IONIC CHARGES

- The ionic charge of an ion is dependent on the number of electrons lost or gained to attain a noble gas configuration.

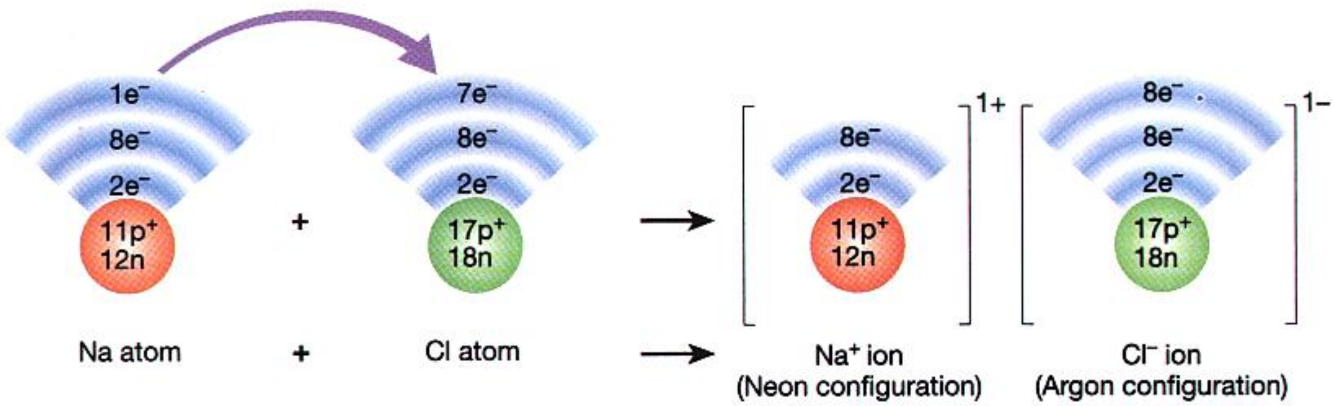
- For most main group elements, the ionic charges can be determined from their group number, as shown below:

		Metals Lose Valence Electrons			Nonmetals Gain Valence Electrons				
Noble Gases		1A (1)	2A (2)	3A (13)	5A (15)	6A (16)	7A (17)		Noble Gases
He	←	**Li^{+}**							
Ne	←	**Na^{+}**	**Mg^{2+}**	**Al^{3+}**	**N^{3-}**	**O^{2-}**	**F^{-}**	→	**Ne**
Ar	←	**K^{+}**	**Ca^{2+}**		**P^{3-}**	**S^{2-}**	**Cl^{-}**	→	**Ar**
Kr	←	**Rb^{+}**	**Sr^{2+}**				**Br^{-}**	→	**Kr**
Xe	←	**Cs^{+}**	**Ba^{2+}**				**I^{-}**	→	**Xe**

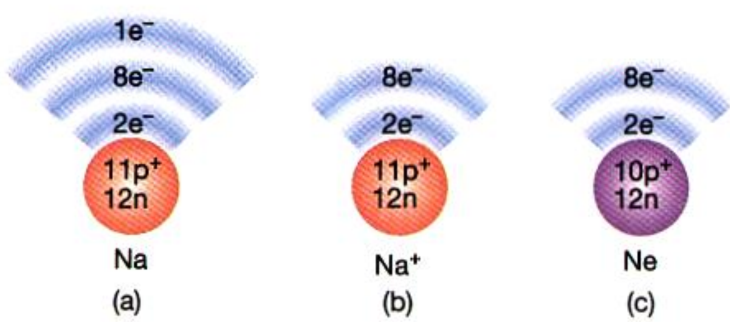
- All other ionic charges need to be memorized and known in order to write correct formulas for the compounds containing them.

IONIC COMPOUNDS

- ***Ionic compounds*** contain ***ionic bonds***, which occur when electrons are ***transferred*** between two atoms.



- ***Ionic bonds*** occur between ***metals*** and ***non-metals***.
- Atoms that lose electrons (***metals***) form positive ions (***cations***).
- Atoms that gain electrons (***non-metals***) form negative ions (***anions***).
- The ***smallest*** particles of ***ionic compounds*** are ***ions*** (not atoms).



