SCH3U Exam Review
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Unit 1: Periodic Table Trends

Atomic Radius: the distance from the centre of the nucleus to the outer most valence ring of the atom.

Trends:

1. Atomic radius *increases* down a column.
    a. The number of energy levels increase as you move down a column thus the electrons are further away from the nucleus creating a larger radius.
2. Atomic Radius *decreases* as you move across a period.
    a. The number of electrons and protons increase as you move across a row, however the number of energy levels (rings) does not.
    b. This leads to a stronger attraction between the protons and the outer electrons, pulling them in a little tighter to the nucleus.

Ionization Energy: the amount of energy required to remove an electron from the outer most shell of an atom.

Trends:

1. Ionization Energy *decreases* down a column.
    a. The number of energy levels increases as you move down a column thus the electrons are further away from the nucleus.
    b. These valence electrons are also SHEILDED from the positive charge of the nucleus by the inner rings of the atom.
    c. This distance coupled with the ELECTRON SHEILDING decreases the attraction of the valence electrons for the nucleus thus making them easier to remove.
2. Ionization Energy *increases* across a period.
    a. The number of electrons and protons increases as you go across a row, however the number of energy levels (rings) does not.
    b. This leads to a stronger attraction between the protons and outer electrons making them harder to remove. (i.e. requires more energy)

Electron Affinity: The amount of energy lost or gained by adding an electron to the valence shell of an electron. (measured in KJ/mol)

- If energy is released, the value is negative and the electron affinity is high
- If energy is required, the value is positive and the electron affinity is low.

Trends:

1. Electron Affinity *decreases* down a column.
    a. There is less attraction due to the electron shielding and larger radius, therefore it is harder to add an electron.
2. Electron Affinity *increases* across a period.
    a. Due to the increase in attraction as you go across a period, it is easier to add an electron to the outer most valence ring.

Ionic and Covalent Compounds

Ionic Compounds: Ionic compounds are formed when charged atoms are attracted by electrostatics. One atom loses an electron (becoming positive) to another atom who gains that electron (thus becoming negative).

- Occur between metal and non-metal.
- Form a crystal lattice structure that is very rigid.
- In a crystal lattice structure, all the ions are surrounded by others of opposite charge.
- Because of this structure, ionic compounds are unable to bend. When you try to bend it, you bring two like charges together, which then repel each other and break the structure apart.
- The ions are tightly packed and have limited motion, mainly vibrational.

Covalent Compounds: are compounds that occur when two non-metals share electrons.

- Tend to be liquids and gases at room temperature.
- Can occur between a metal and non-metal if the ionization energy is too high.
- Carbon, boron and silicon tend to form large covalent networks. This means that many atoms or compounds bind together forming an extremely large entity.

Metallic Bonding: occurs in metal solids.

- Metallic solids are made up of metal atoms. The valence electrons of these atoms are loosely held by the positive nuclei.
- The electrons are free to move because of this.
- The electrons are able to move past each other because the positive nuclei act as a bumper preventing the electrons to repel each other.
- Electricity is free flowing electrons. Therefore, metals are able to conduct electricity because their electrons are able to freely flow through the compound.

Lewis Dot Diagrams: Ionic

1. Print the symbol for the metal in square brackets [ ] with no electrons.
2. Write the charge outside the right square bracket subscripted.
3. Print the non-metal symbol in square brackets with eight electrons around it. (except Hydrogen)
4. Repeat step 2.

Lewis Dot Diagrams: Covalent

1. Identify the centre atom.
2. Add up the number of valence electrons of all the atoms.
3. Draw a single bond from the centre atom to all surrounding atoms.
4. Complete the surrounding atoms “octets”.
5. If there are any atoms remaining, put them around the centre atom as lone pairs.
6. Check the centre atom to see if it’s “octet” is satisfied. If it is not satisfied, than “swing’ a lone pair from the surrounding atoms into a double bond (or triple) with the centre atom. ONLY DONE WHEN THE ELECTRONS RUN OUT!

Polar Covalent Bonds

Polar Covalent Bond: An unequal sharing of electrons that leads to partial charges on the atoms involved.

