Slide 1

Types of Aqueous Reactions

Slide 2

Recognizing types of Reactions

It is possible to predict the possible products of a reaction based on the reactants.

Many of these predictions are easy to make if you understand a few basic things about molecules/atoms.

You, too, can learn to predict reaction products!

Slide 3

A Question

$(NH_4)_2SO_4 + FeCl_3 \rightarrow$

What happens when you mix ammonium sulfate and ferric chloride?

A better question is: what COULD happen?

Slide 4

$(NH_4)_2SO_4$ + $FeCl_3$ →

$(NH_4)_2SO_4$ $FeCl_3$

What kind of molecules are these?

Ionic!

In solution, what form do ionic compounds take?

Ions!

Slide 5

Are these the same?

$(NH_4)_2SO_4$ + $FeCl_3$ → $Fe_2(SO_4)_3$ + NH_4Cl

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)}$ → $Fe_2(SO_4)_{3\,(aq)}$ + $NH_4Cl_{(aq)}$

Kind of depends on what you mean by the same…

The “state of matter” is important. What if they were all solids?

Slide 6

Are these the same?

$(NH_4)_2SO_4$ + $FeCl_3$ → $Fe_2(SO_4)_3$ + NH_4Cl

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)}$ → $Fe_2(SO_4)_{3\,(aq)}$ + $NH_4Cl_{(aq)}$

$(NH_4)_2SO_{4\,(s)}$ + $FeCl_{3\,(s)}$ → $Fe_2(SO_4)_{3\,(s)}$ + $NH_4Cl_{(s)}$

The bottom 2 are definitely different from each other. And only one of them makes any sense!

Slide 7

$(NH_4)_2SO_4$ + $FeCl_3$ → $Fe_2(SO_4)_3$ + NH_4Cl

$(NH_4)_2SO_4$ $FeCl_3$

In solution, what form do ionic compounds take?

NH_4^+ SO_4^{2-} Fe^{3+} Cl^-

What about the products? $Fe_2(SO_4)_3$ NH_4Cl

In solution….

Fe^{3+} SO_4^{2-} NH_4^+ Cl^-

NO DIFFERENCE! NO REACTION!

Slide 8

What are the possible products?

- $(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)}$ →
- $Fe_2(SO_4)_3$ + NH_4Cl
- Why?
- Because + ions can only associate with negative ions and – ions can only associate with positive ions

Slide 9

What are the possible products?

- $(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)}$ →

NH_4^+ SO_4^{2-} Fe^{3+} Cl^-

- $Fe_2(SO_4)_3$ + NH_4Cl
- Are these products solids, liquids, gases, or aqueous?
- Depends! They are ionic, so they could be solids. But, they could also be aqueous. How?
- If they are water soluble solids then they are aqueous! (The reactants are in water already)
- Does it make a difference?
- You bet!!!!

Slide 10

Possible Reactions

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)} \rightarrow Fe_2(SO_4)_{3\,(aq)}$ + $NH_4Cl_{(aq)}$

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)} \rightarrow Fe_2(SO_4)_{3\,(s)}$ + $NH_4Cl_{(s)}$

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)} \rightarrow Fe_2(SO_4)_{3\,(aq)}$ + $NH_4Cl_{(s)}$

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)} \rightarrow Fe_2(SO_4)_{3\,(s)}$ + $NH_4Cl_{(aq)}$

The top one….?

No Reaction – just a bunch of ions in water!

Slide 11

Possible Reactions

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)} \rightarrow Fe_2(SO_4)_{3\,(aq)}$ + $NH_4Cl_{(aq)}$

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)} \rightarrow Fe_2(SO_4)_{3\,(s)}$ + $NH_4Cl_{(s)}$

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)} \rightarrow Fe_2(SO_4)_{3\,(aq)}$ + $NH_4Cl_{(s)}$

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)} \rightarrow Fe_2(SO_4)_{3\,(s)}$ + $NH_4Cl_{(aq)}$

These are all DIFFERENT reactions

How would you know which happens?

You need to know the solubility of the products.