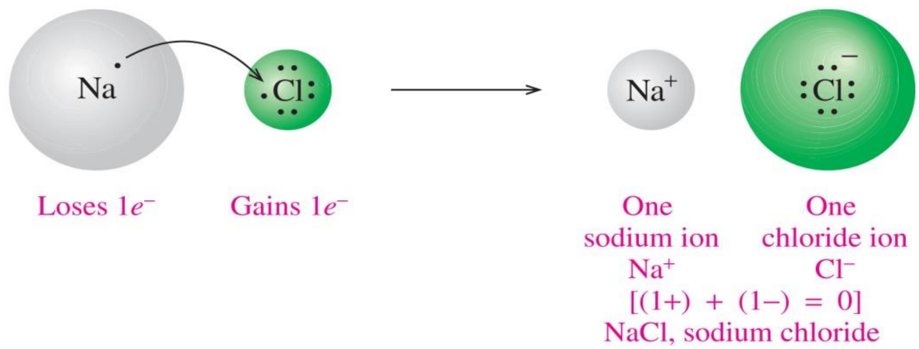
Comparison between sodium atom (a), sodium ion (b) and neon atom (c)

IONIC CHARGES & FORMULAS

- The formula of an ionic compound indicates the number and kinds of ions that make up the ionic compound.

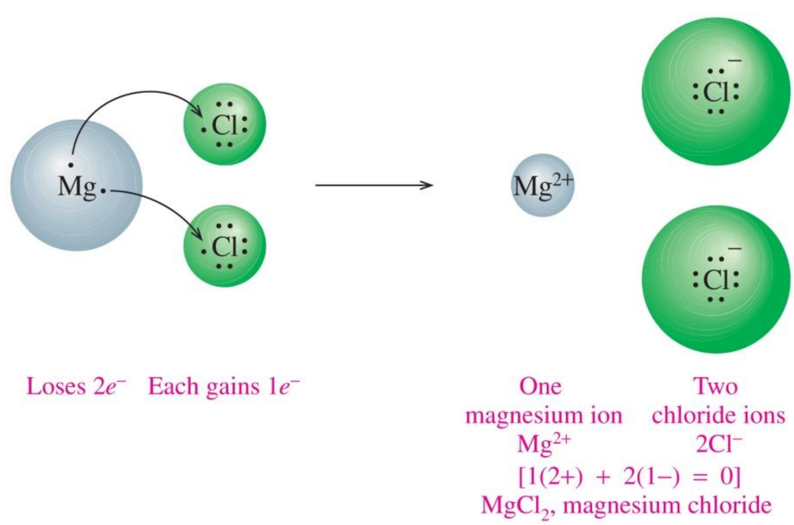
- The sum of the ionic charges in the formula is always zero, which indicates that the total number of positive charges is equal to the total number of negative charges.

- For example, the +1 charge on the sodium ion is cancelled by the –1 charge on the chloride ion, to form a net zero charge.



- When charges between the two ions do not balance, subscripts are used to balance the charges.

- For example, since each magnesium loses 2 electrons, and each chloride gains one electron, 2 chlorides are needed to balance the charge of the magnesium ion. Therefore, magnesium chloride is written as $MgCl_2$.

