Chemistry Study Sheet

Ionic Bonds:

Result from the **transfer** of valence electrons from one cation to another anion. For example NaCl.

Ionic Bond : The electrostatic force that holds oppositely charged ions together in an ionic compound.

Ionic compounds are formed when the amount of valence electrons found in a cation are transferred to the anion to form a stable 8 valence electrons within the anion.

In order to write an ionic compound formula, find the oxidation states of all elements involved, and then criss-cross them from the subscript of its original ion, to the base (that represents the number of ions in the compound) of the other ion. )

Ex. Mg (Magnesium) and Cl (Chlorine)

Mg+2Cl+7 🡪 Mg7Cl2

Properties of Ionic Compounds

The chemical bonds that occur between the atoms in a compound determine many of the physical properties of ionic compounds.

The cations and anions have to form a pattern known as an ionic crystal in which no single anion is next to another anion, and no cation is next to another cation.

The large number of positive and negative ions that exist together in a ratio that is determined by the number of valence electrons transferred from the cation to the anion.

This **strong attraction** between positive and negative ions in an ionic compound result in a crystal lattice. (This is essentially the same thing as an ionic crystal. Ever cation is surrounded by anions, and every anion is surrounded by cations.)

Melting point, boiling point and hardness are determined by the force of these attractions

Ionic bonds are generally strong, therefore a large amount of energy is required to break them apart.

Ionic Compounds have **high melting points and high boiling points**

**They are also hard, rigid and brittle solids** due to the strong attractive forces that keep the ions in place.

(Recap: Ionic compounds have: Strong Attraction b/w anions and cations, high boiling points, high melting points, they are brittle, rigid and hard, because of strong attraction b/w anions and cations.)

The ionic crystal will only crack when like charged particles are repositioned next to each other.

In order to conduct electricity charged particles must be able to move, therefore ionic compounds in a solid state do not conduct electricity because the attractive forces hold ions in place.

Ionic Compounds only **conduct electricity in a liquid state.**

An ionic compound whose aqueous solution conducts an electric current is called an **electrolyte.**

When energy is absorbed during a chemical reaction such as the formation of a bond, the reaction is **endothermic**. If energy is released during a chemical reaction, it called **exothermic**.

The formation of ionic compounds will **always be exothermic.** This is because the attraction of cations for anions near forms a stable system that is lower in energy than the original ions.

Note: If the amount of energy released during the formation of a ionic compound is added to the ionc compound itself, the bonds that hold the negative and positive ions in place will break.

Lattice Energy: The energy required to separate one mole of ions from an ionic compound.

The more negative the lattice energy, the strong the force of attraction between ions in the ionic crystal.

**Lattice energy is related to the size of the ions that are bonded**

Smaller ions generally apply to more negative lattice energies because their nucleus’s are closer.

The opposite applies to larger ions.

(Required knowledge: Ionic radius generally decreases moving across the period, and increases moving across the group.)

Sample question: Which ionic compound has a more negative lattice energy:

Lithium bromide, or lithium chloride.

It would be lithium chloride, this is because when comparing the ionic radius of chlorine to bromine, chlorine is smaller, because lithium is present in both ionic compounds, the distance between the ions in lithium chloride is smaller than lithium bromide, making more lattic energy necessary for that compound.

Lattice Energy is also influenced by charges, generally ions with more positive/negative charges have a more negative lattice energy.

Ionic Compound Formulas:

Formula Unit: The simplest ratio of the ions represented in an ionic compound.

Binary Ionic Compounds: Ionic Compounds composed of one positively charged ion (metal) and one negatively charged ion (nonmetal)

Monatomic Ion : One atom ion.

Determining Charge : In order to determine the charge of an ion, look at its oxidation number, if the atom is a transition metal, the oxidation number or the number of valence electrons that are transferred must be given.

In order to write the formula unit of an ionic compound, pick any nonmetal and any metal. For example: Sulfur and Ca.

