

Regents Chemistry Information Booklet

Student Name:

1- Classification of Matter

Matter is classified as anything that occupies space (has volume) and has mass.

Matter is classified into two different categories: substance and mixtures

1a. Categories of Matter

- Substances → ECAPS

Elements (cannot be broken down by any means). Compounds (two or more substances combined chemically in fixed proportions). A compound is made of two or more elements, each element is made of many atoms that have the same number of protons in their nucleus but a different number of neutrons, we call these atoms isotopes.

The atomic mass of an element (the mass that we see on the periodic table; usually a decimal #) is the weighted average mass of all that element's naturally occurring isotopes. It is calculated as follows:

$(\% / 100 * \text{mass of isotope A}) + (\% / 100 * \text{mass of isotope b} + \dots)$

- Mixtures → Homogeneous (uniform and consistent particle arrangements) or Heterogeneous. When a solute (example: sugar) is fully dissolved in a solvent (example: water), the substance is homogeneous because it appears as one substance, the solute is uniformly mixed with the solvent. NaCl (aq) is homogeneous. Aqueous (Aq) means mixed in water.

1b. Separation Techniques

- Evaporation: is a physical separation technique used to separate the solute from the solvent in solutions
- Electrolysis: is a chemical separation technique used to separate positive ions from negative ions in ionic compounds. Ex: the separations of Na⁺ from Cl⁻ in table salt
- Distillation: is a physical separation technique used to separate liquids that have different boiling points, liquids with lower boiling points evaporate and condense before liquids that have higher boiling points
- Separation funnel: separates immiscible liquids such as separating oil from water. (more dense water sinks and can be drained from the bottom of the funnel, less dense oil floats to the top and can remain in the funnel.
- Chromatography: separating mixtures based on their speed through a medium (ex: colors in gel)
- Filtration: separates insoluble substances from a solvent. Ex: separation of sand from water. We cannot use filtration to separate Salt from water because both the dissolved salt and the water pass through the filter paper.
- Magnet: we use a magnet to separate iron filing from different mixtures

2- Periodic Table Trends and Atomic Radii

- The periodic table is arranged into periods (left to right rows) and groups (vertical columns)
- Elements on the periodic table are organized according to their increasing atomic number which is the number of protons inside their nuclei. ex: each Na atom has 11 protons in its nucleus thus Na is number 11 on the periodic table.
- Elements in the same group usually have the same number of valence electrons which justifies why they react similarly in a chemical reaction. (they lose/gain the same # of electrons).
- Elements in a period have the same # of electron shells (principle energy levels) that orbit their nuclei. The period number of an element is the same as the number of electron shells present in that element. Going down a group makes you go from period 1 to 2 to 3...etc. therefore, the atomic radius increases from 1 shell to 2 shells to 3...to 7 shells. As you increase atomic # down a group, the number of electron shells increase, and the atomic radius increases. The size of each atom's radius is present on table S in picometers. Going to the right of a period, the # of electron shells remain constant because you are in the same period so you must have the same number of shells, but the atom #s (a.k.a # of positive protons) increase. More protons = stronger nuclear charge (stronger positive charge). The power of attraction between the positive nucleus & the negative electrons increases because more protons attract the same number of electron shells; this justifies why the radius of atoms decreases to the right of a period.
- There are four types of elements on the periodic table: Metals which make up 2/3rd of the periodic table elements: metals are located on the left side of the metalloid's staircase; metalloids which are all the elements above and below the staircase except Al and Po; nonmetals which are the elements on the right side of the staircase; and radioactive isotopes which are elements # 83 & above, those elements do not have a stable nucleus so they decay until they achieve stability.

3- Properties of Elements

3a. Metals

- Metals that are more active than Hydrogen will bond with acids to form flammable hydrogen gas in single replacement reaction. See table J to compare the activities of elements.
- At regents 360, we refer to more active metals as "LOW LOW LOW-LEFT LOSERS" because they have LOW ionization energy, LOW electronegativity, located on the LOWER-LEFT side of the periodic table, and love LOSING their valence electrons.
- Ionization energy is the energy needed to remove an electron from a valence electron shell. Electronegativity is the attraction power for electrons.
- As you approach the lower left side of the periodic table, elements became more metallic in nature. Francium is the most metallic element of all times.
- See the table on page three for more information about the properties of metals

Strong Intermolecular forces (IMF)		
HIGH melting and boiling points	They are all solid at STP except mercury (Hg), it is the only liquid metal at STP	They are dense and insoluble in water

They have free mobile Valence electrons			
Malleable and Ductile	Conduct electricity and heat in solid and liquid phases	lustrous	Low Ionization Energy (low energy is needed to remove electrons) Low electronegativity (low attraction for electrons, SEE TABLE S) LIKELY TO LOSE ELECTRONS

Types of Metals		
<p>Alkali Metals:</p> <p>1- Very Active, never alone, always bonded with other elements.</p> <p>2- Bond with group 17 elements to form salts.</p> <p>3- They have an oxidation # +1</p>	<p>Alkaline earth metals:</p> <p>1- Very active, never alone always bonded with other elements</p> <p>2- Bond with group 17 elements to form salts.</p> <p>3- They have an oxidation # of +2</p>	<p>Transitional metals:</p> <p>1- Extend from group 3 to the metalloids.</p> <p>2- They form colorful solutions when mixed with water</p> <p>3- They usually have multiple positive oxidation states</p>

3b. Nonmetals

- At regents 360, we refer to the most active nonmetals as High High Gainers because they have high ionization energies, high electronegativities and love gaining electrons.
- As you approach the upper right side of the periodic table, elements become more nonmetallic in nature. Fluorine is the most nonmetallic element of all times.
- Nonmetals lack luster, brittle, mostly gases with high entropy (random arrangement of particles) and are poor conductors of heat and electricity.
- All elements that are gases at STP (see table A) are classified as nonmetals. Here is a fun way to remember which elements are gases at STP: "Hydrogen likes to have F.O.N with Chlorine and Group 18.
- There are a few solid nonmetals, but Bromine is the only liquid nonmetal.
- Nonmetal generally have low IMF therefore low mp/bp and low densities, high electronegativities and high ionization energies (see table S).
- Group 17 nonmetals are classified as halogens. they combine with metals to form salts. Ex: NaCl
- Group 18 nonmetals are known as noble gases. (they do not react with other elements because they have stable electron configuration (octet stability = chemical stability= 8 electrons in the last electron shell)

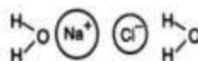
3c. Semimetals (Metalloids)

These are elements which possess properties of both metals and nonmetals.

They are semiconductors, so they are mostly used in electronics to prevent power surges.

3d. Random Facts

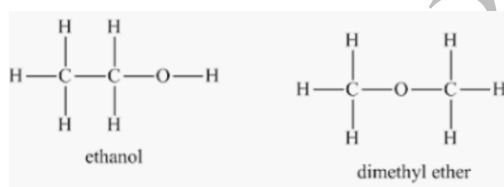
- When combined chemically, metals and nonmetal form ionic compounds which dissociate in water to conduct electricity in aqueous form.



- To prevent rusting, manufacturers combine iron with nickel or magnesium to make it more resistant to corrosion.
- Aluminum or zinc form a self-protective coating when oxidized to protect themselves from corrosion.
- Density of elements can be calculated by dividing the element's mass by its volume. See table S for densities.

3e. Allotropes vs. Isomers

- Allotropes: different forms of an element, ex: oxygen exists as O₂ (oxygen gas) and O₃ (Ozone gas); carbon exists as charcoal, diamonds, and graphite.
Allotropes have different structures therefore different physical and chemical properties
- Isomers: Compounds which have the same chemical formula, same percentage of elements make up both compounds but different functional groups (see table R) thus different structures, thus different physical and chemical properties.



Example of isomers:

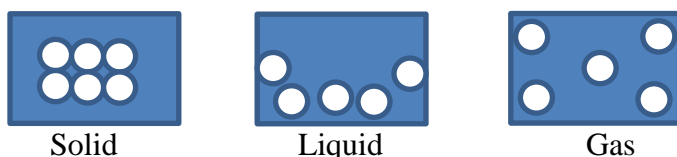
4- Physical Change

4a. Physical Change and Physical Properties

- Physical change is a reversible process. It does not result in the creation of new substances. It is merely a change in the appearance of a substance.
- Common key words used to describe physical change/physical properties: phase change, crush, dissolve, conductivity, density, solubility, malleability, luster, magnetism, and any other description of an object's appearance.