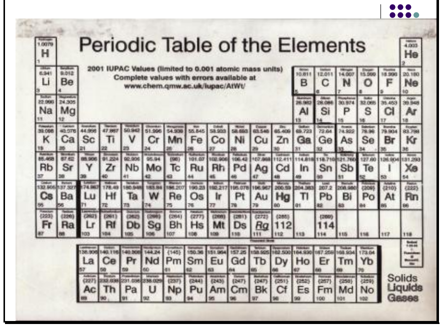
Slide 12

Predicting Solubilities

Solubility is very complicated, but there are a few simple rules that help with a large number of compounds:

1. Group 1A salts are all soluble.
2. All salts containing nitrates, ammonium, chlorate, perchlorate, and acetate are soluble.
3. All Cl, Br, I salts are soluble EXCEPT for Ag, Pb, and Hg_2^{2+} salts
4. All sulfates are soluble EXCEPT Pb, Ca, Sr, Hg_2^{2+} and Ba
5. Metal hydroxides are INSOLUBLE except for those of Group 1A and Ca, Sr, and Ba
6. All salts containg phosphate, carbonate, sulfite, and sulfide are insoluble EXCEPT for those of Group 1A and NH_4^+

Slide 13



Slide 14

Based on those rules…

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)}$ → $Fe_2(SO_4)_{3\,(aq)}$ + $NH_4Cl_{(aq)}$

$(NH_4)_2SO_{4\,(aq)}$ + $FeCl_{3\,(aq)}$ → No Reaction

Slide 15

Precipitation Reactions

This is an example of a "double-replacement reaction" – swapping cations between anions.

This is also an example of a "precipitation reaction" – forming a solid from a reaction performed in aqueous solution. (Although no precipitation reaction occurred in this instance due to solubility of the products.)

Slide 16

Another Example

$Fe_2(SO_4)_{3\,(aq)} + Ca(NO_3)_{2\,(aq)} \rightarrow$???

What's the first thing to consider?

What is the nature of the reactants?

They are ionic!

Slide 17

Another Example

$Fe_2(SO_4)_{3\,(aq)} + Ca(NO_3)_{2\,(aq)} \rightarrow$???

What's the first thing to consider?

What is the nature of the reactants?

They are ionic!

$Fe^{3+}_{(aq)} + SO_4^{2-}{}_{(aq)} + Ca^{2+}_{(aq)} + NO_3^{-}{}_{(aq)} \rightarrow$???

Slide 18

Possible Reactions

$Fe^{3+}_{(aq)} + SO_4^{2-}{}_{(aq)} + Ca^{2+}_{(aq)} + NO_3^{-}{}_{(aq)} \rightarrow$???

Double Replacement!

$Fe^{3+}_{(aq)} + SO_4^{2-}{}_{(aq)} + Ca^{2+}_{(aq)} + NO_3^{-}{}_{(aq)} \rightarrow Fe(NO_3)_3 + CaSO_4$

Next thing to consider…?

Soluble or insoluble?

Slide 19

Possible Products

$Fe^{3+}_{(aq)} + SO_4^{2-}{}_{(aq)} + Ca^{2+}_{(aq)} + NO_3^{-}{}_{(aq)} \rightarrow Fe(NO_3)_3 + CaSO_4$

$Fe(NO_3)_3$???

Soluble! (All salts containing nitrates, ammonium, chlorate, perchlorate, and acetate are soluble.)

$CaSO_4$???

Insoluble! (All sulfates are soluble EXCEPT Pb, Ca, Sr, Hg_2^{2+} and Ba)

Slide 20

Actual Reaction…

$Fe_2(SO_4)_{3\,(aq)} + Ca(NO_3)_{2\,(aq)} \rightarrow Fe(NO_3)_{3\,(aq)} + CaSO_{4\,(s)}$

- This is a precipitation reaction and a double-replacement reaction!

$Fe_2(SO_4)_{3\,(aq)} + 3\ Ca(NO_3)_{2\,(aq)} \rightarrow 2\ Fe(NO_3)_{3\,(aq)} + 3\ CaSO_{4\,(s)}$

Slide 21

Net Ionic Equation

Sometimes, to simplify the expression down to its most important elements, rather than write the full chemical equation it is distilled down to a **net ionic equation**

We ignore **spectator ions** (dissolved ions that don't change in the reaction) and write only the species that change.