Sulfur has a +6 charge/oxidation number, and Calcium has a +2 charge. In order to write the formula unit which will not include charges, simply criss cross the charges so that there will be 2 sulfur ions, and 6 calcium ions.

Ca6S2

However, the ionic compound must always be expressed in the lowest possible ratio, therefore the final formula is:

Ca3S

(Never write ones)

The same applies for ionic compounds containing polyatomic ions. Simply treat the polyatomic ion as a normal ion, and criss cross its charge, however, if there are multiple polyatomic ions **you have top indicate that by putting a parenthesis around the original polyatomic ion, and then outside the parenthesis placing the base.**

Ex:

Cl and NH4

Both of these have a +/- 1 charge therefore they will be written as such:

NH4Cl

Naming Ionic Compounds:

There are only a few rules in naming ionic compounds:

1. The cation always goes first just like in the formula unit.
2. A monatomic anion is always written with the suffix –ide.
3. A polyatomic anion/cation name is never manipulated.
4. When there is a transition metal involved, you must indicate the oxidation number of that transition metal in parenthesis after the cation as a roman numeral.

An **oxyanion** is a polyatomic ion composed of an element bonded to one or more oxygen atoms. Oxyanions are the same in charge and have the same nonmetal, but differ in the amount of oxygens that they are bonded to.

An **oxyanion** with a generally greater number of oxygen atoms is named with the root of the nonmetal and then the suffix –ate.

An **oxyanion** bonded to fewer oxygen atoms is named using the root of the nonmetal plus the suffix –ite.

Example:

Nitrate Nitrite

NO3- NO2-

Writing Lewis dot structures for ionic compounds:

The lewis dot structure for ionic compounds includes:

-The respective charges for both ions (written as a superscript) (NOTE: For the anion or the second ion bonded, write the number in front of the charge for example 2- as opposed to -2, remember that these always appear outside the brackets.)

-The number of ions found in the specific bond expressed as a coefficient. (For the anion, always express the coefficient outside of the brackets.

Always include 8 dots for **any** type of anion found in the ionic bond, indicate this **within the brackets**. (For helium only indicate 2 dots) 🡪 These represent valence electrons, because the anion always has a negative charge an receives electrons, it will always be the one to get the stable octet.

Basic stuff that will screw a lot of people over on the test:

Chemical Bond : The bond that holds two atoms together.

Pseudo-noble gas formation : When the anion in an ionic bond gains 8 valence electrons by receiving electrons from groups 1A-4A.

Metallic Bonding:

Metals form lattices in a solid state (just like ions).

Metals in a metallic compound neither lose or gain electrons.

Instead, in the crowded conditions, the outer energy levels of the metals bonded overlap, resulting in a sea of electrons.

The **electron sea model** suggests that all metal atoms contribute their valence electrons to a sea of electrons composed of electrons that can move freely from one atom to the next. These electrons are called **delocalized electrons**.

The Metallic Bond is the attraction of a metallic cation for delocalized electrons.

Properties of Metallic Compounds

Melting points vary but in general have moderately high boiling and melting points. More often than not, the **melting points** are **lower** than the **boiling points** because the **cations and electrons are mobile in metal**. Therefore it does not require as much energy for the atoms to be able to move past each other (melting), than it does for the atoms to be separated from the group of cations and electrons (boiling).

Metallic Compounds are malleable and ductile, this is due to the mobility of their electrons and cations.

Metals are also generally durable, this is because though the cations and electrons are mobile, they are still strongly attracted to one another and aren’t easily removed from the metal.

The mobility of delocalized electrons keeps the metallic bonds intact.

Metallic Compounds are also good conductors of both heat and electricity. They are such because the delocalized electrons easily move heat from one part of the metal to another. Also because the electrons are mobile, they are good conductors of electricity.

The same delocalized electrons interact with light, absorbing and releasing photons, thereby creating the property of luster in metals.

