Chemistry 3A

Introductory General Chemistry

S. M. Halloran

Fresno City College

- Word Equations
- Chemical Equations
- Balancing Equations
- Predicting Reactions: Single & Double Displacement Reactions
- Writing Chemical Equations for Reactions in Solution
  - Complete Chemical Equations
  - Complete Ionic Equations
  - Net Ionic Equations
- Oxidation/Reduction
  - How to understand them and to apply them

Chemical Reactions & Conservation Laws

- Chemical reactions are expressed in a way just as mathematics uses equations: to show something is connected at a time before with something that comes after
- A reaction is an expression of substances that react—the **reactants**—and the result of the reaction—the **products**

Reactants → Products

- The expression of these is about the conservation laws of mass (matter), energy, and electric charge

Equations in Words

- Chemical reactions are expressed essentially as equations, and the expression can be in words
- Silver metal solid is reacted with sulfur with the result it forms silver sulfide
  - If you have ever seen tarnish on silver metal set, you see silver sulfur compound
  - When natural gas (methane) is burned in a Bunsen burner, the gas reacts with oxygen (O_2) in the air to produce carbon dioxide (CO_2) and water (H_2O)
- A problem with equations as words is that they can not indicate the quantitative nature of reactions

See also **https://youtu.be/vTq4sgGd2QU**

Figure 10.1.2: You can see dark spots of tarnished (silver sulfide) forming on this ring as it reacts with sulfur compounds in the air. (Credit: CK-12 Foundation; Source: CK-12 Foundation; License: CK-12 Curriculum Materials license)

$$\text{Methane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$

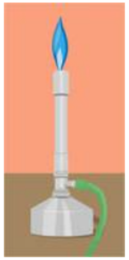


Figure 10.1.3: A Bunsen burner is commonly used to heat substances in a chemistry lab. Methane is reacted with oxygen to form carbon dioxide and water. (Credit: CK-12 Foundation; Source: CK-12 Foundation; License: CK-12 Curriculum Materials license)

Steps in Chemistry = Recipe

Yum! Yum!

- Your book wants to point out that steps in laboratory work such as a chemical synthesis align with a metaphor of making a delicious preparation of shrimp gumbo

Figure 10.2.1 (Credit: jons2 at pdphoto.org; Source: http://commons.wikimedia.org/wiki/File:Shrimp_gumbo.jpg(opens in new window); License: Public Domain)

- Chemists are probably good in the kitchen as well because they must carefully, meticulously gather the ingredients and then “cook” them in the correct way to get an exacting result

Chemical Equations

- Chemical equations start with using chemical formula which express matter as the symbols of the elements of the **Periodic Table**
- The reactants and products as compounds are a sketch of the reaction. This is the **skeleton equation**

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

- The next step is to indicate the physical state of the compounds:

$$CH_4\ (g) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (l)$$

- (s) = solid, (l) = liquid, (g) = gas,
  (aq) = compound in aqueous phase

Symbols in Chemical Equations

It is important to understand the symbols used in chemical reaction equations

Table 10.2.1: Symbols Used in Chemical Equations

Symbols	Description
$+$	Used to separate multiple reactants or products.
$\rightarrow$	Yield sign; separates reactants from products.
$\rightleftharpoons$	Replaces the yield sign for reversible reactions that reach equilibrium.
(s)	Reactant or product in the solid state.
(l)	Reactant or product in the liquid state.
(g)	Reactant or product in the gaseous state.
(aq)	Reactant or product in an aqueous solution (dissolved in water).
$\overset{\text{Pt}}{\rightarrow}$	Formula written above the arrow is used as a catalyst in the reaction.
$\overset{\Delta}{\rightarrow}$	Triangle indicates that the reaction is being heated.

See also ***https://youtu.be/ZcF8E8aAOGs***

Balancing Equations

- With the skeleton equation up on the whiteboard, it is time to apply the laws of the conservation of mass (electric charge & energy can be examined later)
- One of the most important beginnings in chemistry is writing reactions (equations) that show the formation of compounds from their elemental forms

practice exercises calculating the enthalpy of formation is often seen with this

- The elemental form of carbon is C (*s*). The elemental form of hydrogen is H_2 (*g*)
- From these two, the formation of methane CH_4 (*g*) is shown

$$\mathbf{C}\,(s) + \mathbf{H_2}\,(g) \rightarrow \mathbf{CH_4}\,(g)$$

- The equation is missing mass balance however, which it must have on both sides of the reaction arrow
- There is one atom of C on both sides, so that's OK, but there are 2 atoms of H on left side, but 4 atoms on the right. The left needs 2 more atoms, which can be done by doing the following:

$$\mathbf{C}\,(s) + \mathbf{2\ H_2}\,(g) \rightarrow \mathbf{CH_4}\,(g)$$

Details of Balancing Equations

- In the previous slide, a **2** was added to the H_2 (*g*) reactant to achieve the balance of atoms on both sides of the reaction arrow

$$\mathbf{C}\ (s) + \mathbf{2\ H_2}\ (g) \rightarrow \mathbf{CH_4}\ (g)$$

- This **2** and *any whole number/integer* placed in front of a chemical compound in a chemical equation is called a **coefficient**

This is also term in mathematics for a number placed in front of a variable in an algebraic expression

- Putting a coefficient in front of a compound has the effect of multiplying the number of each atom in the compound by the number of the coefficient, and if the atom is subscripted, by that subscript number too
- 2 H_2 = 2 x 2 H = 4 H atoms
  2 CO_2 = 2 x 1 C = 2 C atoms; 2 x 2 O = 4 O atoms
  4 H_2O = 4 x 2 H = 8 H atoms; 4 x 1 O = 4 O atoms
  3 MnO_4 = 3 x 1 Mn = 3 Mn atoms; 3 x 4 O = 12 O atoms

Steps To Balancing Equations

1. Set up the skeleton equation: write down the correct chemical formulas on each side of the reaction arrow, the reactants and products
2. Identify all atoms and count the number of atoms on the reactant and product side

If polyatomic ions are unchanged on both sides, count it as a unit

3. Begin the balancing by focusing on elements that occur in only one compound on one side

For instance, oxygen often in appears in more than one compound on one side the arrow: do not start with balancing oxygen atoms

4. Balancing can ONLY be done by using coefficients: changing subscripts changes the chemical nature of a compound, and this is not permitted
5. When satisfied of completion, verify the numbers of atoms on both sides of reaction arrow
6. Make sure coefficients have lowest possible ratio, while remaining integers

Practice Balancing Equations

1. Set up the skeleton equation: write down the correct chemical formulas on each side of the reaction arrow, the reactants and products
2. Identify all atoms and count the number of atoms on the reactant and product side

   If polyatomic ions are unchanged on both sides, count it as a unit

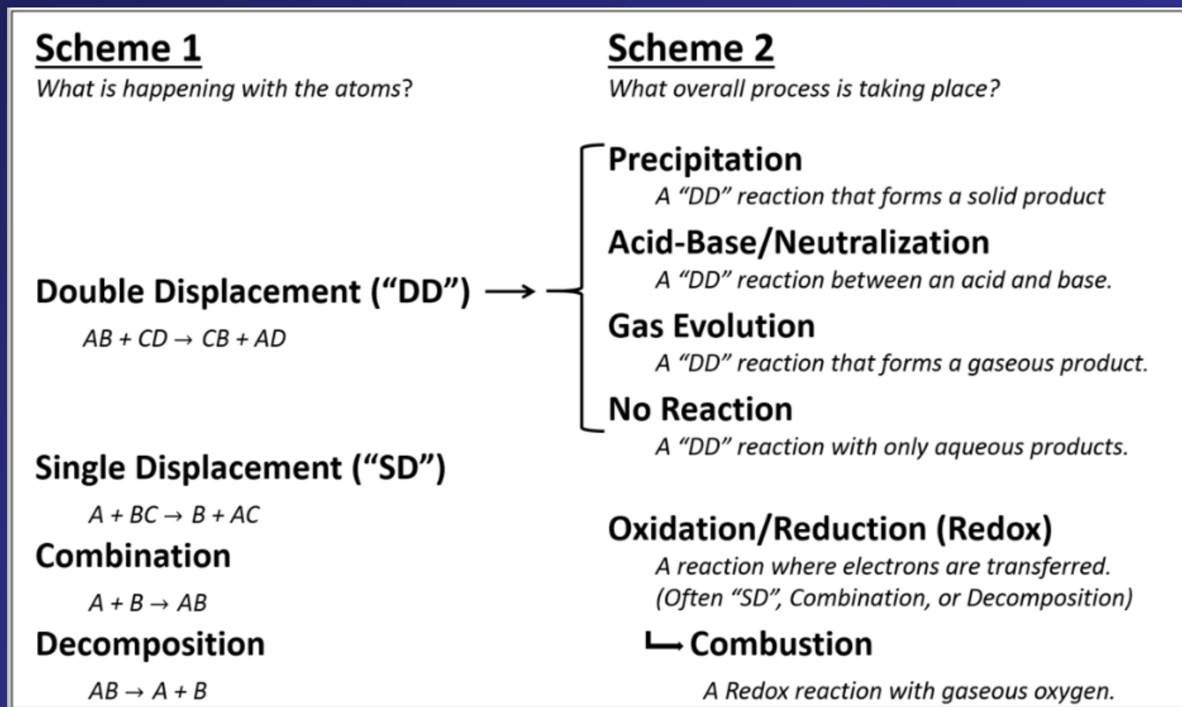
3. Begin the balancing by focusing on elements that occur in only one compound on one side

   For instance, oxygen often in appears in more than one compound on one side the arrow: do not start with balancing oxygen atoms

4. Balancing can ONLY be done by using coefficients: changing subscripts changes the chemical nature of a compound, and this is not permitted
5. When satisfied of completion, verify the numbers of atoms on both sides of reaction arrow
6. Make sure coefficients have lowest possible ratio, while remaining integers

Types of Reactions

- Chemical reactions will fall under certain types/categories on one level, and then on an additional level with question "what is happening with atoms?"
- On an additional level, the question is "what process is happening?"
- You will learn to identify these types or categories of reactions