Slide 22

Writing the Net Ionic Equation

$Fe_2(SO_4)_{3\,(aq)} + 3\ Ca(NO_3)_{2\,(aq)} \rightarrow$
$2\ Fe(NO_3)_{3\,(aq)} + 3\ CaSO_{4\,(s)}$

Rewrite with aqueous ionic species as ions:

$2\ Fe^{3+}_{(aq)} + 3\ SO_4^{2-}{}_{(aq)} + 3\ Ca^{2+}_{(aq)} + 6\ NO_3^-{}_{(aq)} \rightarrow 2\ Fe^{3+}_{(aq)} + 6\ NO_3^-{}_{(aq)} + 3\ CaSO_{4\,(s)}$

Cancel things that appear on both sides!

$3\ SO_4^{2-}{}_{(aq)} + 3\ Ca^{2+}_{(aq)} \rightarrow 3\ CaSO_{4\,(s)}$
$SO_4^{2-}{}_{(aq)} + Ca^{2+}_{(aq)} \rightarrow CaSO_{4\,(s)}$

Slide 23

$3\ SO_4^{2-}{}_{(aq)} + 3\ Ca^{2+}_{(aq)} \rightarrow 3\ CaSO_{4(s)}$

Advantages of the Net Ionic Equation:

1. It is complete – all changes are spelled out.
2. It is concise – only things that actually change in the reaction are shown.
3. It is the simplified recipe – if I want to make $CaSO_4$, does it matter what I use as the source of SO_4^{2-}?

Slide 24

Precipitation Reactions

It should be possible for you to recognize the possibility of a precipitation reaction:

1. Recognize the reactants are ionic
2. The ionic species are aqueous
3. Double replacement products are, therefore, possible
4. One or more of the possible products is insoluble.

Slide 25

Clicker Question #2

Write the correct balanced equation for the following reaction:

$K_2SO_{4\,(aq)} + Pb(NO_3)_{2\,(aq)} \rightarrow$

A. $K_2SO_{4\,(aq)} + 2\ Pb(NO_3)_{2\,(aq)} \rightarrow PbSO_{4\,(aq)} + 2\ KNO_{3\,(aq)}$
B. $K^+_{(aq)} + NO_3^-{}_{(aq)} \rightarrow KNO_{3\,(s)}$
C. $K_2SO_{4\,(aq)} + 2\ Pb(NO_3)_{2\,(aq)} \rightarrow PbSO_{4\,(s)} + 2\ KNO_{3\,(s)}$
D. $Pb^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow PbSO_{4\,(s)}$
E. No Reaction

Slide 26

Predicting Solubilities

Solubility is very complicated, but there are a few simple rules that help with a large number of compounds:

1. Group 1A salts are all soluble.
2. All salts containing nitrates, ammonium, chlorate, perchlorate, and acetate are soluble.
3. All Cl, Br, I salts are soluble EXCEPT for Ag, Pb, and Hg_2^{2+} salts
4. All sulfates are soluble EXCEPT Pb, Ca, Sr, Hg_2^{2+} and Ba
5. Metal hydroxides are INSOLUBLE except for those of Group 1A and Ca, Sr, and Ba
6. All salts containg phosphate, carbonate, sulfite, and sulfide are insoluble EXCEPT for those of Group 1A and NH_4^+

Slide 27

Acid-base reactions

Another type of aqueous chemical reaction is an acid/base reaction – the reaction between an **acid** and a **base**.

There are different types of acids and bases. We rely on the Bronsted-Lowry definition:

- Bronsted-Lowry acid – proton (H^+) donor
- Bronsted-Lowry base - proton (H^+) acceptor

Slide 28

Recognizing acids and bases

Acids are easiest to recognize in this system.

- Bronsted-Lowry acid - proton donor

If you are going to donate a proton, what must be true?

You must have a proton!!!

Slide 29

Which of these compounds is an acid?

HCl

Fe_2SO_4

NaOH

$NaNO_3$

H_2SO_4

Slide 30

Which of these compounds is an acid?

HCl – yep, it has a proton

Fe_2SO_4 – nope, it has no proton

NaOH – nope, it has no proton! (This is the trick question, hydroxide is a unit, there isn't a separate proton there)

$NaNO_3$ - nope, it has no proton

H_2SO_4 – yep, it has a proton – two, in fact.

Slide 31

Recognizing acids and bases

Bases are a bit harder to recognize in this system.