(Recap: Metals are/have: High boiling points, high melting points, malleable, ductile, lustrous, good conductors of electricity, good conductors of heat.) 🡪 These properties are all due to the delocalized electrons.

Due to the nature of the metallic bond, it is easy to introduce other elements into a metallic crystal, forming an alloy which is a mixture of elements that have metallic properties.

Alloys most commonly form when the elements involved are either similar in size or the atoms of one element are considerably smaller than the atoms of the other.

Thus, there are two types of alloy:

Substitutional Alloys, and Interstitial Alloys

**Substitutional alloys** have atoms of the original metallic solid substituted by metal atoms of a similar size.

**Interstitial Alloys** are formed when small holes called interstices in a metallic crystal are filled with smaller atoms.

(The difference between the two is that in one, atoms of the original metallic solid are actually substituted, while in the other, spaces that appear in the original metallic solid are simply filled, there is no substitution.)

The most common interstitial alloy is **carbon steel:** Holes in the iron crystal are filled with carbon, and thus the physical properties of iron are changed. Iron is generally soft and malleable, but the carbon makes it harder, stronger, and less ductile that the original iron.

In substitutional alloys, such as sterling silver, the silver atoms replace portions of the original crystal, and thus the properties of both copper and silver are formed.

(Recap: Two types of alloys: Substitutional and Interstitial, in substitutional alloys, the physical properties of both metals involved, while in interstitial alloys, the physical properties of the original metallic solid are changed due to the introduction of the new metal.)

**Covalent Bonding 🡪** the scary one

Covalent Bond: The bond that results from the **sharing** of valence electrons.

In this type of bond, valence electrons are considered to be part of an outer energy level that both atoms are involved in.

Covalent bonding generally happens between nonmetallic elements.

A **molecule** is formed when two or more atoms are bonded covalently.

Hydrogen, Fluorine, Chlorine, Bromine, and Iodine all exist as naturally occurring diatomic molecules. This is because the diatomic molecules are more stable than the individual atoms.

Example: In a fluorine diatomic molecule, both atoms have 7 valence electrons, and both need one additional to form an octet.

Now this is confusing:

As the two fluorine atoms approach each other there is a resulting repulsive force 🡪 This occurs between like charged electrons, and like charged protons.

An attractive force also occurs between the **protons of one fluorine** and the **electrons of the other** atom.

They continue to move closer until that attractive force is at its maximum attraction.

Thus the **attractive forces** balance the **repulsive forces**.

🡪 At this balance point, the atoms bond covalently. Thus the molecule is formed.

In the end, each fluorine atom will have three lone pairs of electrons and one single bonding pair. (The lone pairs are not shared between electrons.)

In lewis dot structures:

-A single bond (share of two electrons) = H—H

-A double bond (share of four electrons) = H—H

-A triple bond (share of 6 electrons) = H—H

-A quadruple bond (share of 8 electrons)

-Halogens in the 7A group will always form single bonds with nonmetals.

-Group 6A elements will form double bonds with nonmetals.

-Group 5A elements will generally form triple bonds with each other.

-Group 4A elements will generally form quadruple covalent bonds.

(Use these rules to help write lewis dot structures for covalent compounds)

Sigma Bonds – (single covalent bonds) a bond that occurs when the electron pairs are shared in an area centered between the two atoms. A sigma bond results if the atomic obritals overlap end to end, concentrating the electrons in a bonding orbital between two atoms.

Bonding orbital 🡪 A localized region where bonding electrons will most likely be found.

Pi Bond – When parallel orbitals overlap to share electrons,

**Any multiple bond is composed of one sigma bond, and at least one pi bond.**

**The formation of an ionic compound is generally exothermic (releases energy)**

**Breaking an ionic compound is generally endothermic, absorbing energy.**

When naming binary molecules you use prefixes mono-, di-, tri-, tetra-, penta-, in accordance with the number of atoms present in the compound and not their individual charges.

One Rule: Never use the mono prefix on the first nonmetal.

Also 🡪 Be reasonable if something doesn’t pronounce fluently, like mono-oxide, don’t use that.