4b. Particles Diagrams

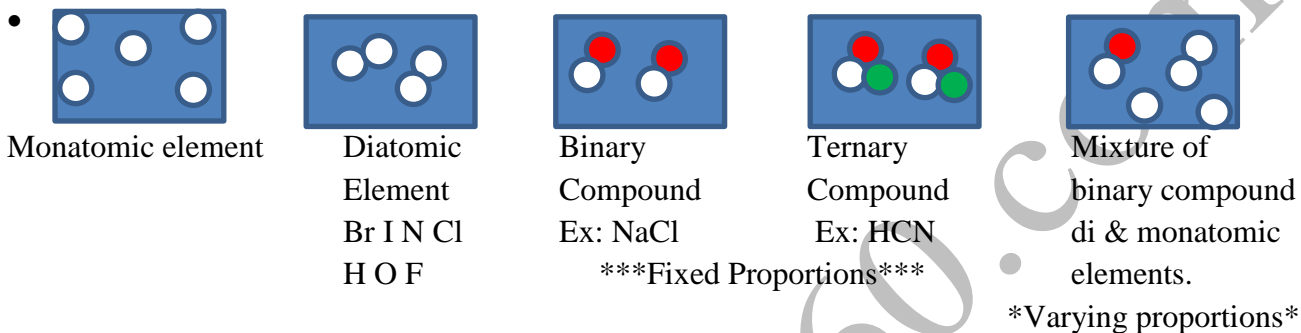
- Particle diagram may be used to represent a phase change



- Melting is changing a solid to a liquid
- Freezing is changing a liquid to a solid
- Evaporation is changing a liquid to a gas

- Condensation is changing a gas to a liquid
- Sublimation is changing a solid to a gas (ex: dry ice sublimates because it has weak intermolecular forces. $\text{CO}_2(s) \rightarrow \text{CO}_2(g)$)
- Deposition is changing a gas to a solid (ex: Snow formation. $\text{CO}_2(g) \rightarrow \text{CO}_2(s)$)

- Particle diagrams may be used to represent substances or mixtures



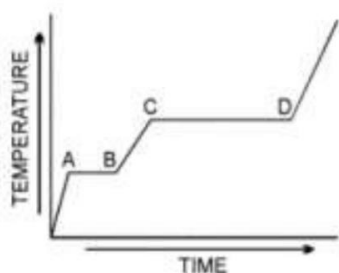
4c. Table H

Table H presents the relationship between vapor pressure and temperature

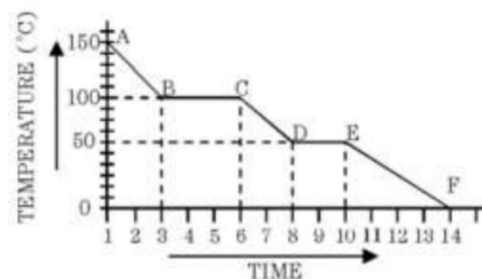
- There is an indirect relationship between vapor pressure of a substance and its boiling point; the boiling point of a substance is the point at which the dotted line on table H intersect with the curve of the substance. Water boils at a temperature of 100 degrees therefore, water has a lower vapor pressure than ethanol (the alcohol present in alcoholic beverages) which boils at a temperature of 80 degrees.
- The dotted line is atmospheric pressure, boiling occurs when the vapor pressure of a liquid equals or exceeds surface pressure.
- Substances that boil at high temperatures have stronger intermolecular forces than substances that boil at lower temperatures. The higher the electronegativity difference between the particles that make up a substance, the higher is the substance's intermolecular forces. See table S for the electronegativity value for each element. The electronegativity difference between H and F in the compound HF is higher than between H and I in the compound HI therefore HF has stronger intermolecular forces and a higher boiling point (67°F) than HI (-31°F).
- Temperature is a measure of the average kinetic energy of particles that make up a substance (how fast particles move). The faster particles move, the higher is the temperature.
- Zero degrees kelvin is the only temperature at which there are no movement of particles. Zero kelvin is known as absolute zero. 0K? cool. See table T for Celsius to Kelvin conversions

4d. Heating and Cooling Curves

- Heating and cooling curves are used to represent phase changes and changes in temperature.
- In a heating curve, we notice horizontal lines that represent phase changes. During phase changes, the kinetic energy (temperature) remains constant, but potential energy changes. This is because heat is needed to break apart the bonds that keep atoms together.
- Phase changes represent a change in the potential energy of a substance. If the phase change is endothermic, the potential energy increase. If the phase change is exothermic, the potential energy decreases.
- Endothermic phase changes: these are phases that require some sort of energy (heat, light, or electrical energy), for example: melting, boiling, and sublimation. Exothermic phase changes release energy, for example freezing, condensation, depositing.
- a phase equilibrium occurs in sealed containers when the rate of evaporation equals the rate of condensation or when the rate of freezing = the rate of melting.
- In a heating curve, we also notice vertical slanted lines which represent a rise in temperature (average kinetic energy) or a drop in temperature in cooling curves.
- Systems in nature will spontaneously favor lower enthalpy (exothermic reactions) and higher entropy. Enthalpy is heat and entropy is the level of disorder and random arrangement of particles in a system. Gases have high entropy and have no fixed shape or volume (they uniformly occupy any size container); the lower the temperature the lower is the entropy, solids have a definite shape, definite volume, and very low entropy.
- Energy will always flow from hotter substances to colder substance. This is the direction of heat under normal conditions.



Heating Curve



Cooling Curve

- On the heating curve shown, the horizontal line segments A-B and C-D represent the phase changes melting and vaporization respectively. The three other slanted lines represent phases of matter (solid, liquid, and gas).
- Point A represent the melting point. The temperature when a substance **begins** to melt. Point C represent the vaporization (boiling) point, the temperature when a substance begins to **boil**.

- On the cooling curve shown, the horizontal line segments BC and D-E represent the phase changes condensation and freezing respectively. The three other slanted line segments A-B, C-D, and EF represent the phases of matter gas, liquid, and solid respectively. Point B represent the condensation point which equals the melting point which is 100 degrees Celsius.
- During a phase change, the temperature does not change. According to the graph, the substance begins condensing at 3 minutes and stops condensing at 6 minutes, this means that the substance takes 3 minutes ($6 - 3 = 3$) to condense. Point D and E represent the freezing point and melting points respectively. The substance begins to freeze at 8 minutes and freezes completely at 10 minutes. The substance therefore takes 2 minutes to freeze. The line segment BC is longer than the line segment DE because more energy is required to boil or condense a substance than to melt or freeze a substance. The heat of vaporization (H_f) of water according to table B is 2260joules/1 gram, while the heat of fusion (H_f) of water is only 334joules/1gram.

4e. Calculating Heat

- We can calculate the heat needed to be absorbed or released to change the phase of a substance by following the formulas on table T. $q = H_f \cdot m$ for melting/freezing, and $q = H_v \cdot m$ for vaporization/condensation. The formula $q = mC(T_{\text{final}} - T_{\text{initial}})$ is used to measure the amount of heat needed to be absorbed to raise the temperature of a substance or the heat needed to be released to drop the temperature of a substance by a certain amount of degrees.
- Table B shows the H_f , H_v , and C values **of water**. If you are not given a specific heat capacity value “C”, or an “ H_f ” or “ H_v ” values in a question, you may use these values to perform the necessary calculations and answer the question.
- A negative heat value indicates that energy is released, a positive heat value indicates that energy is absorbed. All exothermic reactions should yield a negative heat value; all endothermic reactions should yield a positive heat value. See the heat of reaction delta H on table i to determine if a reaction is endothermic or exothermic. Endothermic reactions will have the term energy on the reactants side (left side of the arrow), exothermic reactions will have the term energy on the products side (the right side of the arrow).
- Heat is measured in unit joule or calorie. Refer to table C for unit conversion factors or table D for unit abbreviations. The standard international (SI) units are liters, meters, and grams. Table C examples: 5 millimeters = $5 \cdot 10^{-3}$ meters; 5 kilometers = $5 \cdot 10^3$ meters; 5 kilometers = $5 \cdot 10^6$ millimeters because the difference between millimeters 10^{-3} and kilometers 10^3 is 10^6 .

4f. Significant Figures

Rounding answers to the correct number of significant figures can be done using the P.A. rule

- P stands for present decimal. If a number has a decimal present, we approach the number from the left side and consider the first nonzero digit we see, and all the other digits that follow it to be significant. Example: 0.000120 is three significant figures.

- A stand for absent decimal. If a number does not have a decimal, we approach the number from the right side and consider the first nonzero digit we see, and all the other digits that follow it to be significant. Example 12000 is only two significant figures
- When dividing or multiplying significant two numbers, always round your answer to the least number of significant figures possible. Example 12.0 (three significant figures) / 5 (one significant figure) = 2.4 but we round our answer to one significant figure because out of the two numbers that we divided, the least number of significant figures is one. Our final answer will be rounded 2.

5- Atomic Structure

5a. Atomic Theories

- Dalton: the atom is a positive hard sphere which contains indivisible and identical particles.
- JJ Thomson: conducted the cathode ray tube experiment which lead him to discover that electrons are scattered in the atom like plum is scattered in pudding. He referred to his model of the atom as the plum pudding model.
- Rutherford: used alpha particles to conduct the gold foil experiment which lead him to discover that the atom is mostly empty space and the atom has a small, dense, positive nucleus that makes up most of the atom's mass.
- Bohr: proposed the planetary model. The atom contains a nucleus that is surrounded by electrons which are located in orbits (shells / principle energy levels). The electron capacities of the first 3 electron shells are 2,8,18 respectively. $(2n)^2$ rule where n is the electron shell's number. $2(1)^2= 2$ electrons is the capacity of the first shell).
- Quantum mechanical theory (modern atomic theory) proposed the cloud model of the atom. the atom contains a nucleus surrounded by electrons that are located in orbitals (orbitals are locations where electrons are most likely to be found). High energy electrons are likely to be located in furthest from the nucleus while low energy electrons are likely to be located closer to the nucleus.

When electrons gain energy, they move to higher energy levels temporarily. Electrons release the energy they gained as they return from high energy excited states to low energy ground states. The energy released is in the form of light, each atom produces a unique light spectrum, we can identify what elements are present in other planets by using a spectroscope. Ex: sodium ground electron configuration is 2-8-1 (see periodic table), sodium's excited electron configuration occurs when an electron from the second to last shell moves to the last shell; the new electron configuration is 2-7-3.