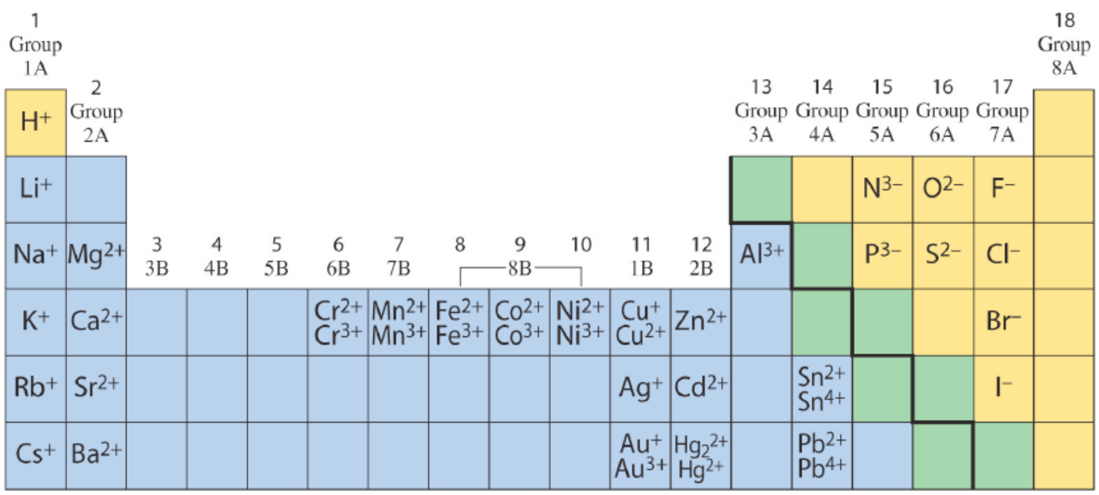


NAMING & WRITING IONIC FORMULAS

- When writing ionic formula, knowing the charge of the ions are important since the net charge on the compound must be zero.

- Some elements produce only one ion (Type I) while others produce two or more ions (Type II).

- Differentiating between type I and II ions is important, since the naming system is different for each. Shown below are the common ions of each type:



- Note that most main-group elements are type I (except Sn and Pb), and most transition elements are type II (except Ag, Cd and Zn).

NAMING & WRITING IONIC FORMULAS

Binary Ionic Compounds (Type I):

- ***Binary*** compounds contain only ***two elements***.
- ***Type I*** are those ***cations*** that form ***only one ion***.
- In these compounds, ***charges of the cations must equal the charges of the anions*** since the ***net charge is zero***.
- ***Subscripts*** are used to ***balance*** the charges between cations and anions.

Name	**Sodium bromide**	**Potassium sulfide**
Ions	Na^+, Br^-	K^+, S^{-2}
Formula	$NaBr$	K_2S

When naming ionic compounds:

- Name the ***cation*** first, the ***anion*** last.
- The ***cation*** name is the same as the name of the ***metal*** it forms from.
- The ***anion*** name takes the ***root of non-metal*** and the ending ***"-ide".***

$MgCl_2$	**magnesium chloride**
NaI	**sodium iodide**
AlF_3	**aluminum fluoride**

Examples:

1. Write formulas for the following ionic compounds:

calcium chloride: _____________

sodium sulfide: _______________

2. Name the following ionic compounds:

Na_3P : _____________________

$BaCl_2$: ______________________

NAMING & WRITING IONIC FORMULAS

Binary Ionic Compounds (Type II):

- ***Type II*** ions are those cations that ***form more than one ion***.
- When naming compounds formed from these ions, include the ***ionic charge as Roman numeral***, in parentheses, after the metal's name.

? –1 **$FeCl_2$**	**+2 –1** **$FeCl_2$**	**Iron(II) chloride**
? –1 **$FeCl_3$**	**+3 –1** **$FeCl_3$**	**Iron(III) chloride**
? –2 **Cu_2O**	**+1 –2** **Cu_2O**	**Copper(I) oxide**
? –2 **CuO**	**+2 –2** **CuO**	**Copper(II) oxide**

Examples:

1. Name each of the following Type II ionic compounds:

   $SnCl_2$: ___________________

   Cu_2S: _____________________

   Fe_2O_3 ____________________

2. Write formulas for each of the following compounds:

   Tin(II) bromide: ___________________

   Lead (IV) oxide: ____________________

   Nickel (III) sulfide:__________________

POLYATOMIC IONS

- Some ionic compounds contain ***polyatomic ions***, an ion composed of ***several atoms bound together***.

- Some common polyatomic ions are:

NH_4^{+}	ammonium	OH^{-}	hydroxide
NO_3^{-}	nitrate	CN^{-}	cyanide
SO_4^{2-}	sulfate	$C_2H_3O_2^{-}$	acetate
PO_4^{3-}	phosphate	HCO_3^{-}	bicarbonate
CO_3^{2-}	carbonate		

- When writing formulas for compounds containing polyatomic ions, treat the polyatomic ion as one group.

