

# Week 3 • Predicting Reactions

## TEACHING IDEAS AND RESOURCES

### Overview:

I split up teaching this skill into the various reaction types. Each type requires information to be memorized, mini-skills to be mastered, and examples to be experienced. Here is how to approach **Week 3**.

- Read through these **Teaching Ideas** about this section of the AP Free Response exam and how they are graded.
- Review the materials that I use to teach **Double Replacement** reactions. I spend the most time on this first set of skills because some of them impact the writing of reaction equations in general.
- Review the materials that I use to teach **Oxidation-Reduction** reactions. I spend the second largest amount of time on these reactions. This topic appears in varied places in the AP chemistry course.
- Review the materials that I use to teach **Lewis Acid-Base** reactions. There are several different applications of Lewis Acids and Bases in predicting reactions.
- Review the materials that I use to teach **Complex Ion** formation and reactions involving complex ions. These are probably the strangest topics for students and appear confusing but they actually make a lot of sense and are easy to grasp.
- Practice doing these reactions. Ask questions and make comments in the **Discussion Board** about the Week 3 Resources and Week 3 AP Questions.
- Read the **Week 3 Reaction Types handout** I give to students to get some details about each type of reaction.
- Read the **Week 3 Labs and Demonstrations** handout that discusses ways to illustrate the four types of reactions.
- These labs might be good for your **Laboratory Write-Up Assignment** due Week 5. Maybe some of them are new to you or you used to do one and want to resurrect it or maybe fine tune a lab you already do. I am also very interested in improvements you can suggest for the way I do the labs.
- Review the **various handouts** and **practice worksheets** that I use to teach this subject.

### Example Questions:

Here are examples I used in Week 1. More appear in the AP questions for this week.

**Example 1:** *"Solutions of silver nitrate and potassium chromate are mixed."*

Ex 1	
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**Possible Question:** How might a student visually tell the two reactants apart?

**Possible Question:** What is the oxidation number of Cr in the chromate ion?

**Example 2:** *"Solid sodium carbonate is added to a solution of acetic acid."*

Ex 2	
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**Possible Question:** How would a student observe that this reaction is complete?

**Possible Question:** If equimolar amounts of the reactants are used, will the pH of the product solution be less than 7, 7, or greater than 7?

## Grading:

These questions are scored on a 5 point scale. The three questions count as 15 points. This score will count for 10% of the free response portion of the exam.

1 point	Writing the <b>reactants</b> in chemical terms.
2 points	Predicting the <b>products</b> of the chemical reaction.
1 point	<b>Balancing</b> the equation.
1 point	Answering a <b>question</b> about the reaction.
<b>5 points</b>	<b>Total</b>

## Philosophy of Teaching Predicting Reactions:

### Organize into Four Types of Reactions:

The number of different kinds of reactions that might be presented is overwhelming, especially to students. I treat the reactions as falling into four categories.

Double Replacement	Two ionic compounds change partners. One or both of the new combinations drives the reaction forward.
Redox	One chemical gains electrons while another loses electrons.
Lewis Acids	One chemical accepts a pair of electrons to make a coordinate covalent bond. The other chemical “donates” the pair of electrons used for the bond.
Complex Ions	A central ion is surrounded by ligands.

Other people treat learning reactions differently. I like fewer categories rather than more categories, but I strongly believe in the idea of thinking about these reactions based on the *chemistry* involved rather than just *patterns*. This focus will help students learn the chemistry they need to answer the question that accompanies each equation. I believe that the way some students have been taught to answer these problems focuses more on patterns and types rather than the underlying chemistry and this led to the new format for the question.

These categories are the focus of several laboratory activities that I outline in the **lab handout**. Some of these are done in first year chemistry and may need to be revisited. Chemical reactions encountered throughout the year provide possible predicting reaction practice. I frequently stop and say, “How would this appear as a predicting reactions question?”