5b. The Structure of the Atom

- Inside every atom there are subatomic particles those are protons, electrons, and neutrons. their notations are located on table O. The number at the bottom of the notation represents the charge of each particle, the number above the notation represents the mass.
- ePNA#: The atom # = # of protons inside the nucleus = the nucleus charge = the # of electrons. For all elements on the periodic table, the # of electrons (negatively charged particles) = the # of protons (positively charged particles). This indicates that all unbonded monatomic elements or diatomic elements (Br.I.N.Cl.H.O.F) are neutral and have an oxidation number (charge) of zero.
- Except for noble gases (group 18 elements), all atoms on the periodic table are not chemically stable. Atoms become chemically stable when they achieve an octet which is the presence of 8 valence electrons (8 electrons in the last electron shell that surrounds the nucleus). Nonmetals will gain electrons (gain negatively charged particles to become more negative and achieve an octet); metals will lose electrons (lose the negatively charged particles to become more positive). Generally, when bonded, metals have positive charges, nonmetals have negative charges. The periodic table shows the potential charges of atoms when bonded, example: Fe can be +2 or +3. This means that Fe can lose 2 electrons or 3 electrons to achieve an octet.
- When an atom loses or gains electrons, it becomes an ion. The net charge of ions = # of Protons - # of electrons ($P_{\#} - e_{\#} = C$). REMEMBER: The # of protons will always equal the atom # (protons are prisoners chained to the atom #). Ex: the # of electrons in Fe^{+2} ion is $26 - e = +2$. $e = 24$. Positive charge = Fewer e; negative charge = Additional e.
- Atoms bond to achieve chemical stability. There are three types of bonding: ionic bonding is between a metal and a nonmetal, Covalent bonding is between two or more nonmetals, a hydrogen bond is a form of a covalent bond that occurs when hydrogen bonds with F, O, or N. a hydrogen bond is a very strong bond. The third type of bonding is metallic.

6- Naming and Writing Formulas

6a. Types of Formulas

There are 3 types of formulas

- Molecular formula which contains the symbols and the # of moles of each atom, for example: N_2 . This can be calculated by using the following formula:

$$\left\{ \frac{\text{given molecular mass}}{\text{e.m.m.t empirical mass}} \right\} \times \text{given empirical formula} \dots M/E * E$$
 - Empirical formula is the molecular formula simplified. It can be calculated as follows (molecular formula / greatest common factor = empirical formula)
 Empirical formula can also be calculated if given the % composition of each atom.
 A compound contains 53% Al and 47% O by mass. What is the empirical formula of this compound?
 $\rightarrow 53 \text{ grams of Al} = 53 / (27, \text{ the mass of Al}) = 2 \text{ moles} \rightarrow 47 \text{ grams of O} = 47 / (16, \text{ mass of O}) = 3 \text{ moles}$
 $\rightarrow \text{the formula is } Al_2O_3$
 - Structural formula: this formula shows how atoms are bonded to each other. For example: $N \equiv N$
- Emmt is just a fun way to calculate the mass of a compound: **e**lement symbols, **m**ass of each element * how **m**any atoms of that element are present = **T**otal mass of the compound.

Example: N₂O.

E	M	M	T
N	14	2	28
O	16	1	16
			28+16 =44grams/1 mole

6b. Naming and Writing Formulas for Ionic Compounds

Writing Rules

Naming Rules

Binary ionic → 2 elements	Ternary ionic → 3 or more elements
Always end in IDE	Table E determines endings
If a metal has multiple charges, reverse crisscross to identify the charge of the metal and use roman numerals Fe ₂ O ₃ Iron III Oxide	If a metal has multiple charges, reverse crisscross the number outside the parentheses and use roman numerals to determine the charge of the metal Fe ₂ (SO ₄) ₃ Iron III Sulfate
If metal has multiple charges but there is no number to crisscross, the charge of the nonmetal is the roman numeral for the metal FeO Iron II Oxide	If a metal has multiple charges but there is no number outside the parenthesis to crisscross, the charge of the polyatomic ion (from table E) is the roman numeral of the multi-charged metal. Fe ₂ SO ₄ Iron II Sulfate

Binary ionic → 2 elements		Ternary ionic → 3 or more elements	
If the metal has a single charge the periodic table will determine its charge		If the metal has a single charge the periodic table will determine its charge	
Roman numeral determines the charge of the metal		Roman numeral determine the charge of the metal	
Periodic table determines the charge of the nonmetal...take the first one on top		Table E, determine the charge of the polyatomic ion	
		Always put table E polyatomic ion in between parenthesis	
Magnesium Fluoride		Magnesium nitrate	
Ions formula	Mg ⁺² F ⁻¹	Ions formula	Mg ⁺² (NO ₃) ⁻¹
Compound	MgF ₂	Compound	Mg(NO ₃) ₂
Iron III oxide		Iron II Sulfate	
Ions formula	Fe ⁺³ O ⁻²	Ions formula	Fe ⁺² (SO ₄) ⁻²
Compound	Fe ₂ O ₃	Compound	Fe(SO ₄)

6c. Naming and Writing Formulas for Molecular Compounds

Naming Rules

Molecular
Use prefixes before every element Never use the prefix mono before the first element
Prefixes Mono-di-tri-tetra-penta-hexa-hepta-octa-nona
N ₂ O DiNitrogen monoxide
NO ₂ Nitrogen dioxide → Notice I didn't give nitrogen the prefix mono because it is the first element.

Writing Rules

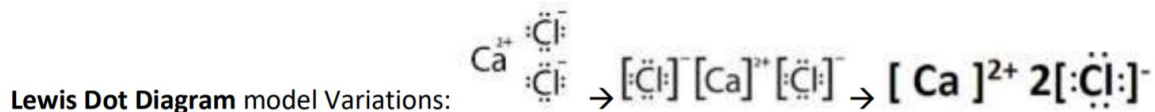
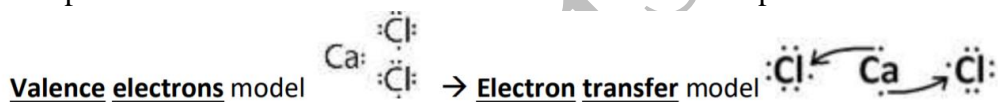
Molecular
Prefixes become the subscripts after the element symbol
Examples: Carbon tetrachloride CCl ₄ Dinitrogen monoxide N ₂ O

6d. Drawing Lewis Dot Diagrams of Atoms, Ions, and Molecular Compounds

- Atoms: to draw the Lewis diagram of an atom, all you need is a symbol and valence electron dots. Ex: **Na•**
- Ion: to draw the Lewis diagram of an ion, you need a symbol, a charge, and 8 valence electron dots for nonmetals, you do not need valence electron dots for metals. A negative charge means an atom gained additional electrons; a positive charge means the atom lost its valence electrons.

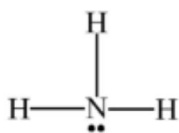


- Ionic compounds: these compounds are made of metals and nonmetal. You must crisscross the charges then follow BVICS, (brackets, valence electrons (8 for nonmetals, non for metals), ionic charges, coefficients as represented by the subscripts in the compound formula, and symbols). A Lewis dot diagram of an ionic compound must show that electrons were transferred from the metal who lost its valence electrons to the nonmetal who gained those valence electrons. Example: Calcium Chloride. Ion formula: $\text{Ca}^{+2}\text{Cl}^{-1} \rightarrow$ Compound formula: CaCl_2

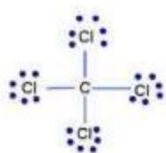


- Molecular Compounds: these compounds are made of nonmetals only. You must align valence electrons of atoms appropriately so that they bridge with valence electrons of other atoms to achieve chemical stability (an octet). Exceptions to the octet rule is the element hydrogen which only needs two electrons to achieve chemical stability. A single bond between two atoms represents a pair of electrons, one electron comes from each atom. in a covalent bond, electrons are shared between two nonmetals. Atoms in molecular compounds are arranged to form different structures. Bent structure is for water, Linear structure is formed when two or three atoms bond. Trigonal pyramidal structure is formed when four atoms bond, tetrahedral structure is formed when five atoms bond.

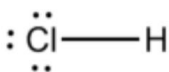
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Trigonal Pyramidal Structure of Nitrogen trihydride. Nitrogen has 5 valence electrons, 3 of those valence electrons will mate with 3 hydrogen atoms by building a bridge (forming a bond). This helps nitrogen gain access to 3 more electrons from the hydrogens and achieve chemical stability for nitrogen, (8 valence electrons). At the same time, Hydrogen has 1 valence electron so that 1 electron will mate with one of the 3 nitrogen's mating electrons to achieve chemical stability for hydrogen (2 valence electrons).



Tetrahedral structure of carbon tetrachloride. Carbon has four valence electrons so it will bond four times to achieve chemical stability (octet rule). Chlorine has 7 valence electrons so it will bond once to achieve chemical stability (octet rule).



Linear structure of Hydrochloric Acid. Chlorine has 7 valence electrons, so it will bond once. Hydrogen has 1 valence electron, so it will bond once. Reminder, hydrogen only needs one additional electron to achieve chemical stability, therefore hydrogen is most likely to only bond once.

6e. Organic Chemistry (Tables P,Q, and R).