Combination Reactions

$$A + B \rightarrow AB$$

- This reaction type is about ***synthesis***: two substances combine to form a single (third) compound
- Sodium metal solid reacts with chlorine gas to form the solid sodium chloride

$$2\ Na\ (s) + Cl_2\ (g) \rightarrow 2\ NaCl\ (s)$$

Oxide Formation: A Combination Reaction

- Oxygen O_2 in the air reacting with numerous elements is perhaps one of the most important significant of reactions

$$2\ Mg\ (s) + O_2\ (g) \rightarrow 2\ MgO\ (s)$$

This reaction was one of your recent lab experiments!!

$$S\ (s) + O_2\ (g) \rightarrow SO_2\ (g)$$

$$2\ S\ (s) + 3\ O_2\ (g) \rightarrow 2\ SO_3\ (g)$$

$$2\ Fe\ (s) + O_2\ (g) \rightarrow 2\ FeO\ (s)$$

$$4\ Fe\ (s) + 3\ O_2\ (g) \rightarrow 2\ Fe_2O_3\ (s)$$

$$4\ K\ (s) + O_2\ (g) \rightarrow 2\ K_2O\ (s)$$

Decomposition Reactions

AB → A + B

- The **decomposition reaction** type appears to be the very opposite of the combination reaction
- One substance (shown as "parts" of A and B) is "decomposed" to form two or more substances/compounds
- Many of these reactions require an input of energy—an "activation energy"—for the reaction to proceed

Figure 10.4.2.2: *Mercury (II) oxide is a red solid. When it is heated, it decomposes into mercury metal and oxygen gas.* (Credit: Ben Mills (User:Benjah-bmm27/Wikimedia Commons);

$2\ HgO\ (s) \rightarrow 2\ Hg\ (l) + O_2\ (g)$

A compound decomposes to elements in this reaction

Decomposition Reactions

- Decomposition reaction include compounds decomposing to compounds:

$$CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$$

- Here metal oxides are formed from their hydroxides

$$2\ NaOH\ (s) \rightarrow Na_2O\ (s) + H_2O\ (l)$$

- Acids like the unstable carbonic acid will decompose quickly to nonmetal oxides and water

$$H_2CO_3\ (aq) \rightarrow CO_2\ (g) + H_2O\ (l)$$

- Potassium chlorate is a well-known decomposition reaction

$$2\ KClO_3\ (s) \rightarrow 2\ KCl\ (s) + H_2O\ (l)$$

- The electrolysis of water is a well-known decomposition

$$H_2O\ (l) \rightarrow H_2\ (g) + O_2\ (g)$$

Combustion Reactions

$$X + O_2\ (g) \rightarrow XO_n + \ldots$$

- **Combustion reactions** always involve gaseous oxygen (O_2) as a reactant
- Humans learned the hard way with the Hindenberg airship disaster that hydrogen gas (H_2) also gets combusted

$$2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$$

- Are you a gas griller? Here is the combustion of propane:

$$C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$$

Among the many important things to learn in chemistry, one of these is just how much oxygen is involved in much of chemistry

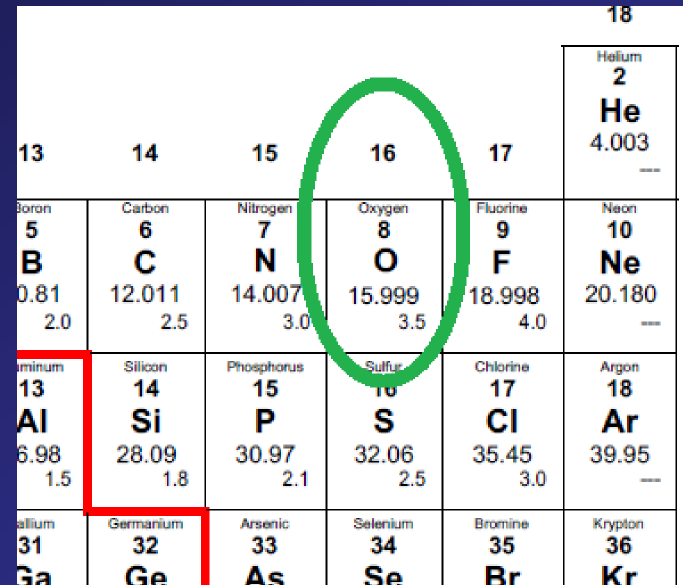


Figure 10.4.3.2: Explosion of the Hindenberg. (Credit: Courtesy of Gus Pasquerella/US Navy; Source:

Single Replacement Reactions

- Silver tarnishes. The metal reacts with hydrogen sulfide (H_2S) which is always present in small amounts in the air

$2\ Ag\ (s) + H_2S\ (g) \rightarrow Ag_2S\ (s) + H_2\ (g)$

- This is an example of a **single-replacement reaction**

$$\mathbf{A + BC \rightarrow AC + B}$$

- One (usually metal) **element A** replaces a similar (also metal) **element B** in a compound
- If **A** is a nonmetal, it will replace the nonmetal that is **B** with **C**

Single Replacement Reactions

- Metal Replacement

$Mg\ (s) + Cu(NO_3)_2\ (aq) \rightarrow$
$Mg(NO_3)_2\ (aq) + Cu\ (s)$

Mg more reactive than Cu and replaces it

- Hydrogen Replacement

$Zn\ (s) + 2\ HCl\ (aq) \rightarrow$
$ZnCl_2\ (aq) + H_2\ (g)$

Acidic proton is replaced by active metal

Water is not usually considered an acid, but some metals are so reactive that they replace the H in H_2O to form the metal hydroxide

$2\ Na\ (s) + 2\ H_2O\ (l) \rightarrow 2\ NaOH\ (aq) + H_2\ (g)$

Figure 10.4.4.2: Zinc metal reacts with hydrochloric acid to give off hydrogen gas in a single-displacement reaction. (Credit: User: Chemicalinterest/Wikimedia Commons; Source: http://commons.wikimedia.org/wiki/File:Zn_reaction_with_HCl.JPG(opens in new window); License: Public Domain)

Figure 10.4.4.3: Sodium metal reacts vigorously with water, giving off hydrogen gas. A large piece of sodium will often generate so much heat that the hydrogen will ignite. (Credit: User:Ajhalls/Wikimedia Commons;

Single Replacement Reactions

- Halogen Replacement

The interesting thing about this reaction is that diatomic chlorine molecule replaces bromine ion to produce chloride ion and diatomic bromine molecule (it actually takes electrons away from Br in an oxidation reaction)

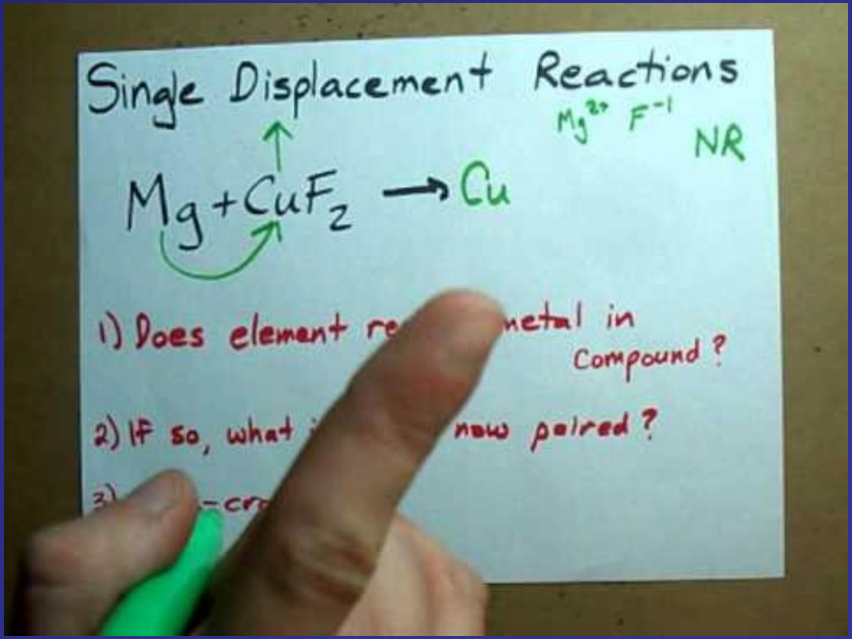
$$Cl_2\ (g) + 2\ NaBr\ (aq) \rightarrow 2\ NaCl\ (aq) + Br_2\ (l)$$

Cl_2 will do the same to NaI.

Fluorine (F_2) will do the same to NaCl or NaBr or NaI.

And Br_2 will do the same to NaI. All these elements are in same Group (column), and they show a reactivity based on the period (row) they are in

Reactivity decreases with increasing period



Double Replacement Reactions

AB + CD → AD + CB

- This reaction is when the cations and anions of two ionic compounds exchange places to form two different ionic compounds
  - **A** & **C** are the cations, **B** and **D** are the anions
  - These reactions happen when one of the products is a **precipitate**, **gas** or a **molecule** like **H_2O**
- Precipitate

$2\ KI\ (aq) + Pb(NO_3)_2\ (aq) \rightarrow$
$2\ KNO_3\ (aq) + PbI_2\ (s)$

Figure 10.4.5.2: Formation of lead iodide precipitate. (Credit: Paige Powers - "Lead Iodide"; Source:

Pb^{2+} and I^- ions strongly bind and form a solid

Double Replacement Reactions

- Gas Formation

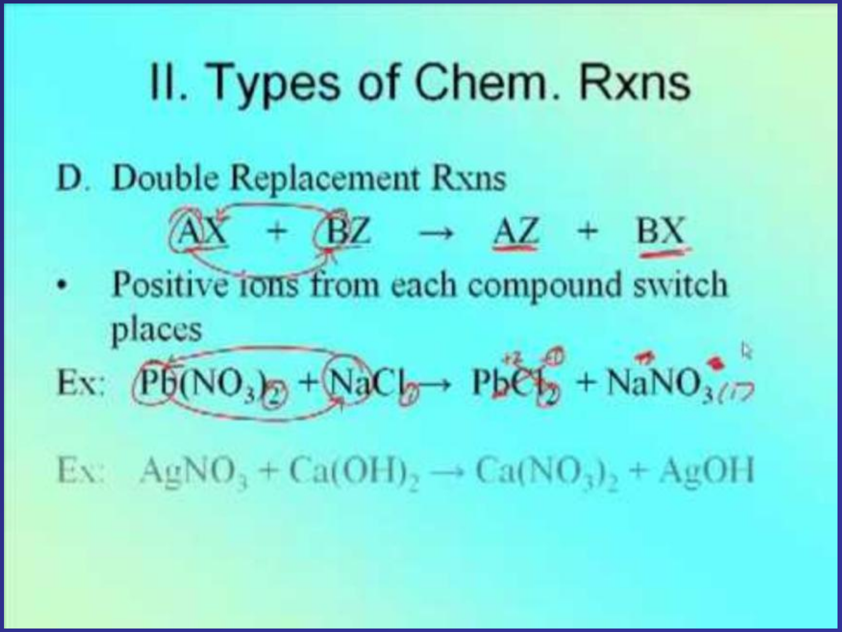
Na_2S (*aq*) + 2 HCl (*aq*) → 2 NaCl (*aq*) + H_2S (*g*)

The H_2S gas bubbling out actually drives the action further to the right, promoting the reaction

- Molecular Compound Formation

HCl (*aq*) + NaOH (*aq*) → NaCl (*aq*) + H_2O (*l*)

This is actually a classic reaction of a **strong acid** with a **strong base** to produce a **salt** and **water**



Predicting Single Displacement Reactions

- The hydrogen and halogen single replacement reactions were discussed a couple of slides back
- For ionic compounds, elements of the same type will replace elements of that type
  - Metal cations/elements replace metal cations
  - Nonmetal anions/elements replace nonmetal anions
- $FeCl_2 + Ca \rightarrow ?$

  $CaCl_2 + Fe$
- $CaBr_2 + F_2 \rightarrow ?$

  $CaF_2 + Br_2$
- $FeI_2 + Cl_2 \rightarrow ?$

  $FeCl_2 + I_2$
- $AlPO_4 + Mg \rightarrow ?$

  $Mg_3(PO_4)_2 + Al$

Predicting Double Displacement Reactions

- There are patterns and trends that can be seen with these reactions making them predictable

1. Exchange anions and cations on reactant side to form new product compounds
2. Since charge magnitude may differ on cations or anions (1+ and 2+, or 1– and 2–) between reactants and products, formulas may need to be corrected (remember the criss-cross method)
3. Balance the equation

- Consider barium chloride ($BaCl_2$) and lithium sulfate (Li_2SO_4) next

Predicting Double Displacement Reactions

1. *Exchange anions and cations on reactant side to form new product compounds*

$BaCl_2\ (aq) + Li_2SO_4\ (aq) \rightarrow BaSO_4 + LiCl$

- In specifying the products, it is vital to follow write the correct chemical formula for compounds and not worry about balance of atom numbers on both sides of arrow at the moment
- Polyatomic ions (like sulfate SO_4^{2-}) despite being composed of multiple elements must be maintained as the groups of elements they are which gives them their identity as polyatomic ions as they are connected by covalent and not ionic bonds, despite being ions as a group of nonmetal elements

Predicting Double Displacement Reactions

2. *Since charge magnitude may differ on cations or anions (1+ and 2+, or 1– and 2–) between reactants and products, formulas may need to be corrected (remember the criss-cross method)*

$BaSO_4$, LiCl

- The exchanged ions Ba^{2+} with SO_4^{2-}, and of Li^{+} with Cl^{-}, created no problem with matching charges of the ions: 1+ matched with 1– and 2+ matched with 2–
- Had this matching not of ions by charge state not occurred, determining correct subscripts to monatomic ions or polyatomic ions would have been necessary

Predicting Double Displacement Reactions

3. *Balance the equation*

$BaCl_2$ (*aq*) + Li_2SO_4 (*aq*) → $BaSO_4$ (s) + 2 LiCl (*aq*)

- At this point, the compounds are determined. Subscripts to monatomic or polyatomic ions will not be changed
- To achieve balance, coefficients will be determined (added) to the compounds to obtain the necessary result
- In the current reaction, it was only necessary to balance with two of the LiCl ionic compound

Practice: Predicting Double Replacement

Predict mixing of: $RbOH\ (aq) + CoCl_2\ (aq)$

1. *Exchange cations/anions*

$$RbOH\ (aq) + CoCl_2\ (aq) \rightarrow RbCl + Co(OH)$$

Note that hydroxide is parenthesized to show as a polyatomic ion it should not have its elements separated

2. *Adjust formulas (subscripts) of products to correct for ion charge states*

$$RbOH\ (aq) + CoCl_2\ (aq) \rightarrow RbCl + Co(OH)_2$$

RbCl needed no correction, but cobalt is a 2+ ion and hydroxide (OH) is a 1– ion, so two OH groups were needed

3. *Balance equation and add phase indicators*

$$RbOH\ (aq) + CoCl_2\ (aq) \rightarrow RbCl\ (aq) + Co(OH)_2\ (s)$$

No coefficients were found to be needed since there was a balance of atoms/elements on both sides. The only thing added was the phase indicators, in particular to show that a precipitate in cobalt(II) hydroxide was formed

Practice: Predicting Double Replacement

Predict mixing of: $SrBr_2$ (*aq*) + $Al(NO_3)_3$ (*aq*)

1. *Exchange cations/anions*

$SrBr_2$ (*aq*) + $Al(NO_3)_3$ (*aq*) → $Sr(NO_3)$ + $AlBr$

Note there is NO subscripts on the product side: only one of each of the cations and anions, keeping in mind nitrate is polyatomic ion so parentheses are put around it to note that, keep track of it

2. *Adjust formulas (subscripts) of products to correct for ion charge states*

$SrBr_2$ (*aq*) + $Al(NO_3)_3$ (*aq*) → $Sr(NO_3)_2$ + $AlBr_3$

The compounds have now been "criss-crossed" to reflect correct subscripts of monatomic/polyatomic numbers

3. *Balance equation*

3 $SrBr_2$ (*aq*) + 2 $Al(NO_3)_3$ (*aq*) → 3 $Sr(NO_3)_2$ + 2 $AlBr_3$

The balance of this equation was selected to be the most difficult of reactions. Usually balancing is not difficult but this was a big test. There was one of each of Sr and Al on both sides, so balancing does not start there. With $(NO_3)_3$ reactant and $(NO_3)_2$ product, do a criss-cross of coefficients [2 $(NO_3)_3$] and [3 $(NO_3)_3$]. And reactant Br_2 and product Br_3, do same: [3 Br_2] and [2 Br_3] and it all works!

Precipitation Reactions

AC + BD → AD (*s*) + BC

- Precipitations are double displacement reactions with one of products being insoluble

  $AgNO_3\ (aq) + K_2Cr_2O_7\ (aq) \rightarrow Ag_2Cr_2O_7\ (s) + KNO_3\ (aq)$

This is an unbalanced equation (can you balance it?)

- Not all mixing of compounds in solutions results in precipitations
- Mixing KBr and NaCl solutions results in no reaction except K^+, Na^+, Br^-, and Cl^- ions in solution, all at half original concentrations

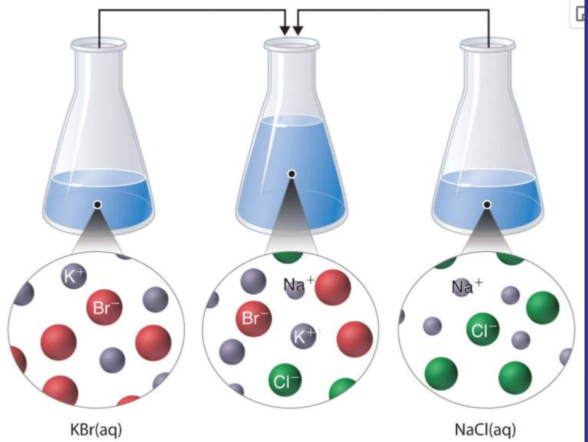


Figure 10.5.1.1: The Effect of Mixing Aqueous KBr and NaCl Solutions. Because no net reaction occurs, the only effect is to dilute each solution with the other. (Water molecules are omitted from molecular views of the solutions for clarity.)

Practice: Precipitation Reactions

The 3 steps of the double displacement add a 4th step (actually at the 3rd step) to determine if a product is a precipitate according to solubility rules on your Green Sheet

- $Sr(OH)_2$ (*aq*) is added to $FeCl_2$ (*aq*)

1. *Exchange ions*

$$Sr(OH)_2\,(aq) + FeCl_2\,(aq) \rightarrow SrCl + Fe(OH)$$

2. *Check exchanged ion charges, fix subscripts*

$$Sr(OH)_2\,(aq) + FeCl_2\,(aq) \rightarrow SrCl_2 + Fe(OH)_2$$

Both metals are 2+ ions, both anions 1–

3. *Check solubility (precipitate) and note the phases*

$$Sr(OH)_2\,(aq) + FeCl_2\,(aq) \rightarrow SrCl_2\,(aq) + Fe(OH)_2\,(s)$$

Metal hydroxides like Fe are insoluble

4. *Balance by finding coefficients*

$$Sr(OH)_2\,(aq) + FeCl_2\,(aq) \rightarrow SrCl_2\,(aq) + Fe(OH)_2\,(s)$$

Balancing not needed

Practice: Precipitation Reactions

- K_3PO_4 solid added to mercury(II) perchlorate, $Hg(ClO_4)_2$ (*aq*)

1. *Exchange ions*

K_3PO_4 (*aq*) + $Hg(ClO_4)_2$ (*aq*) → $K(ClO_4)$ + $Hg(PO_4)$

2. *Check exchanged ion charges, fix subscripts*

K_3PO_4 (*aq*) + $Hg(ClO_4)_2$ (*aq*) → $KClO_4$ + $Hg_3(PO_4)_2$

K^+ matches ClO_4^-, 3 Hg^{2+} needed for 2 PO_4^{3-}

3. *Check solubility (precipitate) and note the phases*

K_3PO_4 (*aq*) + $Hg(ClO_4)_2$ (*aq*) → $KClO_4$ (*aq*) + $Hg_3(PO_4)_2$ (*s*)

Chlorates are soluble, phosphates of mercury are not

4. *Balance by finding coefficients*

2 K_3PO_4 (*aq*) + 3 $Hg(ClO_4)_2$ (*aq*) → 6 $KClO_4$ (*aq*) + $Hg_3(PO_4)_2$ (*s*)

Start with add 3 K to product side, then 3 Hg to reactant side; then see that 3 x 2 = 6 perchlorates are on reactant side, so put 6 in front of $KClO_4$; also need 2 PO_4 groups, so put 2 in front of K_3PO_4, verify that there are now 2 x 3 = 6 K on reactants and you see 6 x 1 = 6 K on products. DONE!