**Basic Skills:** (*for most of you, this is a 1<sup>st</sup> year task... however, students may need to re-learn these skills*)

- **Memorize Ion Formulas and Charges:**

I print out flash cards for students. I have them begin by memorizing 40 common ions. I then add another 20 whose names are based on the first 40 (e.g., sulfite and bisulfate are easier after they memorize sulfate, etc.). We practice these ions using an **Ion Bee** (where all students stand in a line and get to sit down when they get an ion correct) and **Ion Bingo**. We quiz on these ions 10 at a time and also with the Mega-Ion Quiz in which all 60 ions must be correct (actually, they can get one wrong because 29.5 points rounds up to 30).

- **Write Formulas for Ionic Compounds:**
- **Name Ionic Compounds:**
- **Name Acids:**
- **Name Molecular (nonmetal) Compounds:**
- **Name Organic Compounds:**

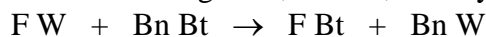
### • Translating words to chemical symbols:

The first task is to recognize the reactant chemicals involved and write them down in chemical terms. This involves knowing how to recognize the names of ionic compounds, nonmetal compounds, acids, and organic chemicals. It also requires students to recognize the phases of chemicals. This is best achieved through practice and experience. I have created a worksheet with words copied from years and years of past exams.

### Double Replacement Reactions:

#### The Pattern:

I teach this pattern of reaction using Fred, Wilma, Barney and Betty.



Fred replaces Barney and Barney replaces Fred. Treating double replacement as two couples is a common idea, but I find that using the Flintstones is a little less “sexually charged” than couples in general. I actually have Flintstones action figures that I use for this.

Note: I also teach single replacement, synthesis and decomposition as patterns of reactions, but not as organizational tools for predicting reactions.



#### Driving Forces:

The big idea with double replacement is that one of the **new combinations** of ions **prevents** the ions from **reforming** the original compounds. This may be because the new combination is so stable it drops out of solution as a **precipitate**, forms a **weak electrolyte** that remains dissolved but doesn't dissociate into ions, or forms a **gas** that bubbles out of the solution.

Driving Force	How Do You Recognize It?
precipitate	You must <b>memorize</b> the <b>solubility rules</b> . Any compound formed from two ions can be recognized as soluble (written as separate ions) or as a precipitate (written as a molecule).
gas formed	You must <b>memorize</b> the <b>combinations</b> that decompose into gases. You must also <b>memorize</b> the <b>gases</b> that form. For example, when you see $H_2SO_3$ as a product, you must know it decomposes into $H_2O$ and $SO_2$ gas.
weak electrolyte	You must <b>memorize</b> the short list of <b>strong acids</b> so you will recognize all the weak acids that dissolve, but do not dissociate into ions. The weak base, ammonia, $NH_3$ , is also in this category. It exists in water primarily as $NH_3(aq)$ and only slightly forms the ions, $NH_4^+ + OH^-$ .

Note: Students sometimes confuse the idea of solubility with reactivity. They think that if a molecule does not dissolve, for example,  $Mg(OH)_2(s)$ , it cannot react. These substances exist primarily in the undissociated state, but do break up into ions a **little**. It will probably react **slowly**, but it **will** react.

#### Molecular, Ionic, and Net Ionic Equations:

Because the equation must be balanced, I work with my students to write the **molecular equation** first. We balance the molecular equation. We annotate each chemical with the phase [(s), (l), (g), or (aq)]. Although this is not necessary for the question, it will **greatly** help with writing the ionic equation and is important training for the students.

The **ionic equation** uses the student's knowledge of solubility as well as of weak and strong electrolytes. Students need to understand the difference between a substance that dissolves and dissociates into ions and a substance that dissolves, but does not dissociate into ions to any great extent. These skills are important for

predicting the reactions in double replacement reactions, but will be needed for all equations since any **aqueous** species must be written as separate ions if they exist primarily in that form.

As an example, the **strong** acid, HCl(aq) should be written  $\text{H}^+ + \text{Cl}^-$ . The **weak** acid, acetic acid,  $\text{HC}_2\text{H}_3\text{O}_2(\text{aq})$  will be written in the undissociated form,  $\text{HC}_2\text{H}_3\text{O}_2$ . Acetic acid dissolves, but does not dissociate. Species labeled (s), (l), or (g) are **never** written as separate ions... only (aq), for example, “solid potassium chloride” is  $\text{KCl}(\text{s})$  and **can** dissolve in water to form ions, but does not exist as separate ions as a solid.