- When naming organic compounds, we use table R to determine the prefix that represents the number of carbons in the longest continuous carbon chain of the given structural formula.
- Table Q determines the ending of the prefix if a hydrocarbon compound was given (a compound that is made of hydrogen and carbon only). Saturated hydrocarbons are made of hydrogen and carbon only and all carbons are single bonded with each other. These hydrocarbons are classified as alkanes and their names end in ane, example: propane is made of three carbons and all those carbons are single bonded. In an unsaturated hydrocarbon, there is at least one double or one triple bond between the two carbons. These carbons are classified as alkenes or alkynes. Example: 2-pentene represent a five carbon chain that contains a double bond located after the second carbon.
- Table R determines the name of a functional group and carbon chains' endings. Example: if a hydrocarbon chain has an OH extension, it is classified as an alcohol and its name would have an ol ending. If the hydrocarbon has an OOH extension, we refer to it as an organic acid. And the name would have an oic acid ending. Example: Propanol is a three carbon chain with an OH at the end of it.

More details on the next page...

Naming of Organic Compounds

Step 1: given the structural formula, find the longest continuous carbon chain. Count the number of carbons in that chain and use table P to assign the appropriate prefix for the chain. For example: if the chain has four carbons, it would begin with the prefix But

Step 2: Number the carbons in the longest continuous carbon chain, starting from the end of the chain that is closest to the attached functional group, or double or triple bond if they exist.

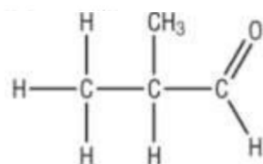
Step 3: If a functional group exists identify the location of the functional group and the prefix for that functional group according to table R examples.

For example, a CH_3 functional group has the prefix methyl, an alcohol group is anol, an aldehyde is anal, a ketone is anone, an organic acid is anoic acid...etc.

Step 4: If a double or triple bond exist, identify the location of that double or triple bond and refer to table Q to identify the proper prefix used to describe the bond. For example: a double bond has the prefix ene, a triple bond has the prefix yne.

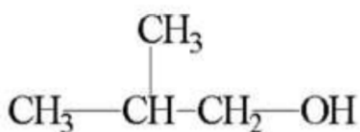
Naming Example:

2 Methyl Propanal

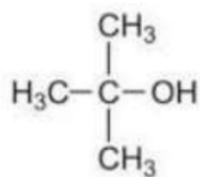


- 1- The longest continuous carbon chain is made of 3 carbons which have the prefix prop. According to table P. The chain has an aldehyde functional group attached, according to table R, the proper prefix for an aldehyde is anal.
- 2- There is another functional group attached to the second carbon, it is CH_3 , which has the prefix 2, methyl. The number two must be inserted is to identify the location of this methyl group.
- 3- All together, the name is **2 methyl propanal**.

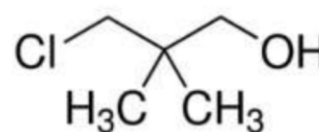
More Ex:



2 methyl, 1 Propanol



2 methyl, 2 propanol



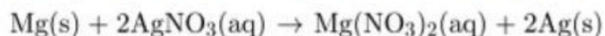
2,2 dimethyl, 3 chloro, 1 propanol

7- Chemical Change

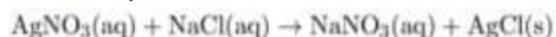
7a. Types of Chemical Reactions

- Synthesis: 2 elements react to synthesize a new compound
- Single replacement \rightarrow when a single element and a single compound react. {must use table J to see if reaction takes place spontaneously} \rightarrow only happens if the single element is more active than the element bonded in the compound.

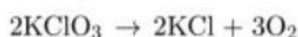
ONLY positive REPLACES positive, ONLY negative replaces negative.



- Double replacement \rightarrow starts with 2 different compounds



- Decomposition \rightarrow starts with a compound that decomposes to simpler types of matter



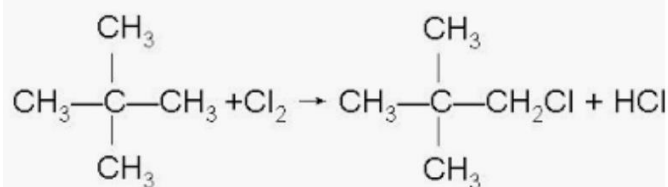
- Electrolysis: using electrical energy to generate a chemical change. Ex: decomposing compounds
- Neutralization: is following titration procedures to combine an acid (table K) and a base (table L) to yield salt + water. $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$. Titration helps us calculate the concentration (molarity) of solutions.

7b. Types of Organic Reactions

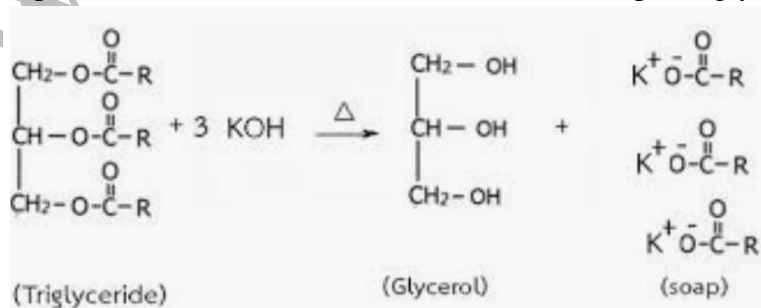
- Addition: when an alkene reacts with a group 17 molecule



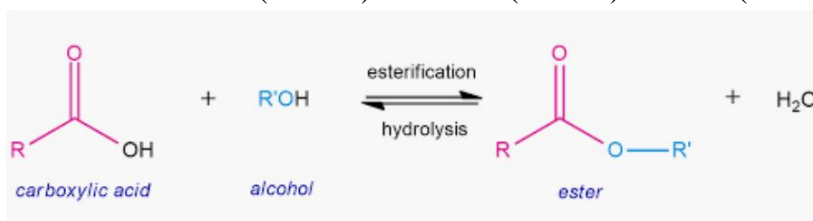
- Substitution: when an alkane reacts with a group 17 molecule



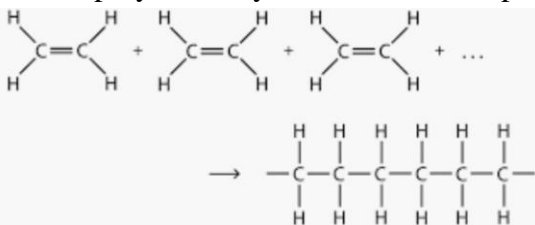
- Saponification: when Fat + Base (table L) \rightarrow Soap and glycerol (R-OH)



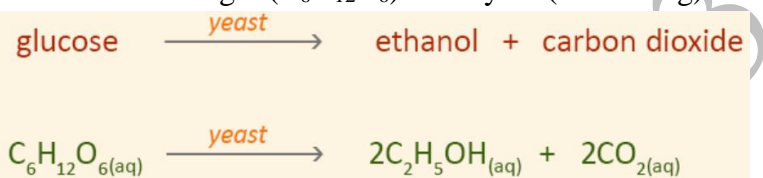
- Esterification: Acid (table K) + alcohol (table R) \rightarrow ester (table R) + water



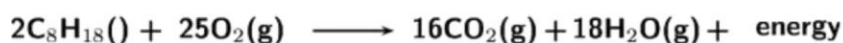
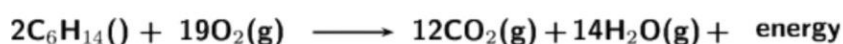
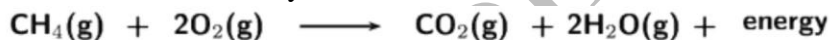
- Polymerization: $n(\text{CH}) \rightarrow (\text{CH})_n$: Repeating hydrocarbon chain. Cellulose and protein are natural polymers. Nylon is an artificial polymer



- Fermentation: Sugar (C₆H₁₂O₆) + Enzyme (ase ending) \rightarrow alcohol + CO₂

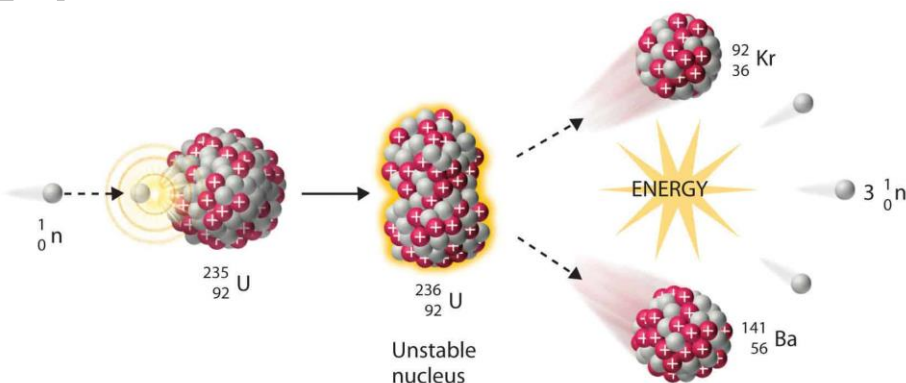


- Combustion: when a hydrocarbon + O₂ \rightarrow CO₂ + H₂O

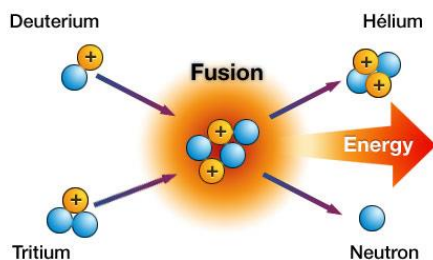
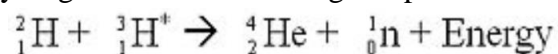


7c. Types of Nuclear Reactions

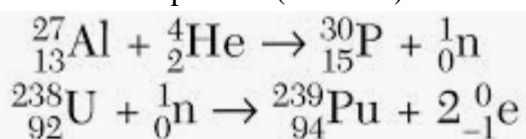
- All nuclear reactions are exothermic (they release lots of energy)
- Fission: when a uranium-235 atom is bombarded with a neutron to release lots of energy



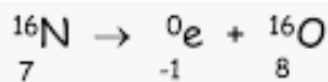
- Fusion: When two light hydrogen isotopes combine to form a heavier helium isotope and release lots of energy. This process requires lots of energy because it is difficult to combine two positive hydrogen nuclei. Like charges repel and both hydrogen nuclei are positive.