Practice: Precipitation Reactions

- Solid sodium fluoride added to aqueous solution of ammonium formate

1. Exchange ions

$NaF\ (aq) + NH_4OOCH\ (aq) \rightarrow Na(OOCH) + (NH_4)F$

2. Check exchanged ion charges, fix subscripts

$NaF\ (aq) + NH_4OOCH\ (aq) \rightarrow NaOOCH + NH_4F$

All ions found to be 1+ or 1−

3. Check solubility (precipitate) and note the phases

$NaF\ (aq) + NH_4OOCH\ (aq) \rightarrow NaOOCH\ (aq) + NH_4F\ (aq)$

Everything soluble, no precipitate

4. Balance by finding coefficients

$NaF\ (aq) + NH_4OOCH\ (aq) \rightarrow NaOOCH\ (aq) + NH_4F\ (aq)$

No balancing was required in this step either

Practice: Precipitation Reactions

- Mix aqueous solutions of calcium bromide and cesium carbonate

1. *Exchange ions*

$$CaBr_2\ (aq) + Cs_2CO_3\ (aq) \rightarrow Ca(CO_3) + CsBr$$

2. *Check exchanged ion charges, fix subscripts*

$$CaBr_2\ (aq) + Cs_2CO_3\ (aq) \rightarrow CaCO_3 + CsBr$$

The exchange actually matched +2 with -2, and +1 with -1

3. *Check solubility (precipitate) and note the phases*

$$CaBr_2\ (aq) + Cs_2CO_3\ (aq) \rightarrow CaCO_3\ (s) + CsBr\ (aq)$$

Carbonates are insoluble for Ca^{2+}, bromides are soluble for Cs^{+}

4. *Balance by finding coefficients*

$$CaBr_2\ (aq) + Cs_2CO_3\ (aq) \rightarrow CaCO_3\ (s) + 2\ CsBr\ (aq)$$

Inspection shows Ca and CO_3 are balanced; add 2 to CsBr to balance Cs atom. The addition caused a balance of Br to the left; DONE

Acid-Base Neutralization

- **Acids** react with **bases** to **neutralize** each other with the products being a **salt** and **water**

$$NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$$

$$2\ NaOH\ (aq) + H_2SO_4\ (aq) \rightarrow Na_2SO_4\ (aq) + 2\ H_2O\ (l)$$

Neutralize nitric acid with calcium hydroxide

1. *Exchange ions*

$$Ca(OH)_2\ (aq) + HNO_3\ (aq) \rightarrow Ca(NO_3)\ (aq) + H_2O\ (l)$$

2. *Check exchanged ion charges for product, fix subscripts*

$$Ca(OH)_2\ (aq) + HNO_3\ (aq) \rightarrow Ca(NO_3)_2\ (aq) + H_2O\ (l)$$

3. *Balance by finding coefficients*

$$Ca(OH)_2\ (aq) + 2\ HNO_3\ (aq) \rightarrow Ca(NO_3)_2\ (aq) + 2\ H_2O\ (l)$$

Ca has one on both sides; next check H atom: 2 on right, 2 x 1 + 1 = 3 on left, so try putting 2 for HNO_3 making it 4 H atoms on left, so put 2 in front of H_2O on right; check N atoms: 2 on left and 2 on right (good); check O atoms: 2 x 1 + 3 x 2 = 8 on left, same check is good on right, DONE!

Practice: Acid-Base Neutralization

- Placeholder here

Gas-Evolving Reactions

- Reactions of acids with carbonate salts usually produce a gas like carbon dioxide

$$2\ HNO_3\ (aq) + Na_2CO_3\ (aq) \rightarrow 2\ NaNO_3\ (aq) + CO_2\ (g) + H_2O\ (l)$$

$$H_2SO_4\ (aq) + CaCO_3\ (aq) \rightarrow CaSO_4\ (aq) + CO_2\ (g) + H_2O\ (l)$$

$$2\ HCl\ (aq) + CaCO_3\ (aq) \rightarrow CaCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$$

In the figure on the right, CO_2 created or released by the addition of the carbonate salt (Na_2CO_3) to dilute acid (HCl) in the tube on the stand (left) leaves the tube through the stopper into tubing that connects to a stoppered tube on the right which contains calcium hydroxide [$Ca(OH)_2$] on the right. Because $CaCO_3$ is insoluble, a precipitate forms making the solution milky

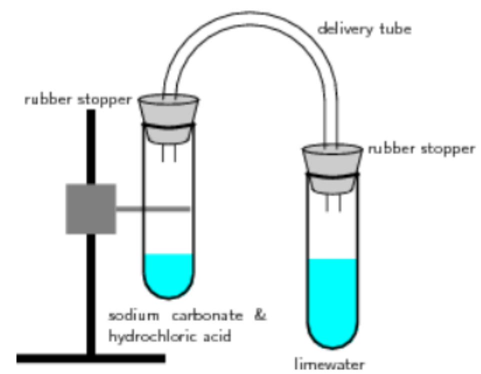


Figure 10.5.2.1: Reaction of acids with carbonates. In this reaction setup, lime water (water + calcium hydroxide) is poured into one of the test tubes and sealed with a stopper. A small amount of hydrochloric acid is carefully poured into the remaining test tube. A small amount of sodium carbonate is added to the acid, and the tube is sealed with a rubber stopper. The two tubes are connected. As a result of the acid-carbonate reaction, carbon dioxide is produced and the lime water turns milky.

Gas-Evolving Reactions

Reactant Type	Intermediate Product	Evolved Gas	Example
sulfide	none	H_2S	2 HCl (*aq*) + K_2S → H_2S (*g*) + 2 KCl (aq)
carbonates, bicarbonates	H_2CO_3	CO_2	2 HCl (*aq*) + K_2CO_3 → H_2O (*l*) + CO_2 (*g*) + 2 KCl (*aq*)
sulfites, bisulfites	H_2SO_3	SO_2	2 HCl (*aq*) + K_2SO_4 → H_2O (*l*) + SO_2 (*g*) + 2 KCl (*aq*)
ammonia	NH_4OH	NH_3	NH_4Cl (*aq*) + KOH → H_2O (*l*) + NH_3 (*g*) + 2 KCl (aq)

Gas-Evolving **REDOX** Reactions

- Reactions of acids with metals also generate a gas:

$$2\ HCl\ (aq) + Zn\ (s) \rightarrow H_2\ (g) + ZnCl_2\ (aq)$$

This reaction is really a case of **oxidation** of the metal, and the **reduction** of protons to form hydrogen gas

Oxidation is when electrons are *taken from* an element/atom in a reaction (the **redox reaction**)

Reduction is when electrons are *gained by* an element/atom in a reaction (the redox reaction)

The two **half-reactions**

$$Zn \rightarrow Zn^{2+} + 2\ e^- \ (\textit{oxidation})$$

$$2\ H^+ + 2\ e^- \rightarrow H_2\ (\textit{reduction})$$

Writing Equations for Reactions

- For the precipitation reaction of barium chloride mixed with sodium sulfate, it is written to illustrate the essence of the reaction. Call this the **complete chemical equation**

$$BaCl_2\ (aq) + Na_2SO_4\ (aq) \rightarrow BaSO_4\ (s) + NaCl\ (aq)$$

- But what really happens when compound is dissolved in an aqueous solution is that compounds do not exist in this way but are dissolved ions in solution.

  Call this the **complete ionic equation**

$$Ba^{2+}\ (aq) + 2\ Cl^{-}\ (aq) + 2\ Na^{+}\ (aq) + SO_4^{2-}\ (aq)$$
$$\rightarrow BaSO_4\ (s) + 2\ Cl^{-}\ (aq) + 2\ Na^{+}\ (aq)$$

Writing Equations for Reactions

$$Ba^{2+}\ (aq) + 2\ Cl^{-}\ (aq) + 2\ Na^{+}\ (aq) + SO_4^{2-}\ (aq)$$

$$\rightarrow BaSO_4\ (s) + 2\ Cl^{-}\ (aq) + 2\ Na^{+}\ (aq)$$

1. On the reactant and product sides, ions (Ba^{2+}, Cl^{-}, Na^{+}, SO_4^{2-}) are present in dissolved form which you have determined to be the true form
2. Stoichiometry is preserved: $BaCl_2$ was solid, and it yielded Ba^{2+} and 2 Cl^{-}. The same for Na_2SO_4: SO_4^{2-} and 2 Na^{+}
3. Ba^{2+} and SO_4^{2-} are not shown on the product side as ions because they form a solid (precipitate) which is $BaSO_4$
4. Na^{+} and Cl^{-} do not participate in the reaction and in effect cancel each other on both sides of the chemical equation or reaction arrow: these are **spectator ions**
5. The reaction can instead be written as a **net ionic equation** as follows, removing the spectator ions:

$$Ba^{2+}\ (aq) + SO_4^{2-}\ (aq) \rightarrow BaSO_4\ (s)$$

Practice Writing Equations

For the precipitation reaction with $AgNO_3$ and $CaCl_2$, write the three equations: complete chemical equation, complete ionic equation, and net ionic equation

complete chemical equation

$$2\ AgNO_3\ (aq) + CaCl_2\ (aq) \rightarrow 2\ AgCl\ (s) + Ca(NO_3)_2\ (aq)$$

complete ionic equation

$$2\ Ag^+\ (aq) + 2\ NO_3^-\ (aq) + Ca^{2+}\ (aq) + 2\ Cl^-\ (aq) \rightarrow$$
$$2\ AgCl\ (s) + Ca^{2+}\ (aq) + 2\ NO_3^-\ (aq)$$

net ionic equation

$$Ag^+\ (aq) + Cl^-\ (aq) \rightarrow AgCl\ (s)$$

Write the three equations (if needed) for Na_2SO_4 and NH_4I

complete chemical equation

$$Na_2SO_4\ (aq) + NH_4I\ (aq) \rightarrow NaI\ (aq) + (NH_4)SO_4\ (aq)$$

complete ionic equation

$$2\ Na^+\ (aq) + SO_4^{2-}\ (aq) + NH_4^+\ (aq) + I^-\ (aq) \rightarrow$$
$$2\ Na^+\ (s) + SO_4^{2-}\ (aq) + NH_4^+\ (aq) + I^-\ (aq)$$

net ionic equation

none: everything soluble, all spectator ions

Writing Equations for Reactions

- Precipitation reactions are often used to test the presence of ions in solution
- Adding $BaCl_2$ to a solution and the formation of a precipitate could indicate sulfate (SO_4^{2-}) ion presence
- Since Ba^{2+} can precipitate with other counter anions, how do we know it is $BaSO_4$? It is when the unknown solution is first acidified: all precipitates of barium dissolves in dilute acid but not the sulfate
- Halides of silver can be distinguished by the color of their precipitates when $AgNO_3$ added to an acidified unknown halide (image right)