In order to write lewis dot structures, find the number of electrons of the entire compound, divide by two, and thus you get the number of disposable bonds. Then apply the bonds so that each atom is connected. Then, apply lone pair electrons wherever necessary to create a stable octet within all atoms present in the compound. Hydrogen atoms can only be bonding with sigma bond (single bonds).

Electronegativity and bond Character:

Electron Affinity: The tendency for an atom of an element to accept a valence electron.

Decreases by groups

Increases by periods

Side Note: Electronegativity/Affinity is not measured by quantity on the metric scale, but rather in “Linus Paulings” units.

Noble Gas Electronegativities are not given on the periodic table of elements because they do not generally form compounds.

Bond Character:

The type of **individual bond** is determined by electronegativity difference. If the difference is:

< .40 🡪 Non-Polar Covalent

>.40 and <1.70 🡪 Polar Covalent

>1.70 🡪 Ionic

Polar Covalent Bonds result in the unequal sharing of valence electrons.

🡪 When you have a polar covalent bond, you always have a **partial negative charge** and **partial positive charge.**

These partial charges are represented by some weird greek looking squiggly thing that looks like an 8 with a super script that is either +/-.

8+ H --- Cl 8-

This represents a polar covalent bond because the electronegativity of chlorine is larger than hydrogen it will attract more valence electrons subsequently giving it a more negative charge that is represented by the greek thingy.

Polarity can also be measured for entire molecules.

Characteristics: Non-Polar Molecules are not generally attracted to a magnetic field.

Polar Molecules are attracted to a magnetic field because they have a denser field of electrons at the negative partial charge.

In order to determine polarity on a molecular level, you have to look at whether or not the molecular structure of the molecule (…sorry for the repeated use of “molecule”) is symmetrical or not. Note that it is not enough o simply claim that because there is an even number of bonds, that the molecule is non-polar.

In order for there to be non-polarity the attraction of the partial charges at any point, must be the same at the opposite side of the center. Therefore, a water molecule is polar, because its structure shows the two bonds at an angle, therefore it is not symmetrical, where as Carbon Tetrachloride is symmetrical and thus non-polar.

(Recap for symmetry: All you need to know is this: The Symmetry DOES NOT DEPEND on charge distribution, but the charge distribution depends on symmetry.)

**Solubility**: This is simple: Two Basic Rules:

Polar Molecules and Ionic Compounds are soluble in polar substances.

Non-polar Molecules dissolve in non-polar substances.

**Finally, The Properties of Covalent Compounds**

Attraction between atoms of molecules is strong (intramolecular forces)

Attraction between molecules is weak (intermolecular forces)

(inter < intra)

For Non-polar molecules🡪 Dispersion Forces (also called “induced dipole” NOT TO BE CONFUSED WITH DIPOLE-DIPOLE FORCE) Dispersion forces are fundamentally the same thing as single bonds, when the outer energy level of two atoms overlaps resulting in a centralized electron density.

For Polar-Molecules🡪 Dipole-Dipole Forces (Attraction between oppositely charged ends of the dipoles)

For Polar Molecules that include Hydrogen covalently bonded to either Flourine, Nitrogen, or Oxygen (FON) 🡪 Hydrogen Bonds (These bonds are similar to dipole dipole forces but are stronger and only found between the hydrogen atom of one dipole and the non-hydrogen atom of another.

In Lewis Dot Structures the hydrogen bond is represented by a dotted line that looks kind of like a barcode except each line is of equal thickness. (…? Google it if your confused).

Intermolecular Forces Determine the majority of the properties of covalent compounds.

MP’s and BP’s are low, (because the intermolecular forces are weak)

They are usually soft (due to low intermolecular forces)

Covalent Compounds also form within a crystal lattice only with less attraction between molecules. The shape of this lattice is determined by the type of intermolecular force (Dispersion, Dipole, Hydrogen Bonds)

Shit, I’m done.