- The mass, energy, and charge of all matter is conserved in chemical reactions but in nuclear reactions, the mass of the products is usually less than the mass of the reactants because mass was converted to energy.
- Artificial transmutation: occurs when a particle from table O is accelerated to bombard with a radioactive isotope to produce a new isotope. We call this artificial transmutation because we accelerated a particle (artificial) to create a change (mutation).



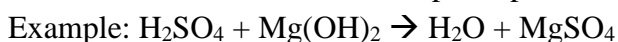
- Natural mutation: when an atom naturally decays and transforms into a new atom. this is natural because it occurs naturally without human interference.



8- Balancing Equations

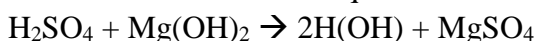
8a. Balancing Chemical Reactions

- Prior to conducting any stoichiometric calculations all equations must be balanced using coefficients. The previously shown combustion reactions are all examples of balanced reactions.
- Here are some tips for balancing reactions. Always leave lonely elements to be the last ones balanced, you may always use a .5 decimal to balance a single element, then multiply the entire reaction by 2, to change the decimal to a whole number.
- Another tip for balancing reactions. If polyatomic ions exist on both side of a reaction, balance them as one ion and do not attempt to split them.



Since OH exists on one side of the reaction and H₂O exists on the other, this reaction can be rewritten as $\text{H}_2\text{SO}_4 + \text{Mg}(\text{OH})_2 \rightarrow \text{H}(\text{OH}) + \text{MgSO}_4$

Now we can balance the equation as follows:



Notice the numbers of SO_4 s and OHs are equal one both sides of the reaction, also the number of Hs and Mgs are equal as well. There are 2 OHs on the reactant side, and 2 OHs on the product side (the coefficient 2 distributes to the entire compound $\text{H}(\text{OH})$).

8b. Balancing Nuclear Reactions

- When balancing a nuclear reaction remember that the total charge of the reactants must equal the total charge of the products (the charge is the number that is located at the bottom of the element's notation). Additionally, the total mass of the reactants must equal the total mass of the products (the mass is the number located at the top of the element's notation). See previously shown natural transmutation and artificial transmutation examples.

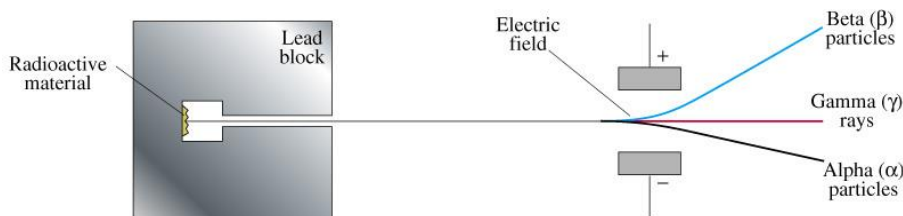
8c. Balancing Charge

- When balancing an equation, the net charge of the reactants must equal the net charge of the products and the total number of moles (coefficients) for each atom in the reactants must equal the total number of moles for each atom in the products.
 Ex: $3\text{Ca}^0 + 2\text{Al}^{3+} \rightarrow 3\text{Ca}^{2+} + 2\text{Al}^0$ notice the net charge on the reactant's side is +6 [(2*3) + 0] and the net charge on the products side is also +6 [(3*2) + 0]. The # of moles for Al is 2 on each side of the reaction and the # of moles of Ca is 3 on each side of the reaction. This equation is balanced.

9- Nuclear Chemistry

9a. Introduction to Nuclear Chemistry

- Table N and table O are used to answer most nuclear chemistry questions. table O shows the most commonly decayed particles alpha, beta, and gamma. Alpha has the most mass therefore the lowest penetration power. Alpha also has the highest charge therefore the most ionization power. Gamma on the other hand has the lowest charge and the lowest mass therefore it has the highest penetration power and the lowest ionization power (it is not harmful to humans)
- Radioactive particles can be separated when they pass through an electric field that contains positive and negatively charged plates. The positive plate attracts negatively charged particles such as beta; the negatively charged plate attracts positively charged particles such as alpha. Gamma and neutron go through the field without being attracted to anything because they are both neutrally charged.
- In order of increasing penetration power, particles are organized as alpha, beta, gamma. In order of increasing ionization energy, particles are organized as gamma, beta, alpha



- Isotopes which number of protons does not equal their number of neutrons are known to have unstable radioactive nuclei. We can calculate the number of neutrons of an isotope by using the man formula. Mass – Atom# = neutron #, for example there are 143 neutrons in the nucleus of U-235. $235 - 92 = 143$. The # of an atom on the periodic table = the number of protons in that atom's nucleus. U-235 has an unstable nucleus because the # of protons (92) does not equal the number of neutrons (143).
- Radioactive isotopes will decay naturally by releasing one of the particles present on table O. - Different radioactive isotopes release different particles as shown on table N.

9b. Characteristics of Fission and Fusion Reactions

	Fusion	Fission
Description	Combines two light hydrogen isotopes to make a heavier helium atom	Bombards a neutron with a uranium isotope to split the uranium nucleus into two lighter nuclei
Reactions	${}^2_1\text{H} + {}^3_1\text{H} \rightarrow {}^4_2\text{He}$ ${}^2_1\text{H} + {}^3_1\text{H} \rightarrow {}^4_2\text{He} + {}^1_0\text{n}$ $4{}^1_1\text{H} \rightarrow {}^4_2\text{He} + 2{}^0_{-1}\text{e}$	${}^{235}_{92}\text{U} + {}^1_0\text{n} \rightarrow [{}^{236}_{92}\text{U}] \rightarrow \begin{cases} {}^{141}_{54}\text{Xe} + {}^{90}_{38}\text{Sr} + 2{}^1_0\text{n} \\ {}^{144}_{55}\text{Cs} + {}^{89}_{37}\text{Rb} + 2{}^1_0\text{n} \end{cases} + 3{}^1_0\text{n}$
Requirement	High temperature is required to initiate the reaction. Millions of degrees kelvin.	Requires less energy to take place
Energy produced in KJ	Releases the most energy. More than a 100 million kilojoules for every mole of helium produced.	Releases less energy than fusion reactions but much more than any chemical reaction. 1 kg of U-235 releases more energy than 20,000 kg of coal.
Waste	Clean process, does not produce waste	Lots of radioactive waste is produced
Environmental concerns	Not possible on earth due to the massive amount of energy produced.	Requires large storage facilities to store radioactive waste for long periods of time until declared safe to dispose into the environment
Uses	Sun and Stars	Helps us generate electricity Nuclear Bomb

9c. Uses of Radioactive isotopes

- X-Rays are very similar to Gamma Rays
- Radioactive isotopes that are used medically have short half-lives, this aids in their quick elimination from the body
- The ratio of U-238 to Pb-206 is calculated for geological dating to determine the age of earth.
- C-14 traces can be used to determine the age of fossils
- Gamma rays can be used to kill microorganisms and sterilize canned foods
- U-233, U-235, and Pu-239 are all used as nuclear fuels or fissionable isotopes.
- A person can protect themselves against nuclear radiation if they hide in steel or high-density concrete containers.

- Boron and Cadmium are commonly used in control rods to help control the fission process by regulating the number of neutrons available for fission.
- Heavy water and graphite are used as moderators to control the speed of neutrons inside a nuclear reactor.
- N-16 is a gas waste product produced from fission reactions.
- Liquid sodium and heavy water are used as coolants to control the temperatures of a nuclear reactor.
- C-14 is used for fossil dating
- I-131 is used to treat thyroid disorders
- Co-60 is used to treat Cancer
- Tc-99 is used to diagnose brain cancer
- Am-241 is used in smoke detectors because it has a long half-life
- Cs-137 is used to irradiate canned food to sterilize it and increase its shelf life.

9d. Half-life calculations

0: Number of Half Lives	0: Time associated with number of half lives	Og: Original Mass of radioactive isotope	1: Original Fraction of isotope prior to decaying
1	The time for 1 Half-life is often given or found on table N. it may also be unknown		1/2 of the original mass remains after 1 half life
2			$\frac{1}{2} * \frac{1}{2} = \mathbf{1/4}$ of the original mass remains after 2 half lives
3			$\frac{1}{2} * \frac{1}{2} * \frac{1}{2} = \mathbf{1/8}$
4			$\frac{1}{2} * \frac{1}{2} * \frac{1}{2} * \frac{1}{2} = \mathbf{1/16}$
5		Remaining Mass	

Example 1

If $\frac{1}{8}$ of an original sample of krypton-74 remains unchanged after 34.5 minutes, what is the half-life of krypton-74?