Figure 10.6.1: The three common silver halide precipitates: AgI, AgBr and AgCl (left to right). The silver halides precipitate out of solution, but often form suspensions before settling. (CC BY-SA 3.0; Cychr).

Practice: Net Ionic Equations

- Show the balanced net ionic equations

K_2CO_3 (*aq*) + $SrCl_2$ (*aq*) → $Sr(CO_3)$ + KCl

$2\ K^+$ (*aq*) + CO_3^{2-} (*aq*) + Sr^{2+} (*aq*) + $2\ Cl^-$ (*aq*) →
$SrCO_3$ (*s*) + $2\ K^+$ (*aq*) + $2\ Cl^-$ (*aq*)

CO_3^{2-} (*aq*) + Sr^{2+} (*aq*) → $SrCO_3$ (*s*)

$FeSO_4$ (*aq*) + $Ba(NO_3)_2$ (*aq*) → $BaSO_4$ + $Fe(NO_3)$

Fe^{2+} (*aq*) + SO_4^{2-} (*aq*) + Ba^{2+} (*aq*) + $2\ NO_3^-$ (*aq*) →
$BaSO_4$ (*s*) + Fe^{2+} (*aq*) + $2\ NO_3^-$ (*aq*)

SO_4^{2-} (*aq*) + Ba^{2+} (*aq*) → $BaSO_4$ (*s*)

Oxidation-Reduction

- Oxidation is a process that is attributed to observations of corrosion ("rusting") and combustion
- Reduction was a term applied to converting metal ores (salts of metals) to the pure metal, which involved a reduction in the mass of the ore
- A copper wire in a silver nitrate solution will take on fuzzy or furry appearance and the solution colored blue

$Cu\ (s) + 2\ AgNO_3\ (aq) \rightarrow Ag\ (s) + Cu(NO_3)_2\ (aq)$

$Cu\ (s) + 2\ Ag^+\ (aq) + 2\ NO_3^-\ (aq) \rightarrow$
$2\ Ag\ (s) + Cu^{2+}\ (aq) + 2\ NO_3^-\ (aq)$

$Cu\ (s) + 2\ Ag^+\ (aq) \rightarrow 2\ Ag\ (s) + Cu^{2+}\ (aq)$

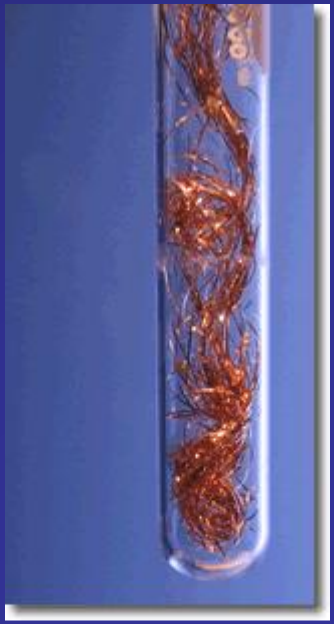


Figure 10.7.1: *Reaction of copper wire in a silver nitrate solution.*

LEO / GER

Need a way to remember that elements losing electrons is oxidation and elements (usually **OXYGEN**) gaining the electrons is reduction?

"LEO goes GER"

Loss of **E**lectrons is **O**xidation
Gain of **E**lectrons is **R**eduction

"OIL RIG"

Oxidation **I**s **L**oss, **R**eduction **I**s **G**ain

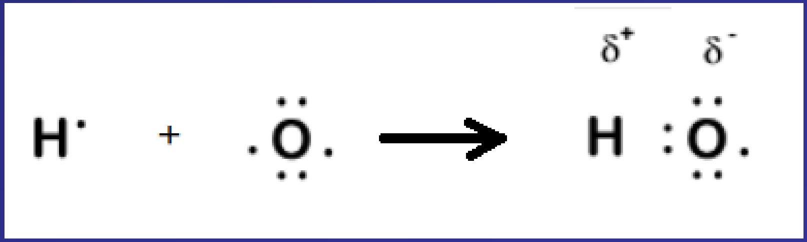
$Zn\ (s) + Fe^{2+}\ (aq) \rightarrow Zn^{2+}\ (aq) + Fe\ (s)$

Zinc gets oxidized and iron(II) gets reduced in the reaction.
Zinc is a reducing agent and iron(II) is an oxidizing agent

Electronegativity & Oxidation

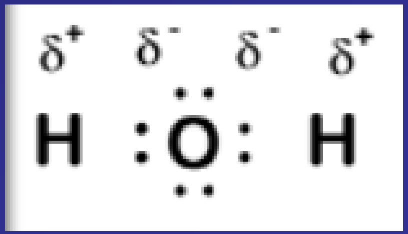
- Electronegativity is both a qualitative and quantitative indicator of how two elements with a covalent (shared electrons) bond will have one element pulling to itself more strongly those bonding electrons compared to the other element
- While fluorine (F) is the most electronegative of all elements, oxygen (O) is 2nd and its chemistry is vital for life
- Oxygen's ability to pull electrons to itself from other elements like hydrogen (H) gives it a special place in chemistry to refer to its properties as oxidation

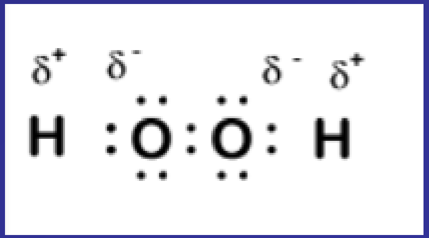
Element	Electronegativity
Fluorine (F)	3.98
Oxygen (O)	3.44
Chlorine (Cl)	3.16
Nitrogen (N)	3.04
Bromine (Br)	2.96
Hydrogen (H)	2.20



Oxidation Number

- Ionic compounds are truly formed by ions in which one, typically metal element loses one or more electrons to become positively ionized (a cation) while another, typically nonmetal element gains one or more electrons to become negatively ionized (an anion)
- The transfer of electrons forms the ionic compounds which differs from two nonmetal elements that form covalent bonds that range from have no bond dipole (no polarity) to having a significant bond dipole (noticeable polarity)
- In water molecule, the hydrogen (**H**) atom "loses" its one electron and oxygen (**O**) atom "gains" that electron: **H** is oxidized to a **+1 oxidation number.** **O** has an oxidation number of **–2**: this is because it takes one electron each from the two H atoms it bonds to





- In hydrogen peroxide (**H_2O_2**, **H–O–O–H**), both **O** atoms have oxidation number of **–1** because they can take an electron from one H atom each

Oxidation Number Rules

- There are a set of rules in determining oxidation numbers for elements in compounds

1. The oxidation number of a pure element, whether monatomic, diatomic or polyatomic, is **0** (zero)
2. Oxidation number of a monatomic ion is equal to the charge of the ion. Na^+ is **+1**, Cl^- is **−1**, Mg^{2+} is **+2**, O^{2-} is **−2**

Element	Oxidation Number
Na	0
H_2	0
O_2	0
P_4	0

Ionic Compound	Ions	Charge	Oxidation Number
NaCl	Na^+	+1	+1
	Cl^-	-1	-1
Mg_3N_2	Mg^{+2}	+2	+2
	N^{-3}	-3	-3

Oxidation Number Rules

3. The oxidation number of hydrogen (H) is almost always **+1** when it is a compound with a nonmetal element. It is **–1** when a compound of a metal cation as a metal hydride.