- Bronsted-Lowry base - proton acceptor

If you are going to accept a proton, what must be true?

You must have available electrons – H^+ is a cation, it needs electrons.

Slide 32

Strong vs. weak

Electrolytes (ionic species) in general, and acids and bases in particular, are considered to come in two types: **strong** and **weak**.

This is not what it sounds like! It has nothing to do with how powerful the solutions is.

A **strong electrolyte** is one that completely dissociates into its ions in water. A **weak electrolyte** is one that only partially dissociates into its ions in water.

Slide 33

Aqueous Acids & Bases

The key to the aqueous chemistry of acids and bases is WATER! (Huge surprise, I know! ☺)

Water, among its many interesting properties, is also **amphiprotic**. Water is both an acid and a base!

H_2O – It has a proton and it has excess electrons on the oxygen.

Slide 34

General Acid/Base Reaction

H-A + X-B → H-B + X-A

(where X and A are any chemical species, H is the proton, and B is any basic species)

Slide 35

Products of Acid/Base reactions

All acid/base reactions have the same two net products: water & a salt!

$HCl + NaOH \rightarrow H_2O + NaCl$

$H_2SO_4 + LiOH \rightarrow H_2O + LiHSO_4$

$HCl + NH_3 \rightarrow Cl^- + NH_4^+$ Where's the water?

Slide 36

The role of water

NH_3 is a weak base – ammonia.

Any weak base in water will accept a proton from water.

$NH_{3\,(aq)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

$NH_{3\,(aq)} + H_2O_{(l)} \rightarrow NH_4OH_{(aq)}$

Ammonia and ammonium hydroxide are used interchangeable.

Slide 37

The role of water

$HCl + NH_3 \rightarrow Cl^- + NH_4^+$

$NH_{3\,(aq)} + H_2O_{(l)} \rightarrow NH_4OH_{(aq)}$

$HCl + NH_4OH \rightarrow H_2O + NH_4Cl$

This is the reason for the old Earth Science myth that all bases have OH^-
They DO, as long as they are in water!

Slide 38

Gas Evolution Reactions

Just what it sounds like: reactions that create a gas as a product.

These reactions can be difficult to identify. A couple guidelines:

1. Sulfides tend to create gas products: $H_2S(g)$
2. Carbonates and bicarbonates (CO_3^{2-} and HCO_3^-) form compounds that break down into gases: H_2CO_3 (aq) breaks down into H_2O (l) and $CO_2(g)$.
3. Sulfites and bisulfites (SO_3^{2-} and HSO_3^-) form compounds that break down into gases: H_2SO_3 breaks down into $SO_2(g)$
4. Ammonium compounds (NH_4^+) can form compounds that break down into $NH_3(g)$: $NH_4OH(aq)$ breaks down into $H_2O(l)$ and $NH_3(g)$.

Slide 39

Gas Evolution Reaction - Examples

$2\ HCl(aq) + K_2S(aq) \rightarrow$

The key here is to see the sulfide. In solution, both of these compounds are ionic and break down into ions:

$2\ H^+ (aq) + 2\ Cl^- (aq) + 2\ K^+(aq) + S^{2-}(aq) \rightarrow$
So, it is still a double replacement reaction:

$2H^+ (aq) + 2Cl^- (aq) + 2\ K^+(aq) + S^{2-}(aq) \rightarrow H_2S + 2KCl$

Slide 40

Gas Evolution Reaction - Examples

2 HCl(aq) + K_2S(aq) → H_2S + 2KCl

But, like with the precipitation reactions, we need to determine what kind of compound the products are.

KCl?

Ionic, so it dissociates!

H_2S?

Actually, it is Jumping Jack Flash!
It's a Gas! Gas! Gas!

Slide 41

Gas Evolution Reaction - Examples

2 HCl(aq) + K_2S(aq) → H_2S (g) + 2KCl(aq)

Or, if you want to write it as a net ionic equation:

$2H^+$ (aq) + $2Cl^-$ (aq) + 2 K^+(aq) + S^{2-}(aq) → H_2S(g) + $2K^+$(aq) + $2Cl^-$(aq)

Cancel the "spectators"

$2H^+$ (aq) + S^{2-}(aq) → H_2S(g)

Slide 42

Gas Evolution Reaction - Examples

$2H^+$ (aq) + S^{2-}(aq) → H_2S(g)

This really shows you a whole class of gas evolution reactions. Take any acid (source of H^+) and any soluble sulfide (source of S^{2-}) and you get H_2S (g)!

