**Thermochemistry Reference Sheet**

 **Q = mC∆T**

|  |  |  |
| --- | --- | --- |
| **Variable** | **Meaning** | **Common Units** |
| Q | Energy being transferred | Joules, kJ, calorie, Calorie |
| m | Mass of sample | grams, kilograms |
| C | Specific Heat | J/g°C, kJ/g°C |
| ∆T | Change in Temperature | °C |

Note:

* Temperature is in CELSIUS not Kelvins for this topic!
* Q can be positive or negative, energy absorbed or energy released
* 1 kJ = 1000 J
* 1 calorie = 4.18 J
* 1 Calorie = 1 kcal = 1000 calories
* Specific Heat values can vary a little bit from source to source. That is ok. Depends on the exact type that was measured. Feel free to add other specific heats you come across to this list!

**Common Specific Heat Values**

|  |  |  |
| --- | --- | --- |
| **Substance** | **J/g°C** | **cal/g°C** |
| Aluminum | 0.90 | 0.23 |
| Copper | 0.39 | 0.093 |
| Ethanol | 2.44 | 0.58 |
| Glass | 0.50 | 0.12 |
| Gold  | 0.13 | 0.031 |
| Graphite | 0.71 | 0.17 |
| Ice | 2.09 | 0.50 |
| Iron | 0.45 | 0.11 |
| Lead | 0.13 | 0.031 |
| Mercury | 0.14 | 0.033 |
| Silver | 0.24 | 0.057 |
| Steam | 1.87 | 0.45 |
| Water (liq) | 4.18 | 1.00 |
| Wood | 1.8 | 0.42 |

**Calorimetry**

* The fundamental idea behind calorimetry is that sometimes we cannot measure things directly, we need to measure them indirectly.
* If we know that the energy INTO a system equals the energy OUT of a system then we can set up an equation where we know the amount of energy being released is equal in number but opposite in sign to the energy being absorbed.
* Remember that the negative sign just means OPPOSITE sign to the other Q value, not necessarily an actual negative value.
* Remember that Tfinal substance 1 and Tfinal substance 2 are the same! They reach the same ending temperature

**For a metal object dropped into water, i.e., a typical calorimetry problem**

Therefore…

Same as…

**R-39**

**Heating and Cooling Curves**

This is an example of water – can be done with any substance but the temperature values will be different. Also please note that the slope and length of lines are not drawn to scale. It is traditional to just draw the heating curve. A cooling curve would just be the opposite direction!

Cooling Curve

100°C

0°C

Q = mCsteam∆T

|  |
| --- |
| **Latent Heat Values for H2O**Lfus = 334 Lvap = 2260 *Freezing point**Boiling point*Q = mLvaporizationQ = mLfusionQ = mCice∆TQ = mCliquid∆THeating Gas 🡪Cooling Gas 🡨Heating Liquid 🡪Cooling Liquid 🡨Heating Solid 🡪Cooling Solid 🡨Vaporizing 🡪Condensing 🡨Melting 🡪Freezing 🡨 |

**Calorimetry Problems Involving Phase Changes**

If a calorimetry problem involves phase changes you need to be careful to use the right equations in your calorimetry calculations – you may not use Q = mC∆T on both sides, you may need to use Q = mL for part of it!

*Example:* 75g of ice (0°C) are dropped into 300g of water (25°C). What is the final temperature of the mixture when the ice
 is finished melting?

You would need to use Q=mL for the ice melting values and Q = mC∆T for the water values

mLfus = - (mCliq∆T)
(75g)(334 ) = -(300g)(4.18)(Tf - 25°C)

**Energy of Formation**

∆H°*rxn* = Σn∆H°*formation*(Products) - Σn∆H°*formation* (Reactants)

*n = number of moles of each product or reactant*

**Hess’s Law**

|  |  |
| --- | --- |
| Reverse a reaction | Change the algebraic sign on ∆H° |
| Multiply a reaction by some factor | Multiply the algebraic sign on ∆H° |

The ∆H°rxn for a reaction can be found by adding together other reaction steps that add up to the reaction of interest.

**Bond Energies**

|  |  |  |  |
| --- | --- | --- | --- |
| **Breaking Bonds**  | Absorbs/Takes Energy | ENDOthermic | ∆H° = + |
| **Making Bonds** | Releases Energy | EXOthermic | ∆H° = - |