Compound	Element	Oxidation Number
HCl	H	+1
	Cl	-1
H_2S	H	+1
	S	-2

4. Oxidation number of oxygen (O) is usually **–2** with metals and most nonmetals. Exceptions are in peroxide (–O–O–) compounds: **–1**, and if bonded to fluorine in very highly reactive oxygen difluoride (OF_2): **+2**

Compound	Element	Oxidation Number
MgO magnesium oxide	Mg	+2
	O	-2
Na_2O sodium oxide	Na	+1
	O	-2
Na_2O_2 sodium peroxide	Na	+1
	O	-1

Oxidation Number Rules

5. Sum of all oxidation numbers in any compounds is zero

This rule and prior rules allows determination of oxidation number of other elements

Mn_2O_7 is the example:
From the **O**=**−2** rule, with **7 O** atoms for **−14** total electrons, and with **2 Mn** atoms, the **Mn oxidation number** = **+14**/**2** = **+7**

Compound	Element	Oxidation Number	Number of Atoms	Total
Mg_3N_2	Mg	+2	3	+6
	N	-3	2	-6
			SUM	0
Mn_2O_7	Mn	+7	2	+14
	O	-2	7	-14
			SUM	**0**
Cl_2O_3	Cl	+3	2	+6
	O	-2	3	-6
			SUM	**0**

Oxidation Number Rules

6. Sum of all oxidation numbers in a polyatomic ion is equal to charge of the ion

$$\sum \text{element oxidation number } \times \text{number of element atoms} = \text{charge on compound}$$

$Cr_2O_7^{2-}$ is the example:

Cr *ON* × Cr atoms +
O *ON* x O atoms = **−2**

Apply the **O**=**−2** rule, and substitute the # of Cr and O atoms:

Cr *ON* × 2 + −2 x 7 = **−2**

Now get the Cr *ON* with algebra:

Cr *ON* = **+6** and O *ON* = **−2**

Compound	Element	Oxidation Number	Number of Atoms	Total
NO_3^-	N	+5	1	+5
	O	-2	3	-6
			SUM	-1
$Cr_2O_7^{2-}$	Cr	+6	2	+12
	O	-2	7	-14
			SUM	**-2**
SO_4^{2-}	S	+6	1	+6
	O	-2	4	-8
			SUM	**-2**

The Copper Wire Sulfur Powder Experiment

- Remember Lab Expt 4a "Law of Constant Composition"
- Copper wire reacting with sulfur was a redox reaction
- Copper atoms lose one electron each (oxidation)
- Sulfur atoms gain two electron each (reduction)

Half-reactions

:: $2\ Cu \rightarrow 2\ Cu^{+} + 2\ e^{-}$

:: $S + 2\ e^{-} \rightarrow S^{2-}$

$2\ Cu\ (s) + S\ (s) \rightarrow Cu_2S\ (s)$

Practice: Oxidation Numbers

Assign oxidation numbers (ON) to EACH ATOM according to rules

a. Cl_2

i. Pure elements cannot take electrons from each other:
Cl *ON* = 0

b. GeO_2

i. When oxygen is a/the anion in the compound especially against a metal or metalloid, it is almost always **−2**
O *ON* = −2

ii. Now apply
$\sum \text{element oxidation number} \times \text{number of element atoms} = \text{charge on compound}$

Ge ON x Ge atoms + O ON x O atoms = charge
substitute:
Ge ON x 1 + −2 x 2 = 0

Ge *ON* = +4

Practice: Oxidation Numbers

Assign oxidation numbers (ON) to EACH ATOM according to rules

a. $Ca(NO_3)_2$

i. The *ON* of a metal ion particularly in Group 1 and 2 elements is the charge of its ion. Ca^{2+} → Ca *ON* = +2

ii. NO_3 is a polyatomic ion with total –1 charge (thus NO_3^- is how it is represented), so the sum of the oxidation numbers on three O atoms and one N atom must be –1.

iii. Oxygen has an *ON* = –2 according to rules, and there are three of them (O *ON* = –2). This totals a –6 charge.

iv. To determine the nitrogen *ON*, it must be part of the polyatomic ion NO_3^-, so N *ON* + (–6) = –1, or N *ON* = +5

Solving for **N *ON*** another way—the better way:

Ca *ON* x Ca atoms + **N *ON*** x N atoms +
O *ON* x O atoms = compound charge

1 x (+2) + **N *ON*** x 2 + (-2) x 6 = 0

N *ON* = —[1 x (+2) + (-2) x 6] / 2 = –[2 + (–12)] /2 = +5

More Practice: Oxidation Numbers

Assign oxidation numbers (ON) to EACH ATOM according to rules

d. H_3PO_4

i. *ON* of hydrogen in a compound with nonmetal monatomic or polyatomic group is +1 . H *ON* = +1

ii. Oxygen has an *ON* = –2 according to rules: O *ON* = –2

iii. Now solve for the P *ON*

H *ON* x H atoms + **P *ON*** x P atoms + O *ON* x O atoms = compound charge

(+1) x 3 + **P *ON*** x 1 + (-2) x 4 = 0

P *ON* = —[(+1) x 3 + (-2) x 4] = –[3 + –8] = **+5**

e. MgO

i. One rule states that metal cations of Groups 1 and 2 have ON which is the charge of the ion. Mg2+ → Mg *ON* = +2

ii. Oxygen has an *ON* = –2 according to rules: O *ON* = –2

If There Is Oxidation, There is Reduction

- Just as there is conservation of mass/matter and of energy, there is a conservation of electric charge
- Electric charge and oxidation numbers/states must balance on both sides of the chemical reaction equation
- Where there is an oxidation, there is a reduction
- Where there is element/atom oxidized, another element/atom is reduced

$$2\,\mathrm{Na} + \mathrm{Br_2} \longrightarrow 2\,\mathrm{NaBr}$$

$$2\underset{0}{Na} + \underset{0}{Br_2} \rightarrow 2\underset{+1-1}{NaBr}$$

What’s Being Oxidized? What’s Reduced?

$$\begin{array}{lccccc} & C & + & O_2 & \longrightarrow & CO_2 \\ \text{ON} & 0 & & 0 & & +4\ \ -2 \end{array}$$

$$6H^{+}(aq) + 2\underset{+7}{Mn}O_4^{-}(aq) + 5H_2\underset{-1}{O_2}(l) \rightarrow 2\underset{+2}{Mn}^{2+}(aq) + 5\underset{0}{O_2}(g) + 8H_2O(l)$$

Combustion Reactions

Make it balanced

Strategy

1. If combustion, then CO_2 (if carbon present) and H_2O are products!
2. Balance the oxidized main atom (C, H, etc) on products' side
3. Balance additional non-O atom (like H) on products' side
4. Count O atoms on products' side
   a. If is an ODD number of O atoms, multiply everything by 2
   b. If EVEN number (before or after doing [a]), put coefficient in front of O_2 on reactants

- $H_2\ (g) + O_2\ (g) \rightarrow\ ?$

$$2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$$

- $C_3H_8\ (g) + O_2\ (g) \rightarrow\ ?$

$$C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$$

- $C_6H_{14}\ (l) + O_2\ (g) \rightarrow\ ?$

$$2\ C_6H_{14}\ (l) + 19\ O_2\ (g) \rightarrow 12\ CO_2\ (g) + 14\ H_2O\ (g)$$

Redox Reactions as Two Half-Reactions

With oxidation-reduction (redox) reactions, they are special in the way that they are expressed as two half-reactions

- Because there is not only mass but charge (shown as "e^-" or electrons), both mass (atom kind & number) and charge must balance (be conserved) on both sides of the arrow
- You normally write the oxidation (loss of electrons, the electron-producing) half-reaction first
- Follow with the reduction (gain of electrons, electron-consuming) half-reaction after
- $Mg\ (s) + Cl_2\ (g) \rightarrow MgCl_2\ (s)$
  - Write oxidation reaction: $Mg \rightarrow Mg^{2+}$
  - Balance it with charge (electrons): $Mg \rightarrow Mg^{2+} + 2\ e^-$
  - Write reduction reaction: $Cl_2 \rightarrow 2\ Cl^-$
  - Balance it with charge (electrons): $Cl_2 + 2\ e^- \rightarrow 2\ Cl^-$
  - Check if the electrons cancel in the half-reactions: yes
  - Write the full reaction: $Mg + Cl_2 \rightarrow Mg^{2+} + 2\ Cl^-$
  - Associate with original reaction: $Mg + Cl_2 \rightarrow MgCl_2$

More Examples of Two Half-Reactions

- $Cu\ (s) + AgNO_3\ (aq) \rightarrow Cu(NO_3)_2\ (aq) + Ag\ (s)$
  - Write oxidation reaction: $Cu \rightarrow Cu^{2+}$
  - Balance it with charge (electrons): $Cu \rightarrow Cu^{2+} + 2\ e^-$
  - Write reduction reaction: $Ag^+ \rightarrow Ag$
  - Balance it with charge (electrons): $Ag^+ + 1\ e^- \rightarrow Ag$
  - Check if the electrons cancel in the half-reactions: no, need to multiply silver reduction by 2: $2\ Ag^+ + 2\ e^- \rightarrow 2\ Ag$
  - Write the full reaction: $Cu + 2\ Ag^+ \rightarrow Cu^{2+} + 2\ Ag$
  - Associate with original reaction: $Cu + 2\ AgNO_3 \rightarrow Cu(NO_3)_2 + 2\ Ag$

More Examples of Two Half-Reactions

- This redox is more complex, and it will explain why it requires an acid solution to make it happen
- MnO_4^- (*aq*) + Fe^{2+} (*aq*) + H^+ (*aq*)→ Mn^{2+} (*aq*) + Fe^{3+} (*aq*) + H_2O (l)
  - Write oxidation reaction: $Fe^{2+} \rightarrow Fe^{3+}$
  - Balance it with charge (electrons): $Fe^{2+} \rightarrow Fe^{3+} + 1\ e^-$

Reduction is different: Mn ON goes from +7 to +2

  - Write reduction reaction: $MnO_4^- \rightarrow Mn^{2+}$

Why don't we write it as $Mn^{7+} \rightarrow Mn^{2+}$?

  - Balance ON with charge (electrons): $MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$

Have to write H^+ and H_2O in reaction. O atoms required because of MnO_4^-

  - Check if the electrons cancel in the half-reactions: no, need to multiply Fe oxidation by 5: $5\ Fe^{2+} \rightarrow 5\ Fe^{3+} + 5\ e^-$
  - Write the full reaction: $5\ Fe^{2+} + MnO_4^- + 8\ H^+ \rightarrow 5\ Fe^{3+} + Mn^{2+} + H_2O$

More Examples of Two Half-Reactions

- This redox also happens in acid solution necessarily
- HNO_3 (*aq*) + Cu (*s*) + H^+ (*aq*)→ NO_2 (*g*) + Cu^{2+} (*aq*) + H_2O (*l*)
  - Write oxidation reaction: Cu → Cu^{2+}
  - Balance it with charge (electrons): Cu → Cu^{2+} + 2 e^-

Reduction is different: N *ON* goes from +5 to +4

- Write reduction reaction: HNO_3 → NO_2
- Balance ON with charge (electrons): HNO_3 + H^+ + e^- → NO_2 + H_2O

Have to write H^+ and H_2O in reaction. O atoms required because of HNO_3

- Check if the electrons cancel in the half-reactions: no, have to multiply the reduction half-reaction by two: 2 HNO_3 + 2 H^+ + 2e^- → 2 NO_2 + 2 H_2O

- Write the full reaction: Cu (s) + 2 HNO_3 + 2 H^+ → Cu^{2+} + 2 NO_2 + 2 H_2O

Activity Series

- Both silver (Ag) and sodium (Na) can be oxidized ($Ag \rightarrow Ag^+ + e^-$, $Na \rightarrow Na^+ + e^-$)
- But Na metal will react with cold water and Ag metal is completely unreactive
- The activity series for elements that show this redox reactivity has been drawn to the Periodic Table on the next slide.
- There are explanations/reasons for the patterns of the reactivity but they are beyond scope of this course

Figure 10.9.1.1: On the left, sodium reacts with water. On the right, silver in the form of cups does not react with water.