Slide 43

So, in 10 seconds or less!

$HCl(aq) + Li_2S(aq) \rightarrow H_2S(g) + LiCl(aq)$

$H_2SO_4(aq) + FeS(aq) \rightarrow H_2S(g) + FeSO_4(aq)$

Slide 44

Another tricky type of reaction:

Oxidation-reduction reactions ("redox reactions") are barely chemical reactions at all.

What is a chemical reaction?

A process in which bonds are broken and made so that atoms change partners.

Slide 45

Consider 2 molecules

FeO and Fe_2O_3

Are they different?

Yes.

What's the difference?

Iron (II) oxide vs. Iron (III) oxide The Oxidation State is different.

Slide 46

What is an "oxidation state"?

The simplest way to think about an "oxidation state" is that it is the charge the atom has or could have if you separated it from the atoms it is bonded to.

Slide 47

Are you stuck with your oxidation state?

Asked a different way: If you are iron in FeO, are you stuck being Fe^{2+} forever?

In fact, you can change oxidation states as often is you like. But, there's a catch…

How do you change oxidation states?

Add or subtract electrons. Fe^{2+} has 1 more electron than Fe^{3+}

Slide 48

What does this reaction look like?

$Fe^{2+} \rightarrow Fe^{3+} + 1\ e^{-}$

Is this a "real" reaction?

Depends on what you mean by "real" and by reaction. Something changed, but no atoms were rearranged so it isn't like the other reactions we've seen before. And, you might ask, what happens to the electron?

Slide 49

This is an "electrochemical" reaction

$Fe^{2+} \rightarrow Fe^{3+} + 1\ e^-$

It's a special kind of process, part electrical and part (barely) chemical. The atom changes oxidation state and creates an electron. The electron can do useful work (power your Ipod) or chemical work (change the oxidation state of something else).

Slide 50

Electrons come, electrons go

$Fe^{2+} \rightarrow Fe^{3+} + 1\ e^-$

$Mn^{5+} + 3\ e- \rightarrow Mn^{2+}$

When electrons "go", it is called an "oxidation". When electrons "come", it is called a "reduction".

[It's easiest to remember that a "reduction" reduces the charge on the ion (oxidation state).]

Slide 51

Like acids and bases…

Oxidation and Reduction always happens simultaneously:

Oxidation half-reaction: $Fe^{2+} \rightarrow Fe^{3+} + 1\ e^-$

Reduction half-reaction: $Mn^{5+} + 3\ e- \rightarrow Mn^{2+}$

Full reaction: $3\ Fe^{2+} + Mn^{5+} \rightarrow 3\ Fe^{3+} + Mn^{2+}$

WTFDYGT?????????????

Slide 52

Chemical reactions don't have electrons

Oxidation and Reduction half-reactions balance so that no NET electrons remain

Oxidation gives you 1 e-: $Fe^{2+} \rightarrow Fe^{3+} + 1\ e^-$

Reduction needs 3: $Mn^{5+} + 3\ e- \rightarrow Mn^{2+}$

$3 \times (Fe^{2+} \rightarrow Fe^{3+} + 1\ e^-)$
$+\ \underline{Mn^{5+} + 3\ e- \rightarrow Mn^{2+}}$
$3\ Fe^{2+} + Mn^{5+} + 3e^- \rightarrow 3\ Fe^{3+} + Mn^{2+} + 3e^-$

$3\ Fe^{2+} + Mn^{5+} \rightarrow 3\ Fe^{3+} + Mn^{2+}$

Slide 53

Is it always that easy?

Of course NOT!

We'll talk about these reactions in much greater detail later. For now, we just want to recognize one when we see it. We'll figure out how to balance them after the snow melts.

Slide 54

Determining oxidation states

1. Free elements are always 0 (Cu, Fe, H_2)
2. The oxidation state of a monoatomic ion is just its charge (Cu^{2+}, Fe^{3+}, S^{2-})
3. The sum of the oxidation of all atoms in a compound or ion is the charge of the compound or ion (H_2O, NH_4^+)
4. Metals are always positive. Group 1A metals are ALWAYS +1. Group 2A metals are ALWAYS +2. (NaCl, $Ca(OH)_2$)
5. Halogens (Group VIIA) are 95% of the time -1.
6. Chalcogenides (Group VIA) are 95% of the time -2.