Electronegativity: a measure of the tendency of an atom to attract a bonding pair of electrons.

To determine the polarity of a covalent bond, you must:

1. Draw the two atoms involved in the bond.
2. Find the electronegativity for each element on the periodic table.
3. Find the ΔEN for the elements. (This value is always positive.)
4. Analyze the results as follows:
    a. $\Delta EN \geq 0.5$ Polar Covalent Bond
    b. $\Delta EN < 0.5$ Pure Covalent Bond (non-polar)
    c. $\Delta EN \geq 1.7$ Ionic Compound

Polar Molecules: Polar molecules are covalent compounds that have a slightly positive charge at one end of the molecule, and a slightly negative charge at the other end of the molecule.

Non-polar Molecules: Non-polar molecules are covalent compounds with no charged ends.
When determining the polarity of a molecule, you must consider two things:

1. The ΔEN
2. The shape of the molecule.

Molecular Shapes

Central Atom Bonds	**Central Atom Lone Pairs**	**Structure**	**Name**	**Symmetry**
2	0		Linear	Symmetrical
3	0		Planer Triangular	Symmetrical
4	0		Tetrahedral	Symmetrical

3	1		Trigonal Pyramidal	Asymmetrical
2	2		Bent	Asymmetrical
1	2		Linear	Asymmetrical

Unit 2: Types Of Chemical Reactions

Synthesis Reaction: Two or more atoms or compounds come together to form one complex compound.

Ex: A + B → AB or AB + C → ABC

- Metal oxides in water form metal hydroxides
- Non-metal oxides in water form acids.

Decomposition Reaction: A compound breaks down to form atoms or simpler compounds.

Ex: AB → A + B or ABC → AB + C

Combustion: the reaction of a substance with oxygen, producing oxides and release energy in the form of heat or light.

Ex: Hydrocarbon + oxygen → carbon dioxide and water

Single Displacement: one element replaces another element from a compound in a chemical reaction.

Ex: Z + AB → ZB + A (metals)

Y + AB → AY + B (non-metals)

- Metals always replace metals.
- Non-metals always replace non-metals.

Double Displacement: elements in different compounds exchange places with each other.

Ex: AB + CD → AD + CB

There are three main types of Double displacement reactions:

1. A reaction forming a precipitate
   a. *Precipitate:* an insoluble solid that is formed by a chemical reaction of two soluble compounds
   b. Ex: $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
2. A neutralization occurs.
   a. *Neutralization:* occurs when an acid and base are combined, producing a salt and water.

   b. Results in the H+ ions from the acid and the OH- ions from the base forming water.
   c. Any Salt (i.e. ionic compound)
   d. Ex: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
3. A gas is formed.
   a. A double displacement reaction that forms a gas often involves two steps:
      i. Double displacement of the reactants forming two new products
      ii. Rapid decomposition of one product into a gas and water.
   b. Ex: $Na_2CO_3 + HCl \rightarrow NaCl + H_2CO_3 \rightarrow H_2O + CO_2\,(g)$
   c. Ex: $NH_4Cl + KOH \rightarrow KCl + NH_4OH \rightarrow H_20 + NH_3(g)$
   d. Ex: $Li_2SO_3 + HNO_3 \rightarrow LiNO_3 + H_2SO_3 \rightarrow H_2O + SO_2(g)$
   e. Ex: $MgS + H_2SO_4 \rightarrow MgSO_4 + H_2S(g)$

Law of Conservation of Mass: states that:

- Matter cannot be created or destroyed.
- Atoms in the reactants are rearranged in the products but are equal on both sides.
- Total mass of the reactants is always equal to total mass of the products.

Unit 3: Quantities in Chemical Reactions

Isotope: atoms of the same element that have a different number of neutrons.

- This causes the mass of these atoms to be different
- The isotopes are taken into consideration when determining the mass of an element

Isotopic Abundance: the relative amount in which each isotope is present in an element (%)

Average Atomic Mass: the mass of an element determined by taking the average mass of all the isotopes of that element (this is the mass found on the periodic table, measured in atomic mass units, aam)

Formula: aam = mass1 (%1) + mass2 (%2)

The Mole: the amount of a substance that contains the same number of elementary particles as the number of atoms in 12.00g of Carbon-12.

