longrightarrow 2\,CO_{(g)}$
In this case, don't worry about which is correct. Your teacher should mark them both correct.

12.2 Decomposition Reactions

Where a compound breaks apart into elements/other compounds (Opposite of synthesis)
$AB \longrightarrow A + B$

- **Binary Compound** breaks down into metal and non-metal
- **Compound with a polyatomic ion** breaks down into simple compound and element
- A **Compound with a polyatomic ion** can also break down into two compounds

Examples

- $2\,H_2O_{(l)} \longrightarrow 2\,H_{2(g)} + O_{2(g)}$

Note: Polyatomic ions with a halogen F, Cl, Br, I, At will follow the second rule. A quick way is the oxygen (which is in all of those halogen polyatomics) will be where it gets cut off

$KBrO_3 \longrightarrow KBr + O_2$

Everything else will follow the third rule.

12.3 Combustion Reactions

Reaction with compound and oxygen to form an oxide for both compounds and releases energy (don't quote me on this)

Combustion reactions often go under synthesis, Metals combine with oxygen. It is only combustion only if it releases a ton of energy

$2\,Mg_{(s)} + O_{2(g)} \longrightarrow 2\,MgO_{(s)}$ Yes

$2\,Fe_{(s)} + O_{2(g)} \longrightarrow 2\,FeO_s$ No

Combustion also sometimes involves fuel, made up of carbon and oxygen (hydrocarbons)

These form carbon dioxide and water.

Without enough oxygen, compounds with carbon has incomplete combustion which leads to carbon monoxide and water.

12.4 Single Displacedment

Aka replacement reactions

Where it displaces a element or compound is displaced by different element.

Two types

- Metal replacing metal cation $A + BC \longrightarrow AC + B$
- Non-metal (usually halogen) replacing an anion (-) in a compound $DE + F \longrightarrow DF + E$

Eg:

- One metal replaces another metal
  $Cu_{(s)} + 2\,AgNO_{3(aq)} \longrightarrow Cu(NO_3)_{2(aq)}\,2\,Ag_{(s)}$

- Hydrogen is treated as a metal
  $Mg_{(s)} + 2\,HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2g}$

- Water is treated as an ionic compound
  $2\,Na_{(s)} + 2\,H_2O_{(l)} \longrightarrow 2\,NaOH_{(aq)} + H_{2(g)}$

When analyzing single displacement,

- Treat hydrogen as a metal
- Treat acids (HCl) as ionic compounds in the form of H^+Cl^-
- Treat water as ionic, $H^+(OH)^-$

A reactive metal will displace or replace any metal in a compound that is below it in the activity series (Bottom right of periodic table)
$Fe_{(s)} + CuSO_{4(aq)} \longrightarrow FeSO_{4(aq)} + Cu_{(s)}$
Iron is more reactive than Copper

The same idea follow for non-metals (especially halogens)
$Cl_{2(g)} + 2\,KBr_{(aq)} \longrightarrow 2\,KCl_{(aq)} + Br_{2(l)}$
For halogens

$$F > Cl > Br > I$$

12.5 Double displacement

Double displacement involves the exchange of cations between two ionic compounds usually in aqueous solution
$AB + CD \longrightarrow CB + AD$
Example:
$NaCl_{(aq)} + AgNO_{3(aq)} \longrightarrow AgCl_{(s)} + NaNO_{3(aq)}$
Not all ionic compounds will undergo double displacement
It is double displacement when:

- A solid forms
- A gas is produced
- A molecular compound is formed

Many double displacement involves the formation of a precipitate

$Na_2CO_{3(aq)} + 2\,HCl_{(aq)} \longrightarrow 2\,NaCl_{(aq)} + H_2CO_{3(aq)}$

BUT

$H_2CO_{3(aq)} \longrightarrow H_2O_{(l)} + CO_{2(g)}$

because carbonic acid is unstable.

Would really be

$Na_2CO_{3(aq)} + 2\,HCl_{(aq)} \longrightarrow 2\,NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

$NH_4Cl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + NH_4OH_{(aq)} \longrightarrow NaCl_{(aq)} + NH_{3(g)} + H_2O_{(g)}$

Neutralization reactions are a type of double displacement which produced water.

These involve the reaction of an acid with a base to form water and an ionic compound

$HNO_{3(aq)} + NaOH_{(aq)} \longrightarrow NaNO_{3(aq)} + H_2O_{(l)}$