Finally, students eliminate the **spectator ions**. The **net ionic equation** is what is required for these questions. The students may rewrite the equation, but simply striking out the spectator ions is sufficient.

## Oxidation-Reduction Reactions:

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The second category of reactions I teach is oxidation-reduction. I begin with simple reactions in which a neutral element becomes an ion and/or an ion becomes a neutral element.  $\text{CuCl}_2 + \text{Al}^\circ \rightarrow \text{AlCl}_3 + \text{Cu}^\circ$  is a good example. We begin with the blue crystals that contain  $\text{Cu}^{2+}$  ions. We make a solution (noting color changes because of complex ions). We add a little piece of aluminum foil and observe. Brown  $\text{Cu}^\circ$  forms on the surface of the aluminum metal and the aluminum metal disappears as it goes into solution.



We look at this reaction and talk about the difference between the  $\text{Cu}^{2+}$  ion and the  $\text{Cu}^\circ$  metal it turns into. The ion must have gained electrons. In the same way the  $\text{Al}^\circ$  loses electrons to become  $\text{Al}^{3+}$  ions in the solution. I tell the students that this is such an important type of reaction that we give it its own name. **Oxidation** means losing electrons and **Reduction** means gaining electrons (**LeO** the lion says **GeR**). **Oxidation** obviously gets its name from combining with oxygen even though an element does not need to combine with oxygen to be oxidized. I have been told that **reduction** gets its name from the idea that metal **ore** (iron compounds, for example) get reduced to a smaller volume of the **metal**. The size is reduced. Students realize quickly that the **charge** is reduced, too.

Nonmetals can undergo oxidation and reduction, also.  $\text{Fe}^\circ + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$  (rust). Fe loses electrons (gets oxidized) to become  $\text{Fe}^{3+}$  while O gains electrons (gets reduced) to become  $\text{O}^{2-}$ . We note that if a **neutral element** shows up as a reactant, the reaction **MUST** be redox. We add that idea to our “bag of tricks”.

Later, we generalize our ideas to include **apparent charges** (oxidation numbers) instead of just ion charges. We also look at chemicals that have a great tendency to change (oxidizing agents and reducing agents). Noting that different metals have different tendencies to gain electrons (get reduced) leads to concept of reduction potentials and electrochemistry. We look at all of this more closely in Week 5. I list some common oxidizing agents and reducing agents (and what they turn into) on the “Stuff I Am Supposed To Know for the AP Test... But Don’t Know Yet” sheet.

## Lewis Acid-Base Reactions:

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Every reaction can be classified as oxidation-reduction (oxidation numbers change) acid-base (oxidation numbers do not change). The Arrhenius and Brønsted-Lowry definitions, however do not cover all situations. When a covalent bond is formed, a pair of electrons is shared between two atoms. Usually both of these come from the same atom. However, when hydrogen phosphide (phosphine) is added to boron trifluoride gas, we have an atom (P) with a lone pair of electrons and an atom (B) with an empty orbital. These two atoms can form a covalent bond  $\text{H}_3\text{P}:\text{BF}_3$  by sharing a pair of electrons, but both electrons forming the bond originally belonged to the P atom. It is still a covalent bond and is called a **coordinate covalent bond**.

Gilbert Lewis (Lewis dot structures) called the phosphine,  $\text{PH}_3$ , a **Lewis base** (an electron-pair donor).  $\text{BF}_3$  was a **Lewis acid** (an electron-pair acceptor). Mnemonic: "Have Pair Will Share – Lewis Base." This is a strange looking reaction, but does show up every other year or so on the predicting reaction question.

$\text{CO}_2$  reacts with  $\text{H}_2\text{O}$  to form  $\text{H}_2\text{CO}_3$ ; this is another example of a Lewis acid-base reaction. One of the electron pairs forming the  $\text{C}=\text{O}$  double bond moves to the O atom leaving an empty orbital on the carbon atom. (Lewis acid). Another species with a lone pair (Lewis base) forms a coordinate covalent bond. The example shown is  $\text{CO}_2 + \text{OH}^- \rightarrow \text{HCO}_3^-$  (bicarbonate).