- (1) 11.5 min (3) 34.5 min
 (2) 23.0 min (4) 46.0 min

0	0	Og	1
1			1/2
2			1/4
3	34.5 minutes		1/8

Answer: 34.5 minutes are per 3 half-lives. 34.5 minutes / 3 = 11.5 minutes are per 1 half-life.
No need to complete the mass column since this question has nothing to do with mass.

Example 2:

How much time is needed for 200 grams of radon-222 to decay to 3.125 grams?

0	0	Og <input type="text" value="200"/>	1
1	Table N: one half life of radon = 1599 years	100	1/2 of the original mass
2	2 * 1599	50	1/4 of the original
3	3 * 1599	25	1/8 of the original
4	4 * 1599	12.5	1/16 of the original
5	5 * 1599	6.25	1/32 of the original
6	6 * 1599 = 9594 years pass	3.125g mass remain	1/64 of the original

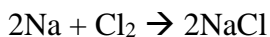
10- Stoichiometry

10a. Concentrations of Elements in Compounds

- % composition (see table T formula) is used to calculate the percentage of an element in a compound. Mass of part / mass of whole * 100%.
Ex: the percent of Ag in AgCl = $(108 / (108 + 35)) * 100 = 76\%$
- When calculating the percent composition of water in a hydrate, you must consider the part as the complete mass of water. You can calculate that by multiplying the total number of moles (coefficients) of water by the mass of water which is 18 grams per mole. The mass of the whole hydrate is the mass of the compound + the mass of the water. Example: what is the percent composition of water in copper sulphate pentahydrate?
CuSO₄ * 5H₂O.
Mass of water is 5 * 18g/mol = 90g
Mass of whole = 250g
% of water = $(90/250) * 100 = 36\%$

10b. Mole to Mole Calculations

- Mole to mole questions, plug the question's information into the given balanced equation to create a fraction between what is known and what is unknown then cross multiply.
Ex: Given the following balanced reaction, how many moles of Cl are needed to produce 5.4 moles of NaCl?



Answer: $\frac{1}{x \text{ Cl}} = \frac{2}{5.4 \text{ NaCl}}$ notice we have equation info on top, and question info at the bottom.

X = 5.4 / 2 = 2.7 moles.

10c. Mole to Gram Calculations

- Use table T formula.
Ex: how many grams are in 5.4 moles of NaCl?
1 mole of any substance equals to its mass therefore 1 mole of NaCl = 23+35 = 58g
5.4 moles of NaCl = 5.4* 58g = 313.2grams
- For mass to mass calculations, simply plug the given information into the given reaction and solve algebraically.

10d. Mole to Liters Calculations

- Avogadro's theory: One mole of any gas at STP = 22.4L
Ex: how many liters are in 2.7 moles of Cl gas at STP?
$$\frac{2.7 \text{ mol Cl}}{X \text{ liters}} = \frac{1 \text{ mol Cl}}{22.4 \text{ liters}}$$
Notice we have question information on the left and Avogadro's information for accepted values. Also notice: whatever units you have as denominators remain as denominators, whatever you have as numerators remain as numerators. DO NOT CHANGE THE ORDER.

10e. Mole to Particles Calculations

- Avogadro's theory: one mole of any gas at STP = 6.02×10^{23} particles
Ex: How many particles are in 2.7 moles of Cl gas at STP?
$$\frac{2.7 \text{ mol Cl}}{X \text{ particles}} = \frac{1 \text{ mol Cl}}{6.02 \times 10^{23} \text{ particles}}$$

10f. Mole to Heat Calculations

- Use table I
Ex: how much heat is released when one mole of aluminum oxide is synthesized?
See table I, the formula is $4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$ $\Delta H = -3351 \text{ KJ}$
This means that when 2 moles of Aluminum oxide is synthesized, 3351 kilojoules of heat is released. A negative delta H means heat is released. A positive delta H means heat is absorbed.
Set a fraction and cross multiply

$$\frac{1 \text{ mol Aluminum Oxide}}{X \text{ kJ}} = \frac{2 \text{ moles Aluminum Oxide}}{3351 \text{ kJ}}$$

X = 1675.5 kJ are released upon the synthesis of 1 mole of Al_2O_3

- Avogadro's theory: at STP if any two gases have the same number of moles, those gases must have the same number of molecules and must occupy the same volume. even if their masses and densities are different. For example: 5 Liters of Argon gas contain the same number of molecules as 5 liters of Neon gas.
- Always convert everything to moles then convert back to the desired unit. Be sure to arrange the units appropriately in your fractions.

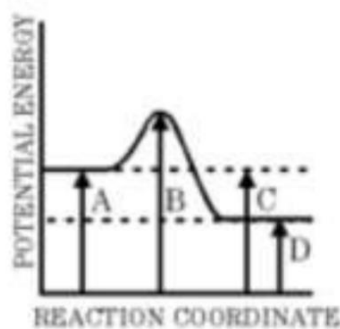
11- Thermochemistry

11a. Introduction to Thermochemistry

- Thermochemistry is the study of energy associated with chemical reactions.
- Energy can be in the form of heat, light, or electricity.
- Energy will always flow from hot substances to cooler substances.
- BARF: Break Absorb, release form. Breaking a bond requires absorption of energy while forming a bond results in a release of energy. When bonds are formed compounds are more chemically stable and their potential energy is lower.
- Table I shows the amount of energy released or absorbed for all compounds

11b. Potential Energy Diagrams

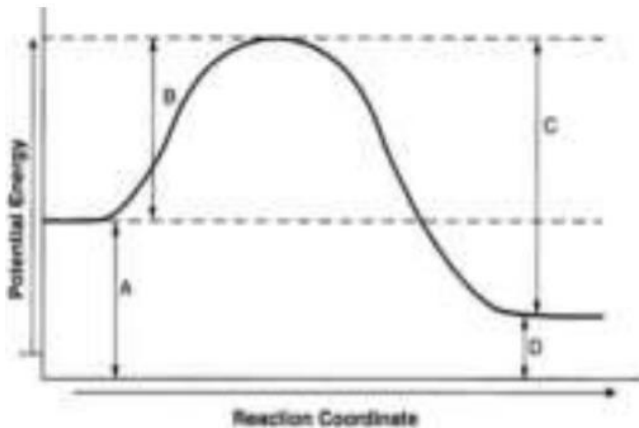
- A potential energy diagram gives us information about the energy released or absorbed in a chemical reaction in joules, kilojoules, or calories.
- It indicates the heat of the reaction, the potential energy of reactants and products, and the activation energy of reactants and products.
- The heat of a reaction is abbreviated delta H. It can be calculated by using the pepper formula ($\Delta H = P_{ep} - P_{er}$). P_{ep} is the potential energy of the product, P_{er} is the potential energy of the reactants.
- If delta H is negative, the reaction is exothermic (the substance releases heat to make its surroundings hotter while itself gets colder), if delta H is positive, the reaction is endothermic (the substance absorbs energy to make its surrounding colder while itself gets hotter).
- You may also read the reactions on table I in the reverse direction (from right to left) but if you do that, you must reverse the charge of delta H.
- The products of endothermic reactions have a higher potential energy than the reactants.



In the first potential energy diagram's picture, line A covers the same distance as line C. Both lines represent the potential energy of the reactants (PER) which is the distance between the X axis and the first horizontal line of the potential energy diagram.

Line B is the potential energy of the activation complex (PEAC). It is the distance between the x axis and the peak of the diagram.

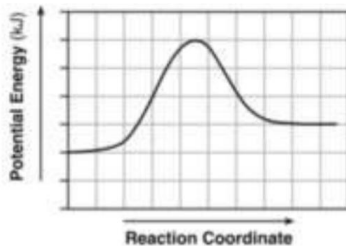
Line D is the potential energy of the products. It is the distance between the x axis and the second horizontal line.



In the second potential energy diagram's picture, line B represents the distance between the first horizontal line and the peak, this line measures the forward activation energy (FAE). The FAE is the energy needed to drive the reaction in the forward direction to favor the production of the products.

Line C represents the distance between the second horizontal line and the peak. This line measures the reverse activation energy (RAE). The RAE is the energy needed to drive the reaction in the reverse direction and favor the production of the reactants.

Delta H can be calculated by subtracting line A from Line D. ($D - A =$ the heat of the reaction) that is pep-per.



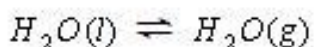
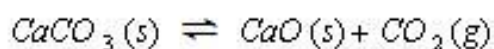
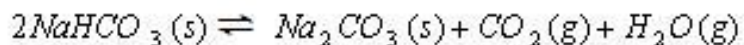
Practice problem. If each box in the third potential energy diagram's picture represents 30Kj, calculate per, peac, pep, fae, rae, and delta H. Determine if the reaction is endothermic or exothermic.

Answers: PER = 60, PEAC = 180, PEP = 90, FAE = 120, RAE = 90, and Delta H = $90 - 60 = +30$, because delta H is positive, the reaction is endothermic.

12 Kinetics and Equilibrium

12a. Equilibrium

- A system is at equilibrium when 1- The concentration of the reactants and the concentration of the products remain constant. 2- The rate of the forward reaction = the rate of the reverse reaction.
- Systems at equilibrium can be presented by equations which have a double headed arrow.
Example:



- When a battery (voltaic cell) is at equilibrium, the cell will have an electrical potential of zero (it will not generate any voltage) and the battery is considered dead.
- All phase changes are considered at equilibrium. This is known as phase equilibrium.
- All saturated solutions are at equilibrium. A solution is considered saturated when the maximum amount of a solute is dissolved in a specific amount of a solvent. This is known as concentration equilibrium. An example of a solute is salt, and example of a solvent is water. Adding more solute to a saturated solution will not affect its concentration.
- The pH scale measures the ratio of H^+ ions to OH^- ions. When a solution has a pH of 7, it is at equilibrium. The # of H^+ ions = the # of OH^- ions. A pH of 7 represents a neutral solution.