potassium nitrate	**1+ 1–** $K_{?}\ (NO_3)_{?}$	KNO_3
calcium hydroxide	**2+ 1–** $Ca_{?}\ (OH)_{?}$	$Ca(OH)_2$
ammonium acetate	**1+ 1–** $(NH_4)_{?}\ (C_2H_3O_2)_{?}$	$NH_4C_2H_3O_2$
sodium sulfate	**1+ 2–** $Na_{?}\ (SO_4)_{?}$	Na_2SO_4
copper(II) nitrate	**2+ 1–** $Cu_{?}\ (NO_3)_{?}$	$Cu(NO_3)_2$

Examples:

1. Write formulas for the following polyatomic compounds:

   sodium carbonate: _________________

   ammonium sulfide: _________________

   magnesium bicarbonate: __________________

POLYATOMIC IONS

- ***Polyatomic*** ionic compounds are named by naming the cation first, followed by the polyatomic ion.

Na_3PO_4	**sodium phosphate**
NH_4Br	**ammonium bromide**
$CuNO_3$	**copper (I) nitrate or cuprous nitrate**
$Pb(CO_3)_2$	**lead (IV) carbonate or plumbic carbonate**

<u>Examples:</u>

2. Name the following polyatomic compounds:

$Mg(OH)_2$: ________________

NaCN: __________________

$Fe_2(SO_4)_3$: ________________

THE MOLE CONCEPT

- Chemists find it more convenient to use ***mass relationships*** in the laboratory, while ***chemical reactions*** depend on the ***number of atoms*** present.
- In order to relate the ***mass and number of atoms***, chemists use the SI unit ***mole*** (abbreviated ***mol***).
- The number of ***particles*** in a ***mole*** is called ***Avogadro's number*** and is ***$6.02x10^{23}$***.

1 mol of H atoms………contains: **6.02×10^{23}** H atoms

1 mol of H_2 molecules………contains: **6.02×10^{23}** H_2 molecules
2 x($6.02x10^{23}$) H atoms

1 mol of **H_2O** molecules….contains: $6.02x10^{23}$ H_2O molecules
2 x ($6.02x10^{23}$) H atoms
1 x ($6.02x10^{23}$) O atoms

1 mol of Na^+ ions………..contains: **6.02×10^{23}** Na^+ ions

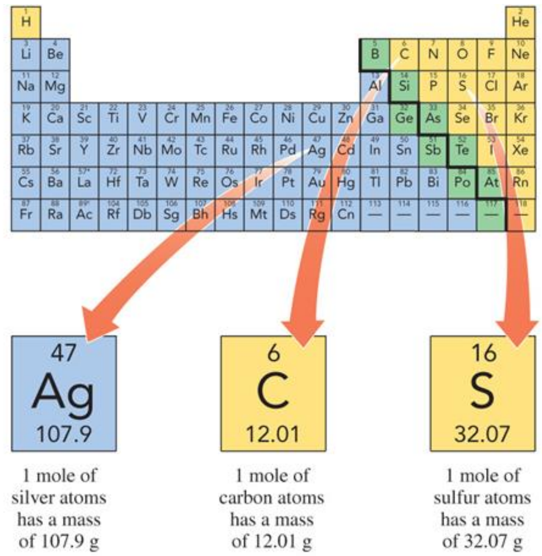
- The **atomic mass** of one atom expressed in **amu** is numerically the same as the **mass of one mole** of atoms of the element expressed in **grams**.

Element	Mass of one atom	Mass of one mole of atoms
H	1.01 amu	1.01 grams
Mg	24.30 amu	24.30 grams
Cl	35.45 amu	35.45 grams

MOLAR MASS

- The mass of one mole of a substance is called ***molar mass*** and is measured in ***grams***.

- Molar mass of elements are listed on the periodic table.



- The mass of one mole of a compound is the sum of the atomic masses of all elements in the compound.