Formula: $N = n \times N_A$

- N is number of particles
- n is number of moles
- N_A IS Avogadro's constant.

Molar Mass: the mass of one mole of a substance

Formula: MM = MassE1 + MassE2 + …. Etc.

Calculating Mass

Formula: m = n × MM

- m is the overall mass in grams
- n is the number of moles
- MM is the molar mass

Law of definite proportions: the elements in a chemical compound are always present in the same proportions by mass.

Mass Percent: the mass of an element in a compound, expressed as a percent of the total mass.

Percentage Composition: refers to the relative masses of each element in a compound.

Formula: %E = (Mass of element/Mass of compound) × 100%

Finding the Empirical Formula

Empirical Formula: the lowest whole number ratio of the element in a compound.

Steps:

1. Assume 100 grams.
2. Change Percentages into grams.
3. Find the moles of elements.
4. Set up the compound ratio and divide but the lowest number.
5. Write out the empirical formula.

Finding the Molecular Formula

Molecular formula: describes the number of atoms that makes up each element in the formula a.k.a. the actual formula.

- Using Empirical Formula and Molar Mass
  1. Find the molar mass of the empirical formula.
  2. Divide MM_{MF} by MM_{EF}
  3. Multiply across by the moles.
- Using Empirical Formula and Number of Atoms

    1. Find Moles using $N = n \times N_A$.
    2. Multiply across by moles.
- Using Molar Mass and Percentage Composition
    1. Assume the molar mass.
    2. Convert the percentages into decimals.
    3. Multiply the decimals by the molar mass to find the mass in grams.
    4. Find the moles by dividing mass in grams by the elements molar mass.
    5. Write out the Molecular Formula.

Hydrated Ionic Compounds

Hydrate: ionic compound that has a specific number of water molecules incorporated into their structure.

Steps to finding the amount of water molecules:

1. Find the mass and molar mass of the ionic compound.
2. Find the mass and molar mass of the water molecules.
3. Find the moles of both the ionic compound and the water molecules.
4. Set up the compound ratio and divide but the lowest number.
5. Simplify, if possible.
6. Write out the whole molecular formula.

Stoichiometry: the study of the relative quantities of reactants and products in a chemical reaction.

Mole Ratios: the relationships between moles in a balanced chemical equation.

Stoichiometric Coefficients: the coefficients in a balanced chemical equation.

Stoichiometric Amounts: reactants present in the mole ratios that correspond exactly to the mole ratio predicted by the balanced chemical equation.

Limiting Reactant (Reagent): the reactant that is completely used up in a chemical reaction.

Excess Reactant (Reagent): the reactant that remains after the reaction is over.

Determining the Limiting Reactant

Steps:

1. Find the moles of both reactants using a formula involving the mole.
2. Divide each compound's moles by the coefficient.
3. Whichever compound has the lower amount after the division, they are the limiting reactant.

Excess Reactant Remaining

Steps:

1. Calculate the number of moles of both the reactants.
2. Find out which reactant is limiting.
3. Use a stoichiometric ratio to find the moles of the excess reactant.
4. Subtract the moles found using the formula from the moles found using the ratio.
5. Use the final moles to find the mass remaining.

Calculating the Percentage Yield

Theoretical Yield: the amount of product predicted by stoichiometric calculations.

Actual Yield: the amount of product that is obtained in an experiment.

Formula: Percentage Yield = (Actual/Theoretical) × 100%

Unit 4: Solutions

Net Ionic Equations

Steps:

1. Write the balanced chemical equation.
2. Write all the aqueous ionic compounds as ions.
3. Write strong acids as ions.
4. All solids, liquids, gases and weak acids are written in their molecular form.
5. When you are finished you have your total ionic equation.
6. Ions that appear on both sides of the equation are called spectator ions and are cancelled.
   a. *Spectator ions:* ions that are not really involved in a chemical reaction and remain unchanged as a result of the reaction.
7. Check your resulting equation to ensure that both sides have the same number of atoms and the same net charge.