Slide 55

Ionic compounds

For an ion or an ionic compound, the oxidation state is easy: it's the charge on the ion.

$FeCl_3$ – Fe must be +3 because there are 3 Cl^- ions stuck to it.

MnO_2 – Mn must be +4 because there are 2 O^{2-} stuck to it.

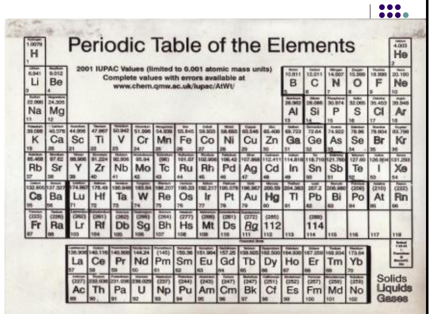
Slide 56

Covalent compounds

Here is where it is a little tricky. The atoms in a covalent compound don't have a real charge on them. BUT, they do have a potential charge if you pulled them all apart and the electrons they share get split up.

CO – carbon must be +2 since O wants to be -2
CO_2 – carbon must be +4 since there are 2 O that want to be -2 each

Slide 57



Slide 58

What's the oxidation state of the atoms?

$SrBr_2$

Br is a halogen – it must be -1 when bonded to a metal

Sr must be +2

SO_3

O is usually -2, which means S must be +6

Slide 59

What's the oxidation state of the atoms?

CO_3^{2-}

O is usually -2, which means C must be +4 since the entire molecule is -2

NO_3^-

O is usually -2, which means N must be +5 since the entire moecule is -1

Slide 60

KI

What's the oxidation state of K?

A. +1
B. -1
C. +2
D. -2
E. Cannot be determined.

Slide 61

KI

K is a Group IA metal. It is +1 ALWAYS!

I is, therefore, -1 (as it usually is for halogens) because the whole molecule must add up to ZERO.

Slide 62

MnO

What is the oxidation state of Mn?

A. +2
B. +4
C. +1
D. -1
E. -2

Slide 63

MnO

Mn is not in one of our 1st two columns. So we can't know its oxidation state directly.

HOWEVER, oxygen is Group VIA and is almost always -2.

So, if O is -2, Mn must be +2 for the whole molecule to be neutral.

Slide 64

$2\ CuO + 2\ FeO \rightarrow Fe_2O_3 + Cu_2O$

Redox reaction? You bet!

How can you tell? Two things are changing oxidation state. (There must always be an oxidation and a reduction.)

What are the oxidation states?

Slide 65

$2\ CuO + 2\ FeO \rightarrow Fe_2O_3 + Cu_2O$

CuO - Copper is
+2
How do you know?
Because O is -2

Cu_2O – Copper is…
+1
How do you know?
Because O is -2

Slide 66

As soon as you see one thing changing oxidation state…

…there is another! (Yoda 1980)

There has to be an oxidation and a reduction.

$2\ CuO + 2\ FeO \rightarrow Fe_2O_3 + Cu_2O$
FeO – Fe is +2 (because O is -2)
Fe_2O_3 – Fe is +3 (because O is -2)

Slide 67

2 CuO + 2 FeO → Fe_2O_3 + Cu_2O

Cu goes from +2 (CuO) to +1 (Cu_2O) while Fe goes from +2 (FeO) to +3 (Fe_2O_3)

So Fe is giving up an electron (going from +2 to +3) while Cu is gaining an electron (going from +2 to +1).

Slide 68

Clicker Question

Which of the following is a redox reaction:

A. $Li_{(s)} + O_{2\,(g)} \rightarrow Li_2O_{\,(s)}$
B. $Pb(NO_3)_{2\,(s)} + Na_2SO_{4\,(s)} \rightarrow PbSO_{4\,(s)} + 2\,NaNO_{3\,(s)}$
C. $Mg_{(s)} + Br_{2\,(l)} \rightarrow MgBr_{2\,(s)}$
D. A and C
E. A, B, and C.