Mass of one mole of H_2O

2 mol H = 2 (1.01 g) = 2.02 g
1 mol O = 1 (16.00 g) = 16.00 g
18.02 g ← **Molar Mass**

Mass of one mole of $Ca(OH)_2$

1 mol Ca = 1 (40.08 g) = 40.08 g
2 mol O = 1 (16.00 g) = 32.00 g
2 mol H = 2 (1.01 g) = 2.02 g
74.10 g ← **Molar Mass**

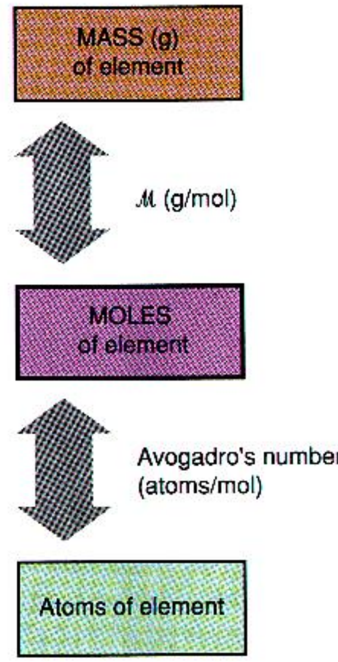
Examples:
Calculate the molar mass of each compound shown below:

1. Lithium carbonate (Li_2CO_3)

2. Salicylic acid ($C_7H_6O_3$)

CALCULATIONS USING THE MOLE CONCEPT

When solving problems involving ***mass-mole-number*** relationships of elements or compounds, we can use:



- The ***molar mass*** to convert between mass and moles.
- ***Avogadro's number*** ($6.02\text{x}10^{23}$) to convert between moles and number of particles.

<u>Examples:</u>

1. How many **moles** of iron are present in 25.0 g of iron?

$$25.0\text{ g Fe }\left(\frac{1\text{ mole}}{55.85\text{ g}}\right) = 0.448\text{ mol Fe}$$

2. What is the mass of 5.00 mole of water?

3. How many magnesium **atoms** are present in 5.00 g of Mg?

$$5.00\ \cancel{\text{g}}\ \text{Mg}\left(\frac{1\ \cancel{\text{mol}}}{24.30\ \cancel{\text{g}}}\right)\left(\frac{6.02\text{x}10^{23}\text{ atoms}}{1\ \cancel{\text{mol}}}\right) = 1.24\text{x}10^{23}\text{ atoms Mg}$$

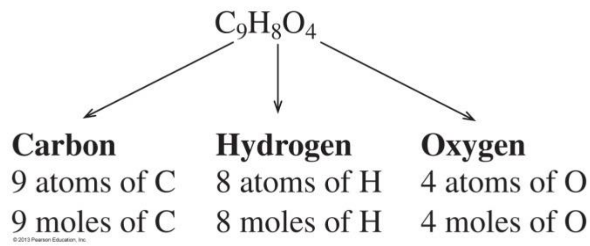
4. How many molecules of HCl are present in 25.0 g of HCl?

5. What is the mass (in grams) of $8.5\text{x}10^{18}$ atoms of carbon?

MOLES OF ELEMENTS IN A FORMULA

- The subscripts in a chemical formula of a compound indicate the number of atoms of each type of element. For example, in a molecule of aspirin, $C_9H_8O_4$, there are 9 carbon atoms, 8 hydrogen atoms and 4 oxygen atoms.

- The subscript also indicates the number of moles of each element in one mole of the compound. For example, one mole of aspirin contains 9 moles of carbon atoms, 8 moles of hydrogen atoms and 4 moles of oxygen atoms.



- Using the subscripts from the aspirin formula, one can write the following conversion factors for each of the elements in 1 mole of aspirin:

$$\frac{\textbf{9 moles C}}{\textbf{1 mole } C_9H_8O_4} \qquad \frac{\textbf{8 moles H}}{\textbf{1 mole } C_9H_8O_4} \qquad \frac{\textbf{4 moles O}}{\textbf{1 mole } C_9H_8O_4}$$

Examples:

1. Determine the moles of C atoms in 1 mole of each of the following substances:

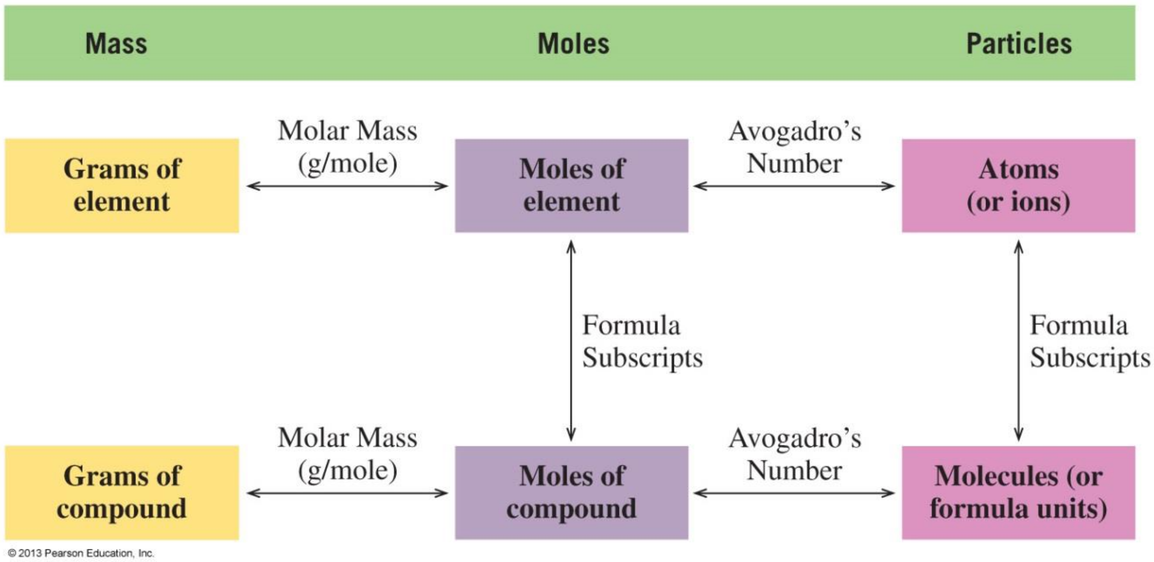
   a) Acetaminophen used in Tylenol, $C_8H_9NO_2$

   b) Zinc dietary supplement, $Zn(C_2H_3O_2)_2$

2. How many carbon atoms are present in 1.50 moles of aspirin, $C_9H_8O_4$?

3. A patient is restricted to intake of 120 mg of sodium in their diet. How many grams of salt (NaCl) can he ingest daily without exceeding this limit?

SUMMARY OF MASS-MOLE CALCULATIONS



ELECTROLYTES & NON-ELECTROLYTES

- Solutions can be characterized by their ability to conduct an electric current. Solutions containing ***ions*** are ***conductors*** of electricity and those that contain ***molecules*** are ***non-conductors***.

- Substances that dissolve in water to form ions are called ***electrolytes***. The ions formed from these substances conduct electric current in solution, and can be tested using a conductivity apparatus (diagram below).

- Electrolytes are further classified as ***strong electrolytes*** and ***weak electrolytes***.

- In water, a ***strong electrolyte*** exists ***only as ions***. Strong electrolytes make the light bulb on the conductivity apparatus glow brightly. ***Ionic substances*** such as NaCl are ***strong electrolytes***.

$$NaCl\ (s) \xrightarrow{H_2O} Na^+\ (aq) + Cl^-\ (aq)$$

(only ions present after solution)

- Solutions containing ***weak electrolytes*** contain only a ***few ions***. These solutions make the light bulb on the conductivity apparatus glow dimly. Weak acids and bases that dissolve in water and produce few ions are ***weak electrolytes.***

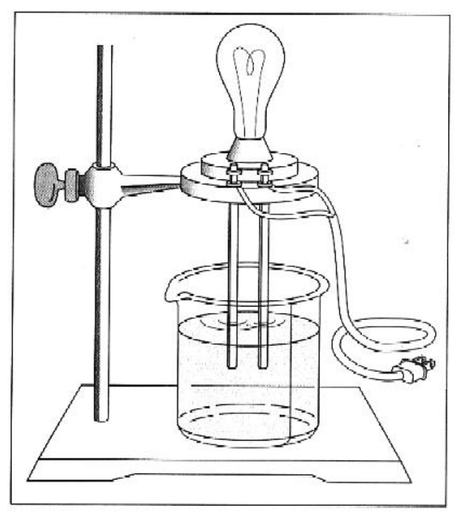
$$HF\ (aq) \underset{\text{Recombination}}{\overset{\text{Dissociation}}{\rightleftarrows}} H^+\ (aq) + F^-\ (aq)$$