Nonmetal oxides ( $\text{CO}_2$ ,  $\text{SO}_2$ ,  $\text{SO}_3$ ,  $\text{NO}_2$ ,  $\text{P}_4\text{O}_{10}$ ) form acids in water ( $\text{H}_2\text{CO}_3$ ,  $\text{H}_2\text{SO}_3$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{HNO}_2$ ,  $\text{H}_3\text{PO}_4$ , respectively). In the same way, metal oxides ( $\text{MgO}$ ,  $\text{CaO}$ ,  $\text{Na}_2\text{O}$ ) all form bases in water (e.g.,  $\text{Mg}(\text{OH})_2$ ,  $\text{Ca}(\text{OH})_2$ ,  $\text{NaOH}$ , respectively).

Taking this one step further, a Lewis acid and Lewis base can form a salt:  $\text{CO}_2 + \text{CaO} \rightarrow \text{CaCO}_3$ .

### Complex Ions:



When a few drops of water are added to the  $\text{CuCl}_2(\text{s})$  shown at the left, the color becomes green. When more water is added, the solution takes on the characteristic peacock blue color associated with  $\text{Cu}^{2+}(\text{aq})$ . This blue color is due to the  $\text{Cu}^{2+}$  ion surrounded by polar  $\text{H}_2\text{O}$  molecules. The green color is  $\text{Cu}^{2+}$  surrounded by chloride ions and water molecules.

blue:  $\text{Cu}(\text{H}_2\text{O})_4^{2+}$  tetraaquo copper(II) ion  
 green:  $\text{Cu}(\text{H}_2\text{O})_2\text{Cl}_2$  diaquodichloro copper(II) complex  
 yellow:  $\text{CuCl}_4^{2-}$  tetrachloro copper(II) ion



The surrounding species ( $\text{H}_2\text{O}$  or  $\text{Cl}^-$  in this case) are called "**ligands**" and can be polar molecules ( $\text{H}_2\text{O}$ ,  $\text{NH}_3$ ) or anions ( $\text{Cl}^-$ ,  $\text{OH}^-$ ,  $\text{CN}^-$ ,  $\text{S}_2\text{O}_3^{2-}$ , etc.) As a rule of thumb, you place twice the number of ligands around an ion as the charge on the central ion.  $\text{Ag}^+$  gets two ligands,  $\text{Cu}^{2+}$  gets four ligands, etc. This is not always correct, but it will do.

Notice in the Silver one-pot reaction shown at the right, the silver ion can form many precipitates and complex ions. When ammonia is added,  $\text{Ag}(\text{NH}_3)_2^+$  is formed. Since the complex is more stable than the precipitate,  $\text{AgCl}(\text{s})$ , the solid dissolves. Silver ion reacts with  $\text{S}_2\text{O}_3^{2-}$  to form the complex ion,  $\text{Ag}(\text{S}_2\text{O}_3)_2^{3-}$ .  $\text{AgCl}_2^-$  showed up on the 2006 AP exam.

This picture comes from a tutorial page by David Brooks. It is worth a visit.

[<http://dwb.unl.edu/dwb/Meetings/Oct-8-99/ComplexIonsTutor.html>]

Silver One Pot Reactions				
0.1 M $\text{AgNO}_3$	1 M $\text{NaHCO}_3$	0.2 M $\text{Na}_3\text{PO}_4$	1.0 M $\text{NaOH}$	1 M $\text{NaCl}$
$\text{Ag}^+$	$\text{Ag}_2\text{CO}_3$	$\text{Ag}_3\text{PO}_4$	$\text{Ag}_2\text{O}$	$\text{AgCl}$
$\text{Ag}(\text{NH}_3)_2^+$	$\text{AgBr}$	$\text{Ag}(\text{S}_2\text{O}_3)_2^{3-}$	$\text{AgI}$	$\text{Ag}_2\text{S}$
6.0 M $\text{NH}_3$	0.1 M $\text{KBr}$	1.0 M $\text{Na}_2\text{S}_2\text{O}_3$	0.1 M $\text{KI}$	0.1 M $\text{Na}_2\text{S}$