Concentration: the amount of solute per quantity of solvent.

Concentration can be calculated four different ways:

1. Concentration as a mass/volume percent
   a. m/v % = (mass of solute in grams/ volume of solution in mL) × 100%
2. Concentration as mass/mass percent
   a. m/m% = (mass of solute in grams/mass of solution in grams) × 100%

3. Concentration as volume/volume percent
    a. v/v% = (volume of solute in mL/ volume of solution in mL) × 100%
4. Concentration in Parts per Million and Parts per Billion
    a. Ppm = (mass of solute in grams/ mass of solution in grams) × 10^6
    b. Ppb = (mass of solute in grams/ mass of solution in grams) × 10^9

<u>Calculating Concentration Using Moles</u>

Formula: C = n/V

- C represents concentration in mol/L.
- n represents moles in mol.
- V represents volume in L.

<u>The pH Scale:</u> a scale used by chemists to determine the concentration of an acid.

- pH stands for "the power of hydrogen"
- The scale ranges from 0 to 14, with 7 being neutral.
- Every unit on the scale represents a tenfold effect on the concentration.

Formula: $pH = -\log (H^+)$ or $pH = -\log (H_3O^+)$

<u>Strong Acid:</u> an acid that completely dissociates in water.

There are 6 strong acids:

- HCl Hydrochloric acid
- HBr Hydrobromic Acid
- HI Hydroiodic Acid
- HNO_3 Nitric Acid
- H_2SO_4 Sulfuric Acid
- $HClO_4$ Perchloric Acid

<u>Weak Acid:</u> An acid that dissociates very slightly in water.

<u>Strong Base:</u> a base that dissociates completely in water.

- All oxides and hydroxides of the *alkali metals* are strong bases.
- All oxides and hydroxides of the *alkaline earth metals* are strong bases.

<u>Weak Base:</u> a base that dissociates very slightly in water.

Conjugate Acid-Base Pairs

- An acid is a substance from which a proton (H^+) is removed. (I.e. a proton donor.)
- A base is a substance that can remove a proton from an acid. (i.e. a proton acceptor)

Conjugate Acid: The particle that is formed when a proton is accepted.

Conjugate Base: The particle that remains when a proton is removed.

General Equation: $HA + B^- \rightarrow A^- + HB$

A B CB CA

Dilution

Stock Solution: a concentrated solution that is diluted to lower the concentration for actual use.

Formula: $C_1V_1 = C_2V_2$

Unit 5: Gases

Boyle's Law

Boyle found that volume and pressure are inversely related.

- $P \alpha 1/V$
- P ↑ V ↓ and vice versa
- As pressure goes up, volume goes down.
- Temperature and Amount are held constant.

Formula: $P_1V_1 = P_2V_2$

Charles' Law

Charles found that volume and temperature are directly related if you hold pressure and amount constant.

Formula: $T_1/V_1 = T_2/V_2$

Gay-Lussac's Law

Gay-Lussac found that pressure and temperature are directly related if you hold volume and amount constant.

Formula: $P_1/T_1 = P_2/T_2$

STP: means standard temperature and pressure, which is 101.3 kPa and 0°C.

SATP: standard ambient temperature and pressure, which is 100 kPa and 25°C.

Combined Gas Law

When all the formulas are put together, we know that pressure and volume are directly related to temperature and inversely related to each other, if you hold amount constant.

Formula: $P_1V_1/T_1 = P_2V_2/T_2$

Avogadro's Law

Avogadro found that volume and amount are directly related if you hold temperature and pressure constant.

Formula: $n_1/V_1 = n_2/V_2$

Dalton's Law of Partial Pressures

The total pressure of a mixture of gases is the sum of the pressures of each of the individual gases.

The Ideal Gas Law

If you want to not hold anything constant so that you can perform an experiment at any conditions, you would use the ideal gas law:

Formula: $PV = nRT$

- P represents pressure measured in kPa
- V represents volume measured in L
- n represents the number of moles
- T represents temperature measured in K
- R represents a constant 8.314.