(few ions present after solution)

- Substances that ***do not form any ions*** in solution are called ***non-electrolytes***. With these solutions the bulb on the conductivity apparatus does not glow. ***Covalent molecules*** that dissolve in water but do not form ions, such as sugar, are ***non-electrolytes***.

$$C_{12}H_{22}O_{11}\ (s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}\ (aq)$$

(no ions present after solution)



ELECTROLYTES & NON-ELECTROLYTES

Classification of Solutes in Aqueous Solutions

Type of Solute	Dissociation	Particles in Solution	Conducts Electricity?	Examples
Strong electrolyte	Complete	Ions only	Yes	Ionic compounds such as NaCl, KBr, $MgCl_2$, $NaNO_3$; NaOH, KOH; HCl, HBr, HI, HNO_3, $HClO_4$, H_2SO_4
Weak electrolyte	Partial	Mostly molecules and a few ions	Yes, but poorly	HF, H_2O, NH_3, $HC_2H_3O_2$ (acetic acid)
Nonelectrolyte	None	Molecules only	No	Carbon compounds such as CH_3OH (methanol), C_2H_5OH (ethanol), $C_{12}H_{22}O_{11}$ (sucrose), CH_4N_2O (urea)

Examples:

1. Identify the predominant particles in each of the following solutions and write the equation for the formation of the solution:

   a) NH_4Br

   b) Urea (CH_4N_2O)

   c) HClO (weak acid)

DOUBLE REPLACEMENT REACTIONS

- Double replacement reactions can be subdivided into one of the following subgroups:

1. ***Precipitation:***
In these reactions one of the products formed is an insoluble solid called a ***precipitate***. For example, when solutions of potassium chromate, K_2CrO_4 , and barium nitrate, $Ba(NO_3)_2$, are combined an insoluble salt barium chromate, $BaCrO_4$, is formed.

$$K_2CrO_4\ (aq) + Ba(NO_3)_2\ (aq) \rightarrow \underbrace{BaCrO_4\ (s)}_{\text{Precipitate}} + 2KNO_3\ (aq)$$

These reactions will be further discussed in Chapter 8

2. ***Neutralization:***
The most important reaction of acids and bases is called **neutralization**. In these reactions an acid combines with a base to form a **salt and water**. For example:

$$\underset{\text{acid}}{HCl\ (aq)} + \underset{\text{base}}{NaOH\ (aq)} \longrightarrow \underset{\text{salt}}{NaCl\ (aq)} + \underset{\text{water}}{H_2O\ (l)}$$

Salts are ***ionic*** substances with the ***cation*** donated from the ***base*** and the ***anion*** donated from the ***acid***. In the laboratory, neutralization reactions are observed by an increase in temperature (exothermic reaction).

3. ***Unstable product:***
Some chemical reactions ***produce gas*** because one of the products formed in the reaction is ***unstable***. Two such products are listed below:

Carbonic acid	$H_2CO_3\ (aq) \rightarrow CO_2\ (g) + H_2O\ (l)$
Sulfurous acid	$H_2SO_3\ (aq) \rightarrow SO_2\ (g) + H_2O\ (l)$

When either of these products appears in a chemical reaction, they should be replaced with their decomposition products.

$$2\ HCl\ (aq) + Na_2CO_3\ (aq) \rightarrow 2\ NaCl\ (aq) + \underset{\textbf{(unstable)}}{H_2CO_3\ (aq)}$$

$$2\ HCl\ (aq) + Na_2CO_3\ (aq) \rightarrow 2\ NaCl\ (aq) + CO_2\ (g) + H_2O\ (l)$$

SOLUBLE & INSOLUBLE SALTS

- Many ionic solids ***dissolve*** in water and are called ***soluble salts***. However, some ionic solids ***do not dissolve*** in water and do not form ions in solution. These salts are called ***insoluble salts*** and remain solid in solution.