“Yellow Sheet”

Reacts with cold H_2O, replacing hydrogen

Reacts with steam (not cold H_2O), replacing hydrogen

Does not react with H_2O, but with acid (H^+)

Unreactive with water or all but strongest acids (H^+)

Reactive nonmetals: $F_2 > Cl_2 > Br_2 > I_2$

within colored blocks, numbers 1, 2, 3, ... indicate order of reactivity from most to least

1	2											13	14	15	16	17	18
Hydrogen 1 H 1.008 2.1																	Helium 2 He 4.003 ---
Lithium 3 Li 6.94 1.0 1	Beryllium 4 Be 9.012 1.5											Boron 5 B 10.81 2.0	Carbon 6 C 12.011 2.5	Nitrogen 7 N 14.007 3.0	Oxygen 8 O 15.999 3.5	Fluorine 9 F 18.998 4.0 1	Neon 10 Ne 20.180 ---
Sodium 11 Na 22.99 0.9 6	Magnesium 12 Mg 24.31 1											Aluminum 13 Al 26.98 1.5 2	Silicon 14 Si 28.09 1.8	Phosphorus 15 P 30.97 2.1	Sulfur 16 S 32.06 2.5	Chlorine 17 Cl 35.45 3.0 2	Argon 18 Ar 39.95 ---
Potassium 19 K 39.10 0.8 2	Calcium 20 Ca 40.08 1.0 5	Scandium 21 Sc 44.96 1.3	Titanium 22 Ti 47.88 1.5	Vanadium 23 V 50.94 1.6	Chromium 24 Cr 52.00 1.6 4	Manganese 25 Mn 54.94 1.5	Iron 26 Fe 55.85 1.8 5	Cobalt 27 Co 58.93 1.8 1	Nickel 28 Ni 58.69 1.8 2	Copper 29 Cu 63.55 1.9 1	Zinc 30 Zn 65.39 1.6 3	Gallium 31 Ga 69.72 1.6	Germanium 32 Ge 72.61 1.8	Arsenic 33 As 74.92 2.0	Selenium 34 Se 78.97 2.4	Bromine 35 Br 79.90 2.8 3	Krypton 36 Kr 83.80 3.0
Rubidium 37 Rb 85.47 0.8	Strontium 38 Sr 87.62 1.0 4	Yttrium 39 Y 88.91 1.2	Zirconium 40 Zr 91.22 1.4	Niobium 41 Nb 92.91 1.6	Molybdenum 42 Mo 95.94 1.8	Technetium 43 Tc (98) 1.9	Ruthenium 44 Ru 101.07 2.2	Rhodium 45 Rh 102.91 2.2	Palladium 46 Pd 106.42 2.2	Silver 47 Ag 107.87 1.9 3	Cadmium 48 Cd 112.41 1.7 6	Indium 49 In 114.82 1.7	Tin 50 Sn 118.71 1.8 3	Antimony 51 Sb 121.76 1.9	Tellurium 52 Te 127.60 2.1	Iodine 53 I 126.90 2.5 4	Xenon 54 Xe 131.29 2.6
Cesium 55 Cs 132.91 0.7	Barium 56 Ba 137.33 0.9 3	Lanthanum 57 La 138.91 1.1 *	Hafnium 72 Hf 178.49 1.3	Tantalum 73 Ta 180.95 1.5	Tungsten 74 W 183.84 1.7	Rhenium 75 Re 186.21 1.9	Osmium 76 Os 190.23 2.2	Iridium 77 Ir 192.22 2.2	Platinum 78 Pt 195.08 2.2 4	Gold 79 Au 196.97 2.4 5	Mercury 80 Hg 200.59 1.9 2	Thallium 81 Tl 204.38 1.8	Lead 82 Pb 207.20 1.8 4	Bismuth 83 Bi 208.98 1.9	Polonium 84 Po (209) 2.0	Astatine 85 At (210) 2.2	Radon 86 Rn (222) 2.4
Francium 87 Fr (223) 0.7	Radium 88 Ra (226) 0.9	Actinium 89 Ac (227) 1.1	Rutherfordium 104 Rf (267) ---	Dubnium 105 Db (268) ---	Seaborgium 106 Sg (271) ---	Bohrium 107 Bh (272) ---	Hassium 108 Hs (270) ---	Meitnerium 109 Mt (276) ---	Darmstadtium 110 Ds (281) ---	Roentgenium 111 Rg (280) ---	Copernicium 112 Cn (285) ---	Nihonium 113 Nh (284) ---	Flerovium 114 Fl (289) ---	Moscovium 115 Mc (288) ---	Livermorium 116 Lv (293) ---	Tennessine 117 Ts (294) ---	Oganesson 118 Og (294) ---

*lanthanide	Cerium 58 Ce 140.12 1.1	Praseodymium 59 Pr 140.91 1.1	Neodymium 60 Nd 144.24 1.1	Promethium 61 Pm (145) 1.1	Samarium 62 Sm 150.36 1.2	Europium 63 Eu 151.97 1.1	Gadolinium 64 Gd 157.25 1.2	Terbium 65 Tb 158.93 1.1	Dysprosium 66 Dy 162.50 1.2	Holmium 67 Ho 164.93 1.2	Erbium 68 Er 167.26 1.2	Thulium 69 Tm 168.93 1.3	Ytterbium 70 Yb 173.04 1.1	Lutetium 71 Lu 174.97 1.1
**actinide	Thorium 90 Th 232.04 1.3	Protactinium 91 Pa 231.04 1.5	Uranium 92 U 238.03 1.4	Neptunium 93 Np (237) 1.4	Plutonium 94 Pu (244) 1.3	Americium 95 Am (243) 1.3	Curium 96 Cm (247) 1.3	Berkelium 97 Bk (247) 1.3	Californium 98 Cf (251) 1.3	Einsteinium 99 Es (252) 1.3	Fermium 100 Fm (257) 1.3	Mendelevium 101 Md (259) 1.3	Nobelium 102 No (258) 1.3	Lawrencium 103 Lr (262) ---

Practice: Activity Series

- Use the information from the Activity series shown in the drawn Periodic Table to make predictions about reaction
- Al (s) + $Zn(NO_3)_2$ (aq) → ?

Yes: 2 Al (s) + 3 $Zn(NO_3)_2$ (aq) → 2 $Al(NO_3)_3$ + 3 Zn (s)

- Ag (s) + HCl (aq) → ?

No: Ag will not react with dilute acid

- $FeCl_2$ (aq) + Zn (s) → ?

Yes: Zn is more reactive than Fe and can be oxidized

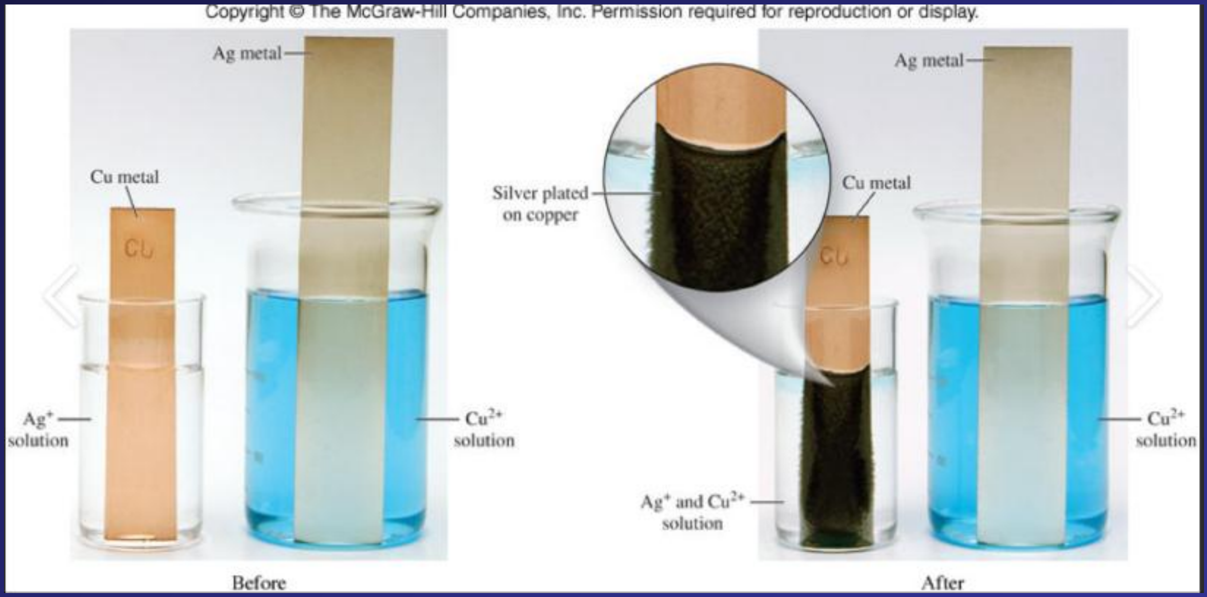
$FeCl_2$ + Zn → Fe + $ZnCl_2$

- HNO_3 (aq) + Au (s) → ?
- No: gold unreactive with even acids of moderate strength

Chemistry Related to Batteries

- Batteries used in electronic devices are basically containers of chemicals that react to produce electrons that enable electric current. These reactions are redox reactions
- Reactions like the oxidation of copper in a silver solution are the basis for the chemistry:

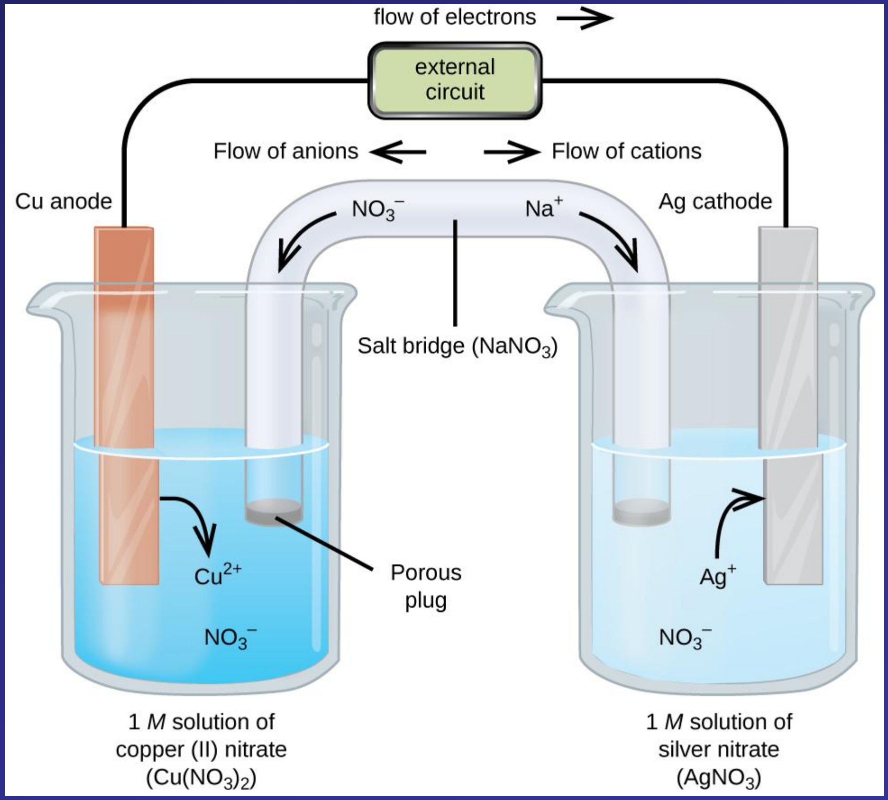
$$2\ Ag^{+}\ (aq) + Cu\ (s) \rightarrow 2\ Ag\ (s) + Cu^{2+}\ (aq)$$



Silver-Copper Galvanic Cell

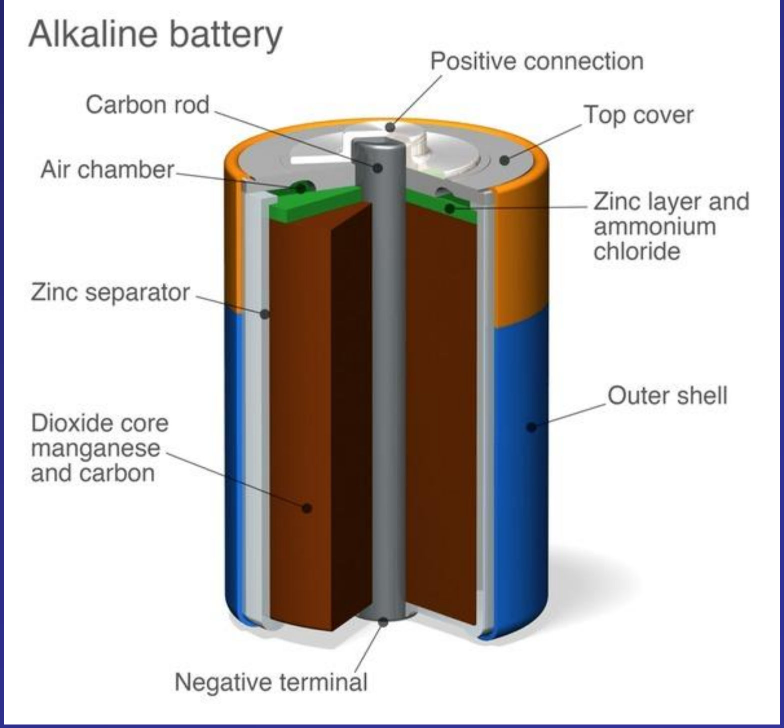
- Instead of wasting the electron flow from copper metal to silver solution, set up a compartmentalized solid copper/copper solution and solid silver/silver solution and connect these as a circuit with a wire to a device to be powered, and "salt bridge" of sodium nitrate

- The electrode where reduction occurs is the **cathode** (electrons gained)
- The electrode where oxidation occurs is the **anode** (electrons lost)



Dry Cell Battery

- Your nonrechargeable A, AA, AAA batteries are dry cell type
- Zn case is anode (neg pole), graphite rod is cathode (pos pole)
- A paste of either MnO_2, NH_4Cl, $ZnCl_2$ or KOH electrolyte (alkaline type)



- Cathode:

$2\ MnO_2\ (s) + 2\ NH_4^+\ (aq) + 2\ e^- \rightarrow Mn_2O_3\ (s) + H_2O\ (l) + 2\ NH_3\ (aq)$

- Anode:

$Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2\ e^-$

Storage Batteries

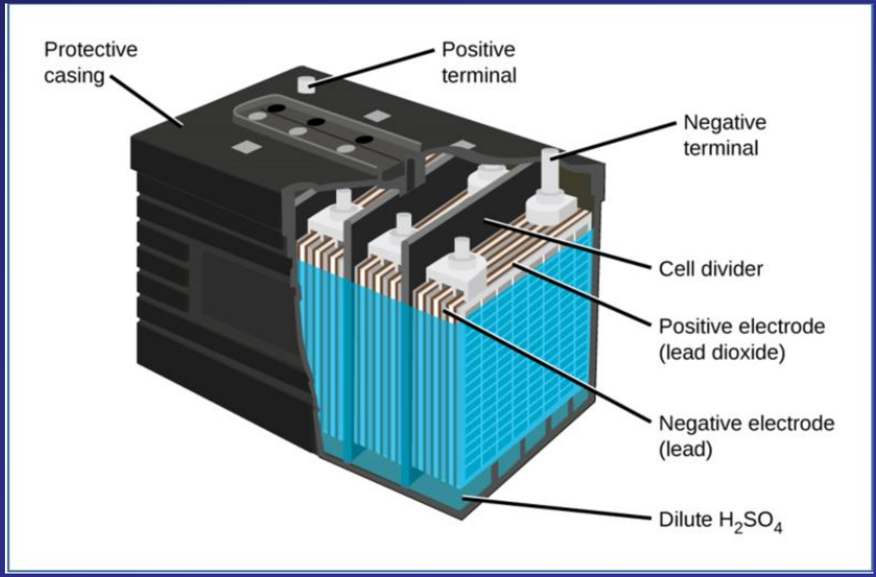
- The classic type is the rechargeable lead-acid automobile battery. Anode is grid of Pb-Sb or Pb-Ca alloy packed with spongy lead; cathode in lead(IV) oxide, in sulfuric acid electrolyte

- Cathode:

PbO_2 (*s*) + 4 H^+ (*aq*) + 2 e^- → $PbSO_4$ (*s*) + 2 H_2O (l) + 2 NH_3 (*aq*)

- Anode:

Pb (s) + SO_4^{2-} (*aq*) → $PbSO_4$ (*s*) + 2 e^-



Fuel Cells

Works by continuously streaming H_2 and O_2 gas in the presence of a catalyst in a KOH electrolyte

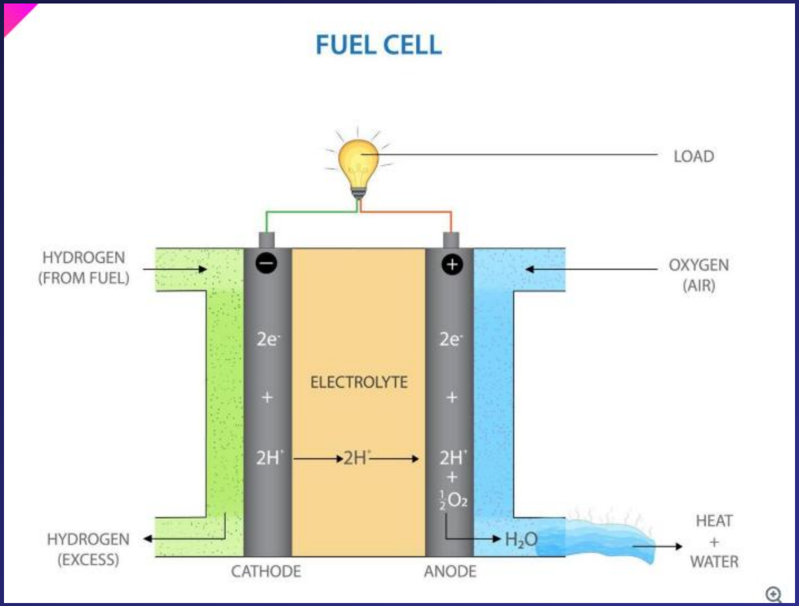
- Cathode:

$O_2\ (g) + 2\ H_2O\ (l) + 4\ e^- \rightarrow 4\ OH^-\ (aq)$

- Anode:

$2\ H_2\ (g) + 4\ OH^-\ (aq) \rightarrow 4\ H_2O + 4\ e^-$

- Net: $2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (l)$



Corrosion

- Falls in category of "undesirable" redox reactions
- The rusting of iron (Fe) is one such reaction

$$Fe\ (s) \rightarrow Fe^{2+}\ (aq) + 2\ e^{-}$$

$$H_2O\ (l) + \tfrac{1}{2}\ O_2\ (g) + 2\ e^{-} \rightarrow 2\ OH^{-}\ (aq)$$

- Further oxidation to Fe^{3+} occurs, and this reacts with OH– ions to form iron(III) oxide [Fe_2O_3] and also iron(III) hydroxide [$Fe(OH)_3$]

Prevention?

- Put paint, grease, plastic to cover the metal and prevent O_2 from getting access
- Put Zn or Mg on iron surface: they are more readily oxidized and Fe keeps its integrity. **Sacrificial anodes** in water heaters work this way