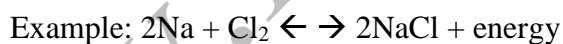
12b. Reaction Rates

- Several factors can affect a system that is at equilibrium, for example: TCCSNP: Temperature, concentration, catalyst, surface area, nature of reactants, and pressure.
- A catalyst does not cause an equilibrium to shift right to favor the production of the products or shift left to favor the production of the reactants. It only helps the reaction reach equilibrium faster by changing colliding particles' orientations and providing alternate collision pathways to lower forward and reverse activation energies and increase collision rate
- The collision theory states that a reaction takes place when reacting particles collide with sufficient energy and proper orientation. The more reactants collide effectively, the faster is the reaction rate (with more collisions, you increase reaction rates and decrease reaction time).

- Increasing temperature: temperature is a measure of the average kinetic energy (particles' velocity), the higher the temperature the faster is the motion of the particles, fast moving particles have a higher chance of collision with enough energy.
- Increasing concentration: more particles per unit area increases the chance of collision.
- Increasing surface area: particles in powder form occupy more space per unit area than particles in lump form, the particles of a substance that is in powder form have a larger surface area and a higher chance of collisions.
- The nature of reactants plays an important role in affecting reaction rates. Ionic compounds are more likely to dissociate into ions in solution form and change its concentration more than molecular compounds which do not dissociate. A higher concentration of ions leads to a higher chance of collision between the particles.
- Increasing pressure forces a certain number of gas particles to occupy a smaller space (smaller volume). The existence of many particles at the same time in a small area results in a higher chance of collision. There is an indirect relationship between pressure and volume.

12c. Le Chatellier Principle

- Le Chatellier Principle Helps us predict the effect of applying a stress on a system at equilibrium
- Increasing the pressure on a reaction made entirely of gases will cause the reaction to shift to the side with the least **total** number of moles (coefficients). Decreasing pressure will cause reaction to shift to the side with the greatest number of moles.
- Increasing temperature on an exothermic reaction will cause the reaction to shift towards the reactants. Increasing the temperature on an endothermic reaction will cause the reaction to shift towards the products thus increase the concentration of the products.
- When you add a substance to a reaction that is at equilibrium, the equilibrium will shift away from wherever that substance is located on the given equilibrium reaction. When you remove a substance from a reaction that is at equilibrium, the equilibrium will shift towards wherever that substance is located on the given reaction.



Adding Na will cause the equilibrium to shift to the right thus favoring the consumption of Na and Cl, and the production of NaCl and energy.

Removing Na will cause the system to shift to the left to replenish Na.

13- Properties of Solutions

13a. Concentration of solutions (molarity, ppm, table G)

- There are three types of solutions: unsaturated, saturated, and supersaturated. A saturated solution is at equilibrium, it has the maximum amount of solute dissolved in a specific amount of solvent. An unsaturated solution has less than the maximum amount of solute dissolved. A

supersaturated solution has more than the maximum amount of solute dissolved (the additional solute settles at the bottom as a precipitate).

- Table G shows different curve lines indicating the saturation points of solutions. Some of the curves show an increase in solubility with temperature. This direct relationship is only valid when solids dissolve in water. Other curves on table G show a decrease in solubility with an increase in temperature, this indirect relationship is valid when gases dissolve in water.
- Table G is used to determine the number of grams of solute (solid or gas) that is needed to saturate a 100mL of water at different temperatures. Any amount of gram below the curve would indicate that the solution is unsaturated, any number of grams above the curve indicate a supersaturated solution, any amount of solute on the curve indicate a saturated solution.

Example: How many grams of solute would settle if 60 grams of NaNO_3 is dissolved in 50 grams of water at 50 degrees Celsius?

Answer: the question states that 60 grams of solute dissolve in 50 grams of water, this is the same ratio as dissolving 120 grams of the solute in 100 grams of water.

According to table G, 115 grams of NaNO_3 are needed to saturate a 100 grams of water at 50 degrees Celsius.

If given 120 grams, then 120-115 would yield 5 grams to precipitate and settle at the bottom of the solution.

- Molarity and ppm are two units used to measure the concentration of a solution. The formulas for calculating them is on table T. for example. What is the concentration of a 1 liter solution that contains 2 moles of NaCl ? The answer is: molarity = moles/liters. therefore $2/1 = 2$ molar is the concentration of this solution.

Example 2: what is the concentration in ppm of a solution that contains 2 moles of NaCl dissolved in 1254mL of water?

$\text{Ppm} = (\text{mass of solute} / \text{mass of solution}) * 1000,000$

Answer: Mass of 2 moles of NaCl using the mole to gram conversion formula on table T = $58 * 2 = 116\text{grams}$.

1mL of water = 1 gram. Therefore we have 1254grams of water.

The mass of the solution is $116\text{grams} + 1254\text{grams} = 1370\text{g}$

$\text{Ppm} = (116/1370) * 1000,000 = 84.68\text{ppm}$.

13b. Colligative properties

- The higher the concentration of dissolved particles, the lower is the freezing point (the temperature at which a substance begins to turn from liquid to solid), and the higher is the boiling point (the temperature at which a substance begins to turn from liquid to gas)
- Electrolytes are substances that dissociate into charged ions in water. Examples of electrolytes include acids from table K, bases from table L, and ionic compounds (See table F for exceptions)
- The more ions a substance dissociates into, the lower is the freezing point of its solution and the higher is its boiling point. $\text{Mg}(\text{OH})_2$ dissociates into 1 mole of Mg^{+2} and 2 moles of OH^- ions in water. LiOH on the other hand will only dissociate into 1 mole of Li^+ and 1 mole of OH^- ions.

The freezing point of a solution containing 1 mole of $\text{Mg}(\text{OH})_2$ is lower than the freezing point of a solution containing 1 mole of LiOH .

- Sulfuric acid H_2SO_4 is a much better conductor of electricity than many compounds.
- When compounds containing transitional metals dissolve in water, colored solutions are formed. Example: copper makes green/blue solutions.

13c. Solubility (Polarity) of Molecules vs. Solubility of Bonds

- Polarity (also known as solubility in water) of a molecule is determined using the SNAP rule. A symmetrical molecule is nonpolar, an asymmetrical molecule is polar. We can determine the symmetry of a compound by drawing its structural formula. If the elements on the right side of the structure are the same as the elements on the left, and the elements on the top side of the structure are the same as the elements at the bottom, the molecule is symmetrical. Bent structures such as SH_2 and H_2O , as well as trigonal pyramidal structures such as NH_3 are always asymmetrical Polar. Polar means they dissolve in water.

Example: CH_4 is symmetrical and nonpolar but CH_3Br is asymmetrical and polar.



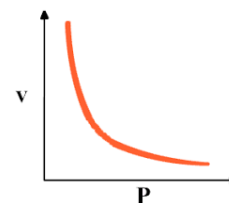
- Polarity of a bond is based on the electronegativity difference of the atoms that the bond is in between. You can see the electronegativity value of each element on table S.
- If the electronegativity difference is equal to or below 0.4 the bond is considered nonpolar covalent. If the difference is above 0.4 the bond is polar covalent. If the bond is between a metal and a nonmetal, it is classified polar ionic (see table F for ionic compounds exceptions). Example: in the CH_3Br compound shown, the molecule is polar and will dissolve in water despite having three nonpolar C-H bonds and only one polar C-Br bond.

14- Properties of Gases

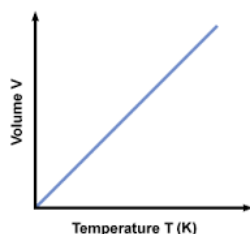
- Hydrogen likes to have FON with chlorine and group 18. Those are all the periodic table gases.
- Gases have a high entropy that is gas particles are arranged randomly.
- Gases have no definite shape and unlike liquids or solids, they have no definite volume.

14a. Pressure, Volume, and Temperature Relationships

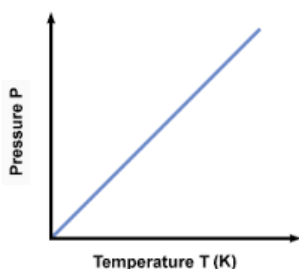
- Boyles Law: pressure and volume have an inverse relationship. If pressure increases, volume decreases. $P_1V_1=P_2V_2$.
- Increasing the pressure to reduce the volume does not impact the number of molecules of a gas sample, this only changes the amount of space they occupy and reduces the gas entropy.



- Charles Law: temperature is directly proportional to volume. As temperature increases, the kinetic energy of the particles increases causing them to move faster and occupy more space (larger volume). $\frac{V_1}{T_1} = \frac{V_2}{T_2}$



- Gay-Lussac's Law: temperature is directly proportional to pressure. As temperature increases, the kinetic energy of the particles increases causing them to collide more with each other and the container they are located in. $\frac{P_1}{T_1} = \frac{P_2}{T_2}$



- You may use the combined gas law formula on table T to calculate pressure, volume, or temperature of a gas.

14b. Kinetic Molecular Theory of Ideal Gases

- No attraction between gas particles
- No fixed volume, particles occupy a negligible volume
- No net energy is lost or gained upon collision of gas particles. All collisions are elastic
- Particles move in random, constant, straight-line motion
- Real gases behave like ideal gases if their pressure is lowered and their temperature is increased. (beach condition)

15- Acids and Bases

15a. Properties

- Acids are located on table K
 - They have a high concentration of H^+ ions which makes them sour
 - BA H AD: **B**ases **A**ccept H^+ ions, **A**cids **D**onate H^+ ions
 - They have a pH that is less than 7 and greater than 0
 - They are electrolytes (conduct electricity) because they dissociate into positive and negative ions that carry charge
 - Acids react with metals that are more active than hydrogen (see table J) to release flammable hydrogen gas and corrode metals. Ex: $Zn + HCl \rightarrow ZnCl_2 + H_2(g)$
 - Examples: vinegar, coffee, CO_2 , wine, and citrus fruits
-
- Bases are located on table L
 - They have a low concentration of H^+ ions but a high concentration of OH^- ions which makes them bitter
 - They have a pH greater than 7 and less than or equal to 14
 - They are electrolytes (conduct electricity) because they dissociate into positive and negative ions that carry charge.
 - Bases are NOT alcohols. Bases have a metal attached to an OH ion (except NH_3), alcohols have a carbon attached to an OH functional group. Also be careful, NH_3 is a base that is often written as H_3N . Lastly, be careful, although vinegar's formula CH_3COOH ends with an OH, it is not a base, it is an acid and its accepted formula is $HC_2H_3O_2$
 - Example: green vegetables, broccoli, spinach, asparagus, celery, avocado, bleach, soap

	Acids	Bases (basic=alkaline)
pH (measure of H^+ ion concentration)	Most H^+ ion concentrations	Least H^+ ion concentration)
pOH (measure of OH^-)	Least OH^- ion concentration	Most OH^- ions concentration
PH Range	$0 \leq pH < 7$ Most acidic pH	$14 \geq pH > 7$ Most basic is pH 14

15b. The Relationship Between pH and H^+ Ion Concentration

- If the pH changes by 1, the H^+ ion concentration changes by 10 folds, if it changes by 2, the H^+ ion concentration changes by 100. If it changes by 3, the H^+ concentration changes by 1000. The difference between initial and final pH determine how many zeros you will put after the number one for H^+ ion concentration.

Example: a solution has a pH of 3, a student added 5mL of a substance that increased the solution's hydrogen ion concentration by 100 folds. What is the new pH of this solution?

Answer: and increase in H^+ ion concentration means the solution becomes more acidic therefore, the solution will have a lower pH. The H^+ concentration changed by two zeros therefore the New pH must be 1.

- $pH = -\log[H^+]$. $pOH = -\log[OH^-]$
- $10^{-pH} = H^+$ concentration
- $pH + pOH = 14$
- Neutral solution, $pH = pOH = 7$
- Acid solutions, $[H^+] > [OH^-]$; $pH < 7$
- Basic solutions, $[H^+] < [OH^-]$; $pH > 7$

15c. Titration

- Titration is also known as neutralization
- It is the process of adding an acid to a base to make salt and water
- Titration result in a solution that has a pH of 7
- The formula for titration is on table T. $MaVa = MbVb$. The letter M stands for molarity
- We use the titration formula to determine the concentration of an unknown substance
Ex: If 3mL of a 6 molar NaOH was used to neutralize 5mL of an acid with an unknown concentration. What is the concentration of this acid?

Answer: $MaVa = MbVb$

$$Ma(5) = (6)(3)$$

$Ma = 18/5 = 3.6$ molar. This acid is very concentrated and should be handled with care

15d. Classifications of Acids

- There are 6 acids which ionize (dissociate) completely in solutions form those are (HCl, HBr, HI, H_2SO_4 , HNO_3 , $HClO_4$). Lab technicians may choose to use Sulfuric acid because it ionizes into two moles of hydrogen making it a stronger acid than HCl which only ionizes into 1 mole of Hydrogen.
- There are 9 strong bases which ionize (dissociate) completely in solutions form, those are any group 1 metals combined with OH, and CBS (Calcium, barium and strontium) combined with OH).
- Acids: they start with a hydrogen positive ion. a hydrogen atom has 1 proton and 1 electron (remember epNA#?). When a hydrogen atom becomes a positive ion, it loses that one electron (loses the negative charge) yielding only that positive one charge which represent the charge of the remaining proton. Conclusion: H^+ ion = a proton
- Types of acids: Monoprotic acids, ex: HCl (contain 1 proton); Diprotic acids, ex: H_2SO_4 (contain two protons); Triprotic, ex: H_3PO_4 (contains three protons).

15e. Indicators (table M)

- Table M shows the pH range of which certain indicators are not useful.
- An indicator can be used to estimate a possible pH of an unknown solution
- The color on the left of table M represents a solution with a pH lower than the given range.
- The color on the right represents a solution that has a pH above the given range.

Ex: A solution that has an unknown concentration appears yellow in methyl orange. To better determine the concentration of this solution we used a litmus paper which turned red upon contact with the solution. What is a possible pH for this solution?

a- 4.45 b- 5.2 c- 3.4

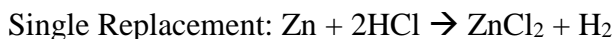
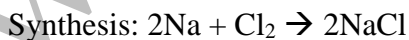
Answer: choice A: methyl orange is yellow only when the pH of the solution is above 4.4. The second test with the litmus paper revealed that the pH is less than 4.5. Litmus paper is only red when the pH is less than 4.5. Conclusion: the unknown solution has a pH greater than 4.4 and less than 4.5.

16- Redox Reactions

16a. Properties of Redox Reaction

- Redox reactions are reduction and oxidation reactions
- Redox reactions are interdependent meaning oxidation cannot occur without reduction and vice versa.
- Oilep Riger: **O**xidation **i**s **l**osing **e**lectrons at the **p**roducts' side; **R**eduction **i**s **g**aining **e**lectrons at the **r**eactants' side
- When a compound is decomposed to its prime elements, a redox reaction takes place. Synthesis of a new compound from its prime elements is also an example of a redox reaction. Lastly, single replacement reactions are examples of redox reactions.

Examples:



16b. Assigning Oxidation Numbers

- Lonely monatomic and diatomic elements have an oxidation # of zero
- If a bonded element has only one charge on the periodic table, its oxidation # will always be that charge. Example: bonded group 1 elements have an oxidation# of +1; bonded group 2 elements have an oxidation # of +2
- Bonded fluorine has an oxidation number of -1.
- Bonded oxygen has an oxidation # of -2
- In H₂O₂, Oxygen has a an oxidation number of -1
- In OF₂, Oxygen has an oxidation number of +2
- The sum of the oxidation numbers of all atoms in a compound must equal zero unless the compound has an overall charge. If a compound has an overall charge, the sum of the total oxidation #s must equal that over all charge,

Ex: Fe₂O₃

The oxidation # of one bonded Oxygen atom is -2, therefore the total oxidation # for oxygen = -6
 The sum of all oxidation numbers of all atoms that make up a compound must = zero therefore,
 Fe₂ + -6 = zero. Fe₂ = +6 therefore the oxidation # of one Fe atom is +3. You always state the oxidation number per one atom only.

Ex PO₄⁻³

Bonded oxygen again = -2, the total = -8

P + -8 = -3, P = +5

16c. Writing Balanced Redox Reactions

- Remember Oilep, Riger
- The number of electrons gained must always equal the number of electrons lost.
- Always use the balanced reaction to write the redox half reactions

Example: write the redox half reactions for the following two reactions



Oxidation half reaction: $2\text{Na}^0 \rightarrow 2\text{Na}^{+1} + 2 \text{ electrons}$

Reduction half reaction: $\text{Cl}_2^0 + 2 \text{ electrons} \rightarrow 2\text{Cl}^{-1}$

Oxidation half reaction: $\text{Zn}^0 \rightarrow \text{Zn}^{+2} + 2 \text{ electrons}$

Reduction half reaction: $2\text{H}^{+1} + 2 \text{ electrons} \rightarrow \text{H}_2^0$

17- Electrochemistry

	Voltaic	Electrolytic
Purpose	Converts chemical energy to electrical energy Used to make batteries	Converts electrical energy to chemical energy. Used for plating jewelry
Spontaneity	Spontaneous	Not Spontaneous because it requires an external power source
Composition	<p>Two metal electrodes from table J one of them serves as a positive cathode that is reduced by gaining electrons (Red Cat, riger); the other more active metal serves as a negative anode that is oxidized by losing electrons (Oil An Ox).</p> <p>The mass of the anode decreases, the mass of the cathode increases.</p> <p>Electrons always flow from the more active metal (the anode) to the less active metal (the cathode).</p> <p>Electrons flow through a conductive wire from the anode to the cathode</p> <p>If two half cells exist separately to contain ionic solutions, a salt bridge is used to transfer ions between the cells in both directions.</p> <p>If a cell reaches equilibrium, it dies.</p>	<p>Same as voltaic cell but in electrolytic cells, the anode is positive and connects to the positive end of the battery; the cathode is negative and connects to the negative end of the battery</p> <p>The cathode gets plated and increases in mass, the anode loses electrons and decreases in mass.</p> <p>Electrons always flow from the more active metal (the anode) to the less active metal (the cathode).</p> <p>Electrons flow through a conductive wire from the anode to the cathode.</p> <p>If two half cells exist separately to contain ionic solutions, a salt bridge is used to transfer ions between the cells in both directions.</p>