1. Chemists use a set of ***solubility rules*** to ***predict*** whether a salt is ***soluble or insoluble***. These rules are summarized below:

SOLUBLE	**NO_3^-**	**No exceptions**
	Na^+, K^+, NH_4^+	**No exceptions**
	Cl^-, Br^-, I^-	**Excepts for those containing Ag^+, Pb^{2+}**
	SO_4^{2-}	**Except for those containing Ba^{2+}, Pb^{2+}, Ca^{2+}**
INSOLUBLE	**S^{2-}, CO_3^{2-}, PO_4^{3-}**	**Except those containing Na^+, K^+, NH_4^+**
	OH^-	**Except those containing Na^+, K^+, Ca^{2+}, NH_4^+**

Examples:

1. Use the solubility rules to determine if each of the following salts are soluble or insoluble:

   a) K_3PO_4 ______________________

   b) $CaCO_3$ ______________________

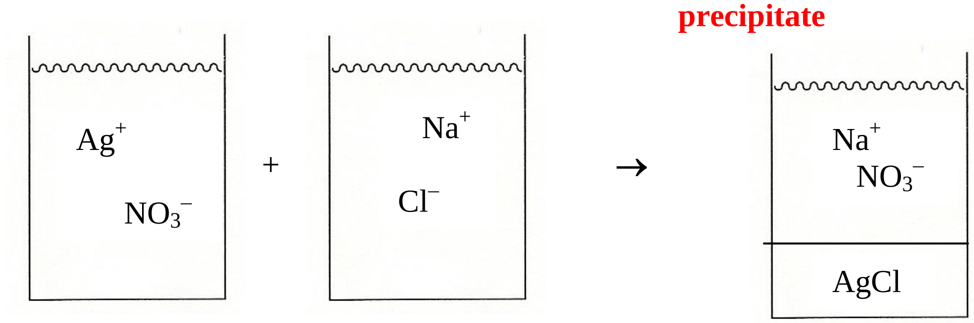
FORMATION OF A SOLID

- Solubility rules can be used to predict whether a solid, called a ***precipitate***, can be formed when two solutions of ionic compounds are mixed.

- A solid is formed when two ions of an insoluble salt come in contact with one another.

- For example, when a solution of $AgNO_3$ is mixed with a solution of NaCl , a white insoluble salt AgCl is produced.

$AgNO_3$ (aq) + NaCl (aq) → AgCl (s) + $NaNO_3$ (aq)
precipitate



- Double replacement reactions in which a precipitate is formed are called ***precipitation*** reactions.

- The solubility rules can be used to predict whether a precipitate forms when two solutions of ionic compounds are mixed together. The stepwise process is outlined below:

  1. Write the ***molecular equation*** for the reaction by predicting the products formed from the combination of the reactants. Use the solubility rules to determine if any of the products are insoluble. Label all the states and balance the equation.

  2. Using the molecular equation above, write the ***complete ionic equation*** by breaking all of the soluble compounds into their corresponding ions; leave the precipitate compound together as molecular.

  3. Using the complete ionic equation above, write the ***net ionic equation (NIE)*** by cancelling all ions that appear as the same on both sides of the equation. These ions are called ***spectator*** ions.

  4. If no precipitate forms in step 2, write "NO REACTION" after the arrow and stop.

PRECIPITATION REACTIONS

- For example, the reaction of $AgNO_3$ and NaCl, can be predicted as shown below:

Step 1: $AgNO_3$ (aq) + NaCl (aq) → ?????????

$AgNO_3$ (aq) + NaCl (aq) → AgCl (**?**) + $NaNO_3$ (**?**)

$AgNO_3$ (aq) + NaCl (aq) → AgCl (**s**) + $NaNO_3$ (**aq**) (**molecular equation**)

Step 2: Ag^+ + NO_3^- + Na^+ + Cl^- → AgCl (**s**) + Na^+ + NO_3^- (**complete ionic eq**)

Step 3: -Ag^+ + ~~NO_3^-~~ + ~~Na^+~~ + Cl^- → AgCl (**s**) + ~~Na^+~~ + ~~NO_3^-~~

Net Ionic Equation: $Ag^+ + Cl^- \rightarrow AgCl\ (s)$

precipitate

Examples:

Predict the products for each reaction shown below and write molecular, complete ionic and net ionic equations. If no reaction occurs, write "No Reaction" after the arrow.

1. Na_2SO_4 (aq) + $Pb(NO_3)_2$ (aq) → ???????

Step 1:

Step 2

:

Step 3: